3 ACID, BASE AND SOLUTION

Acids

An acid is any hydrogen-containing substance that is capable of donating a proton (hydrogen ion) to another substance.

As per Robert Boyle, Acid is,

1) Having a sour taste
2) Being corrosive and
3) Chemical that changes the colour of certain vegetable dyes, such as litmus.
4) The value of pH is < 7

Based on their occurrence, they are divided into two types- Natural and Mineral acids.

Natural Acids: These are obtained from natural sources, such as fruits and animal products.

e.g. lactic, citric, and tartaric acid etc.

Mineral Acids: Mineral acids are acids prepared from minerals.

e.g. Hydrochloric acid (HCl), Sulphuric Acid (H_2SO_4), and nitric acid (HNO_3) etc.

Base

A base is a molecule or ion able to accept a hydrogen ion from an acid.

Properties of base

- They change the colour of red litmus paper to blue.
- They undergo chemical reactions with acidic substances to form salts.
- They typically accept H^+ ions (or protons) from suitable donor compounds.
- They hold hydroxide ions that are either partially or completely displaceable.
- The value of pH is >7

Physiological importance of Acid and Base

An acid is any hydrogen containing substance that is capable of donating a proton (hydrogen ion) to another substance A base is a molecule or ion able to accept a hydrogen ion from an acid.

At pH at this level is ideal for may biological processes, one of the most important being the oxidation of blood. Also many of intermediates of biochemical reactions in the body becomes ionized at neutral pH, which causes the utilization of intermediates to be more difficult.

Definition of pH

The full form of pH is **Potential of Hydrogen**. pH is known as the negative logarithm of H+ ion concentration. Hence the meaning of the name pH is explained as the strength of hydrogen. pH describes the concentration of the hydrogen ions in a solution and it is the indicator of acidity or basicity of the solution. The pH value on a **pH-scale** varies from 0 to 14.

The concentration of hydronium ion is conveniently expressed on a logarithmic scale. This scale is known as the pH scale. pH of acids and bases is defined as the negative logarithm (with base 10) of activity of hydrogen ion (H+).

pH of Acid And Base

The pH of the solution has a range from 0 to 14.

- Solutions with a pH value varying from 0 to 7 on the pH scale are called acidic solutions.
- Solution with a pH value ranging from 7 to 14 is known as basic solutions.
- On a pH scale, solutions with a **potential of hydrogen** value equal to 7 are known as neutral solutions.

Solutions with a pH-value of 0 are considered to be extremely acidic. Additionally, the acidity decreases as the pH value increase from 0 to 7 while solutions with a pH value equal to 14 are known to be strongly alkaline solutions. The acid and base intensity depends on the number of H^+ and OH^- ions produced. Acids furnishing more H+ ions are known to be strong acids, and vice versa.

Importance of pH

- A living organism can withstand only a limited range of pH changes, and any more pH adjustments will make life difficult. For example: in the case of acid rain, the pH of the water is less than 7. It increases the pH of river water as it flows into a river which hinders the survival of marine life.
- The human stomach contains hydrochloric acid, which helps the digestion of the food. Because the stomach releases so much HCl we feel a lot of pain and discomfort during indigestion. It would be minimised by using antiacids.
- Often the bacteria present in our mouth lower the pH of our mouth by generating acids by food particle degradation. Therefore, by

maintaining the pH, we are told to clean our mouths with toothpaste that are basic to prevent their decay.

- In the case of bee-sting, we feel a lot of pain when the bee injects the methanoic acid through its sting. Therefore, we are usually recommended to apply baking soda or other mild bases to the surface because it helps to maintain the surface pH.

pH scale

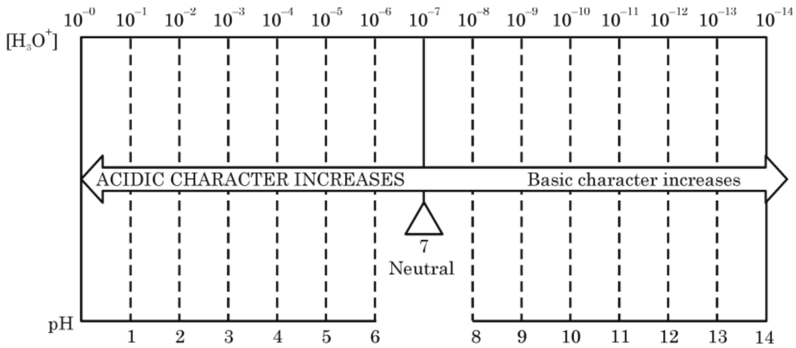
pH Scale : According to this scale, pH ranges from 0 to 14. Any solution having equal concentration of $[H_3O^+]$ and $[OH^-]$ is neutral is nature. If a solution contains either of H_3O^+ or OH^- in greater concentration, then the solution becomes acidic or basic, A solution having a concentration of $[H_3O^+$ ions more than 1×10^{-7} mol/L is said to be acidic. Similarly the solution containing the concentration of H_3O^+ ions less than 1×10^{-7} mol /L is known as basic Acidic or basic solutions can be distinguished on the basis of their pH values.

For an acidic solution, $[H_3O^+] > 10^{-7}$, hence pH < 7

For a neutral solution, $[H_3O^+] = 10^{-7}$, hence pH = 7

For a basic solution, $[H_3O^+] < 10^{-7}$, hence pH > 7

pH scale can be diagrammatically represented as :



pH Scale

pH of some common substances is given in the table as under :

Substance	pH value
1 M NaOH	14
Human Blood	7.4
Milk	6.8
1 M HCl	0
Soft drink	3.0
Vinegar	2.4 –3.4
Black coffee	5.0

Solubility Product : If a sparingly soluble salt like AgCl is stirred with water, only a small amount of it goes with solution while the most of the salt remains undissolved. But whatever, amount of salt dissolved, it gets completely dissociated into ions. In other words, there exists a dynamic equilibrium between undissolved solid salt and ions which it furnished in solution. In general, for an electrolyte $A_x B_y$,

The equilibrium may be represented as :

$$\underset{\text{(Undissolved solid)}}{A_xB_y} \rightleftharpoons \underset{\text{(ions)}}{xA^{y+} + yB^{x-}}$$

The solubility product for $A_x B_y$ may be written as :

$$K_{sp} = [A^{y+}]^x \times [B^{x-}]^y$$

where x and y represents the number of ions in the formula of electrolyte.

Thus,

Solubility **product may be** defined as the product of molar **concentrations of** its ions in a saturated **solution each raised** to a power equal to **number of ions produced** on dissociation **of one mole of** electrolyte.

For example, solubility product of $BaSO_4$ may be represented as :

$$BaSO_4\,(s) \rightleftharpoons Ba^{2+} + SO_4^{2-}$$

$$Ksp = [Ba^{2+}][SO_4^{2-}]$$

Applications of pH

1. pH tells the acidity or basicity if pH lies between 0–7 the solution is acidic and pH lies between 8 –14 solution is basic. The solution pH 7 are called neutral.
2. The concentration of solution can be checked by pH e.g if a solution of pH = 2 is diluted with water so that volume becomes double then the pH is pH = 2 means $[H_3O^+] = 10^{-2}$ M

   After dilution to double the volume $[H_3O^+] = \dfrac{10^{-2}\,M}{2} = 5 \times 10^{-3}$ M

   $$pH = -\log(5 \times 10^{-3}) = 3 - \log 5 = 3 - 0.699 = 2.3$$

   It means pH increases on dilution.
3. pH plays an important role in detecting the completion of the acid – basic reaction. Indicator changes its colour with the change in H^+ ion concentration at the end point of the acid base titration /reactions.

4. pH plays an important and crucial role in metabolism process in human body.
5. ph has important role in agriculture. Soil is often tested for determine whether acidic or basic fertilizers are required for a particular crop.
6. Food preservation also needs a definite pH value
7. It has great importance in biochemical reactions such as digestion of food. The human blood has pH 7.4. If its pH changes by 0.2 units death results.

Industrial Applications of pH

1. Hydrogen ion concentration helps in controlling the various metallurgical process in the extraction of metals and also helps in the quality control.
2. pH plays important role in pharmaceutical industry in the preparation of drugs.
3. In sugar industry also maintaining of proper pH is must for the crystallization of sugar.

pH of Different Body Fluids

Although the pH of blood ranges from 7.35-7.45, the pH of other body fluids is different. pH indicates the level of H+ ions, where low pH indicates too many H+ ions and high pH indicates too many OH- ions. If the pH levels drop below 6.9, it can lead to coma. However, different body fluids have different pH values.

The pH of saliva is ranges from 6.5 to 7.5. After swallowing, the food reaches the stomach where upper and lower parts of stomach have different pH values. The upper part has a pH of 4−6.5, while the lower part is highly acidic with a pH of 1.5−4.0. It then enters the intestine which is slightly alkaline, with a pH of 7−8.5. Maintaining the pH values of different regions is critical for their function.

Role of pH in Human System

pH balances act a role to keep the body function optimally in human. The ideal pH of the body is somewhat alkaline, which enables certain biochemical reactions like oxygenating the blood.

A pH at this level is ideal for several biological processes, one of the most significant being the oxygenation of blood. Lot of the intermediates of biochemical reactions in the body become ionized at a neutral pH that causes the consumption of these intermediates to be more difficult.

Acids and bases are significant in living things because most enzymes can do their work only at a certain level of the acidity. Cells secrete acids and bases to maintain the proper pH for enzymes to work. For example, every time you digest food, acids and bases are at work in your digestive system.

Our bodies live/die at a cellular level and the cells must maintain alkalinity in order to function and stay alive. An acidic state causes lack of oxygen at a cellular level.

A pH of < 7.4 is sub-optimal and provides the flawless environment for bacteria, mold and viruses to develop.

Electronic Concept of Oxidation and Reduction

(*a*) Oxidation: According to this concept, Oxidation may be defined as a process in which an atom or an ion loses one or more electrons. It is also known as **de-electronation.**

This loss of electron either increases the positive charge or decrease the negative charge of the atom or the ion.

For example,

(*i*) Loss of electrons results in increase in positive charge.

$$Na \rightarrow Na^{+} + e^{\ominus}$$

$$Mg \rightarrow Mg^{2+} + 2e^{\ominus}$$

$$Fe^{2+} \rightarrow Fe^{3+} + e^{\ominus}$$

$$Sn^{2+} \rightarrow Sn^{4+} + 2e^{\ominus}$$

(*ii*) Loss of electrons results in decrease in negative charge

$$MnO_4^{2-} \rightarrow MnO_4^{-} + e^{\ominus}$$

$$S^{2-} \rightarrow S + 2e^{\ominus}$$

$$2Cl^{-} \rightarrow Cl_2 + 2e^{\ominus}$$

(*b*) Reduction: According to electronic concept, reduction may be defined as the process in which an atom or ion gains one or more electrons. It is also known as **electronation.**

This gain of electron either decreases the positive charge on increases the negative charge of the atom or the ion. For example,

(*i*) Gain of electrons results in decrease in positive charge:

$$Fe^{3+} + e^{\ominus} \rightarrow Fe^{2+}$$

$$2Hg^{2+} + e^{\ominus} \rightarrow Hg_2^{+}$$

$$Sn^{4+} + 2e^{\ominus} \rightarrow Sn^{2+}$$

(*ii*) Gain of electrons results in increase in negative charge:

$$Cl_2 + 2e^{\ominus} \rightarrow 2Cl^-$$

$$MnO_4^- + e^{\ominus} \rightarrow MnO_4^{2-}$$

$$S + 2e^{\ominus} \rightarrow S^{2-}$$

Preparation of Various Standard Solutions

A standard solution is a solution with a known concentration.

Preparation of Standard Solution

(a) The mass of solute needed is calculated and weighted.
(b) The solute is dissolved in same distilled water in a beaker.
(c) The solution is transferred into a volumetric flask.
(d) More distilled water is added to obtain the required volume.

For example: If you are preparing sodium hydroxide (atomic mass of Na-23u, O-16u, H-1u) at 1 M of NaOH you need to dissolve 40 gms of NaOH is one litre of water.

Preparation of Solutions of Common Use in Laboratory

The solutions used in the laboratory for analysis have definite concentration and normalities and are prepared in distilled water. Common solutions of acids, bases and salts are given below:

(a) BASES :

Ammonium Hydroxide Solution dilute: Dilute 355 ml. of the conc. solution

(Sp. gr. 0.88, 15N) to one litre with water.

Calcuium Hydroxide (Lime water) : Shake 2-3 gm. of slaked lime with one litre of water.

Decant off or siphon the clear liquid.

Potassium Hydroxide Solution: Dissolve 310 gm. of the pure sticks

(about 90% KOH) is distilled water and dilute to one litre.

Sodium Hydroxide: Dissolve 220 gm. of ordinary (= 90%) NaOH in water and dilute to one litre.

ACIDS : Approximate Normality

Dilute Acetic Acid: Dilute 285 ml. of glacial acetic acid

(Sp. gr. 1.05, 17 N) to one litre with water.

Dilute Hydrochloric Acid: Dilute 500 ml. of conc. acid

(Sp. gr. 1.16, 10 N) to one litre with water.

Dilute Sulphuric Acid: Pour 140 ml. of the conc. acid (Sp. gr. 1.84, 36N) slowly and with constant stirring into 500 ml. of water. Cool the dilute solution to one litre with water.

Dilute Nitric Acid: Dilute 310 ml. of the conc. acid (Sp. gr. 14.2, 16 N) to one litre with water.

(b) SALT SOLUTIONS:

Ammonium Oxalate : Dissolve 35 gm. of the salt in one litre of water.

Ammonium Sulphate : Dissolve 132 gm. of the salt per litre.

Barium Chloride: Dissolve 61 gm. of the salt in one litre of water.

Bromine Water: Shake 11 ml. of liquid bromine with water and dilute to one litre.

Calcium Chloride: Dissolve 55 gm. of the commonly available hydrated salt in one litre of water.

Chlorine Water: Pass chlorine (from solid $KMnO_4$ and conc. HCl) through water till a saturated solution is obtained.

Cobalt Nitrate: Dissolve 44 gm. of the salt per litre of water.

Ferric Chloride: Dissolve 13.5 gm. of the salt in one litre of water.

Ferrous Sulphate: Dissolve 140 gm. of the hydrated salt in one litre of water containing 10ml. of conc. H_2SO_4.

Iodine Solution: Dissolve 20 gm. of solid potassium iodide in 30-40 ml. water and then add 12.7 gm. iodine. Shake well and make the volume of one litre.

Lead Acetate: Dissolve 95 gm. of the salt in 200 ml. water by heating and adding conc. acetic acid till a clear solution is obtained.

Magnesium Sulphate: Dissolve 62 gm. of hydrated salt in one litre of water.

Mercuric Chloride. Dissolve 27 gm. of the salt in sufficient hot water and dilute to one litre with water.

Potassium Chromate: Dissolve 49 gm of the salt in one litre of water.

Potassium Cyanide: Dissolve 32.5 gm. of the salt in one litre of water.

Potassium Ferricyanide: Dissolve 55 gm. of the salt in one litre of water

Potassium Ferrocyanide: Dissolve 53 gm. of the salt in one litre of water.

Potassium Iodide Solution: Dissolve 83 gm. of the salt in one litre of water

Potassium Permanganate: Dissolve 3.2 gm. of the salt in one litre of water

Potassium Thiocyanate: Dissolve 49 gm. of the salt in one litre of water.

Silver Nitrate: Dissolve 17 gm. of the salt in one litre of distilled water and store in coloured bottles

Preparation of Standard Solutions

Solutions of accurately known strength are called standard solutions. A standard solution contains a known weight of reagent in a definite volume of solution. Molecular weight and atomic weight of commonly used chemicals has been shown in Table.

Molar solution

Molar solution is one, which contains one molecular weight of the reagent in one litre of the solution. Molarity is expressed as M.

Normal solution

Normal solution is one, which contains one equivalent weight of the reagent in one litre of the solution. Normality is expressed as N.

Equivalent weight of acid = Molecular weight/ No. of replaceable H ions.

Table 3.1: Name, formula, molecular weight and equivalent weight of commonly used chemicals.

Name	Formula	Mo. wt. (g/mol)	Eq. wt. (g/mol)
Hydrochloric acid	HCl	36.5	36.5
Nitric acid	HNO_3	63	63
Potassium hydroxide	KOH	56	56
Calcium hydroxide	$Ca(OH)_2$	74	37
Potassium dichromate	$K_2Cr_2O_2$	294	49
Sodium thiosulphate	$Na_2S_2O_3.5H_2O$	248	248
Sodium chloride	$NaCl$	58.5	58.5
Potassium chloride	KCl	74.5	74.5
Iodine	I_2	254	127

Most of the synthetic dyes in generally used as indicators are organic substances of complex structure. Among the most reliable of these indicators are methyl red and phenolphathalein (Table 3.2).

Table 3.2: Common indicators used in animal nutrition laboratory

Indicator	pH range	End point	Preparation
Methyl red	4.4 to 6.3	Pink in acidic medium and colourless in basic medium	For preparing a stock solution, 0.2 g of dye is dissolved in 100 ml of alcohol and filtered
Methyl orange	2.9 to 4.0	Orange in acidic medium and colourless in basic medium	For preparing a stock solution, 0.1 g of dye is dissolved in 100 ml of distilled water, filtered and used
Phenolphthalein	8.3 to 10.0	Pink in basic medium and colourless in acidic medium	For preparing a stock solution, 0.2 g of henolphthalein is dissolved in 110 ml of alcohol and 90 ml of distilled water

There are a few standard solutions which are used for analysis of feed stuffs:

1. N/10 H_2SO_4
2. N/10 NaOH
3. N/10 $KMnO_4$
4. 0.256 N (1.25% (w/v)) H_2SO_4
5. 0.313 N (1.25% w/v) NaOH
6. 40 per cent NaCl (w/v)
7. 3 per cent KNO_3 (w/v)
8. 20 per cent ammonium molybdate (w/v)
9. 50 per cent HCl (w/v)

Certain primary standard solutions are also required for standardization of the above solutions. These are:

1. N/10 Na_2CO_3
2. N/10 $(COOH)_2 . 2H_2O$

Preparation of N/10 H_2SO_4

Equivalent weight of H_2SO_4 = 49 g
Specific gravity = 1.84 g/ml
So, volume of 49 g H_2SO_4 = 26.6 ml

Concentrated H_2SO_4 (reagent grade) is about 97 per cent pure.

Therefore, actual amount of concentrated H_2SO_4 required for 1.0 litre of N/10 H_2SO_4 solution =

$$\frac{100}{97} \times 26.6 = 27.42 \text{ ml}$$

Thus, for 1.0 litre of N/10 H_2SO_4 solution, 2.74 ml of concentrated H_2SO_4 is required.

Procedure

Take 2.74 ml sulphuric acid in a beaker half-filled with distilled water. Transfer the contents and washings to a volumetric flask (1 litre) and make volume up to the mark. Shake well and titrate this solution with 10 ml of 0.1 N Na_2CO_3 using mixed / methyl orange as an indicator. Repeat the titration to get at least three concordant readings.

Standardization

Suppose 10 ml of 0.1 N Na_2CO_3 = 9.5 ml of H_2SO_4

$$V_1N_1 = V_2N_2$$

$$10 \times 0.1 \text{ N} = 9.5 \times N_2$$

$$N_2 = 0.10526$$

To prepare 1 litre N/10 H_2SO_4, the volume of 0.10526 N acid required is $1000 \times 0.1/0.10526 = 950$ ml. Take 950 ml of 0.10526 N acid and dilute it to one litre. Check it again with N/10 Na_2CO_3 for three times. It must neutralize equal volume of N/10 Na_2CO_3 solution. Label it as 0.1 N H_2SO_4.

Precautions

Add H_2SO_4 with the help of a burette.

Never add water to an acid.

Preparation of N/10 NaOH solution

Molecular weight of NaOH	= 40
Acidity (number of replaceable OH group)	= 1
Equivalent weight of NaOH	= 40

Therefore, 4 g of NaOH dissolved in one litre of solution will give N/10 NaOH solution.

Procedure

Weigh quickly 4 g NaOH in a beaker (as it is hygroscopic) and dissolve it in distilled water (preferably CO_2 -free). Transfer the contents and the washings to a volumetric flask (1 litre). Cool and then make volume up to the mark. Shake well and standardize this solution against N/10 oxalic acid using phenolphthalein as an indicator. Label it as 0.1 N NaOH solution.

Preparation of N/10 $KMnO_4$ solution

Dissolve 3.2 g $KMnO_4$ in one litre of distilled water. Boil it for 10-15 minutes and then allow to stand for few days and then filter it through glass wool.

Take 10 ml of N/10 oxalic acid in a beaker. Add 5 ml dilute sulphuric acid, warm it to 60-70ºC and titrate against $KMnO_4$ from the burette till a light pinkish colour appears. Take three concordant readings.

Suppose 10 ml 0.1 N oxalic acid = 9.75 ml of $KMnO_4$

$$V_1N_1 = V_2N_2$$

$$10 \times 0.1\ N = 9.75 \times N_2$$

$$N_2 = \frac{10 \times 0.1\ N}{9.75} = 0.10256$$

To prepare 1000 ml 0.1 N $KMnO_4$ solution the volume of $KMnO_4$ will be taken.

$$\frac{100 \times 9.75 \times 0.1}{10 \times 0.1}$$

Now take 975 ml of prepared $KMnO_4$ solution and make it 1000 ml by adding distilled water.

Note:

Ordinary or even pure distilled water contains traces of organic matter which reduces the $KMnO_4$ solutions. That is why the solution is boiled and kept for some time before standardization.

In the absence of sufficient amount of dilute H_2SO_4 or due to the rapid addition of $KMnO_4$ in titration flask, brown turbidity (manganous oxide) may appear.

Preparation of N/10 Na_2CO_3 solution

Molecular weight of Na_2CO_3 = 106

$Na_2CO_3 = 2\ HCl = 2\ NaCl + H_2O + CO_2$

So, acidity of Na_2CO_3 = 2

Equivalent weight of Na_2CO_3 = 53

Therefore, 5.3 g Na_2CO_3 is required for each litre of solution to make N/10 Na_2CO_3. Na_2CO_3 is hygroscopic, therefore, it must be made perfectly anhydrous before it is weighed out. Quantity of acid/ alkali required for preparation of different molar/normal solutions has been shown in Table 3.3.

Table 3.3 : Acid/ alkali required for preparation of different normal solutions

Normality	95% NaOH (g/l), M=N	36% HCl (ml/l), M=N	98% H_2SO_4 (ml/l), M=2N	<68% HNO_3 (ml/l), M=N
0.1 N	4.2	8.5	2.7	6.5
0.5 N	21.0	42.6	13.7	32.8
1.0 N	42.1	85.3	27.4	65.7
1.5 N	63.1	127.9	41.1	98.5
2.0 N	84.2	170.6	54.8	131.4
2.5 N	105.2	213.2	68.5	164.2
5.0 N	210.5	426.5	137	328.5
10.0 N	421	853	274	657

M = Molar; N = Normal

Procedure

Take 6-7 g of Na_2CO_3 (A.R.) in a nickel crucible and heat it in a hot air oven at about 100ºC for few hours so as to drive out any moisture and to convert any moisture and to convert any preformed $NaHCO_3$ to Na_2CO_3. Cool in a desiccator and weigh exactly 5.3 g dried salt and dissolve it in a little quantity of freshly boiled distilled water. Transfer it to one litre measuring flask and make volume up to the mark. Shake well and label it as 0.1N Na_2CO_3 solution.

Preparation of N/10 oxalic acid

Oxalic acid $(COOH)_2 .2H_2O$ is to be dissolved in one litre of distilled water to get N/10 oxalic acid solution.

Procedure

Weigh accurately 6.3 g $(COOH)_2.2H_2O$ and transfer it to a volumetric flask (1 litre), half-filled with distilled water. Shake well and make the volume up to the mark. Label it as N/10 oxalic acid solution.

Note: If anhydrous oxalic acid $(COOH)_2$ is available then dissolve 4.5 g of the acid in one litre of distilled water to get 0.1 N oxalic acid solution.

Preparation of standardized 0.313 N (1.25%) NaOH solution

Add 13.16 g of NaOH (95% NaOH) in one litre distilled water and shake well. Standardize this solution against known concentration of oxalic acid solution using phenolphthalein as an indicator.

Preparation of standardized 0.256 N (1.25per cent (w/v) H_2SO_4 solution

To prepare 1.25 per cent (w/v) H_2SO_4 solution, 12.5 g of H_2SO_4 (100 per cent) is to be added to distilled water to make the volume 1000 ml.

$$\text{Volume of } H_2SO_4 \text{ be taken} = \frac{12.5 \times 100}{1.84 \times 97} = 7 \text{ ml}$$

Procedure

Add 7.0 ml concentrated H_2SO_4 (specific gravity 1.84 g/ml and 97per cent concentration) in a 1000 cc volumetric flask half-filled with distilled water. Shake well and add distilled water to make volume up to the mark. Standardize this solution against known concentration of Na_2CO_3 using mixed/methyl orange indicator.

Precaution and preservation of standard solutions

The bottle must be kept tightly stoppered to prevent evaporation of solvent. Some solutions must be protected from atmospheric gases. For example, sodium hydroxide solution is affected by atmospheric CO_2.

$$2NaOH + CO_2 = Na_2CO_3 + H_2O$$

But never put a glass stopper on NaOH solution container because NaOH will react with air and glass between stopper and neck of volumetric flask. It will fix permanently and you cannot remove the glass stopper from volumetric flask. $KMnO_4$ solution should be preserved in colour (amber) bottles. The container should be shaken well before the withdrawal of a portion of solution to ensure uniform composition of the solution.

Standard Solution

A standard solution *is a solution with a known concentration*.

In chemistry, a solution is a mixture of two or more substances or elements in which neither substance or element chemically changes. For example, salt water is a solution that contains water (H_2O) and salt (NaCl). Concentration is the amount of dissolved substance (solute) in a solution. In other words, it is the amount of stuff that has been mixed into your liquid. Concentration is usually measured in molarity (M), which is the number of moles of solute per liter.

In a *standard solution*, you know precisely what the concentration of the solute is. Successfully preparing a *standard solution* requires care, because without an accurate measurement, you can't accurately compare the *standard solution* with an unknown solution.

Importance of Standard Solutions

Standard solutions are used for comparison purposes when analyzing unknown solutions. *Standard solutions* can also be added to an unknown solution until a reaction occurs in order to gather information about the unknown solution. For example, you can add a weak base to an unknown acid until the solution changes color in order to determine the unknown solution's exact pH.

Types of Standard Solutions

There are two types of standard solutions known as primary solution and secondary solution.

(a) Primary Standard Solution:

Primary standard solutions the solutions made form primary standard substances. These substances have a high purity which nearly equals 99.9% purity. We can dissolve this substance in a known volume of solvent in order to obtain the primary standard solution. These solutions can involve chemical reactions. Therefore, we can use this reagent to determine the unknown concentration of a solution that undergoes a particular chemical reaction.

Primary standard solution example:

Examples of primary standards for the titration of solutions, based on their high purity, are:

- Standardization of sodium thiosulfate ($Na_2S_2O_3$) solution with potassium bromate ($KBrO_3$).
- Sodium carbonate for the standardization of aqueous acids such as sulfuric acid (H_2SO_4), nitric acid (HNO3), and hydrochloric acid (HCl) solutions.
- Zinc powder (Zn) is used to standardize EDTA solutions.
- Sodium chloride (NaCl) is used as a primary standard for silver nitrate ($AgNO_3$) reactions.
- Sodium carbonate (Na_2CO_3), potassium dichromate ($K_2Cr_2O_7$), and potassium hydrogen phthalate (KHP) are some examples of primary standards.

(b) Secondary Standards Solutions

Secondary standards solutions are the solutions made from secondary standard substances. We prepare these solutions for a specific analytical experiment. We should determine the concentration of these solutions using primary standards. Most of the times, these solutions are useful for the calibration of analytical instruments.

However, the purity of these solutions is less compared to primary standards and the reactivity is high. Due to this high reactivity, these solutions get contaminated easily. Some common examples are anhydrous sodium hydroxide and potassium permanganate. These compounds are hygroscopic.

Difference between Primary and Secondary Standard Solution

Purity

Primary Standard Solution: Primary standard solutions are extremely pure (about 99.9%).

Secondary Standard Solution: Secondary standard solutions are not very pure.

Reactivity

Primary Standard Solution: Primary standards are less or not reactive.

Secondary Standard Solution: Secondary standards are reactive than primary standards.

Water Absorption

Primary Standard Solution: Primary standards are not hygroscopic.

Secondary Standard Solution: Secondary standards are somewhat hygroscopic.

Applications

Primary Standard Solution: Primary standard solutions are used to standardize secondary standards and other reagents.

Secondary Standard Solution: Secondary standard solutions are used for specific analytical experiments.

SI Units of Measurement:

Strength of solution: Strength of a solution refers to the amount of solute in grams present in one litre of solution. **It is expressed in g/litre.** If '*a*' gram of solute is present in 'V' ml of a given solution, then,

$$\boxed{\text{Strength of solution } = \frac{a}{\text{V}} \times 1000 \text{ g/litre}}$$

Molarity (M) of solution : It may be defined as the number of moles of solute dissolved per litre of given solution. **It is represented by 'M'**. It is expressed as mol/litre.

$$\boxed{\text{Molarity (M)} = \frac{\text{No. of moles of solute}}{\text{Volume of solution in litres}}}$$

If '*a*' gram of solute is present is 'V' ml of a solution, Then,

$$\boxed{\text{Molarity (M)} = \frac{a \times 1000}{\text{V} \times \text{Molecular Mass}}}$$

$$\text{No. of moles of solute} = \frac{\text{Given Mass of solute}}{\text{Molecular mass of solute}}$$

Molality (*m*) : It may be defined as the number of moles of solute dissolved per kg (1000 g) of the solvent. **Its unit is mol/kg**.

$$\text{Molality } (m) = \frac{\text{No. of moles of solute}}{\text{Mass of solvent (in kg)}}$$

$$= \frac{\text{No. of moles of solute}}{\text{Mass of solvent in grams}} \times 1000$$

If '*a*' grams of solute is present in '*b*' grams of solvent,

$$\text{Then, molality } (m) = \frac{a \times 1000}{\text{molecular mass } \times b}$$

Mass fraction : It may be defined as the ratio of mass of one component to total mass of all the components present in the solution. For a binary solution containing the components A and B,

$$\text{Mass fraction of A} = \frac{\text{Mass of A}}{\text{Mass of A} + \text{Mass of B}} = \frac{W_A}{W_A + W_B}$$

$$\text{Mass fraction of B} = \frac{\text{Mass of B}}{\text{Mass of} + \text{Mass of B}} = \frac{W_B}{W_A + W_B}$$

Mass fraction has no units.

Mole fraction : It may be defined as the ratio of no. of moles of one component to total number of moles of all component present in the solution.

For a binary solution containing A and B,

$$\text{Mole fraction of A } (x_A) = \frac{n_A}{n_A + n_B}$$

$$\text{Mole fraction of B } (x_B) = \frac{n_B}{n_A + n_B}$$

The sum of mole fractions of all the components present in a solution is always equal to one

i.e. $$x_A + x_B = \frac{n_A}{n_A + n_B} + \frac{n_B}{n_A + n_B} = 1$$

Mole fraction being a ratio, has no units.

Normality : The number of gram equivalents of the solute dissolved per litre of the solution is known as normality . **It is denoted by N.** Mathematically gram equivalent is given by

$$\text{gram equivalent} = \frac{\text{Mass of the substance in grams}}{\text{Equivalent mass}}$$

and normality is given by

$$N = \frac{\text{Strength of solution in gm/litree}}{\text{equivalent mass of solute}}$$

Molar solution: A solution containing one gram mole of the solute dissolved per litre of the solution is known as molar solution.

Use of Standard Solution

Standard solutions are used in analytical chemistry to determine the concentration and/or content of a substance. They have several scientific, medical, and industrial applications.

The Uses of Standard Solutions in Industry

In **industry standard** solutions are primarily employed for quality control and environmental safety compliance purposes.

Standard solutions are often used to calibrate the accuracy and precision of chemical monitoring instruments. For instance, they can help ensure the apparatus responsible for measuring calories and electrochemical and turbidity parameters are working correctly. Standard solutions can also determine or rule out the presence of certain pollutants in water. As a result. they're commonly used to analyse drinking water and wastewater,

Importance of using a Standard Solution

Standard solutions can be used for both the **qualitative** and **quantitative analysis** of substances.

They're a crucial element in titration experiments (more on those below), providing a simple but reliable means of determining concentrations and chemical species. At the very least, standard solutions can help to rule out certain chemical species.

Standard solutions also allow experiments to be repeated easily. Scientists can independently confirm certain experiments that use standard solutions, thus helping to standardise the design and methodology.

Advantages of using Standard Solutions

One of the main advantages of using standard solutions is that they simplify the methodology and instruments used in an experiment. A simple burette and flask are easy to set up and use for titration, for example.

Another advantage of using standard solutions is the relatively low cost. In most cases, there is no need to use expensive instruments when performing experiments that involve standard solutions.

The Use of Standard Solutions in Titrations

A standard solution is used as a titrant in **titration experiments**. It's gradually added via a burette to an analyte solution and slowly allowed to drip down until an endpoint of reaction is reached.

EXERCISE

1. What do you mean by Acid?
2. What do you mean by Base?
3. Write Properties of Acid.
4. Write Properties of Base.
5. Explain Oxidation and Reduction.
6. What is the importance of base and acid?
7. Why is pH important to the human body?
8. How does the pH level affect the human body?
9. Define Solution.
10. What are types of standard solution?
11. What is meant by primary standard solution?
12. What is a secondary standard solution?
13. What are the uses of standard solution?
14. What is the SI unit of standard solution?

❑❑❑