Acid Bases and Salts

Acids, Bases and Salt

The word acid is derived from the Latin word acidus meaning sour to taste. On the other hand, bases are substances which are bitter in taste and change the colour of red litmus to blue. We all are familiar with the sour taste of lemon, curd, vinegar, unripened grapes, green mangoes, tomatoes, etc. They have sour taste because of presence of acids in them. On the other hand, some substances such as washing soda, lime (chuna) and materials like soap which have soapy touch and a bitter taste. These substances are either bases or contain bases.

Indicator

"Indicator is a chemical compound which is added to the solution in very small amount to detect their acidic or basis nature."

As they show colour change in acidic and basic medium, they are also called acid-base indicators. In other words,

"An acid-base indicator is that substance which possesses one colour in acidic medium and a different colour in alkaline medium."

- **Litmus**
  - Litmus is a mixture of water soluble dyes extracted from Lichens, especially Rocella tinctorina. It is the most commonly used indicator to detect acids and bases.
  - Litmus solution is a purple coloured dye and most commonly used in the laboratory.
  - In the neutral solution, it has purple colour. In the acidic solution, it turns red whereas in the basic solution, it turns blue.

In nutshell:

(i) An acid turns blue litmus into red

(ii) A base or an alkali turns red litmus into blue.

Acidic substances turning blue litmus solution into red	Basic substances turning red litmus solution into blue
Vinegar	Baking soda solution
Lemon juice, Orange juice	Washing soda solution
Juice or unripe mangoes	Bitter gourd (karela) extract
Tamarind (imli) juice	Cucumber (kheera) extract

Note: Litmus solution itself is neither acidic nor basic. It is neutral and has a purple colour. To use it as an indicator, it is made acidic or alkaline to give it red or blue colour.

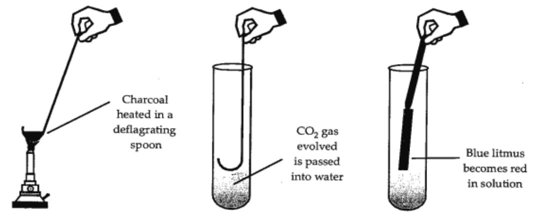
- **Some Other Indicators**
  - **Red cabbage juice:** It is purple in colour in neutral medium and turns red or pink in the acidic medium. In the basic or alkaline medium, its colour changes to green.
  - Turmeric juice: It is yellow in colour and remains as such both in the neutral and acidic medium. In the basic medium, its colour becomes reddish or deep brown.
  - Phenolphthalein: It is also an organic dye. In neutral or acidic solution, it is colourless while in the basic solution, the colour of indicator changes to pink.
  - Methyl orange: Methyl orange is an orange coloured dye and keeps this colour in the neutral medium. In the acidic medium, the colour of indicator becomes red and in the basic medium it changes to yellow.

ACTIVITY CORNER

To show that an aqueous solution of CO_2 is acidic

- Take a piece of charcoal in a deflagrating spoon.
- When it starts burning, immediately lower in a tube containing water. After some time, transfer the solution to another tube.
- Now lower a blue litmus strip into it. What will you observe?
- **Observation:** The litmus paper will acquire a red colour.
- **Conclusion:** This shows that the solution is acidic. Actually, carbon present in coke upon burning has been oxidized to $CO_{2(g)}$. The gas upon passing through water has formed carbonic acid (H_2CO_3). The acid has turned blue litmus red.

$$CO_{2(g)} + H_2O_{(l)} \longrightarrow H_2CO_{3(aq)}$$



ILLUSTRATION

1. What will you observe when:

(i) Red litmus paper is introduced into a solution of sodium sulphate.

(ii) Methyl orange is added to dilute hydrochloric acid.

(iii) A drop of phenolphthalein is added to the solution of lime water.

(iv) Blue litmus is introduced into a solution of ferric chloride.

Sol. (i) It will not undergo colour change because the solution of sodium sulphate (Na_2SO_4) in water is almost neutral. Both NaOH and H_2SO_4 expected to be formed in solution have nearly same strength.

(ii) In the acidic solution, the colour of methyl orange will change to red.

(iii) Lime water contains traces of calcium hydroxide, $Ca(OH)_2$. It is therefore, basic in nature. The colour of phenolphthalein will become pink.

(iv) Ferric chloride ($FeCl_3$) solution on reacting with water will form ferric hydroxide and hydrochloric acid. Since the acid is strong, the blue litmus will change to red.

$$FeCl_{3(aq)} + 3H_2O_{(l)} \rightarrow \underset{(Weak)}{Fe(OH)_{3(s)}} + \underset{(Strong)}{3HCl_{(aq)}}$$

2. What is the colour of phenolphthalein in basic medium and that of methyl orange in acidic medium?

Sol. Phenolphthalein gives pink colour in basic medium whereas methyl orange gives red colour in acidic medium.

- **Olfactory Indicators**

  Indicators giving different odours in acidic and basic medium are called olfactory indicators.

  Some natural olfactory indicators are:

- **Onion odoured cloth strips:** Onions have a characteristic ammoniacal smell which is retained in acidic medium but is destroyed in basic medium.
- **Vanilla essence:** Vanilla has characteristic sweet smell which is retained in acidic medium but is destroyed in basic medium.

ACTIVITY CORNER

To test acids and basses using olfactory indicators.

- Take two clean cloth strips. Smell the strips to check that they have no smell.
- Take a small onion and chop it finely. Place the chopped onion along with strips of cloth in a plastic bag. Tie up the bag tightly and them place in the fridge for 12 hours.
- Take out the strips and check their odour. You will observe that they smell like onion, i.e., a stingy smell.
- Place the strips on stainless steel plates separately. Pour few drops of dilute hydrochloric acid on one strip and few drops of sodium hydroxide solution on the other strip.
- Rinse both cloth strips with fresh water. Again check their odour.
- **Observation:** You will observe that the strip treated with hydrochloric acid no longer smell, but the strip rinsed with sodium hydroxide solution still smells like onion.
- **Conclusion:** It is because the onion essence is basic in nature, so it is neutralized by hydrochloric acid.

ILLUSTRATION

3. What is the name given to the indicators giving different odours in acidic and basic medium?

Sol. Olfactory indicators.

4. What is an acid-base indicator? Give two examples of synthetic acid-base indicators.

Sol. An acid-bases indicators is a substance which has one colour in the acidic medium and a different colour in the basic medium. Two examples of synthetic indicators are phenolphthalein and methyl orange.

5. What are olfactory indicators? Name two substance which can be used as olfactory indicators?

Sol. Olfactory indicators are those substances which have one odour in acidic medium and a different odour in basic medium. For example, onions, vanilla essence and clove oil can act as olfactory indicators.

Acids

- Liebig defined acids as "a compound which contains one or more hydrogen atoms replaceable partially or completely by a metal or a positive radical to produce a salt."

  Acids are substances which have sour taste and turn blue litmus solution to red. Some colour changes shown by various indicators with acids are:

- change the colour of blue litmus solution to red.
- give red colour with methyl orange.

- give colourless solution with phenolphthalein.
- no reaction with red litmus solution.

Basicity or Protocity of an Acid

The number of replaceable hydrogen atoms in one molecule of an acid or the number of H^+ ions produced in aqueous solution by one molecule of an acid is its basicity.

Acids are of three type: 1. Monoprotic acids, 2. Diprotic acids, 3. Polyprotic acids

The basicity of some acids are given in the following lines:

Acid	Basicity
HF, HCl, HBr, HI	1
HOCl, HOBr, HOI	1
$HClO_2$ (chlorous acid)	1
$HClO_3$ (chloric acid)	1
$HBrO_3$ (Bromic acid)	1
HIO_3 (Idic acid)	1
$HClO_4$ (Perchloric acid)	1
HIO_3 (Periodic acid)	1
HNO_3 (Nitric acid)	1
H_3PO_2 (Hypophosphorous acid)	1
H_3BO_3 (Boric acid)	1
H_2CO_3 (Carbonic acid)	2
H_2SO_3 (Sulphurous acid)	2
H_2PO_3 (Orthophosphorous acid)	2
H_2SO_4 (Sulphuric acid)	2
H_3PO_4 (Phosphuric acid)	3
$H_4P_2O_7$ (Pyrophosphoric acid)	4

The polybasic acids undergo complete ionization in several steps, i.e., one H^+ ion is separated from the molecule one at a time. Thus, the number of steps of complete ionization is equal to be the basicity of the acid.

- **Monoprotic Acids (Mono-basic Acids)**

  Acid which provides one proton in aqueous solution is known as monoprotic acid. Example: HCl (hydrochloric acid), HNO_3 (nitric acid), CH_3COOH (acetic acid), HPO_3 (metaphosphoric acid), H_3PO_2 (hypophosphorous acid).

$$HNO_{3(aq)} \rightarrow H^+_{(aq)} + NO^-_{3(aq)}$$

$$CH_3COOH_{(aq)} \rightleftharpoons CH_3COO^-_{(aq)} + H^+_{(aq)}$$

$$HCl_{(aq)} \xrightarrow{Water} H^+_{(aq)} + Cl^-_{(aq)}$$

- **Diprotic Acids (Dibasic Acids)**

  Acids which can donate two protons in aqueous solution are known as diprotic acids. Example: H_2SO_4 (sulphuric acid), $H_2C_2O_4$ (oxalic acid), H_3PO_3 (phosphorous acid), H_2CO_3 (carbonic acid), H_2S (hydrogen sulphide), H_2SO_3 (sulphurous acid), H_2CrO_4 (chromic acid).

$$\underset{\text{Sulphric acid}}{H_2SO_{4(aq)}} \rightarrow 2H^+_{(aq)} + SO_4{}^{2-}{}_{(aq)}$$

$$\underset{\text{Oxalic acid}}{H_2C_2O_{4(aq)}} \rightleftharpoons 2H^+_{(aq)} + C_2O_4{}^{2-}{}_{(aq)}$$

$$\underset{\text{Carbonic acid}}{H_2CO_{3(aq)}} \rightleftharpoons 2H^+_{(aq)} + CO_3{}^{2-}{}_{(aq)}$$

$$\underset{\text{Hydrogen sulphide}}{H_2S_{(aq)}} \rightleftharpoons 2H^+_{(aq)} + S^{2-}{}_{(aq)}$$

- **Triprotic Acids (Tribasic Acids)**

  Acids which can donate three protons in aqueous solution are known as triprotic acids. Example: H_3PO_4 (orthophosphoric acid).

  These acids dissociate as: $H_3A \rightleftharpoons 3H^+ + A^{3-}$

  . i.e., H_3PO_4, etc. H_3PO_3 and H_3PO_2 must also be tribasic but actually they are dibasic and monobasic acids respectively.

$$\underset{\text{Phosphorous acid}}{H_3PO_3} \rightleftharpoons 2H^+ + HPO_3^{2-}$$

$$\underset{\text{Metaphosphorous acid}}{H_3PO_2} \rightleftharpoons H^+ + H_2PO_2^-$$

- **General Characteristics of Acids**
  - Sour taste: As already discussed, almost all acidic substances have a sour taste.
  - Action on litmus paper: Acids turn blue litmus solution red.
  - Action on methyl orange: Methyl orange turns red when 1-2 drops of its solution are added to the solution of an acidic substance.
  - Corrosive nature: Most of the acids are corrosive in nature. They produce a burning sensation on the skin and holes in the clothes on which they fall.

  They also attack metal structure and stone work. Hence, they are never stored in metal containers. They are always stored in containers made of glass or ceramics as they are not attacked by the acids.

ACTIVITY CORNER

To test the given samples with the help of red litmus solution, blue litmus solution, phenolphthalein and methyl orange indicators.

- **Requirements:** HCl, H_2SO_4 HNO_3, CH_3COOH, NaOH, $Ca(OH)_2$, KOH, $Mg(OH)_2$, NH_4OH, red litmus solution, blue litmus solution, phenolphthalein, methyl orange indicator and test tubes.

- **Procedure :**

1. Take each of the above solutions in separate test tubes.

2. Test the nature of these solutions by adding a drop of red litmus solution, blue litmus solution, phenolphthalein and methyl orange indicators.

3. Observe the change in colour and record your observations in observation table.

Sample solution	Red litmus solution	Blue litmus solution	Phenolphthalein indicator	Methyl orange indicator
HCl	No colour change	Red	Colourless	Pink
H_2SO_4	No colour change	Red	Colourless	Pink
HNO_3	No colour change	Red	Colourless	Pink
CH_3COOH	No colour change	Red	Colourless	Pink
NaOH	Blue	No colour change	Pink	Yellow
$Ca(OH)_2$	Blue	No colour change	Pink	Yellow
KOH	Blue	No colour change	Pink	Yellow
$Mg(OH)_2$	Blue	No colour change	Pink	Yellow
NH_4OH	Blue	No colour change	Pink (becomes colourless after sometimes)	Yellow (becomes colouress after sometimes)

- **Conclusion:** Adds turn blue litmus red but have no effect on red litmus. 'Bases rum red litmus blue but have no effect on blue litmus. Phenolphthalein is colourless in acidic medium and turns pink in basic medium. Methyl orange is yellow in basic medium and red in acidic medium.

Following table enlists some common organic acids and their natural sources.

S. No.	Name of the acid	Natural source of the acid
1.	Acetic acid	Vinegar
2.	Citric acid	Orange, lemon
3.	Tartaric acid	Tamarind, grapes
4.	Oxalic acid	Tomato
5.	Lactic acid	Sour milk (curd)
6.	Mechanoic acid	Ant sting, bees sting, hairs of nettle plants
7.	Oleic acid	Olive oil
8.	Hydrochloric acid	Gastric juice
9.	Malic acid	Apples (green)
10.	Stearic acid	Fats
11.	Butyric acid	Rancid butter
12.	Uric acid	Urine
13.	Ascorbic acid (Vitamin C)	Fresh fruits and vegetables
14.	Tannic acid	Bark, wood(of tree) and tea
15.	Folic acid	Green leafy vegetables

Chemical Properties of Adds

- Reactions With Metals

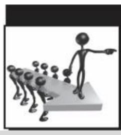
Reactive metals displace hydrogen from dilute acids which is evolved as hydrogen gas and metals salts are formed. Those metals which are more reactive than hydrogen i.e., come before hydrogen in reactivity series can displace hydrogen.

Metal + Dilute acid → Metal salt + Hydrogen

$Zn_{(s)} + H_2SO_{4\ (dil.)} \rightarrow ZnSO_{4(aq)} + H_{2(g)}$

$Mg_{(s)} + 2HCl_{(dil.)} \rightarrow MgCl_{4(aq)} + H_{2(g)}$

$Fe_{(s)} + H_2SO_{4\ (dil.)} \rightarrow FeSO_{4(aq)} + H_{2(g)}$



Do You Know

Metals which can displace hydrogen from dilute acids are known as active metals e.g., Na, K, Zn, Fe, Ca, Mg, etc. The active metals which lie above hydrogen in the activity series are electropositive in nature. Their atoms lose electrons to form positive ions and these are accepted by H^+ ions of the acid. As a result, $H_{2(g)}$ is evolved For example:

$$Zn_{(s)} \rightarrow Zn^{2+}_{(aq)} + 2e^-$$

$$2H^+_{(aq)} + SO_4{}^{2-}_{(aq)} + 2e^- \rightarrow H_{2(g)} + SO_4{}^{2-}_{(aq)}$$

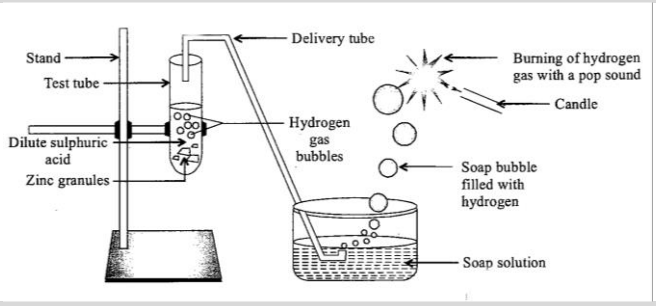
$$Zn_{(s)} + \underset{(Acid)}{2H^+_{(aq)}} \rightarrow Zn^{2+}_{(aq)} + H_{2(g)}$$

This reaction is a redox reaction in which $Zn_{(s)}$ has been oxidized to $Zn^{2+}_{(aq)}$ ions by losing electrons while $H^+_{(aq)}$ ions have been reduced to $H_{2(g)}$ by accepting them.

ACTIVITY CORNER

To show that hydrogen gas is produced when metals react with dilute acids.

- Requirements: Zinc granules, dilute H_2SO_4, test tube, delivery tube, trough, stand and soap solution.



- Set up the apparatus as shown in the figure.
- Take about 5 ml of dilute H_2SO_4 in a test tube and add a few pieces of zinc granules to it.
- Some bubbles will form in the test tube, and the size of zinc granules will decreases. This is due to the reaction of zinc with dilute H_2SO_4 to liberate hydrogen gas.
- Pass this hydrogen gas through a delivery tube into a soap solution. We will observe that soap bubbles filled with hydrogen gas are produced.
- Now bring a lighted candle near a gas-filled bubble. A popping sound is produced due to the burning of hydrogen gas.
- **Observation :** Zinc reacts with dilute H_2SO_4 to give H_2 gas
  $Zn + H_2SO_4 \rightarrow ZnSO_4 + H_2$
- **Conclusion:** All acids (minerals or organic) on reaction with metals produce hydrogen gas. But, metals react with acids at different speeds. For example, metals like sodium, potassium and magnesium reacts rapidly while iron, zinc and aluminium react slowly to evolve hydrogen gas. Also, some metals like copper, silver and gold do not react with acids.

Do You Know

In the laboratory, for the preparation of hydrogen from granular zinc, the following acids (i) conc. H_2SO_4, (ii) Conc. HCl, (iii) Nitric acid cannot be used.

(a) Conc. H_2SO_4 is not used because it act as strong oxidizing agent. It react with nascent hydrogen to produce SO_2 i.e., a part of the acid gets reduced to sulphur dioxide.

$H_2SO_{4(Conc.)} + H_2 \rightarrow 2H_2O + SO_2$

(b) Hydrogen liberated by the action of cone. HCl on zinc will be impure as it will contain fumes of volatile HCl. Moreover, $ZnCl_2$ formed is insoluble in cone. HCl. It will form a coating on zinc and the reaction stops after sometime.

$Zn + 2HCl \rightarrow ZnCl_2 + H_2$

(c) HNO_3 plays a double role. It acts as an acid and as a strong oxidizing agent. The hydrogen (nascent) first formed reduces the nitric acid into various oxides.

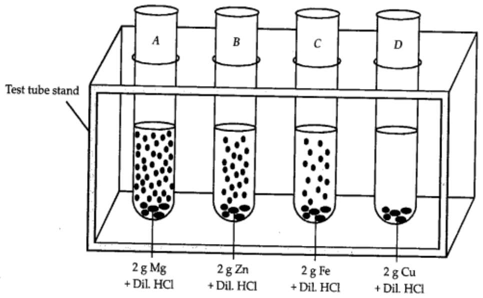
$2HNO_3 + 2H \rightarrow 2NO_2 + 2H_2O$.

- Mn and Mg react with dil. HNO_3 to evolve hydrogen.
- Most of the metals with the exception of noble metal like gold and platinum are attacked by nitric acid. Nitric acid plays a double role in the action of metals, i.e., it acts as an acid as well as an oxidizing agent.
  Metal + $HNO_3 \rightarrow$ Nitrate + H
  $2HNO_3 + 2H \rightarrow 2NO_2 + 2H_2O$
  $2HNO_3 + 6H \rightarrow 2NO + 4H_2O$
  $2HNO_3 + 10H \rightarrow N_2 + 6H_2O$

ACTIVITY CORNER

To test the reactivity of metals with acid by comparing the evolution of hydrogen

- Take four test tubes and label them as A, B, C and D. Place them in a test tube stand. Put small pieces of magnesium, zinc, iron and copper (same mass, say 2 g each) in the tubes A, B, C, D respectively. Now, pour equal amounts of dilute hydrochloric acid (say 5 mL) in each of the four test tubes. The following reactions will take place:



Reaction of different metals with the same acid

$$\underset{\text{Magnesium}}{Mg} + \underset{\text{Hydrochloric acid}}{2HCl} \longrightarrow \underset{\text{Magnesium chloride}}{MgCl_2} + \underset{\text{Hydrogen}}{H_2}$$

$$\underset{Zinc}{Zn} + \underset{\text{Hydrochloric acid}}{2HCl} \longrightarrow \underset{\textit{Zinc}\text{ chloride}}{ZnCl_2} + \underset{\text{Hydrogen}}{H_2}$$

$$\underset{\text{Iron}}{Fe} + \underset{\text{Hydrochloric acid}}{2HCl} \longrightarrow \underset{\text{Iron (II) chloride}}{FeCl_2} + \underset{\text{Hydrogen}}{H_2}$$

$$\underset{\text{Copper}}{Cu} + \underset{\text{Hydrochloric acid}}{2HCl} \longrightarrow \text{No recation}$$

- **Observation:** The rate of evolution of hydrogen gas from each of the test tubes. It is observed that at room temperature.
- Magnesium reacts most vigorously.
- Zinc reacts less vigorously than magnesium.
- Iron reacts slowly.
- Copper does not react at all.

  Hence, the order of reactivity is
  Mg > Zn > Fe > Cu.
- **Conclusion:** All metals do not react with the same acid with the same vigour.

- **Reaction with Metal Oxides**

  Acids react with metal oxides to form salt and water. The reactions are generally endoergic and require heat. As metal oxide react with acids to produce neutral, salts they are called basic oxides.

  Metal oxide + Acid → Salt + water

  $CaO(g) + 2HCl_{(aq)} \rightarrow CaCl_{2(aq)} + H_2O_{(i)}$

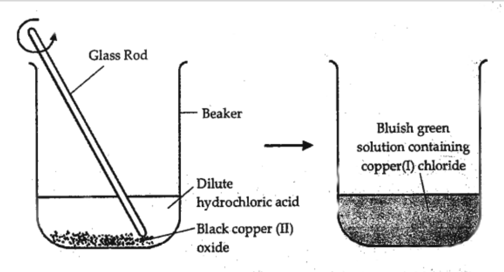
  $MgO_{(s)} + H_2SO_{4(aq)} \longrightarrow MgSO_{4(aq)} + H_2O_{(i)}$

ACTIVITY CORNER

Reaction of metal oxide with hydrochloric acid

- Take a pinch of cupric oxide in a test tube. Then add dilute HCl solution dropwise to the test tube. Gently shake the test tube if required.
- **Observation:** The black colour of cupric oxide slowly disappears and the bluish green solution is formed.
- **Conclusion:** Cupric oxide reacts with HCl to produce cupric chloride and water. This result in loss of black colour of cupric oxide during reaction and cupric chloride formed dissolve in water to produce a bluish green solution.

$$\underset{(Black)}{CuO_{(s)}} + 2HCl_{(dil.)} \longrightarrow \underset{(Bluish\ green)}{CuCl_{2(aq)}} + H_2O_{(l)}$$



Reaction between copper (II) oxide (black) with nydrochloric acid to form copper (II) chloride (bluish green)

- **Reaction with Metal Carbonates and Metal Hydrogen Carbonates (Metal Bicarbonates)**
- Acids react with metal carbonates and metal hydrogen carbonates producing metal salts and $CO_{2(g)}$ is evolved during reaction.

$$\underset{\textit{Calcium carbonate}}{CaCO_{3(s)}} + 2HCl_{(dil.)} \rightarrow \underset{\textit{Calcium chloride}}{CaCI_{2(aq)}} + H_2O_{(aq)} + CO_{2(g)}$$

$$\underset{\textit{Sodium biacarbonate}}{2NaHCO_{3(s)}} + H_2SO_{4(aq.)} \rightarrow \underset{\textit{Sodium sulphate}}{Na_2SO_{4(aq)}} + H_2O_{(aq)} + CO_{2(g)}$$

- All metal carbonates and hydrogen carbonates react with the acids to give the corresponding salt, carbon dioxide gas and water. Hence, we can write the reaction in general as.

  Metal carbonates or hydrogen carbonate + Acid → Salt + CO_2 + H_2O

DO YOU KNOW

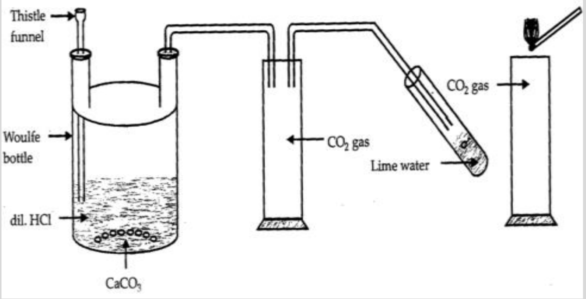
- The egg shells, chalk and marble also contains calcium carbonate as the main constituent. Hence, they react with acid as shown above.
- If a person is suffering from the problem of acidity then he is advise to take baking soda (sodium bicarbonate). Acids and $NaHCO_3$ reacts and person gets relief.

ACTIVITY CORNER

To show that acid reacts with metal carbonate to liberate carbon dioxide.

- **Requirements:** $CaCO_3$ (marble chips), Woulfe bottle, thistle funnel, dil. HCl, gas jar, match box, delivery tube bent at two right angles, limewater.
- **Procedure:**
- Take marble chips in a Woulfe-bottle.

- Set the apparatus as shown in diagram.
- Add dil. HCl with the help of thistle funnel.
- Collect the gas in gas jar by upward displacement of air.
- Bring a burning matchstick near the gas jar and record your observations.
- Pass the gas evolved through lime water and note down your observations.



Reaction of calcium carbonate with dilute hydrochloric acid liberate carbon dioxide gas which turns limewater milky and extinguishes burning matchstick

- **Observation:** The burning matchstick will get extinguished because carbon dioxide is neither combustible nor supporter of combustion. Lime water will turn milky due to formation of insoluble calcium carbonate. It can be used as test for CO_2.
- **Chemical Reactions :**

$$\underset{\substack{\text{Marble}\\\text{(Calcium carbonate)}}}{CaCO_{3(s)}} + \underset{\substack{\text{Hydrochloric}\\\text{acid}}}{2HCl_{(dil.)}} \rightarrow \underset{\substack{\text{(White ppt.)}\\\text{Calcium chloride}}}{CaCl_{2(aq.)}} + \underset{\text{Water}}{H_2O_{(l)}} + \underset{\substack{\text{Carbon}\\\text{dioxdie}}}{CO_{2(g)}}$$

$$\underset{\text{Lime water}}{Ca(OH)_{2(s)}} + \underset{\text{Carbon dioxide}}{CO_{2(g)}} \rightarrow \underset{\substack{\text{(White ppt.)}\\\text{Calcium chloride}}}{CaCO_{3(s)}} + \underset{\text{Water}}{H_2O_{(l)}}$$

- **Conclusion:** Metal carbonate reacts with dilute acid to liberate carbon dioxide. All metal carbonates and hydrogen carbonates react with acids to form corresponding salts, water and carbon dioxide.

Let us take more reactions of metal carbonates with dilute acids.

$$\underset{\substack{\text{Magnesium}\\\text{carbonate}}}{MgCO_{3(s)}} + \underset{\substack{\text{Hydrochloric}\\\text{acid}}}{2HCl_{(dil.)}} \rightarrow \underset{\substack{\text{Magnesium}\\\text{chloride}}}{MgCl_{2(aq.)}} + H_2O_{(l)} + CO_{2(g)}$$

$$\underset{\substack{\text{Zinc}\\\text{carbonate}}}{ZnCO_{3(s)}} + 2HCl_{(dil.)} \longrightarrow \underset{\substack{\text{zinc}\\\text{chloride}}}{ZnCl_{2(aq.)}} + H_2O_{(l)} + CO_{2(g)}$$

$$\underset{\substack{\text{Sodium}\\\text{carbonate}}}{NaCO_{3(s)}} + 2HCl_{(dil.)} \longrightarrow \underset{\substack{\text{Sodium}\\\text{chloride}}}{2NaCl_{2(s)}} + H_2O_{(l)} + CO_{2(g)}$$

- Metal hydrogen carbonates also react with dil. acids to liberate carbon dioxide gas.

$$\underset{\substack{\text{Sodium}\\\text{hydrogen carbonate}}}{NaHCO_{3(s)}} + HCl_{(dil.)} \rightarrow NaCl_{(aq.)} + H_2O_{(l)} + CO_{2(g)}$$

$$\underset{\substack{\text{Sodium}\\\text{hydrogen carbonate}}}{2NaHCO_{3(s)}} + H_2SO_{4(aq.)} \rightarrow \underset{\text{Sodium sulphate}}{NaSO_{4(aq.)}} + H_2O_{(aq)} + CO_{2(g)}$$

$$\underset{\substack{\text{Calcium}\\\text{hydrogen carbonate}}}{Ca(HCO_3)_{2(s)}} + 2HCl_{(dil.)} \longrightarrow CaCl_{2(aq.)} + 2H_2O_{(l)} + 2CO_{2(g)}$$

Only alkali metals and alkaline earth metals are known to form bicarbonates. Other metals do not form bicarbonate salts.

- **Reaction of Acids with Bases**

  Acids and bases react to produce a salt and water. In the reaction their acidity and basicity is destroyed and such reactions are called neutralization reactions. In general, a neutralization may be represented as:

  Acid + Base → Salt + water

  Sodium hydroxide react with hydrochloric acid to produce sodium choride (salt) and water.

  $HCl_{(aq)} + NaOH_{(aq)} \rightarrow NaCl_{(aq)} + H_2O_{(l)}$

  Some other examples of neutralization reaction are:

$$\underset{\text{Nitric acid}}{HNO_{3(aq)}} + \underset{\text{Sodium hydroxide}}{NaOH_{(aq)}} \longrightarrow \underset{\text{Sodium nitrate}}{NaNO_{3(aq)}} + H_2O_{(l)}$$

$$\underset{\text{Sulphuric acid}}{H_2SO_{4(aq)}} + \underset{\text{Sodium hydroxide}}{2NaOH_{(aq)}} \longrightarrow \underset{\text{Sodium sulphate}}{Na_2SO_{4(aq)}} + 2H_2O_{(l)}$$

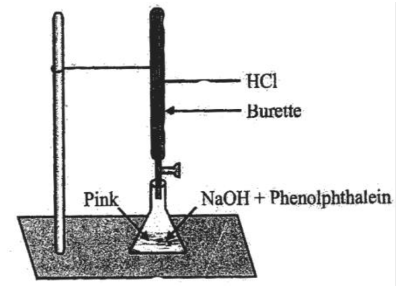
$$\underset{\text{Carbonic acid}}{H_2CO_{3(aq)}} + \underset{\text{Potassium hydroxide}}{2KOH_{(aq)}} \longrightarrow \underset{\text{Potassium carbonate}}{K_2CO_{3(aq)}} + 2H_2O_{(l)}$$

$$\underset{\text{Carbonic acid}}{H_2SO_{3(aq)}} + \underset{\text{Ammonium hydroxide}}{2NH_4OH_{(aq)}} \longrightarrow \underset{\text{Ammonium sulphate}}{(NH)_4SO_4} + 2H_2O_{(l)}$$

ACTIVITY CORNER

To demonstrate neutralization reaction between acid and base.

- **Requirements:** NaOH (1M), HCl (1M) phenolphthalein, burette, pipette, titration flask.
- **Procedure :**
- Take 10 mL of NaOH in titration flask with the help of pipette.
- Add a drop of phenolphthalein and observe the colour of the solution.
- Take HCl in a burette and note down the initial reading.
- Open the stop cock of burette and add HC1 into titration flask slowly drop wise till the pink colour due to NaOH becomes colourless.
- Note down the final reading.



Neutralization reaction between NaOH and HCL

- **Observation:** The pink colour of phenolphthalein changes to colourless.
- **Chemical Reaction:** NaOH + HCl $\longrightarrow$ NaCl+H_2O
- **Explanation:** Phenolphthalein gives pink colour with NaOH (base). When acid is added, NaOH reacts with acid to form NaCl and H_2O. When whole of NaOH reacts with acid, it becomes colourless because neutralization reaction is complete.
- **Conclusion:** NaOH (base) reacts with HCl (acid) to form (salt) and H_2O. It is an example of neutralization reaction.

- **Reaction of Acids with Sulphites and Bisulphites (Hydrogen Sulphites)**

  Dilute acids react with sulphites and hydrogen sulphites (bisulphites) with evolution of sulphur dioxide gas to give salts.

  Sulphite + Acid $\longrightarrow$ Salt + Water + Sulphur dioxide

  Some examples are :

$$\underset{\text{Calcium sulphite}}{CaSO_{3(s)}} + \underset{\text{hydrochloric acid}}{2HCl_{(aq)}} \rightarrow \underset{\text{Calcium chloride}}{CaCl_{2(aq)}} + H_2O_{(l)} + \underset{\text{sulphur dioxide}}{SO_{2(g)}}$$

$$\underset{\text{Sodium sulphite}}{NaSO_{3(s)}} + \underset{\text{hydrochloric acid}}{H_2SO_{4(aq)}} \rightarrow \underset{\text{Sodium sulphate}}{Na_2SO_{4(aq)}} + H_2O_{(l)} + \underset{\text{sulphur dioxide}}{SO_{2(g)}}$$

$$\underset{\text{Sodium bisulphite}}{NaHSO_{3(s)}} + \underset{\text{hydrochloric acid}}{HCl_{(aq)}} \rightarrow \underset{\text{Sodium chloride}}{NaCl_{(aq)}} + H_2O_{(l)} + \underset{\text{sulphur dioxide}}{SO_{2(g)}}$$

$$\underset{\text{Potassium bisulphite}}{KHSO_{3(s)}} + \underset{\text{sulphuric acid}}{H_2SO_{4(aq)}} \rightarrow \underset{\text{Potassium hydrogen sulphate}}{KHSO_{4(aq)}} + H_2O_{(l)} + \underset{\text{sulphur dioxide}}{SO_{2(g)}}$$

- **Reaction of Acids with Metal Sulphides**

  When metal sulphides react with acid, metal salt is formed with evolution of hydrogen sulphide gas characterised by its rotten egg like smell.

  Metal sulphide + Acid $\longrightarrow$ Salt + Hydrogen sulphide

$$\underset{\text{Ferrous sulphite}}{FeS_{(s)}} + \underset{\text{sulphuric acid}}{H_2SO_{4(aq)}} \rightarrow \underset{\text{Ferrous sulphate}}{FeSO_{4(aq)}} + \underset{\text{Hydrogen sulphide}}{H_2S_{(g)}}$$

$$\underset{\text{zinc sulphite}}{ZnS_{(s)}} + \underset{\text{sulphuric acid}}{H_2SO_{4(aq)}} \longrightarrow \underset{\text{Zinc sulphate}}{ZnSO_{4(aq)}} + \underset{\text{Hydrogen sulphide}}{H_2S_{(g)}}$$

Chemical Nature of Acids

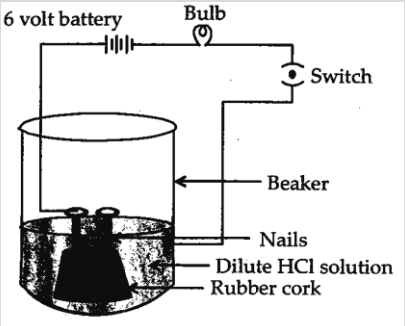
All acids produce hydrogen gas on reaction with metals. This means that hydrogen is common to all acids. It has been observed that the aqueous solutions of all acids conduct electricity. They allow the passage of electric current through them. For example, if electric current passed through aqueous solution of acids, then, current flows through their solutions.

ACTIVITY CORNER

To see what is common in all acids

Requirements: Solutions of hydrochloric acid, sulphuric acid, glucose, alcohol, etc.

Fix two iron nails on a cork and place the cork in a 100 mL beaker. Connect the nails to the two terminals of a 6 volt battery with wires through a bulb and a switch as shown in figure.



Now, pour some dilute HCl in the beaker and switch on the circuit. Note if the bulb glows or not.

Repeat the experiment with dilute H_2SO_4 glucose and alcohol. Does the bulb glow in all cases?

Observation: You will observe that the bulb glows in case of acids (dilute HCl and dilute H_2SO_4). This means that the aqueous solutions of acids conduct electricity. However, the bulb does not glow in case of glucose and alcohol.

Conclusion: This shows that alcohol and glucose do not conduct electricity. Since the electric current is carried by ions, this shows that aqueous solutions of all acids contain ions.

Conduction of electricity through a solution of a substance can take place only if ions are present in the solution. This shows that the acids when dissolved in water produce hydrogen ions (H^+). These H^+ ions do not exist as such in the solution. They combine with water molecules and exist as hydronium ions, H_3O^+ ($H_2O + H^+ \longrightarrow H_3O^+$). As each H^+ ion can combine with a number of water molecules, in the solution, we represent it as $H^+_{(aq)}$. The dissociation of a few acids in the aqueous solution to give $H^+_{(aq)}$ ions along with the corresponding negative ion is represented below:

$$\underset{\text{Hydrochlric acid}}{HCl_{(aq)}} \longrightarrow H^+_{(aq)} + \underset{\text{Chloride ions}}{Cl^-(aq)}$$

$$\underset{\text{Nitric acid}}{HNO_{3(aq)}} \longrightarrow H^+_{(aq)} + \underset{\text{Nitrate ions}}{NO_3^-(aq)}$$

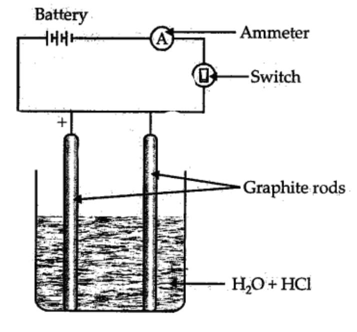
$$\underset{\text{Sulphuric acid}}{H_2SO_{4(aq)}} \rightarrow 2H^+_{(aq)} + \underset{\text{Sulphate ions}}{SO_4^{2-}}$$

$$\underset{\text{Acetic acid}}{CH_3COOH_{(aq)}} \longrightarrow \underset{\text{Acetate ions}}{CH_3COO^-_{(aq)}} + H^+_{(aq)}$$

ACTIVITY CORNER

To show that solution of acid can conduct electricity

- **Requirements:** Dil. HCl, battery, ammeter, connecting wires, switch, two graphic rods.
- Take dil. HCl in container as shown in the diagram.
- Put two graphite rods in dil. HCl and complete the circuit as shown in diagram.
- Switch on the current.
- Observe the reading of ammeter.



- **Observation:** The pointer of ammeter moves showing that current is passing through the solution.
- **Conclusion:** Acids conduct electricity in aqueous solution.

Arrhenius Concept of Acids

In 1984, Svante Arrhenius, a Swedish chemist proposed a theory to define acids and bases known as Arrhenius concept of acids. According to Arrhenius theory. Acids are substances which dissociate in aqueous solution to give hydrogen ions (or hydronium ions). For example, HCl, HNO_3, CH_3COOH and H_2SO_4 are all acids as they dissociate in aqueous medium to give hydrogen ions.

$$\underset{\text{Hydrochloric acid}}{HCl_{(aq)}} + H_2O_{(l)} \longrightarrow \underset{\text{Hydronium ion}}{H_3O^+_{(aq)}} + Cl^-_{(aq)}$$

$$\underset{\text{Nitric acid}}{HNO_{3(aq)}} + H_2O_{(l)} \longrightarrow \underset{\text{Hydronium ion}}{H_3O^+_{(aq)}} + NO^-_{3(aq)}$$

$$\underset{\text{Acetic acid}}{CH_3COOH_{(aq)}} + H_2O_{(l)} \longrightarrow \underset{\text{Hydronium ion}}{H_3O^+_{(aq)}} + CH_3COO^-_{(aq)}$$

- **Hydronium Ions**

Acids dissociate in water to produce hydrogen ions in aqueous solutions. These hydrogen ions are highly unstable and combine with water molecules to form hydronium ions.

$$HCl \longrightarrow \underset{\text{Hydrogen ion}}{H^+} + Cl^-$$

$$H^+ + H_2O \longrightarrow \underset{\text{Hydronium ion}}{H_3O^+}$$

Thus, the properties of an acid are due to $H^+_{(aq)}$ ions or hydronium ions (H_3O^+) which it gives in the aqueous solution.

- **Strong and Weak Acids**

Those acids which are completely ionized in water are called strong acids.

Strong acids produce a large amount of hydrogen ions in aqueous medium. Perchloric acid ($HClO_4$), chromic acid (H_2CrO_4), sulphuric acid (H_2SO_4) and nitric acid (HNO_3) are all strong acids because they are completely ionized in water. When the concentrated solution of an add is diluted by mixing water, then the concentration of hydrogen ions H^+ [or hydronium ions H_3O^+] per unit volume decreases.

$$HCl_{(aq)} + H_2O_{(l)} \longrightarrow H_3O^+_{(aq)} + Cl^-_{(aq)}$$

$$HNO_{3(aq)} + H_2O_{(l)} \longrightarrow H_3O^+_{(aq)} + NO_3^-{}_{(aq)}$$

$$H_2SO_{4(aq)} + H_2O_{(l)} \longrightarrow H_3O^+_{(aq)} + SO_4^{2-}{}_{(aq)}$$

Those acids which are partially ionized in water are called weak acids.

Weak acids produce only a small amount of hydrogen ions in aqueous medium. Acetic acid (CH_3COOH), carbonic acid (H_2CO_3), formic acid (HCOOH) are example of weak acids as they are ionized to a limited extent in aqueous medium.

$$\underset{\text{Acetic acid}}{CH_3CHOOH_{(aq)}} + H_2O_{(l)} \rightleftharpoons H_3O^+_{(aq)} + \underset{\text{Acetate ion}}{CH_3COO^-_{(aq)}}$$

$$H_2CO_{3(aq)} + 2H_2O_{(l)} \rightleftharpoons 2H_3O^+_{(aq)} + CO_3^{2-}{}_{(aq)}$$

$$\underset{\text{Hydrogen cyanide}}{HCN_{(aq)}} + H_2O_{(l)} \rightleftharpoons H_3O^+_{(aq)} + \underset{\text{Cyanide ion}}{CN^-_{(aq)}}$$

Double headed arrow ($\rightleftharpoons$) indicates incomplete ionization in case of weak acids.

Do You Know

- Except HF other halogen acids (HCl, HBr, HI) are strong acids.

- In acids of sulphur, H_2SO_4, H_2SO_5, $H_2S_2O_7$ are strong acids.
- Acid of chromium (H_2CrO_4, $H_2Cr_2O_7$) are strong acids.
- HNO_3 and $HC1O_4$ (perchloric acid) are strong acids.
- HCN, Phenol and all carboxylic acids are weak monoprotic acids.
- H_2CO_3, H_2S and H_3PO_3 are weak diprotic acids.
- $HClO_3$ HNO_3, Conc. H_2SO_4, H_2CrO_4 are strong oxidizing agents.
- HBr and HI are strong reducing agents.
- HCl is the weakest reducing agent.

ILLUSTRATION

6. A solution of acetic acid in water is highly concentrated. Will you call it a strong acid? Explain?

Sol. No, acetic acid will remain a weak acid even if it is highly concentrated because its degree of dissociation (α) is much less than one. This means that in water, only a small amount of the acid exists as ions whatever may be its concentration or amount dissolved in a given volume of water.

7. What are strong and weak acids? In the following list of acids, separate strong acids from weak acids.

Hydrochloric acid, citric acid, acetic acid, nitric acid, formic acid, suphurice acid

Sol. Strong acids are those which ionize almost completely in aqueous solution and hence produce a large amount of H^+ ions and and therefore, conduct electricity to a large extent. Weak acids are those acids which ionize to a small extent in the aqueous solution and hence produce a small amount of H^+ ions, and therefore, conduct electricity to a small extent.

Strong acids: Hydrochloric acid, nitric acid, sulphuric acid.

Weak acids: Citric acid, acetic acid, formic acid.

Concentrate and Dilute Acids

- Mixing an acid with water results in decrease in the concentration of H_3O^+ ions per unit volume. This process is called dilution and the acids thus obtained are called dilute acids. Generally pure acids are used as concentrated acids. Example: H_2SO_4 and HCl. 68% HNO_3 solution in water is used as concentrated HNO_3. In case of weak adds, on dilution, total number of H^+ ions in the solution increases because dissociation of the weak add increases on dilution but even in this case, concentration of H^+ ions per unit volume decreases.
- The process of dissolving an acid into water is a highly exothermic process. During this a lot of heat is produced.

ILLUSTRATION

8. Why do HCl, HNO_3, etc., show acidic characters in aqueous solutions while solutions of compounds like alcohol and glucose do not show acidic character?

Sol. It is because HCl and HNO_3 ionize in aqueous solution whereas ethanol and glucose do not ionize in aqueous solution.

Do You Know

It is always desirable to add acid to water slowly and not water to acid and keeping the solution continuously stirred. If water is added to a concentrated acid, the heat produced may cause the mixture to splash out and may spill on our clothes and body. This may cause serious burns. The glass container may also break due to excessive local heating.

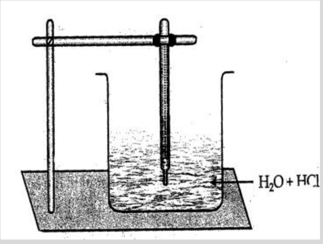
Industrial Uses of Acids

- Hydrochloric Acid (HCl): Used for cleaning metallic items and in dyestuffs, in tanning, and printing industry.
- Sulphuric acid (H_2SO_4): Used in the manufacture of fertilizers, chemicals, paints, dyes, plastics, synthetic fibres, detergents, explosives and car batteries. It is called King of Chemicals.
- Nitric acid (HNO_3): Used for manufacture of fertilizers, explosives (like TNT: trinitro toluene), dyes, plastics and drugs.
- Acetic acid (CH_3COOH): Used as vinegar in cooking and food preservatives.

ACTIVITY CORNER

To show mixing of acid in water is an exothermic process.

- **Requirements:** Water, HCl, HNO_3 H_2SO_4, thermometer, stirrer or glass rod.



- **Procedure:**
- Take water in a beaker.
- Dip thermometer it, hang it and note down the initial temperature.
- Pass HCl gas through water and stir it with glass rod.
- Note down rise in temperature.
- Repeat the experiment with H_2SO_4 (conc.) and HNO_3 (conc.)

- **Observation**: There is rise in temperature in each case.
- Chemical Reactions and Explanation: HCl ionizes in aqueous solution and the process is exothermic, i.e., heat is released.

$$HCl_{(g)} + H_2O_{(l)} \rightarrow H_3O^+{}_{(aq.)} + Cl^-{}_{(aq.)} + Heat\ energy$$

When conc. H_2SO_4 is added to water, the rise in temperature is maximum because it is highly exothermic process.

$$H_2SO_{4(l)} + 2H_2O_{(l)} \rightarrow 2H_3O^+{}_{(aq.)} + SO_4{}^{2-}{}_{(aq.)} + Heat\ energy$$

When HNO_3 is dissolved in water, there is rise in temperature.

$$HNO_{3(l)} + H_2O_{(l)} \rightarrow H_3O^+{}_{(aq.)} + NO_3{}^-{}_{(aq.)} + Heat\ energy$$

- **Conclusion:** Dissolution of acids in water is an exothermic process.

Bases

Bases are substances which have bitter taste, soapy touch and turn red litmus solution to blue. Some colour changes shown by various indicators with bases are:

- change red colour of red litmus solution to blue.
- give yellow colour with methyl orange.
- give pink colour with phenolphthalein.
- no reaction with blue litmus solution.

- **Strong Bases**

The bases which completely ionize in aqueous solution to give hydroxide ions (OH^-) are called strong bases.

Sodium hydroxide (NaOH), potassium hydroxide (KOH), etc. are called strong bases as they completely dissociate in aqueous solution to produce large number of OH^- ions on ionization. Bases (like NaOH, KOH) that are soluble in water are called alkalies.

As they produce large number of ions on ionization, they are also strong electrolytes.

- **Weak Bases**

The bases which ionize to small extent in aqueous solution to give hydroxide ions (OH^-) are called weak bases.

Ammonium hydroxide (NH_4OH), calcium hydroxide [$Ca(OH)_2$] are called weak bases as they only partially (less than 100%) ionize in water. For example,

$$\underset{\text{Ammonium hydroxide}}{NH_4OH_{(aq)}} \xrightleftharpoons{Water} NH^+{}_{4(aq)} + OH^-{}_{(aq)}$$

$$\underset{\text{Calcium hydroxide}}{Ca(OH)_{2(aq)}} \xrightleftharpoons{Water} Ca^{2+}{}_{(aq)} + 2OH^-{}_{(aq)}$$

Weak bases are also weak electrolytes as they ionize only to limited extent in water.

- **Acidity or Hydroxicity of Base**

It is equal to the number of hydroxide ions (OH^-) furnished by a base molecule in its aqueous solution after complete dissociation.

- Mono-acidic bases : They dissociate as :

$$BOH \rightleftharpoons B^+ + OH^-$$

i.e,. $NaOH, KOH, NH_4OH,$ etc.

- Di-acidic bases : They dissociate as :

$$B(OH)_2 \rightleftharpoons B^{2+}\ 2OH^-$$

i.e., $Ca(OH)_2$, $Mg(OH)_2$, etc.

- Tri-acidic bases : They dissociate as :

$$B(OH)_3 \rightleftharpoons B^{3+}\ 3OH^-$$

i.e., $Al(OH)_3$, etc.

Chemical Properties of Bases

- **Reaction of Bases with Metals**

Some metals react with bases to liberate hydrogen gas. For example, zinc reacts with sodium hydroxide and hydrogen gas is evolved. The reaction is:

$$\underset{\text{Sodium hydroxide}}{2NaOH} + \underset{\text{Zinc}}{Zn} \longrightarrow \underset{\text{Sodium zincate}}{Na_2ZnO_2} + \underset{\text{Hydrogen gas}}{H_2}$$

- **Reaction with Acids**
  Bases react with acids to undergo neutralization reaction forming salt and water. For example, sodium hydroxide reacts with hydrochloric acid to form sodium chloride (salt) and water.

$$\underset{\substack{\text{Sodium hydroxide}\\ \text{(Base)}}}{\mathrm{NaOH_{(aq)}}} + \underset{\substack{\text{Hydrochloric acid}\\ \text{(Acid)}}}{\mathrm{HCl_{(aq)}}} \longrightarrow \underset{\substack{\text{Sodium chloride}\\ \text{(Salt)}}}{\mathrm{NaCl_{(aq)}}} + \mathrm{H_2O_{(l)}}$$

$$\underset{\text{Sodium hydroxide}}{\mathrm{2NaOH_{(aq)}}} + \underset{\text{Sulphuric acid}}{\mathrm{H_2}SO_{4(aq)}} \longrightarrow \underset{\text{Sodium } \textit{sulphate}}{\mathrm{Na_2}SO_{4(aq)}} + \mathrm{2H_2O_{(l)}}$$

$$\underset{\text{Sodium hydroxide}}{\mathrm{CaOH_{2(aq)}}} + \underset{\textit{Hydrochloric}\text{ acid}}{\mathrm{HCl_{(aq)}}} \longrightarrow \underset{\text{Calcium } \textit{chloride}}{\mathrm{CaCl_{2(aq)}}} + \mathrm{2H_2O_{(l)}}$$

- **Reaction of Bases with Non-metal Oxides**
  Bases when react with non-metallic oxides such as carbon dioxide then they produce metal salt and water. In the reaction, non-metal oxide reacts with base to form salt which indicates their acidic nature. For example, sodium hydroxide reacts with carbon dioxide to form sodium carbonate and water as:

$$\underset{\text{Sodium hydroxide}}{\mathrm{2NaOH_{(s)}}} + \underset{\textit{Carbon dioxide}}{CO_{2(g)}} \longrightarrow \underset{\textit{Sodium carbonate}}{\mathrm{Na_2CO_{3(aq)}}} + \mathrm{H_2O_{(l)}}$$

$$\underset{\text{Calcium hydroxide}}{\mathrm{Ca(OH)_2}} + \underset{\textit{Carbon dioxide}}{CO_{2(g)}} \longrightarrow \underset{\textit{Calciu carbonate}}{\mathrm{CaCO_{3(aq)}}} + \mathrm{H_2O_{(l)}}$$

- **Reaction of Bases with Heavy Metal Salts** Metal salts react with aqueous solution of bases to produce precipitates of insoluble metallic hydroxides. These reactions are just like double displacement reactions in which metal ions are exchanges between two salts in aqueous solution. In these reactions, metal ions are exchanged between salt and base in aqueous solution. For example,

$$\underset{\textit{Zinc sulphate}}{ZnSO_{4(aq)}} + \underset{\textit{Sodium hydorxide}}{\mathrm{2NaOH_{(aq)}}} \longrightarrow \underset{\textit{Sodium sulphate}}{\mathrm{Na_2SO_{4(aq)}}} + \underset{\substack{\textit{Zink hydroxide}\\ \textit{(White ppt.)}}}{Zn(OH)_2}$$

$$\underset{\textit{Cupper sulphate}}{CuSO_{4(aq)}} + \underset{\textit{Ammonium hydorxide}}{\mathrm{2NH_4OH_{(aq)}}} \rightarrow \underset{\textit{Ammonium sulphate}}{\mathrm{(Na_4)_2SO_{4(aq)}}} + \underset{\textit{Copper hydroxide}}{Cu(OH)_2}$$

$$\underset{\textit{Ferrous chloride}}{FeCl_{2(aq)}} + \underset{\textit{Sodium hydorxide}}{\mathrm{2NaOH_{(aq)}}} \rightarrow \underset{\textit{Sodium chloride}}{\mathrm{2NaCl_{(aq)}}} + \underset{\substack{\textit{Ferrous hydroxide}\\ \textit{(Brown ppt.)}}}{Fe(OH)_2}$$

Chemical Nature of Bases

Like acids, bases also possesses some general characteristic properties. This shows that chemically, they must have sometimes in common. It is observed that bases when dissolved in water produce hydroxide ions in the solution. Thus, this is the common characteristic of all bases. For example,

$$\underset{\textit{Sodium hydorxide}}{\mathrm{NaOH_{(s)}}} + Water \longrightarrow \underset{\substack{\textit{Sodium}\\ \textit{ions}}}{\mathrm{Na^+_{(aq)}}} + \underset{\substack{\textit{Hydroxide}\\ \textit{ions}}}{OH^-_{(aq)}}$$

$$\underset{\textit{Potassium hydorxide}}{\mathrm{KOH_{(s)}}} + Water \longrightarrow \underset{\substack{\textit{Potassium}\\ \textit{ions}}}{K^+_{\mathrm{(aq)}}} + \underset{\substack{\textit{Hydroxide}\\ \textit{ions}}}{OH^-_{(aq)}}$$

$$\underset{\textit{Ammonium hydorxide}}{\mathrm{NH_4OH}} + Water \longrightarrow \underset{\substack{\textit{Ammonium}\\ \textit{ions}}}{NH_4{}^+_{\mathrm{(aq)}}} + \underset{\substack{\textit{Hydroxide}\\ \textit{ions}}}{OH^-_{(aq)}}$$

$$\underset{\textit{Magnesium hydorxide}}{Mg\mathrm{(OH)_{2(s)}}} + Water \longrightarrow \underset{\substack{\textit{Magnesium}\\ \textit{ions}}}{Mg^{2+}_{\mathrm{(aq)}}} + \underset{\substack{\textit{Hydroxide}\\ \textit{ions}}}{2OH^-_{(aq)}}$$

Arrhenius Concept for Bases

According to Arrhenius concept, "bases are the substances which dissociate in aqueous solution to generate hydroxide (OH^-) ions". example, substances such as NaOH, KOH, Mg(OH)₂, Ca(OH)2, NH₄OH, etc. are bases. Bases like NaOH, KOH, etc. which produce single OH^- ion on ionization are called monoacidic bases. Bases like Mg(OH)₂, Ca(OH)₂ which produce two OH^- ions on ionization are called diacidic bases.

$$\underset{\textit{Potassium hydorxide}}{\mathrm{KOH_{(s)}}} \xrightarrow{H_2O} K^+_{\mathrm{(aq)}} + OH^-_{(aq)}$$

$$\underset{\textit{Magnesium hydorxide}}{Mg\mathrm{(OH)_{2(s)}}} \xrightarrow{H_2O} Mg^{2+}_{\mathrm{(aq)}} + 2OH^-_{(aq)}$$

$$\underset{\textit{Calcium hydorxide}}{\mathrm{Ca(OH)_{2(s)}}} \xrightarrow{H_2O} \mathrm{Ca}^{2+}_{\mathrm{(aq)}} + 2OH^-_{(aq)}$$

$$\underset{\textit{Ammonium hydorxide}}{\mathrm{NH_4OH}_{(s)}} \xrightarrow{H_2O} NH_4{}^+_{\mathrm{(aq)}} + OH^-_{(aq)}$$

- **Applications of Arrhenius Concept**

- It explains the acidic nature of aqueous solution of non-metallic oxides, (i.e., CO₂, SO₂, SO₃ ,N₂O₃, N₂O₅, P₂O₃ etc.)

$$CO_2 + H_2O \rightleftharpoons H_2CO_3 \rightleftharpoons 2H^+ + CO_3^{2-}$$

$$SO_3 + H_2O \rightleftharpoons H_2SO_4 \rightleftharpoons 2H^+ + SO_4^{2-}$$

- It explains the basic nature of aqueous solutions of several substances like metal hydroxides, CaO, Na₂O, NH₃, N₂H₄ etc.

$$CaO + H_2O \rightleftharpoons Ca(OH)_2 \rightleftharpoons Ca^{2+} + 2OH^-$$

$$NH_3 + H_2O \rightleftharpoons NH_4OH \rightleftharpoons NH_4^+ + OH^-$$

- Water is amphoteric in nature as it furnishes both H^+ and OH^- in aqueous solution.

$$H_2O \rightleftharpoons H^+ + OH^-$$

- Neutralization reaction: Basically it is a reaction between H^+ and OH^- to form H₂O.

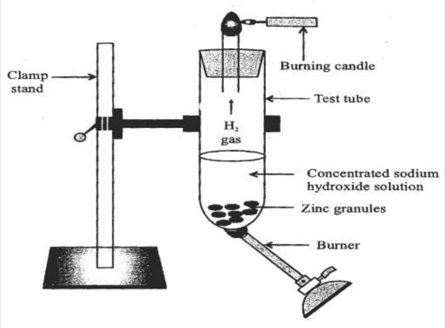
$$H^+_{(aq)} + OH^-_{(aq)} \rightleftharpoons H_2O_{(l)}$$

- Strength of acid/base: It depends upon the tendency of acid or base to furnish H^+ or OH^- in solution.

ACTIVITY CORNER

To study reaction of base with zinc metal.

- Take 2-3 pieces of granulated zinc in a test tube.
- Add about 2 mL of concentrated sodium hydroxide solution. Warm the contents of the test tube.



- **Observation:** It is observed that zinc metal dissolves and the hydrogen gas is evolved which is a combustible gas and can be tested by burning candle or the flame of a burner.
- **Conclusion:** Metals react with bases to give H_2 gas.

Industrial Uses of Bases

- **Sodium hydroxide (NaOH):** Used in manufacture of soap, paper, synthetic fibre (called rayon), petrol refining and as a laboratory reagent. It is also known as caustic soda.
- **Potassium hydroxide (KOH):** Used in alkaline batteries and soft soap known as caustic potash.
- **Calcium hydroxide, $Ca(OH)_2$:** Used in manufacture of bleaching powder, softening of hard water and neutralizing acid in the soil. Also known as slaked lime.
- **Magnesium hydroxide, $Mg(OH)_2$:** Used as an antacid to neutralize excess acid in the stomach and to cure indigestion. Also known as milk of lime.
- **Ammonium hydroxide (NK_4OH):** Used for removing grease stains from clothes.

ILLUSTRATION

9. The following reactions occur in aqueous solution. Predict the products and identify the acids and bases (and conjugate species) in the reaction of –

(a) NH_3 with CH_3COOH

(b) $N_2H_5^+$ with CO_3^{2-}

(c) H_3O^+ with OH^-

(d) HSO_4^- with $HCOO^-$.

Sol. (a) $\underset{Base}{NH_3} + \underset{Acid}{CH_3COOH} \rightarrow \underset{Acid}{NH_4^+} + \underset{Base}{CH_3COO^-}$

(b) $\underset{Acid}{N_2H_5^+} + \underset{Base}{CO_3^{2-}} \longrightarrow \underset{Base}{N_2H_4} + \underset{Acid}{HCO_3^-}$

(c) $\underset{Acid}{H_3O^+} + \underset{Base}{OH^-} \longrightarrow \underset{Acid\ and\ base}{2H_2O}$

(d) $\underset{Acid}{HSO_4^-} + \underset{Base}{HCOO^-} \rightarrow \underset{Base}{SO_4^{2-}} + \underset{Acid}{HCOOH}$

10. What is Arrhenius definition of bases?

Sol. A base is a substance which when dissolved in water gives hydroxide (OH^-) ions in the solution.

11. NH_3 does not hydroxyl group, then why is it a base?

Sol. NH_3 when dissolved in water forms ammonium hydroxide (NH_4OH) which ionizes to give OH^- ions in the solution. Hence, it is a base.

12. What is the difference between a strong base and a weak base. Give two examples. Of each of them.

Sol. **Strong base:** A base which dissociates completely in the aqueous solution to produce a large amount of OH^- ions is called a strong base. For example, sodium hydroxide (NaOH) and potassium hydroxide (KOH).

Weak base: A base which dissociates only partially in the aqueous solution and hence gives only a small amount of OH^- ions is called a weak base. For example, ammonium hydroxide (NH_4OH) and hydroxide, $Ca(OH)_2$.

Salts

A salt is a compound formed from an acid by the replacement of the hydrogen in the acid by a metal. Salts are formed when acids react with bases. Salts can be defined as chemical substances that are formed as a result of neutralization of acid by base. They can be divided into 4 classes:

- **Normal Salts**
  They do not generally contain any replaceable hydrogen atom. They are formed when all replaceable hydrogens of an acid are replaced by metal, e.g., NaCl, KCl etc. they are formed when a strong acid is completely neutralized by a strong base. These salts do not react with water on dissolving. In aqueous solution their pH is 7.0.

e.g. When HCl is neutralized with NaOH, normal salt NaCl is produced.

$HCl + NaOH \rightarrow NaCl + H_2O$

All chlorides, bromides, iodides, sulphates, nitrates of alkali and alkaline earth metals are normal salts.

- **Neutral Salts**

  These salts are formed by the reaction of a weak acid with weak base. e.g. When CH_3COOH is neutralized with NH_4OH, neutral salt CH_3COONH_4 is produced.

  $CH_3COOH + NH_4OH \rightarrow CH3COONH_4 + H_2O$

- **Acidic Salts**

  These salts contain replaceable hydrogen which is released in aqueous solvent. They are formed when a polybasic acid is partially neutralized by a weak base. e.g. $NaHCO_3$, $NaHSO_3$. They are acidic because they retain one or more hydrogen of parent acid.

  E.g. When weak base NH_4OH reacts with strong acid HCl, acidic salt NH_4Cl is formed.

  $NH_4OH + HCl \rightarrow NH_4Cl + H_2O$

  In aqueous solution their pH is less then 7.0. These salts further react with bases to form neutral salts, e.g. NH_4Cl, $(NH_4)_2CO_3$.

Note:

- Ammonium halides (NH_4Cl, NH_4Br, NH_4I), NH_4NO_3, $(NH_4)_2SO_4$ are acidic salts.
- Metal sulphates except of alkali & alkaline earth metals, all metal sulphates are acidic salts.
- Except nitrates of alkali & alkaline earth metals, almost all metal nitrates are acidic salts.

- **Basic Salts**

  These salts are formed by partial neutralization of a polyacidic base with weak acid. They contain replaceable OH^- group that ionizes in aqueous medium. They are basic because they retain one or more hydroxyl group of parent base. In aqueous solution their pH is more then 7.0. These salts further react with acids to form neutral salts.

- Except hydroxides of alkali metals and alkaline earth metals other hydroxides are weak hydroxides.
- Most metal acetates, carbonates, oxalates, cyanides and fluorides are basic salts.

Family of Salts

Various salts having common metal ion can be grouped in a single family. Few families of salts are:

- The salts of 'hydrochloric acid' are called 'chlorides'
- The salts of 'sulphuric acid' are called 'sulphates'.
- The salts of 'nitric acid' are called 'nitrates'.
- The salts of 'carbonic acid' are called 'carbonates'.
- The salts of 'acetic acid' are called 'acetates'.

Family	Sodium salts	Potassium salts
Common ion	Na^+	K^+
Chlorides	NaCl	KCl
Bromides	NaBr	KBr
Nitrates	$NaNO_3$	KNO_3
Sulphates	Na_2SO_4	K_2SO_4
Carbonates	Na_2CO_3	K_2CO_3

ACTIVITY CORNER

To write the formulae of the salts and to identify their acids and bases and the families.

- Write down the formulae of the salts given below:

  (1) Potassium sulphate, (2) Sodium sulphate (3) Calcium sulphate (4) Magnesium sulphate (5) Copper sulphate (6) Sodium chloride (7) Sodium nitrate (8) Sodium carbonate (9) Ammonium chloride.
- Identify the acids and bases from which the above salts may be obtained.
- Salts having the same positive or negative radicals are said to belong to a family. For example NaCl and Na_2SO_4 belong to the family of sodium salts. Similarly, NaCl and KCl belong to the family of chloride salts.
- How many families can you identify among the salts given in this activity?

- **Discussion and Conclusion:**

	Salt	Formula	Acid	Base
1.	Potassium sulphate	K_2SO_4	H_2SO_4	KOH
2.	Sodium sulphate	Na_2SO_4	H_2SO_4	NaOH
3.	Calcium sulphate	$CaSO_4$	H_2SO_4	$Ca(OH)_2$
4.	Magnesium sulphate	$MgSO_4$	H_2SO_4	$Mg(OH)_2$
5.	Copper sulphate	$CuSO_4$	H_2SO_4	$Cu(OH)_2$
6.	Sodium chloride	NaCl	HCl	NaOH
7.	Sodium nitrate	$NaNO_3$	HNO_3	NaOH
8.	Sodium carbonate	Na_2CO_3	H_2CO_3	NaOH
9.	Ammonium chloride	NH_4Cl	HCl	NH_4OH.

- **Families:**

  - On the basis of common acids

  Sulphates: K_2SO_4 Na_2SO_4, $CaSO_4$, $MgSO_4$, $CuSO_4$
  Chlorides: NaCl, NH_4Cl; Carbonates: Na_2CO_3; Nitrates: $NaNO_3$

  - On the basis of common bases.

  Sodium salts: Na_2SO_4, NaCl, $NaNO_3$, Na_2CO_3
  Potassium salts: K_2SO_4; Calcium salts: $CaSO_4$
  Magnesium salts: $MgSO_4$; Copper salts: $CuSO_4$
  Ammonium salts: NH_4Cl

- **Mixed Salts**

  Salts containing more than one cation or anion other than H^+ or H^- ions are called mixed salts, e.g. $CaOCl_2$, $NaKCO_3$, $Na(NH_4)HPO_4$. They are generally produced by simultaneous neutralization of single acid by more than one base or single base by more than one acid.

- **Double Salts**

  When saturated solutions of two salts containing different cations are crystallized in a single solution, crystals obtained are of a double salt. e.g. when K_2SO_4 & $Al_2(SO_4)_3$ are crystallized by mixing their equimolar solutions a double salt $K_2SO_4.Al_2(SO_4)_3.24H_2O$ is formed. These salts give positive tests for all constituent ions,

- **What Do Acids and Bases have in Common?**

  Water itself shows acidic and basic nature to some extent as:

  $$H_2O \rightleftharpoons H^+ + OH^-$$

  But water is overall neutral because H^+ and OH^- ions present are equal in number. When acidic substances are dissolved in water, they increase the concentration of H^+ ions in the solution which makes it overall acidic. However OH^- ions are present in the solution but in less concentration as compared to H^+ ions.

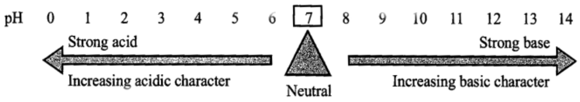
  When basic substances are dissolved in water, they increase the concentration of OH^- ions in the solution which makes it overall basic. However H^+ ions are present in the solution but in less concentration as compared to OH^- ions.

pH Scale

- The hydrogen ion concentration $[H^+]$ in aqueous solution is represented in moles per litre. Since, $[H^+]$ is very small, it is inconvenient to express the acidity or alkalinity of an aqueous solution on the basis of $[H^+]$ ion concentration.
- pH scale is an scale used to major the strength of Arrhenius acids and bases. It was designed by Sorensen in 1909. In pH/ p stands for potenz meaning power in German and H stands for hydrogen ions. As acidic or basic nature is due to active H^+ ions in aqueous medium, pH can also be defined as a scale to represent the activity of H^+ ions in aqueous solution.
- In chemical terms pH can be defined as logarithm of reciprocal of H_3O^+ ions in aqueous solution.

  or $pH = \log\frac{1}{[H_3O^+]}$ OR $[H^+] = 10^{-pH}$

  In pH scale less is the value of pH more is the acidic nature of the substance. In pH scale:
  - all substances having pH value between 0 and 7 are called acids.
  - all substances having pH value equal to 7 are neutral.
  - all substances having pH value between 7 and 14 are called bases.



- **Logarithmic Values for Some Common Digits**

  $\text{Log}_{10}\, 2 = 0.3010$ $\log_{10} 3 = 0.48$ $\log_{10} 4 = 0.60$ $\log_{10} 5 = 0.69$ $\log_{10} 10^x = x$ i.e., $\log_{10} 10^5 = 5$

  If concentration of H^+ ion is more than 1, pH value comes less than 0 and pH scale only include numeric values between 0 to 14 and pH value of a solution cannot be negative. If negative value is calculated for pH than it must be stated that pH value of solution is less than 0, but use of negative sign is omitted.

ILLUSTRATION

13. Calculate Ph value of

(a) 2×10^{-3} M HCl, $2\times10^{-4}\text{M}\,\text{H}_2\text{SO}_4$

Sol. $(a)\ pH = \log\frac{1}{[H^+]}$ $[H^+] = 2\times10^{-3}$

$$pH = \log\frac{1}{2\times10^{-3}} = \log_{10}\frac{10^3}{2} = \log_{10}10^3 - \log_{10}2$$

(b) $2\times10^{-4}\text{M}\,\text{H}_2\text{SO}_4$ $Given, M =$ $2\times10^{-4}\text{M}$

$[H^+] =$ $2\times2\times10^{-4} = 4\times10^{-4}$

$$\text{pH} = \log_{10}\frac{1}{2\times10^{-4}} = \log_{10}\frac{10^4}{4}$$

$= \log_{10}10^4 - \log_{10}4 = 4 - 0.60 = 3.4$

14. Determine the Ph of the solution when hydrogen ion concentration is

$(i) 1.0\times10^{-3}\ M\ HCl, (ii) 1.0\times10^{-5}\ M\ HNO_3$

Sol.

$(\text{i})\ \text{pH} = \text{-log}[\text{H}^+].\ \ [\text{H}^+] = 1.0\times10^{-3}\ \text{M}$

$\therefore\ \text{pH} = \text{-log}(1.0\times10^{-3})$

= - (-3) log 10 = 3 $(\because \log 10 = 1)$

(ii) $[\text{H}^+] = 1.0\times10^{-5}\ \text{M}$.

$\text{pH} = \text{-log}[\text{H}^+] = -\log(1.0\times10^{-5})$

= - (- 5) log 10 = 5

15. Ph of a solution changes from 6 to 5. What changes do not you expect in hydrogen ion concentration?

Sol. We know that Ph = -log[H^+]

If pH is 6, then [H^+]will be $1.0\times10^{-6}\ mol\, L^{-1}$

If pH is 5, then [H^+]will be $1.0\times10^{-5}\ mol\, L^{-1}$

Increase in concentration of H^+ ions

$$= \frac{1.0\times10^{-5}\ \text{mol}\,\text{L}^{-1}}{1.0\times10^{-6}\ \text{mol}\,\text{L}^{-1}} = 10\ \text{times}$$

Therefore, H^+ concentration is increased 10 times when pH changes from 6 to 5.

16. What is the pH of a 0.01 M NaOH solution?

Sol. Since NaOH is completely dissociated, thus gives 0.01 M OH^- ions. Since OH^- ion concentration is known, we can calculate the H^+ ions concentration from ionic product value.

Ionic product of water,

$$[H^+][OH^-] = 1\times10^{-14}\ or\ [H^+] = \frac{1\times10^{-14}}{[OH^-]}$$

$$Now, [OH^-] = 0.01M = 1\times10^{-2}\ M$$

$$\therefore\ [H^+] = \frac{1\times10^{-14}}{1\times10^{-2}} = 10^{-12}\ M$$

$$pH = -\log[H^+] = -\log[10^{-12}] = -(-12)\log 10 = 12.$$

17. Calculate the concentration of H^+ and OH^- of a solution whose (a) pH =0, (b) pH = 14.

Sol. (a) pH = 0, $[H^+] = 10^{-pH} = 10^{-0} = 1$

$$[OH^-] = \frac{K_w}{[H^+]} = \frac{1\times10^{-14}}{1} = 1\times10^{-14}$$

(b) $pH = 14, [H^+] = 10^{-pH} = 10^{-14}$

$$[OH^-] = \frac{K_w}{[H^+]} = \frac{1\times10^{-14}}{10^{-14}} = 1$$

18. Determine the pH of solution when hydrogen ion concentration is

$(a) 1.0\times10^{-9}\ M\ \ (b) 1.0\times10^{-5}\ M.$

Sol. (i) $pH = -\log[H^+].[H^+] = 1.0\times10^{-9}\ M$

$\therefore pH = -\log(1.0\times10^{-9}) = -(-9)\log 10 = 9$

$(\because \log 10 = 1)$

(ii) $[H^+] = 1.0\times10^{-5}\ M.\ \ pH = -\log[H^+]$

$= -\log(1.0\times10^{-9}) = (-5)\log = 5.$

19. Calculate the pH of 10^{-5} M NaOH solution.

Sol. NaOH is a strong base. It ionizes completely in aqueous solution as

$$NaOH \xrightarrow{+Water} Na^+_{(aq)} + OH^-_{(aq)}$$

$\therefore p[OH^-] = [NaOH] = 10^{-5}\ M$

To calculate the pH of the solution, we should know $[H^+]$. This can be calculated by using the formula

$$[H^+][OH^-] = K_w = 10^{-14}\ M^2$$

Or $[H^+]$ $= \frac{10^{-14}}{[OH^-]} = \frac{10^{-14}}{10^{-5}} = 10^{-9} M.\ \ \therefore\ pH = 9$

20. How many grams of NaOH should be dissolved in one litre of the solution to prepare a solution with pH =12?

Sol. pH = 12 means $[H^+] = 10^{-12}\ M$

$$\therefore [H^+]\ = \frac{10^{-14}}{[OH^-]} = \frac{10^{-14} M^2}{10^{-12}\ M} = 10^{-2} M$$

As NaOH ionizes completely as

$$NaOH \xrightarrow{+Water} Na^+ + OH^-$$

$$\therefore [NaOH] = [OH^-] = 10^{-2}\,M = 10^{-2}\,mol\,L^{-1}$$

$$= (10^{-2}\,mol\,L^{-1})(40\,g\,mol^{-1}) = 0.4\,g\,L^{-1}$$

$$(\because \text{ Molar mass of } NaOH$$

$$= 23+16+1 = 40\,g\,mol^{-1})$$

21. Calculate the pH of a solution obtained by dissolving

(i) 0.02 mole of hydrochloric acid in 2 litres of the solution.

(ii) 0.01 mole of sulphuric acid in 2 litres of the solution.

Sol. (i) Molar concentration of HCl acid solution

$$= \frac{0.02\,mol}{2\,L} = 0.01\,mol\,L^{-1} = 10^{-2}\,mol\,L^{-1}$$

As HCl is a strong acid, it completely ionizes as

$$HCl \xrightarrow{+Water} H^+ + Cl^-$$

$\therefore [H^+] = [HCl] = 10^{-12}\,M$. Hence, pH = 2

(ii) Molar concentration of H_2SO_4 solution

$$= \frac{0.01\,mol}{2\,L} = 0.005\,mol\,L^{-1}$$

As H_2SO_4 is a strong acid, it completely ionizes as: $H_2SO_4 \xrightarrow{+Water} 2H^+ + SO^{2-}{}_4$

Thus, 1 molecule of H_2SO_4 gives $2H^+$ ions

$$\therefore [H^+] = 2 \times [H_2SO_4] = 2 \times 0.005 = 0.01\,mol/L$$

$$= 10^{-2}\,mol\,L^{-1}.\,Hence,\,pH = 2$$

- **Dilution Concept**

  If molarity of an acid is M_1 and volume is V_1 , when some water is added in this acid, then volume According to dilution concept,

  $M_1V_1 = M_2V_2$

ILLUSTRATION

22. When 900 mL H_2O is added in 2M, 100 ml HCl, calculate pH value of this acidic solution.

Sol. According to,

$M_1V_1 = M_2V_2$

$V1 = 100$ mL, $V_2 = 100 + 900 = 1000$ mL,

$M_1 = 2$, $M_2 = ?$

$$\Rightarrow 2 \times 100 = M_2 \times 1000$$

$$\Rightarrow M_2 = 2 \times 10^{-1} = 0.2\,M$$

$$[H^+] = 0.2$$

$$pH = \log\frac{1}{0.2} = \log\frac{10}{2} = \log 5 = 0.69$$

Importance of pH in Daily Life

pH is quite useful to us in a number of ways in daily life. These applications are based on the neutralization reactions which we have studied.

- **Comparison of Relative Acidic and Basic Strengths of Acids and Bases**

  The main utility of pH is to know about the acidic and basic strengths of solutions and also to compare these.

  In general, lesser the pH of a solution, more will be its acidic strength, similarly, higher the pH of a solution, more will be its basic strength. For example, pH of vinegar is about 3.0 while that of coffee is nearly 4.5. This means that vinegar is a stronger acid than coffee. Similarly, pH of washing soda solution (Na_2CO_3 solution) is nearly 9.0 and that of lime water containing $Ca(OH)_2$ is 10.5. This indicates that lime water is a stronger base than washing soda solution.

- **pH of Our Digestive System**

  Our stomach produces hydrochloric acid. It helps in the digestion of food without harming our stomach. The stomach produces too much acid and this causes pain and irritation and feeling of burning. This stage is called acidity. To get relief from this pain, we take tablets known as antacids. These contain bases to neutralize the excess acids. Magnesium hydroxide (called milk of magnesia), a mild base is also used for this purpose.

- **pH of Change Leads to Tooth Decay**

  The white enamel coating on our teeth is made up of insoluble calcium phosphate which is quite hard. When the pH in the mouth falls below 5.5, the enamel gets corroded. The bacteria's present in the mouth break down the sugar to lactic acid. The formation of these acids causes decrease in pH. It is therefore, advisable to avoid eating sugary foods and also to keep the mouth clean so that sugar and food particles may not be present. The toothpastes contain in them some basic ingredients and they help in neutralizing the effect of the acids and also in increasing the pH in the mouth.

- **Self-Defence by Animals and Plants Through Chemical Warfare**

  Stings of bees and ants contain methanoic acid (or formic acid). When stung, they cause lot of pain and irritation. The cure is the rubbing of the affected area with soap. Sodium hydroxide (NaOH) present in the soap neutralizes acid injected in the body and thus brings the pH back to its original level. The effect of stings by wasps containing alkali is neutralized by the application of vinegar which is ethanoic acid (or acetic acid). Nettle plants are herbaceous in nature and grow in the wild. These have sharp hairs which contain in them methanoic acid. If they happen to touch the body by accident, their stings are very painful because methanoic acid present gets injected in the body. These are commonly known as stinging nettles. The stung area is rubbed by the leaves of "dock plants' which often grow beside nettle plants.

- **Soil pH and Plant Growth**
  The growth of plants in a particular soil is also related to its pH. Soils with high peat content or iron minerals or with rotting vegetation tend to become acidic. The soil pH can reach as low as 4. The acidic effect can be neutralized by 'liming the soil' which is carried by adding calcium hydroxide, calcium oxide or powdered chalk. Similarly, the soil with excess of lime stone or chalk is usually alkaline. Sometimes, its pH reaches as high as 8.3 and is quite harmful for the plant growth. In order to reduce the alkaline effect, it is better to add some decaying organic matter which is acidic in nature (compost or manure).

Plants	Preferred pH range
Potatoes	4.5 – 6.0
Carrot, sweet potato	5.5 – 6.5
Cauliflower, garlic, tomato	5.5 – 7.5
Onion, cabbage	6.6 – 7.5

ILLUSTRATION

23. What type of pH should the soil have for healthy growth of plants?

Sol. Soil should neither be alkaline nor highly acidic.

24. Why do we feel a burning sensation in the stomach when we overeat? What is the medicine used called? Give one example.

Sol. When we overeat, the burning sensation in the stomach is due to excess hydrochloric acid produced in the stomach. The medicine used to neutralize it is called “antacid”. The most commonly used antacid is magnesium hydroxide (milk of magnesia)

25. Why does tooth decay start when the pH of the mouth is lower than 5.5?

Sol. The tooth enamel is made up of calcium phosphate which is insoluble in water. When pH in the mouth is lower than 5.5, it becomes moderately acidic, the dissolution of calcium phosphate starts, i.e., tooth decay begins.

Some Important Acids, Bases and Salts

Sodium Hydroxide, NaOH

Chemical Formula	NaOH
Common name	Caustic soda
Action on litmus	Turn red litmus blue
Solubility	Soluble in water
(Highly alkaline)	

- **Pure Sodium Hydroxide**
  Commercial sodium hydroxide, besides impurities such as Na_2SO_4, NaCl etc. invariably contains some amount of sodium carbonate due to the absorption of atmospheric CO_2. However, it may be purified by dissolving it in alcohol in which impurities including Na_2CO_3 NaCl, Na_2SO_4 are insoluble. The filtrate on evaporation given off alcohol leaving behind pure solid NaOH.

- **Physical Properties**
  (i) It is a deliquescent, white crystalline solid.
  (ii) It is highly soluble in water and the solution is bitter in taste, corrosive and soapy to touch.
  (iii)It is sparingly soluble in alcohol. However, KOH is fairly soluble in alcohol.
- **Chemical Properties**
  (i) Action of atmosphere: NaOH on exposure to atmosphere absorbs moisture and CO_2 to form Na_2CO_3.

  $2NaOH + CO_2 \longrightarrow Na_2CO_3 + H_2O$

  (ii) Strong alkali nature: Sodium hydroxide is a strong alkali as it ionizes completely in water furnishing OH^- ions.

  $NaOH \longrightarrow Na^+ + OH^-$

  (a) It reacts with acids forming salts.
  Type 1: NaOH reacts with strong acids give normal salts which are neutral salts.

  $NaOH + HCl \longrightarrow NaCl + H_2O$

  $2NaOH + H_2SO_4 \longrightarrow Na_2SO_4 + 2H_2O$

  $NaOH + HNO_3 \longrightarrow NaNO_3 + H_2O$

  Type 2: NaOH reacts with weak acids give basic salt.

  $NaOH + CH_3COOH \longrightarrow CH_3COONa + H_2O$

  (b) It combines with acidic oxides to form salts which may be acidic, basic or neutral.

  NaOH + Acidic oxide $\longrightarrow$ Salt + Water

$$2NaOH + CO_2 \longrightarrow \underset{\text{Sodium carbonate}}{Na_2CO_3} + H_2O$$

$$2NaOH + SO_2 \longrightarrow \underset{\text{Sodium sulphite}}{Na_2SO_3} + H_2O$$

$$2NaOH + 2NO_2 \longrightarrow \underset{\text{Sodium nitrate}}{NaNO_3} + NaNO_2 + H_2O$$

(c) Oxides of Amphoteric metals (aluminium, zinc, tin and lead) dissolve in sodium hydroxide forming corresponding salts. These metals are called amphoteric metals as they form amphoteric oxides.

$$Al_2O_3 + 2NaOH \longrightarrow \underset{\text{Sodium meta-aluminate}}{2NaAlO_2} + H_2O$$

$$ZnO + 2NaOH \longrightarrow \underset{\text{Sodium zincate}}{Na_2ZnO_2} + H_2O$$

$$SnO + 2NaOH \longrightarrow \underset{\text{Sodium stannite}}{Na_2SnO_2} + H_2O$$

$$SnO_2 + 2NaOH \longrightarrow \underset{\text{Sodium stannate}}{Na_2SnO_3} + H_2O$$

$$PbO + 2NaOH \longrightarrow \underset{\text{Sodium plumbite}}{Na_2PbO_2} + H_2O$$

$$PbO_2 + 2NaOH \longrightarrow \underset{\text{Sodium plumbate}}{Na_2PbO_3} + H_2O$$

(iii) **Action on amphoteric metals:** Amphoteric metals (Be, Al, Ga, Zn, Sn and Pb) react with NaOH to liberate hydrogen gas.

$2Al + 2NaOH + 2H_2O \longrightarrow 2NaAlO_3 + 3H_2$

$Zn + 2NaOH \longrightarrow Na_2ZnO_2 + H_2O$

$Sn + 2NaOH + H_2O \longrightarrow Na_2SnO_3 + 2H_2$

$Pb + 2NaOH + H_2O \longrightarrow Na_2PbO_3 + 2H_2$

(iv) Action on ammonium salts: Ammonium salts on heating with sodium hydroxide solution release ammonia gas.

$$NH_4Cl + NaOH \longrightarrow NH_3\uparrow + NaCl + H_2O$$

$$(NH_4)_2SO_4 + 2NaOH \longrightarrow Na_2SO_4 + 2NH_3\uparrow + 2H_2O$$

(v) Action of carbon monoxide: NaOH reacts with carbon monoxide under pressure at 150-200°C to form sodium formate.

$$NaOH + CO \xrightarrow[5-10\,atm.]{150-200^{o}C} \underset{\text{Sodium formate}}{HCOONa}$$

Note: carbon monoxide is neutral gas but it show acidic nature with NaOH. For this reaction its structure is responsible.

(vi) Hydroxides of Amphoteric metals (Be, Al, Ga, Zn, Sn and Pb) are soluble in NaOH, due to complex formation, but other hydroxides do not dissolve in NaOH.

Note:

1. Any mixture of hydroxides of amphoteric metals and other metals can be separated by NaOH. **Example:** Mixture of $Fe(OH)_3$ and $Al(OH)_3$ is separated by NaOH.
2. Some Hydroxide [$Zn(OH)_2$, $Cu(OH)_2$, $Cd(OH)_3$] are soluble in NH_4OH. Example: Mixture of $Al(OH)_3$ and $Zn(OH)_2$ cannot be separated by NaOH, because both are soluble in NaOH. This mixture is separated by NH_4OH.

(vii) Caustic Property: Sodium hydroxide breaks down the proteins to a pasty mass. On account of this property, it is commonly called as Caustic soda. $NaOH_{(aq.)}$, is also called as soda lye.

- **Uses**

  Sodium hydroxide is used as a reagent in the laboratory. Sodium hydroxide is a strong alkali and is used widely in the chemical industry, e.g., soap manufacture, purification of bauxite, manufacture of rayon etc. However, when a base is required to neutralize acids, it is more economical to use sodium carbonate, ammonia or lime. Chlorine gas which is a valuable by-product in NaOH manufacture can be used for the production of chemicals such as polyvinyl chloride (PVC), an important plastic.

Sodium Carbonate (NaCO3.10H2O)

Chemical Formula	Na_2CO_3-$10H_2O$
Common name	Washing soda
Action on litmus	Neutral
Solubility	Sparingly soluble in water (weakly basic)

It is the most important chemical of commercial use. It exists in various forms such as:

(a) Anhydrous sodium carbonate or Na_2CO_3 or soda ash.

(b) Monohydrate sodium carbonate or $Na_2CO_3.H_2O$ or crystal carbonate.

(c) Heptahydrate sodium carbonate or $Na_2CO_3.7H_2O$.

(d) Decahydrate sodium carbonate or $Na_2CO_3.10H_2O$.

The first process for the manufacture of Na_2CO_3 was invented by Le-Blanc. However, this process has become obsolete and replaced by modern methods. The most important method for the preparation of sodium carbonate is Solvay process or ammonia soda process.

- **Solvay Process or Ammonia Soda Process**

  The raw materials used in this process are the common salt (NaCl) or brine solution, ammonia and lime stone. The process in brief, involves the formation of a sparingly soluble sodium bicarbonate by the reaction of sodium chloride and ammonium bicarbonate in aqueous solution. Sodium bicarbonate on heating decomposes into sodium carbonate. The reactions taking place in the process are:

$$NH_3 + H_2O + CO_2 \longrightarrow \underset{\textit{Ammonium blcarbonate}}{NH_4HCO_3}$$

$$NaCl + NH_4HCO_2 \longrightarrow \underset{\textit{Soudium blcarbonate}}{NaHCO_3} \downarrow + NH_4Cl$$

$$\underset{\textit{Filtered and heated}}{2NaHCO_3} \xrightarrow{\Delta} Na_2CO_3 + \underbrace{H_2O + CO_2}_{\textit{Can be used again}}$$

- **Electrolytic Process**

  In the Nelson cell used in the manufacture of sodium hydroxide, carbon dioxide under pressure is blown along with steam. The sodium hydroxide produced then reacts with carbon dioxide to form sodium carbonate. The solution is concentrated and crystallized.

  $$2NaOH + CO_2 \longrightarrow Na_2CO_3 + H_2O$$

- **Physical Properties**

  (i) Sodium carbonate is a white crystalline solid. The decahydrate form on standing in air effloresces and crumbles to powder.

  (ii) On heating, the monohydrate changes into the anhydrous form which does not decompose on further heating even to redness.

  (iii)It is soluble in water with the evolution of considerable amount of heat to show alkaline nature due to hydrolysis.

  $$Na_2CO_3 + 2H_2O \longrightarrow 2NaOH + H_2CO_3$$

- **Chemical Properties**

  (i) Action of acids: It is readily decomposed by acids with evolution of carbon dioxide. The reaction involves two steps.

  $$Na_2CO_3 + HCl \longrightarrow NaHCO_3 + NaCl$$

  $$NaHCO_3 + HCl \longrightarrow NaCl + H_2O + CO_2$$

  (ii) Action of CO_2: On passing CO_2 through the concentrated solution of sodium carbonate, sodium bicarbonate gets precipitated.

  $$Na_2CO_3 + H_2O + CO_2 \longrightarrow 2NaHCO_3$$

  (iii) Action of Silica: The mixture of sodium carbonate and silica on fusion gives sodium silicate.

  $$Na_2CO_3 + SiO_2 \longrightarrow Na_2SiO_3 + CO_2\uparrow$$

  Sodium silicate is called soluble glass or water glass as it is soluble in water.

  (iv) Action of slaked lime: Sodium hydroxide is formed when the solution of sodium carbonate containing sulphur on treatment with sulphur dioxide gives sodium thiosulphate ($Na_2S_2O_3$).

  $$Na_2CO_3 + SO_2 \xrightarrow{H_2O} Na_2SO_3 + CO_2\uparrow$$

  $$Na_2SO_3 + S \longrightarrow Na_2S_2O_3$$

- **Uses**

  Sodium carbonate is used:

  - in the manufacture of glass, sodium silicate, paper, borax, soap powders, caustic soda etc.
  - in laundry and in softening of water as washing soda.
  - in textile and petroleum refining.
  - as a laboratory reagent. The mixture of Na_2CO_3 and K_2CO_3 is used as a fusion mixture. It is used in quantitative analysis to standardize acid solutions. In qualitative analysis, it is used in the detection of acidic radicals especially of insoluble salts.
  - for the preparation of various carbonates of metals.
  - $Na_2CO_3.NaHCO_3.2H_2O$ is sodium sesqui carbonate. It is neither deliquescent nor efflorescent and is used for wood washing.
  - A mixture of Na_2CO_3 + CaS is called black ash.

Sodium Hydrogen Carbonate (NaHCO3)

Chemical Formula	$NaHCO_3$
Common name	Baking soda
Action on litmus	Turn red litmus to blue
Solubility	Sparingly soluble in water (weakly basic)

- **Preparation**

  It is obtained as the intermediate product in the Solvay or ammonia soda process. Normal carbonate can be changed to bicarbonate by passing carbon dioxide through its saturated solution.

  $$Na_2CO_3 + CO_2 + H_2O \longrightarrow \underset{\textit{Sparingly soluble}}{2NaHCO_3}$$

- **Properties**

  - $NaHCO_3$ is a white crystalline solid, sparingly soluble in water.
  - The solution is alkaline in nature due to hydrolysis.

  $$NaHCO_3 + H_2O \rightleftharpoons NaOH + H_2CO_3$$

  Its aqueous solution gives yellow colour with methyl orange but no colour with phenolphthalein and thus is weak base.

- On heating, it releases carbon dioxide and water forming sodium carbonate.

$$2NaHCO_3 \longrightarrow Na_2CO_3 + H_2O + CO_2$$

- **Uses**

  $NaHCO_3$ is used :

  - as a medicine (antacid) to neutralize the acidity in the stomach.
  - in making effervescent drinks.
  - for making baking powder (a mixture of potassium hydrogen tartar ate and sodium bicarbonate).
  - for production of CO_2.
  - in fire extinguishers.

- **Sodium Chloride**

Chemical Formula	NaCl
Common name	Rock salt
Action on litmus	Neutral
Solubility	Soluble in water

- **Occurrence**

  It occurs abundantly in nature. The three major sources are:

  - Sea water which contains about 3% of sodium chloride.
  - Water of inland lakes such as Sambhar lake in Rajasthan.
  - Salt mines.

- **Preparation of NaCl**

  Salty water is subjected to evaporation in open air in fields which leaves behind residue of impure sodium chloride as deposits.

- **Purification of Common Salt**

  Sodium chloride obtained from sea water or from lakes contains many impurities such as sulphates of sodium and magnesium along with chlorides of calcium and magnesium. The chlorides of calcium and magnesium are particularly undesirable on account of their deliquescent nature.

  For its purification common salt is dissolved in minimum quantity of water to get a saturated solution. Insoluble impurities are filtered off. Hydrogen chloride gas is passed through the saturated solution. Due to common ion effect, the ionic product of NaCl increases above the solubility product and therefore NaCl is then thrown out as pure crystals. The impurities being more soluble remain in the mother liquor. The crystals are filtered, washed and dried.

$$NaCl_{(g)} \xrightleftharpoons{Water} Na^+_{(aq.)} + Cl^-_{(aq.)}$$

$$HCl_{(g)} \xrightleftharpoons{Water} H^+_{(aq.)} + Cl^-_{(aq.)}$$

Thus, $[Na^+][Cl^-] > K_{sp_{NaCl}}$

- **Properties**

  - It is white crystalline solid (m.pt. 1073 K), soluble in water (solubility 36/100 g at 239 K). It dissolves in water with absorption of heat.
  - It is insoluble in alcohol.
  - The common salt is the starting material for the preparation of all other sodium compounds e.g., NaOH, Na_2CO_3, $NaHCO_3$, etc. and extraction of sodium.

- **Uses**

  NaCl is used as:

  - an essential constituent of food.
  - a preservative of food articles like fish, meat, etc.
  - for making useful sodium compounds.
  - in 'salting out' of soap, and in making freezing mixture.

  Note: Pure NaCl is not hygroscopic. It shows hygroscopic nature due to impurities.

Calcium Carbonate, CaCO3

Chemical Formula	$CaCO_3$
Action on litmus	Neutral
Solubility	Slightly soluble in water

It exists in nature in huge quantities in various forms such as lime stone, marble, chalk, etc.

- **Preparation**

  It is obtained by passing carbon dioxide through lime water or by adding sodium carbonate solution to $CaCl_2$.

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

$$CaCl_2 + Na_2CO_3 \longrightarrow CaCO_3 + 2NaCl$$

- **Properties**

  - It is a white powder, insoluble in water.
  - It dissolves in water in presence of CO_2 due to formation of calcium bicarbonate.

$$CaCO_3 + H_2O + CO_2 \longrightarrow Ca(HCO_3)_2$$

- **Uses**

  - Precipitated chalk is used in toothpastes and face powders, in medicine for indigestion, in adhesives and in cosmetics.

- Chalk is used in paints and distempers.
- Marble is used for building purposes, in the production of CO_2 in laboratory.
- Limestone is used in the manufacture of quick lime, slaked lime, cement, washing soda and glass.
- As flux in the smelting of iron and lead ores.

Calcium Sulphate

Chemical Formula	$CaSO_4.2H_2O$
Common name	Gypsum
Action on litmus	Neutral
Solubility	Slightly soluble in water

It occurs in nature as Anhydrite ($CaSO_4$) and Gypsum ($CaSO_4.2H_2O$). $CaSO_4.2H_2O$ is also known as alabaster.

- **Preparation**

It is prepared by reacting any calcium salt either sulphuric acid or a soluble sulphate (i.e., Na_2SO_4).

$$CaCl_2 + H_2SO_4 \longrightarrow CaSO_4 + 2HCl$$

$$CaCl_2 + Na_2SO_4 \longrightarrow CaSO_4 + 2NaCI$$

- **Properties**

- It is a white crystalline solid, sparingly soluble in water and solubility decreases as the temperature increases.
- It dissolves in dilute acids. It also dissolves in ammonium sulphate due to the formation of double sulphate, $(NH_4)_2SO_4.CaSO_4.H_2O$.
- Gypsum on heating first changes from monochnic form to orthorhombic form without loss of water. At 120°C, it loses three-fourth of its water of crystallization and forms hemihydrates $(2CaSO_4).H_2O$ or $CaSO_4.(1/2)H_2O$ which is commonly known as Plaster of Paris. It becomes anhydrous at 200°C. The anhydrous form is known as burnt plaster or dead plaster. On strongly heating, it decomposes to give calcium oxide.

$$2[CaSO_4.2H_2O] \xrightarrow{120^{\circ} C} \underset{Plaster\ of\ Paris}{[2CaSO_4.H_2O]} + 3H_2O$$

$$[2CaSO_4.H_2O] \xrightarrow{200^{\circ} C} \underset{Dead\ plaster}{2CaSO_4} + H_2O$$

$$2CaSO_4 \xrightarrow[Strongly]{Heated} 2CaO + 2SO_2 + O_2$$

- A suspension of gypsum when saturated with ammonia and carbon dioxide forms ammonium sulphate, a nitrogenous fertilizer.

$$2NH_3 + CaSO_4 + CO_2 + H_2O \longrightarrow (NH_4)_2SO_4 + CaCO_3$$

It forms calcium sulphide on heating strongly with carbon.

$$CaSO_4 + 4C \longrightarrow CaS + 4CO$$

- **Uses**

It is used:

- For the manufacture of plaster of Paris, cement, ammonium sulphate, sulphuric acid, etc.
- For preparing blackboard chalks.
- In anhydrous form as drying agent.

Calcium Sulphate Hemihydrate

Chemical Formula	$2CaSO_4.H_2O$ or $CaSO_4.\frac{1}{2}H_2O$
Common name	Plaster of Paris
Action on litmus	Neutral
Solubility	Insoluble

- **Preparation**

It is obtained when gypsum is heated at 120°C.

$$2[CaSO_4.2H_2O] \xrightarrow{120^{\circ} C} \underset{\substack{Plaster\ of\ Paris \\ (Calcuim\ sulphate\ hemihydrate)}}{2CaSO_4.H_2O} + 3H_2O$$

- **Properties**

- Plaster of Paris is a white powder.
- Plaster of Paris has the property of setting to hard mass when a paste with water is allowed to stand aside for some time. Slight expansion occurs during the setting as water is absorbed to reform $CaSO_4.2H_2O$ (gypsum). The setting process is exothermic.

$$\underset{Plaster\ of\ Paris}{2CaSO_4.2H_2O} \xrightarrow[H_2O]{Setting} \underset{\substack{Gypsum \\ (Orthorhombic)}}{CaSO_4.2H_2O} \xrightarrow{Hardening} \underset{\substack{Gypsum \\ (Monoclinic)}}{CaSO_4.2H_2O}$$

The setting of Plaster of Paris is catalyzed by sodium chloride and is retarded by borax or alum. Addition of alum to Plaster of Paris gives hard setting. The mixture is known as Keene's cement.

- When Plaster of Paris is heated at 200°C, it forms anhydrous calcium sulphate which is known as dead plaster. It has no setting property as it takes up water only very slowly.

- **Uses**

It is used:

- In making blue paint.
- In laundry for blueing purposes.
- In making wallpaper and blue tinted paper.
- In calico printing.
- In making ornamental casts and idols.

- For making casts for supporting fractured bones (Plaster).

Calcium Oxychloride

Chemical Formula	$Ca(OCl)_2.2H_2O$
Common name	Bleaching powder
Solubility	Soluble

Bleaching powder is a mixed salt of calcium hypochloride [$Ca(OCl)_2.2H_2O$] and basic calcium chloride [$CaCl_2.Ca(OH)_2.H_2O$]. Actually it is not a true compound but simply it is represented by formula $CaOCl_2$.

- **Methods of Preparation**

  It is prepared by action of Cl_2 on dry slaked lime.

  $$Ca(OH)_2 + Cl_2 \longrightarrow CaOCl_2 + H_2O$$

  According of Cliffored, the formation of bleaching powder is explained as:

  $$2Ca(OH)_2 + 2Cl_2 \rightarrow Ca(OCl)_2 + CaCl_2 + 2H_2O$$

  $$CaCl_2 + Ca(OH)_2 + H_2O \rightarrow CaCl_2.Ca(OH)_2.H_2O$$

  $$3Ca(OH)_2 + 2Cl_2 \longrightarrow Ca(OCl)_2 + \underbrace{CaCl_2.Ca(OH)_2.H_2O}_{Bleaching\ powder} + H_2O$$

- **Physical Properties**

  It is pale yellow powder having smell of NH3. It is sparingly soluble in water and gives a milky suspension when dissolved in water. Its aqueous solution gives positive test for Ca^{2+}, Cl^- and OCl.

- **Chemical Properties**

  - **Decomposition:** On standing, it decomposes (auto oxidation) to calcium chloride and calcium chlorate.

  $$6CaOCl_2 \longrightarrow 5CaCl_2 + Ca(ClO_3)_2$$

  On standing chlorine percentage decreases regularly. Due to this germicidal nature is lost. Decomposition of bleaching powder is catalyzed by cobalt chloride.

  $$2CaOCl_2 \xrightarrow{CoCl_2} 2CaCl_2 + O_2$$

  - Reaction with Dilute Acid:

  (a) With limited quantity of acid: With limited quantity of dilute acids (HCl, H_2SO_4 etc.) bleaching powder liberates oxygen.

  $$2CaOCl_2 + H_2SO_4 \longrightarrow CaSO_4 + CaCl_2 + 2HOCl$$

  $$HOCl \longrightarrow HCl + [O]$$

  Evolution of oxygen is responsible for its bleaching action.

  (b) With excess of acid: With excess of acid, bleaching powder liberates chlorine which is called available chlorine. Due to available chloride, bleaching powder acts as disinfectant.

  $$CaOCl_2 + H_2SO_4 \longrightarrow CaSO_4 + H_2O + Cl_2$$

  $$CaOCl_2 + 2HCl \longrightarrow CaCl_2 + H_2O + Cl_2$$

  Cl_2 is liberated with CO_2 also.

  $$CaOCl_2 + CO_2 \longrightarrow CaCO_3 + Cl_2$$

  - **Oxidizing Nature:** Bleaching powder is good oxidizing agent because it gives nascent oxygen. i.e.,

  $$CaOCl_2 + Na_3AsO_3 \longrightarrow Na_3AsO_4 + CaCl_2$$

  $$CaOCl_2 + 2FeSO_4 + H_2SO_4 \longrightarrow Fe_2(SO_4)_3 + CaCl_2 + H_2O$$

  $$CaOCl_2 + KNO_2 \longrightarrow CaCl_2 + KNO_3$$

  $$3CaOCl_2 + 2NH_3 \longrightarrow 3CaCl_2 + 3H_2O + N_2 -$$

  Ethyl alcohol, acetone etc. give chloroform when treated with bleaching powder.

  - **Bleaching Action:** As bleaching powder gives nascent oxygen, it shows bleaching property.

  $$\text{Coloured matter} + [O] \longrightarrow \text{Colourless matter}$$

Bleaching Process

1. The material to be bleached is first passed through solution of NaOH to remove greasy matter.
2. Now material is passed through aqueous solution of bleaching powder and then through very dilutes HCl solution. HCl reacts with bleaching powder to liberate nascent oxygen which bleaches the material.
3. To remove excess of Cl_2, material is dipped in antichlor substances like hypo or sodium bisulphite.

$$2Na_2S_2O_3 + Cl_2 \longrightarrow 2NaCl + Na_2S_4O_6$$

$$NaHSO_3 + Cl_2 + H_2O \longrightarrow NaHSO_4 + 2HCl$$

- **Uses**

  Bleaching powder is used:

  - As bleaching agent in textile, paper and jute industry.
  - As disinfectant.
  - As oxidant.
  - In the preparation of several compounds.
  - For making wool unshrinkable.

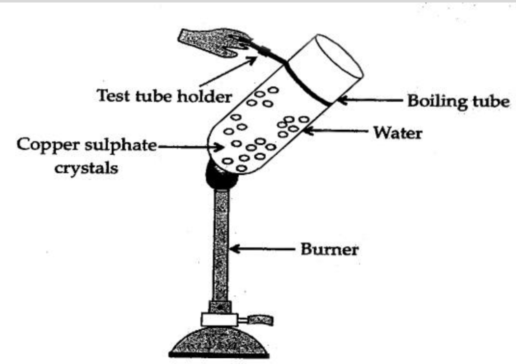
Water of Crystallization

Water of crystallization is the fixed number of water molecules present in one formula unit of a salt. For example, five water molecules are present in one formula unit of copper sulphate. So, the chemical formula of hydrated copper sulphate is $CuSO_4.5H_2O$ Similarly, ten water molecules are present in one formula unit of washing soda. The chemical formula of hydrated sodium carbonate is $Na_2CO_3.10H_2O$. Another important hydrated salt is gypsum which has two molecules of water of crystallization. It has the formula $CaSO_4.2H_2O$.

ACTIVITY CORNER

To test the presence of water of crystallization in copper sulphate crystal

- Take a few crystals of copper sulphate in a dry boiling tube. These have bluish green colour. Heat the tube by holding it with a test tube holder on the flame of the burner.
- You will observe that the colour of copper sulphate after heating becomes white. You may notice water droplets in the holding tube which are obtained from water of crystallization.



- Now, add 2-3 drops of water on the sample of copper sulphate obtained after heating. You will observe that the blue colour of copper sulphate is restored.
- **Observation:** We observe that copper sulphate crystals which seems to be dry contain water of crystallization. When we heat the crystals this water is removed and the salt becomes white. When we again moisten the crystals, the blue colour of crystals again reappears.
- **Conclusion :** The water of crystallization is given out on heating according to the reaction given below:

$$\underset{(Blue)}{CuSO_4.5H_2O} \xrightarrow{Heat} \underset{(White)}{CuSO_4 + 5H_2O}$$

$$\underset{(White)}{CuSO_4 + 5H_2O} \longrightarrow \underset{(Blue)}{CuSO_4.5H_2O}$$

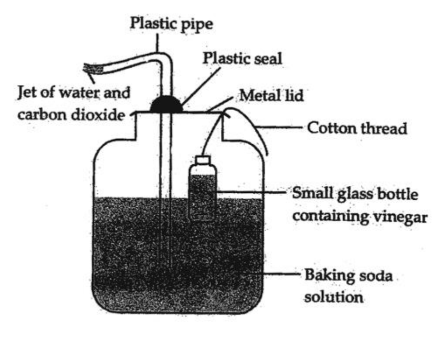
Do You Know

Crystalline salts like $Na_2CO_3.10H_2O$, $CaSO_4.2H_2O, CaSO_4.\frac{1}{4}H_2O, CuSO_4.5H_2O$, etc. which seem to dry contain water of crystallization. Water of crystallization is a fixed number of water molecules present in one formula unit of the salt.

ACTIVITY CORNER

To prepare soda acid fire extinguisher

- **Requirements:** (i) An empty and clean 1/2 litre jam jar with a metal lid, (ii) a small bottle, (iii) a plastic tube about 1 m long and 5 mm diameter, (iv) plasticine, baking soda, vinegar, a cotton thread and water, proceed as follows:
- Remove the metal lid from the empty jam jar and in the middle of it make a hole with the help of a nail and hammer, such that the size of the hole is just equal to the diameter of plastic pipe.
- Pass the plastic pipe through the hole in the lid, such that, when lid is placed on the jar, it is close to the base. Apply plasticine on the upper side of lid to make it air tight.



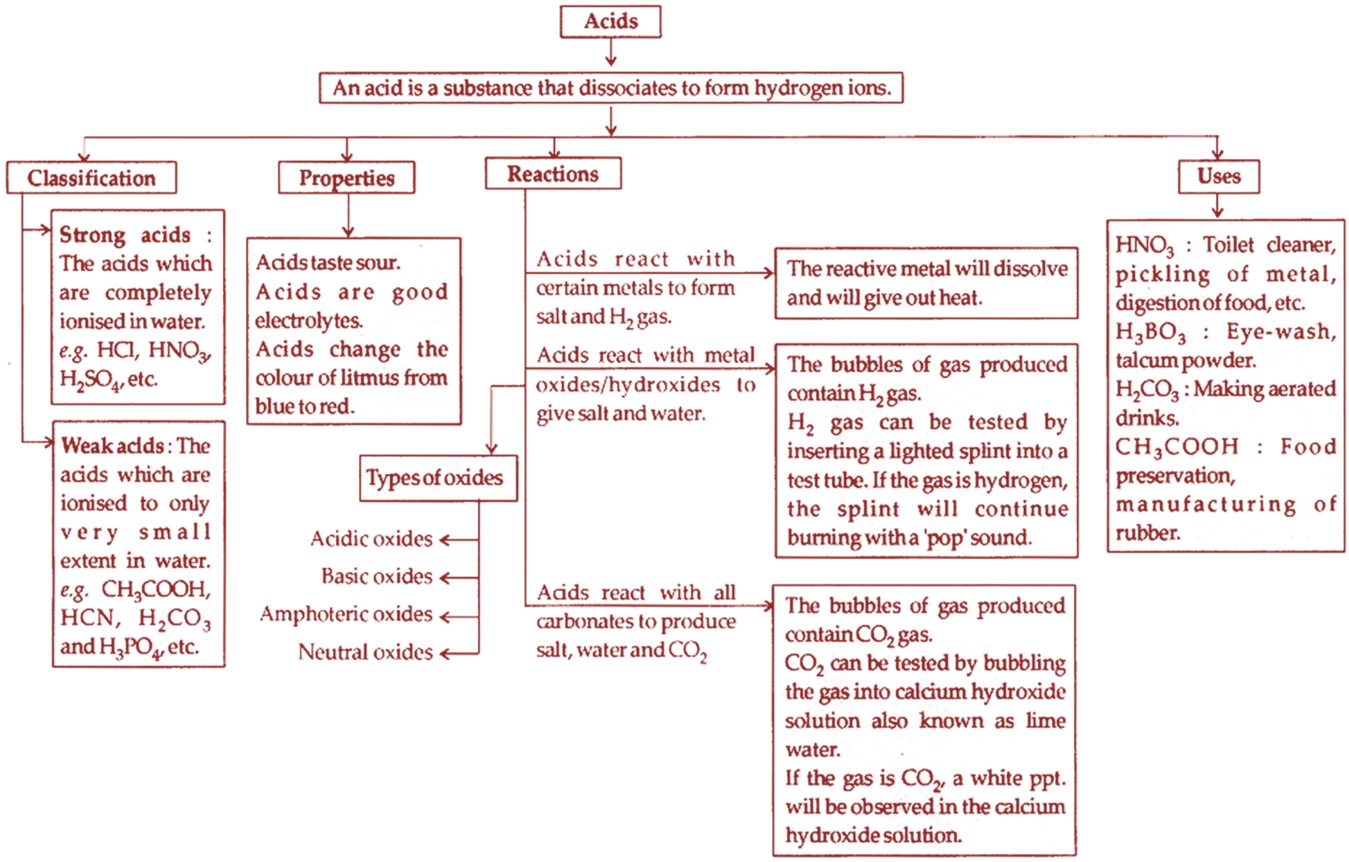
- Make a solution of baking soda by dissolving about 30 g of it in 300 cc of water. Pour the solution in the jar.
- Take a small glass bottle (such as inkpot) and tie a loop around its neck with a long cotton thread. Fill 3/4 of the bottle with vinegar.
- Suspend the glass bottle in the jar by holding the cotton thread such that a part of it is above the baking soda solution.

- Fix the lid on the jar. The lid will hold the cotton thread. Now, your working model of fire extinguishers is ready.
- In order to use it, life the vinegar mixes with baking soda solution, it produces carbon dioxide gas, which forces out the contents in the bottle.

- **Observation:** When sulphuric acid reacts with NaHCO3 to give CO_2.

$$2NaHCO_3 + H_2SO_4 \rightarrow Na_2SO_4 + 2H_2O + CO_2.$$

CONCEPT MAP



An acid is a substance that dissociates to form hydrogen ions.

Strong acids : The acids which are completely ionised in water. *e.g.* HCl, HNO_3, H_2SO_4, etc.

Weak acids : The acids which are ionised to only very small extent in water. *e.g.* CH_3COOH, HCN, H_2CO_3 and H_3PO_4, etc.

Acids taste sour.
Acids are good electrolytes.
Acids change the colour of litmus from blue to red.

Acids react with certain metals to form salt and H_2 gas.

The reactive metal will dissolve and will give out heat.

Acids react with metal oxides/hydroxides to give salt and water.

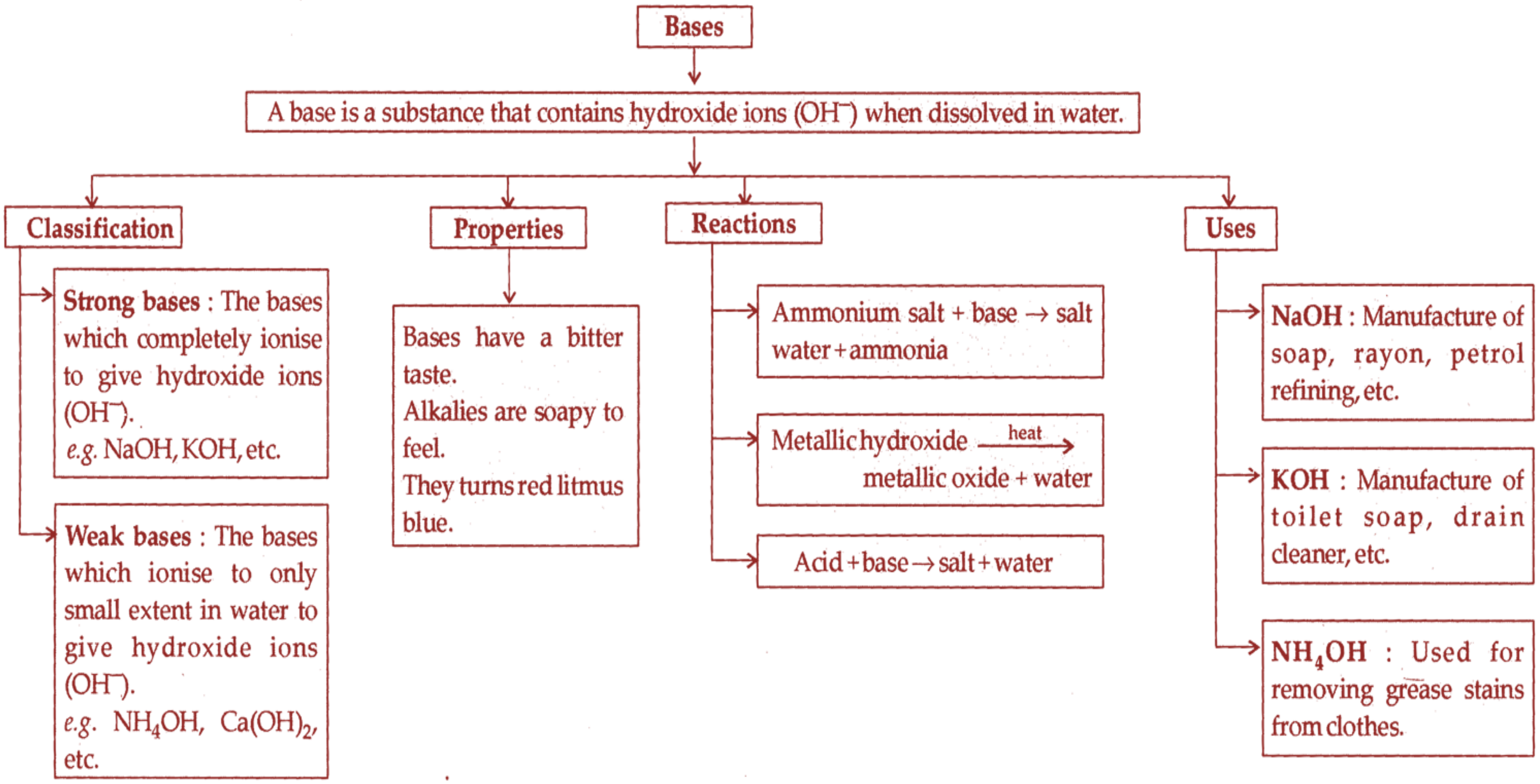
The bubbles of gas produced contain H_2 gas.
H_2 gas can be tested by inserting a lighted splint into a test tube. If the gas is hydrogen, the splint will continue burning with a 'pop' sound.

Acids react with all carbonates to produce salt, water and CO_2

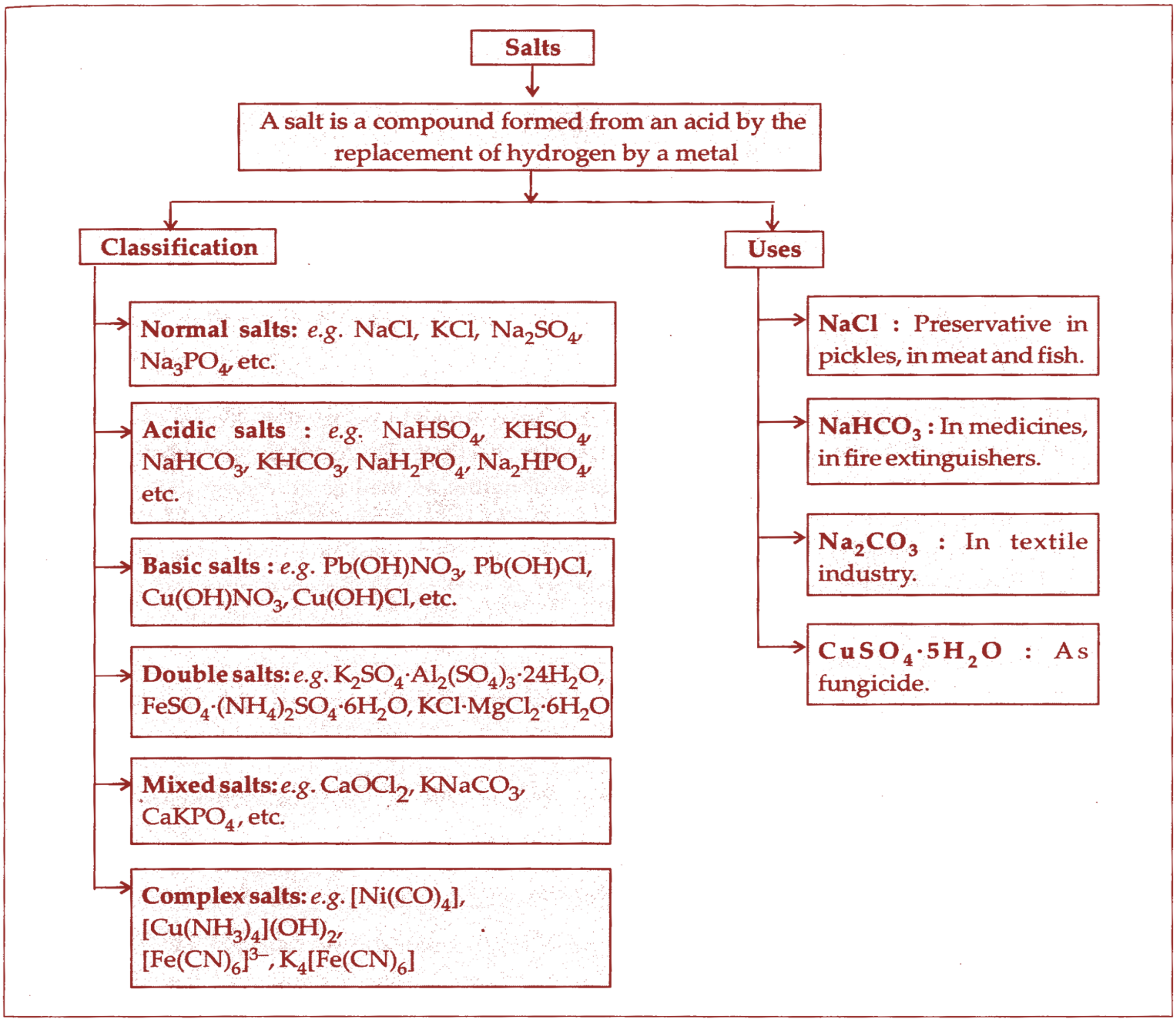
The bubbles of gas produced contain CO_2 gas.
CO_2 can be tested by bubbling the gas into calcium hydroxide solution also known as lime water.
If the gas is CO_2, a white ppt. will be observed in the calcium hydroxide solution.

HNO_3 : Toilet cleaner, pickling of metal, digestion of food, etc.
H_3BO_3 : Eye-wash, talcum powder.
H_2CO_3 : Making aerated drinks.
CH_3COOH : Food preservation, manufacturing of rubber.

CONCEPT MAP


Bases
A base is a substance that contains hydroxide ions (OH^-) when dissolved in water.
Classification
Properties
Reactions
Uses
Strong bases : The bases which completely ionise to give hydroxide ions (OH^-). *e.g.* NaOH, KOH, etc.
Weak bases : The bases which ionise to only small extent in water to give hydroxide ions (OH^-). *e.g.* NH_4OH, $Ca(OH)_2$, etc.
Bases have a bitter taste.
Alkalies are soapy to feel.
They turns red litmus blue.
Ammonium salt + base → salt water + ammonia
Metallic hydroxide $\xrightarrow{heat}$ metallic oxide + water
Acid + base → salt + water
NaOH : Manufacture of soap, rayon, petrol refining, etc.
KOH : Manufacture of toilet soap, drain cleaner, etc.
NH_4OH : Used for removing grease stains from clothes.

CONCEPT MAP



Salts

A salt is a compound formed from an acid by the replacement of hydrogen by a metal

Classification

- **Normal salts:** *e.g.* NaCl, KCl, Na_2SO_4, Na_3PO_4, etc.
- **Acidic salts :** *e.g.* $NaHSO_4$, $KHSO_4$, $NaHCO_3$, $KHCO_3$, NaH_2PO_4, Na_2HPO_4, etc.
- **Basic salts :** *e.g.* $Pb(OH)NO_3$, $Pb(OH)Cl$, $Cu(OH)NO_3$, $Cu(OH)Cl$, etc.
- **Double salts:** *e.g.* $K_2SO_4 \cdot Al_2(SO_4)_3 \cdot 24H_2O$, $FeSO_4 \cdot (NH_4)_2SO_4 \cdot 6H_2O$, $KCl \cdot MgCl_2 \cdot 6H_2O$
- **Mixed salts:** *e.g.* $CaOCl_2$, $KNaCO_3$, $CaKPO_4$, etc.
- **Complex salts:** *e.g.* $[Ni(CO)_4]$, $[Cu(NH_3)_4](OH)_2$, $[Fe(CN)_6]^{3-}$, $K_4[Fe(CN)_6]$

Uses

- **NaCl :** Preservative in pickles, in meat and fish.
- **$NaHCO_3$:** In medicines, in fire extinguishers.
- **Na_2CO_3 :** In textile industry.
- **$CuSO_4 \cdot 5H_2O$:** As fungicide.

ESSENTIAL POINTS For COMPETITIVE EXAMS

Bronsted - Lowry Concept

The Arrhenius concept of acids and bases was extended further by Johannes Bronsted and Thomas M. Lowry (1923).

According to Bronsted-Lowry concept, an acid is a substance which can donate proton (H^+) while a base is a substance which can accept a proton (H^+).

In other words, according to Bronsted-Lowry concept, acid is a proton donor and base is a proton acceptor. For example, in the following reaction, HCl loses a proton and is an acid whereas NH_3 accepts a proton and is a base.

$$HCl_{(aq)} + NH_{3(aq)} \rightleftharpoons NH^+_{4(aq)} + Cl^-_{(aq)}$$

The main advantage of this concept is that it is not restricted to neutral molecules. In this concept acid-base reaction are regarded as proton transfer reactions. There are certain substances such as H_2O, HCO_3^-, HSO_4^-, etc. which are capable of donating as well as accepting the proton. These substances are called amphoteric substances. For example, water behaves as an acid as well as base as illustrated below:

As a base: $\underset{acid}{HCl_{(aq)}} + \underset{base}{H_2O_{(l)}} \rightleftharpoons H_3O^+_{(aq)} + Cl^-_{(aq)}$

As an acid: $\underset{acid}{H_2O_{(l)}} + \underset{base}{NH_{3(aq)}} \rightleftharpoons NH^+_{4\,(aq)} + OH^-_{(aq)}$

Similarly, bicarbonate ion (HCO_3^-) behaves as an acid and a base as:

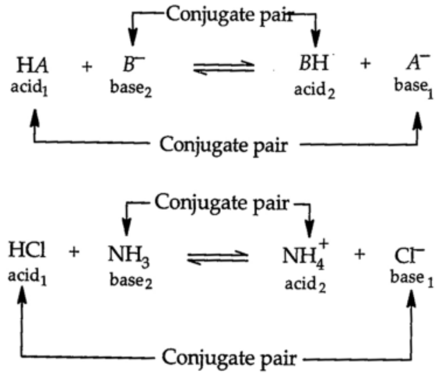
$$\underset{acid}{HCO_{3(aq)}^{-}} + \underset{(base)}{H_2O_{(l)}} \rightleftharpoons NH_{4(aq)}^{+} + CO_{3(aq)}^{2-}$$

$$\underset{acid}{HCl_{(aq)}} + \underset{base}{HCO_{3(aq)}^{-}} \rightleftharpoons H_2CO_{3(aq)} + Cl_{(aq)}^{-}$$

All Arrhenius acids are also Bronsted acids but Arrhenius bases are not Bronsted bases.

Conjugate Acid-Base Pair

In case of acid-base reaction, the reverse reaction is also an acid-base reaction. Every acid forms a conjugate base by the loss of a proton while every base forms a conjugate acid by the gain of a proton (H^+).



For example,

- Add base reactions always proceed in the direction from the stronger to weaker acid-base combination. For example,

$$\underset{\text{stronger acid}}{H_2SO_4} + \underset{\text{stronger base}}{NH_3} \xrightarrow{\text{Water}} \underset{\text{weaker acid}}{NH_4^+} + \underset{\text{weaker base}}{HSO_4^-}$$

Lewis Concept of Acids and Bases

According to Lewis concept, an add is a substance which can accept a pair of electrons, whereas a base is a substance which can donate a pair of electrons i.e., acids are electron acceptors and bases are electron donors.

- Chemical Species which can act as Lewis Acid

  - Electron deficient neutral compounds i.e. molecules in which central atom has incomplete octet, e.g., BF_3, $FeCl_3$, $AlCl_3$, etc.

  - All simple cations, e.g., $H^+, Mg^{2+}, Ag^+, Fe^{3+}$, etc.

  - Molecules having a multiple bond between atoms of different electro negativities, e.g., SO_3, CO_2, etc.

  - In coordinate complexes, metal atoms act as Lewis acid e.g., in $Ni(CO)_,$ Ni acts as Lewis acid.

  - Molecules having a central atom with empty d-orbital's, e.g., SiF_4, $SnCl_4$.

- **Chemical Species which can act as Lewis Base**

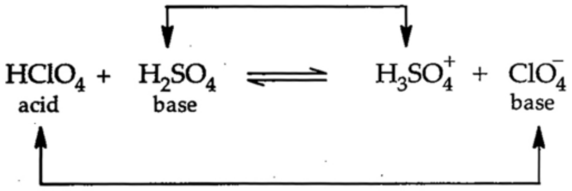
A conjugate pair of acid and base differ by a proton only.

$$Acid \rightleftharpoons Base + H^+$$

- A strong acid gives a weak conjugate base and vice versa.

- If two acids (with respect to water) are mixed together then the weaker acid will act as a base with respect to the stronger acid.

For example,



- Electron rich neutral compounds i.e. molecules with at least one lone pair of electrons, e.g., $\ddot{N}H_3, R\text{-}\underset{\cdot\cdot}{\ddot{O}}\text{-}H, R\text{-}\ddot{N}H_2$

  - All negatively charged ions, e.g., CN^-, OH^-, Cl^-, etc.

Buffer Solutions

Buffer solution is denned as a solution which resists the change in its pH value when small amount of add or base is added to it or when the solution is diluted. Buffer solution has a definite pH value at specific temperature and it does not change on keeping for a long time. Buffers are classified into two categories:

- **Simple Buffers**
  These are the solutions of salt of weak acid and weak base. For example CH_3COONH_4 (ammonium acetate), NH_4CN, etc.
- **Mixed Buffers**
  These are the mixture of two solutions. These are further of two types:
  (i) Acidic buffers
  These are the solutions of a mixture of weak acid and salt of this weak acid with strong base.
  For example $CH_3CHOOH + CH_3COONa$. They have pH value lesser than 7 i.e., pH < 7.
  (ii) Basic buffers

These are the solutions of mixture of a weak base and salt of this weak base with strong acid.

For example $NH_4OH + NH_4Cl$. They have the pH value more than 7 i.e., pH > 7.

Solution of ampholytes e.g. proteins and amino acids also serve as simple buffers.

A mixture of an acid salt and a normal salt of polybasic acid e.g., $Na_2HPO_4 + Na_3PO_4$ are also included in simple buffers.

- **pH of a Buffer Solution**

  pH value of a buffer solution is given by Henderson-Hasselbalch equation

  (i) Acidic buffer: $\boxed{\mathrm{pH} = \mathrm{pK_a} + \log\frac{[\mathrm{Salt}]}{[\mathrm{Acid}]}}$

  where K_a is dissociation constant of acid and $pK_a = -\log K_a$, [Salt] and [Acid] are molar concentrations of salt and acid respectively.

  (ii) Basic buffer: $\boxed{pOH = pK_b + \log\frac{[Salt]}{[Base]}}$

  where K_b, is dissociation constant of base and $pK_b = -\log K_b$, [Salt] and [Base] are molar concentrations of salt and base respectively. By calculating the value of pOH, pH can be determined as

  $\boxed{\mathrm{pH} = 14 - \mathrm{pOH}}$

  pH of buffer varies with temperature because value of K_w, changes with temperature.

- **Buffer Capacity (Buffer Index)**

  Buffer capacity is the number of moles of an acid or base added to change pH of one litre of buffer solution by one unit.

  Hence, Buffer capacity =

  $$\frac{\text{Number of moles of acid or base added per litre of buffer}}{\text{Change in pH}} = \frac{\mathrm{dn}}{\mathrm{d(pH)}}$$

  Buffer capacity is maximum for those buffers which have equimolar concentration of weak acid and its salt or weak base and its salt

  i.e., [Salt] = [Acid] or [Salt] = [Base]

  Hence, pH = pK_a or pOH = pK_a,

- **Blood as a Buffer Solution**

  Blood acts as a buffer solution and maintains a pH level of about 7.4. Any change in pH of blood can produce illness or even causes death.

Oxyacids

Compound in which -OH group is attached to non-metal is known as oxyacids. In these acids ions are derived from - OH group.

Oxyacids of Nitrogen

- Nitrous Acid, HNO_2

  The free acid is unknown. It is known only in solution.

  Properties:

  - Aqueous solution of nitrous acid is pale blue. This is due to the presence of nitrogen trioxide, N_2O_3 The colour fades on standing for some time.

  - It is a weak add and reacts with alkalies to form salts known as nitrites.

  $$HNO_2 + NaOH \longrightarrow NaNO_2 + H_2O$$

  - The arid is unstable and even in cold solution, it undergoes auto-oxidation.

  $$3HNO_2 \longrightarrow 2NO + HNO_3 + H_2O$$

  - **Oxidizing nature:** It acts as an oxidizing agent due to ease with which it decomposes to give nascent oxygen.

  $$2HNO2 \longrightarrow H_2O + 2NO + O$$

  Reducing nature: Nitrous arid also acts as a reducing agent as it can be oxidized into nitric acid.

  Uses: In analytical chemistry, it is used both as an oxidizing and reducing agent.

- **Nitric Acid, HNO_3**

  Chemical properties:

  - It is a very strong arid. It exhibits usual properties of acids. It reacts with basic oxides, carbonates, bicarbonates and hydroxides forming corresponding salts.

  $$CaO + 2HNO_3 \longrightarrow Ca(NO_3)_2 + H_2O$$

  $$Na_2CO_3 + 2HNO_3 \longrightarrow 2NaNO_3 + H_2O + CO_2$$

  $$NaOH + HNO_3 \longrightarrow NaNO_3 + H_2O$$

  - **Oxidizing nature:** Nitric acid acts as a strong oxidizing agent as it decomposes to give nascent oxygen easily.

  - **Oxidation of non-metal (Reaction with non-metal):**

  (i) Sulphur reacts with hot and cone. Nitric acid to give H_2SO_4 and nitrogen dioxide gas i.e., sulphur is oxidized to sulphuric acid.

  $$S + \underset{\text{conc. and hot}}{6HNO_3} \longrightarrow H_2SO_4 + 6NO_2 + 2H_2O$$

  (ii) Carbon is oxidized to carbonic arid.

  $$C + 4HNO_3 \longrightarrow H_2CO_3 + 4NO_2 + H_2O$$

(iii) Phosphorus reacts with hot and conc. nitric acid to give H3P04 and N03 i.e., it is oxidized to orthophosphoric arid.

$$2P + \underset{conc.and\ hot}{10HNO_3} \longrightarrow 2H_3PO_4 + 10NO_2 + 2H_2O$$

(iv) Iodine reacts with hot and cone. HNO_3 to give iodic acid i.e., it is oxidized to iodic acid.

$$I_2 + \underset{hot\ and\ conc.}{10HNO_3} \longrightarrow 2HIO_3 + l0NO_2 + 4H_2O$$

- Action on metals:

Mostly active metals react with arid to give hydrogen. Nitric acid cannot be used in preparation of hydrogen because it is strong oxidizing agent. It react with hydrogen (at the moment of formation) to produce any oxide of nitrogen (NO, N_2O or NO_2).

- **Metals which are above hydrogen in electro-chemical series :**

  1. Magnesium and manganese are the metals that liberate hydrogen with 6 % dilute nitric acid.

$$Mg + \underset{dilute}{2HNO_3} \longrightarrow Mg(NO_3)_2 + H_2$$

$$Mn + \underset{dilute}{2HNO_3} \longrightarrow Mn(NO_3)_2 + H_2$$

2. When metal nitrates are formed and nitric acid is reduced:

(a) Reaction of nitric acid with zinc:

Case 1: Zinc react with 6% diluted HNO_3 to produce zinc nitrate and ammonium nitrate

$$\text{Zinc} + \text{nitric acid } (6\%) =$$
$$\text{zinc nitrate} + \text{ammonium nitrate} + \text{water}$$

$$4Zn + \underset{very\ dilute}{10HNO_3} \rightarrow 4Zn(NO_3)_2 + NH_4NO_3 + 3H_2O$$

Case 2: Zinc reacts with dil. HNO_3 (20%) to form zinc nitrate and nitrous oxide (N_2O).

Zinc + nitric acid (20%) = zinc nitrate + nitrous oxide + water

$$4Zn + \underset{dilute}{10HNO_3} \rightarrow 4Zn(NO_3)_2 + N_2O + 5H_2O$$

Case 3: It reacts with conc. HNO_3 (70%) to form nitrogen dioxide.

Zinc + nitric acid (70%) = zinc nitrate + nitrogen dioxide + water

$$Zn + \underset{conc.}{4HNO_3} \rightarrow Zn(NO_3)_2 + 2NO_2 + 2H_2O$$

(b) Reaction of nitric acid with iron:

Case 1: Iron with very dilute nitric acid forms ammonium nitrate.

Iron + nitric acid (6%) = ferrous nitrate + ammonium nitrate + water

$$4Fe + 10HNO_3 \rightarrow \underset{Ferrous\ nitrate}{4Fe(NO_3)_2} + NH_4NO_3 + 3H_2O$$

Case 2: Iron with dilute nitric acid forms nitrous oxide.

Iron + nitric acid (20%) = ferrous nitrate + nitrous oxides + water

$$4Fe + \underset{dilute}{10HNO_3} \longrightarrow 4Fe(NO_3)_2 + N_2O + 5H_2O$$

Case 3: Iron with cone. HN03 forms nitrogen dioxide (NO_2).

Iron + nitric acid (70%) = ferric nitrate + nitrogen dioxide + water

$$Fe + \underset{conc.}{6HNO_3} \longrightarrow \underset{Ferric\ nitrate}{Fe(NO_3)_3} + 3NO_2 + 3H_2O$$

Iron is rendered passive by highly concentrated nitric acid 80%.

(c) Reaction of nitric acid with tin:

Case 1: Tin forms ammonium nitrate with dilute nitric acid.

Tin + nitric acid (20%) = Stannous nitrate + ammonium nitrate + water

$$4Sn + \underset{dilute}{10HNO_3} \rightarrow \underset{S\tan nous\ nitrate}{4Sn(NO_3)_2} + NH_4NO_3 + 3H_2O$$

Case 2: Tin forms metastannic acid with cone. HNO_3 and evolves nitrogen dioxide.

Tin + nitric acid (hot and cone.) = Meta stannic acid + nitrogen dioxide + water

$$Sn + \underset{hot\ conc.}{4HNO_3} \longrightarrow \underset{metas\tan nic\ acid}{H_2SnO_3} + 4NO_2 + H_2O$$

(d) Reaction of nitric acid with lead:

Case 1: It forms nitric oxide with dilute HNO_3.

Lead + nitric acid (20%) = lead nitrate + nitric oxide + water

$$3Pb + \underset{Dilute}{8HNO_3} \rightarrow 3Pb(NO_3)_2 + 2NO + 4H_2O$$

Case 2: It forms nitrogen dioxide with conc. HNO_3.

Lead + nitric acid (conc.) = lead nitrate + nitrogen dioxide + water

$$Pb + \underset{Conc.}{4HNO_3} \longrightarrow Pb(NO_3)_2 + 2NO_2 + 2H_2O$$

- **Metals which are below hydrogen in the electrochemical series**

(a) Reaction of nitric acid with copper:

Case 1: Copper with cold dil. HNO_3 forms nitric oxide (NO).

Copper + nitric acid (20%) = copper nitrate + nitric oxide + water

$$3Cu + 8HNO_3 \rightarrow 3Cu(NO_3)_2 + 2NO + 4H_2O$$

Case 2: Copper with hot cone. HNO_3 forms nitrogen dioxide (NO_3)

Copper + nitric acid (hot and conc.) = copper nitrate + nitrogen dioxide + water

$$Cu + \underset{Hot\ conc.}{4HNO_3} \longrightarrow Cu(NO_3)_2 + 2NO_2 + 2H_2O$$

(b) Reaction of nitric acid with silver:

Case 1: Silver behaves similarly as copper.

Silver + nitric acid (20%) = silver nitrate + nitric oxide + water

$$3Ag + \underset{dilute}{4HNO_3} \rightarrow 3AgNO_3 + NO + 2H_2O$$

Silver + nitric acid (conc.) = silver nitrate + nitrogen dioxide + water

$$Ag + \underset{Conc.}{2HNO_3} \longrightarrow AgNO_3 + NO_2 + H_2O$$

(c) Reaction of nitric acid with mercury:

Case 1: Mercury with dilute nitric acid forms mercurous nitrate and nitric oxide.

Mercury + nitric acid (20%) = mercurous nitrate + nitric oxide + water

$$6Hg + \underset{dilute}{8HNO_3} \rightarrow \underset{Mercurous\ nitrate}{3Hg_2(NO_3)_2} + 2NO + 4H_2O$$

Case 2: Mercury with conc. HNO_3 forms mercuric nitrate and nitrogen dioxide.

Mercury + nitric acid (conc.) = mercuric nitrate + nitrogen dioxide + water

$$Hg + \underset{conc.}{4HNO_3} \rightarrow \underset{mercuric\ nitrate}{3Hg_2(NO_3)_2} + 2NO_2 + 2H_2O$$

- **Metals which become passive**

  The concentrated nitric acid renders metals like iron, cobalt, nickel, chromium etc. The inertness exhibited by metals under conditions in which chemical activity is expected is known as passivity. For example, iron displaces copper from copper sulphate solution. This property of iron is lost if it is dipped in cone. HNO_3.

- **Metals which do not react**

  Noble metals like gold, platinum, iridium, rhodium, etc., are not acted upon nitric acid. However, these metals dissolve in aqua-regia (3 parts conc. HCl and one part conc. HNO_3). Aqua-regia forms nascent chlorine which attacks metals.

- **Gold:**

$$\left[HNO_3 + 3HCl \rightarrow \underset{Nitrosyl\ chloride}{NOCl} + 2H_2O + 2Cl\right]\times 3$$

$$[Au + 3Cl \longrightarrow AuCl_3]\times 2$$

$$[AuCl_3 + HCl] \longrightarrow HAuCl_4 \times 2$$

$$2Au + 3HNO_3 + 11\,HCl \longrightarrow \underset{Chloroauric\ acid}{2HAuCl_4} + 3NOCl + 6H_2O$$

- **Platinum:**

$$\left[HNo_3 + 3HC1 \longrightarrow NOCl + 2H_2O + 2Cl\right]\times 2$$

$$Pt + 4Cl \longrightarrow PtCl_4$$

$$PtCl_4 + 2HNO_3 \longrightarrow H_2PtCl_6$$

$$Pt + 2HNO_3 + 8HCl \longrightarrow \underset{Chloroplantic\ acid}{H_2PtCl_6} + 2NOCl + 4H_2O$$

Oxyacids of Sulphur

A large number of oxyacids are known in the case of sulphur either in Free State or in the form of salts or both. Oxyacids with S — S links are called thioacids. Acids having sulphur in lower oxidation state belong to -ous series while those having sulphur in higher oxidation state belong to *-ic* series.

For example:

Sulphurous acid ((H_2SO_3)) O.N. of sulphur, + 4.

Sulphuric acid (H_2SO_4) O.N. of sulphur, + 6.

The following are main oxyacids of sulphur:

1. Sulphurous acid series:	**Oxidation number**
(i) Sulphurous acid, H_2SO_3	+4
(ii) Thiosulphurous acid, $H_2S_2O_2$	+1
(iii) Hyposulphurous acid, $H_2S_2O_2$	+3
(iv) Pyrosulphurous acid, $H_2S_2O_7$	+4

2. Sulphuric acid series:	**Oxidation Number**
(i) Sulphuric acid, H_2SO_4	+6
(ii) Thiosulphuric acid, $H_2S_2O_3$	+5, -1
(iii) Pyrosulphuric acid, $H_2S_2O_7$	+5

3. Thionic acid series:

(i) Dithionic acid, $H_2S_2O_6$

(ii) Polythionic acid, $H_2S_nO_6$ (n = 3, 4, 5, 6)

4. Peroxy acid series:	**Oxidation number**
(i) Peroxy monosulphuric acid, H_2SO_5 (Caro's acid)	+6
(ii) Peroxy disulphuric acid, $H_2S_2O_8$ (Marshall's acid)	+6

- **Sulphuric Add (Oil of Vitriol) H₂SO₄**
  Sulphuric acid is considered as the King of chemicals. In ancient days. It was called oil of vitriol as it was prepared by distilling ferrous sulphate (Green vitriol).
  Heat is evolved when dissolved in water, usually bumping occurs. Due to this reason water should not be added to concentrated sulphuric acid for dilution but concentrated sulphuric add should be added slowly to cold water with constant stirring. Due to its great affinity for water, it is used as a dehydrating substance.
  - It decomposes carbonates and bicarbonates into carbon dioxide.
  - It reacts with electropositive metals evolving hydrogen.
  - It displaces more volatile acids forming their metal salts.

$$Ca_3(PO_4)_2 + 3H_2SO_4 \longrightarrow 3CaSO_4 + 2H_3PO_4$$

$$\underset{\text{Calcium oxalate}}{CaC_2O_4} + H_2SO_4 \longrightarrow CaSO_4 + \underset{\text{Oxalic acid}}{H_2C_2O_4}$$

- **Oxidizing nature:** It acts as a strong oxidizing agent.
  (a) Non-metals such as carbon and sulphur are oxidized to their oxides.

$$C + 2H_2SO_4 \longrightarrow CO_2 + 2SO_2 + 2H_2O$$

$$S + 2H_2SO_4 \longrightarrow 3SO_2 + 2H_2O$$

Phosphorus is oxidized to orthophosphoric acid.

$$2P + 5H_2SO_4 \rightarrow 2H_3PO_4 + 5SO_2 + 2H_2O$$

(b) Metals like copper, silver, mercury, etc. react with cone. H_2SO_4 to produce metal salt and SO_2 gas.

$$Cu + 2H_2SO_4 \longrightarrow CuSO_4 + SO_2 + 2H_2O$$

$$Ag + 2H_2SO_4 \longrightarrow Ag_2SO_4 + SO_2 + 2H_2O$$

$$Hg + 2H_2SO_4 \longrightarrow HgSO_4 + SO_2 + 2H_2O$$

- **Dehydrating nature:** H_2SO_4 act as strong dehydrating agent.

$$\underset{\text{Cane sugar}}{C_{12}H_{22}O_{11}} \xrightarrow[(-11H_2O)]{H_2SO_4} 12C$$

$$\underset{\text{Glucose}}{C_6H_{12}O_6} \xrightarrow[(-6H_2O)]{H_2SO_4} 6C$$

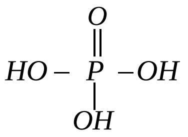
H_2SO_4 absorbs sulphur trioxide forming oleum or fuming sulphuric acid.

$$H_2SO_4 + SO_3 \longrightarrow \underset{\text{Oleum}}{H_2S_2O_7}$$

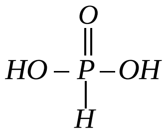
- **It is used:**
  1. In the refining of petroleum.
  2. In storage batteries.
  3. As an important laboratory reagent. In organic chemistry it is used for synthesis of various organic compounds and sulphonation reactions. It is also used as an oxidizing and a dehydrating agent.

Important Points

- AH organic acids (except those having — SO_3H, i.e., sulphonic acids) are weak acids. However halogen substituted acids are stronger one. CCl_3COOH is as stronger as HCl.
- **Mineral acids:** (H_2SO_4, HCl, HNO_3) are strong acids.
- In oxo-acids only those H-atoms are replaceable which are attached on O-atom, i.e.,



$HO-\overset{\overset{O}{\|}}{\underset{\underset{OH}{|}}{P}}-OH$, i.e., H_3PO_4 is tri basic acid.



$HO-\overset{\overset{O}{\|}}{\underset{\underset{H}{|}}{P}}-OH$, i.e., H_3SO_3 is dibasic acid.

- An add salt has at least one replaceable H-atom, $NaHSO_4, NaHS, NaHCO_3, NaHC_2O_4,$ $NaHC_2O_4,\ NaH_2PO_4, Na_2HPO_4, NaH_2PO_3$.
  NaH_2PO_3 is not an acid salt since it does not have replaceable H-atom.
- Basic character of oxides decreases along the period and increases down the group i.e., Basic character decreases along the period:
  $Na_2O > MgO > A1_2O_3 > SiO_2 > P_2O_5 > SO_2 > Cl_2O_3$
  Basic character increases down the group:
  (a) $Li_2O < Na_2O < K_2O < Rb_2O < Cs_2O$
  (b) $OF_2 < Cl_2O < Br_2O < I_2O$
- Oxides of metals are normally basic, oxides of non-metals are normally acidic. CO, N_2O and NO are neutral.

Basic oxides	Acidic oxides	Amphoteric
K_2O, CaO, MgO	$\underset{(Neutral)}{CO}$ and CO_2	ZnO, Al_2O_3, BeO, SnO_2 (All are metal oxides)
CuO, Fe_2O_3 etc.	$\underbrace{NaO, NO}_{(Neutral)}, N_2O_3,$ N_2O_4, N_2O_5	As_2O_3 (metalloid oxide)
All are metal oxides	$F2O, SiO_2, P_2O_3,$ P_2O_5, SO_2 etc.	As_2O_3 (metalloid oxide)

Notes: CO acts as an acid if allowed to react with NaOH at high P and T.

$$CO + NaOH \xrightarrow{P,T} HCOONa$$

CO acts as Lewis base (ligand) in complex formation.

(a) Oxides of non-metals having same oxidation no. of non-metal in their respective oxo-acids are known as acid anhydrides.

(b) The greater the number of oxygen atoms and the more electronegative the atom present in a molecule of Oxo-acid, the stronger the acid. i.e., $HClO_4 > HClO_3$.

Notes: However this rule is not obeyed in oxo-acids of phosphorus.

$$\begin{array}{ccccc} H_3PO_2 & > & H_3PO_3 & > & H_3PO_4 \\ 6.3\times10^{-2} & & 1.5\times10^{-2} & & 7.5\times10^{-2} \end{array}$$

- **Some Acidic Strength Order**

  1. Acidic strength mineral acid:

  $HClO_4 > HI > HBr > H_2SO_4$

  $> HCl > HNO_3 > H_3PO_4 > H_3AsO_4$

  2. Acidic strength halogen acids:

  $HI > HBr > HCl > HF$

  3. Acidic strength of hydride of oxygen family:

  $H_2Te > H_2Se > H_2S > H_2O$

  4. Acidic strength of oxacids of halogen family:

  $HOCl > HOBr > HOI$

  $HC1O_3 > HBrO_3 > HIO_3$

  $HC1O_4 > HClO_3 > HClO_2 > HClO$

  5. Acidic strength of organic acids: Acidic strength of aliphatic acid decreases with increasing number of carbon atom.

  $HCOOH > CH_3COOH > C_2CH_5OOH$

- **Some Basic Strength Order**

  1. $KOH > NaOH > Ca(OH)_2 > NH_4OH$

  2. $NaOH > NH_3 > H_2O$

  3. $LiOH < NaOH < KOH < Rb(OH) < CsOH$

- **Salient Features of pH Concept**

  - pH of a solution decreases as $[H^+]$ is solution increases.

  - For most practical purposes, the pH scale extends form 0 to 14 (at 25°C). A solution of pH = 0 is acidic and pH = 14 is alkaline.

  - The midpoint of the scale at pH = 7 represents neutrality. pH below 7 being increasingly acidic and those above 7 increasingly basic.

  - For any aqueous solution at 25° it must be true that $[H^+][OH^-] = 10^{-14}$

  No matter how acidic or basic a solution might be, it must contain H^+ and OH^- ions and the product of effective molar concentration equal to 10^{-14} or K_w.

  Also, log $[H^+]$ + log $[OH^-]$ = - 14 or,

  - log $[H^+]$ + (- log $[OH^-]$) = 14

  or pH + pOH = 14.

  - At the temperature of human body, about 37°C neutrality occurs at pH 6.8 and pH scale lies between 0 to 13.6.

  pH of some solutions

Substances	pH range
Gastric juice	1.0 – 3.0
Soft drinks	2.0 – 4.0
Lemons	2.2 – 2.4
Vinegar	2.4 – 3.4
Apples	2.9 - 3.3
Urine (human)	4.8 – 8.4
Rain water	6.0
Soda water	Less than 7.0
Milk	6.3 – 6.6
Saliva (human)	6.5 – 7.5
Blood (human)	7.3 – 7.5
Milk of Magnesia	10.5
Sea water	8.5
Black Coffee	5.0
Tomato	4.0
Tears	7.4

- **pH Value of Acidic Mixture**

  When V₁ ml acid (whose normality is N₁) and V₂ ml of other strong acid are mixed together then,

  Resultant normality, $N_R = \dfrac{N_1V_1 + N_2V_2}{V_1 + V_2}$

  pH value of resultant acidic solution $pH = \log \dfrac{1}{N_R}$

  50 mL $\dfrac{N}{10} HCl$ and 50 mL $\dfrac{N}{10} H_2SO_4$ are mixed together $N_R = \dfrac{50\times\frac{1}{10} + 50\times\frac{1}{10}}{100} = \dfrac{10}{100} = 0.1$

  $$pH = \log\frac{1}{N_R} = \log\frac{1}{0.1} = 1$$