

AN INITIATIVE BY IITian

MODULUS
STUDY MATERIAL

TARGET YOUR JEE WITH IITians

CHEMISTRY

Acids, Bases and Salts

www.modulushyd.com

Ph. 7 99 33 74 333 , 7 99 33 75 333 , 7 99 33 76 333

**Do you know how much percentage of marks
a student has to score in IITJEE advanced exam to get
a rank in top 100 ?
a seat getting rank in IITs ?**

Unbelievably they are just 75% and 35% (avg of last 10 years data)
Many students who score 10/10 in boards are not able to score 35% in
JEE advanced but an average student can make it to IIT if the coaching institute has

- ✓ Expert faculty - Excellent teaching
- ✓ Meticulous planning - Long-term strategy
- ✓ Constant motivation - Continuous support
- ✓ Rigorous practice - Accuracy and Speed

Previous Years JEE ADVANCED Rank VS % of Marks

RANK	2010	2011	2012	2013	2014	2015	2016	2017	2018	2019	2020	Avg %	2021 (M. Marks 360)		2021 (M. Marks 360)	
													Marks	Rank	Marks	Rank
1	85.48	91.67	96.01	92.22	92.78	93.06	86.02	92.62	93.61	93.01	88.89	91.40	348	1	196	1544
10	81.39	85.42	88.48	85.83	86.94	83.92	78.49	86.88	86.11	82.80	81.06	84.99	324	8	185	2043
100	70.76	79.79	78.68	76.39	77.50	69.84	68.61	83.33	74.72	72.85	71.72	74.93	323	13	178	2395
500	62.17	72.08	69.36	66.67	68.61	60.32	54.57	77.32	64.44	61.02	59.34	65.08	313	25	160	3542
1000	57.87	68.13	64.46	61.67	63.33	56.75	49.46	72.95	58.59	55.65	53.28	60.22	310	27	154	4137
2000	52.97	62.50	58.33	55.56	57.22	49.21	43.55	67.48	53.06	50.00	46.46	54.21	302	46	145	5005
3000	49.69	59.17	54.41	51.94	53.61	45.44	42.74	63.38	49.17	46.77	42.42	50.79	297	66	135	6180
4000	47.03	56.67	51.72	49.17	51.11	42.66	37.90	60.38	46.39	44.09	39.39	47.86	296	70	131	6750
5000	44.99	54.38	49.51	47.22	48.89	40.48	36.02	57.65	43.89	41.94	37.12	45.64	286	94	126	7516
6000	43.35	52.17	47.55	45.56	47.22	38.69	34.41	55.45	41.94	40.32	35.10	43.85	266	201	121	8494
7000	41.92	51.04	45.83	43.89	45.56	37.10	33.33	53.55	40.28	38.71	33.59	42.25	254	299	109	10848
8000	40.49	49.79	43.38	42.50	44.44	35.91	31.99	51.63	38.61	37.63	32.32	40.79	245	395	102	12874
9000	39.47	48.54	43.14	41.11	43.06	34.52	30.91	50.27	37.22	36.29	30.81	39.58	239	471	100	13215
10000	38.85	47.71	41.91	40.00	41.94	33.33	29.84	48.90	36.11	35.22	29.80	38.51	231	600	85	18547
QUAL%	38.85	47.71	34.55	33.88	35.00	23.81	20.16	22.50	25.00	25.00	17.42	29.44	212	1006	79	21157

**For query Call : 7993376333, 7993375333, 7993374333
Visit : www.modulushyd.com**

Chapter 5

Acids, Bases and Salts



Remember

Before beginning this chapter, you should be able to:

- review the methods of preparation, properties, and uses of acids and bases.
- review the types of salts, their methods of preparation, properties and uses.
- understand neutralisation reaction.

Key Ideas

After completing this chapter, you should be able to:

- classify the acids and bases based on different properties and calculate the equivalent masses of acids and bases.
- learn about the preparation of standard solution and estimation of volumes of solutions (volumetric analysis).
- study the self-ionisation of water and the ionic product of water.
- understand about pH (the hydrogen ion concentration of solutions), calculation of pH and its importance in daily life.
- identify the classification of salts based on the characteristics.
- understand about the laboratory preparation of specific soluble and insoluble salts.
- understand the application of the concepts of acids and bases in qualitative analysis of simple salts.

INTRODUCTION

Acids and bases are two important classes of compounds that we come across in chemistry. Initially, the substances that are sour to taste were termed acids. In fact, the name acid was derived from a Latin word 'acidum' meaning 'sour.' Similarly, bases or alkalies were characterised by their bitter taste and their slippery touch.

The earliest idea regarding acids was that all acids necessarily contain oxygen as one of the elements. Later, further researches proved that the acids are of two categories. The first category of acids called hydracids contains only hydrogen and a non-metal, for example, HCl. The second category of acids that contain oxygen in addition to hydrogen and a non-metal is called oxyacids for example, H₂SO₄. These two types of acids are together called mineral acids as they originate from mineral sources in contrast to the organic acids that have their origin from plant and animal sources.

Acids and bases were basically distinguished by their response to litmus test. Acids turn blue litmus red and bases turn red litmus blue. 'Salts' is another important category of chemical compounds. The term salt is generally used in our daily life for common salt, which is an important component of our diet. In the language of chemistry, however, the term 'salts', covers all electrovalent compounds having positive and negative radicals formed by the reactions of corresponding acids and bases. Therefore, the study of salts is invariably associated with the study of acids and bases.

THEORIES OF ACIDS AND BASES

Svante Arrhenius, for the first time in 1884, proposed a systematic theory to explain the nature of acids and bases.

Arrhenius Theory of Acids and Bases

Postulates

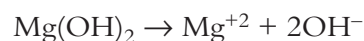
1. An acid is a substance that contains hydrogen and produces H⁺ ions in an aqueous solution.

Examples:



2. A base is a substance that produces OH⁻ ions in the aqueous solution.

Examples:

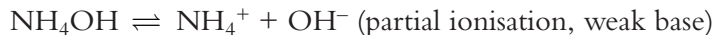


3. All acids that ionise completely and hence, produce more concentration of H⁺ ions are called **strong acids**. All acids that ionise partially and hence, produce less concentration of H⁺ ions are called **weak acids**.

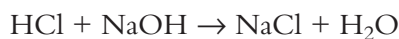
Examples:



4. All bases that ionise completely and hence, produce more concentration of OH⁻ ions in their aqueous solutions are called **strong bases**. All bases that ionise partially and produce less concentration of OH⁻ ions are called **weak bases**.

Examples:

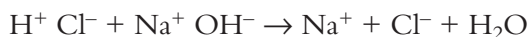
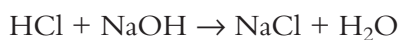
5. Neutralization is defined as the reaction between an acid and a base to give salt and water.

Example:

Acid Base Salt

Since the strong acid and the strong base are completely ionised in aqueous solutions, they will be present in the form of ions. During the reaction, the anion of acid and cation of the base do not undergo any net change.

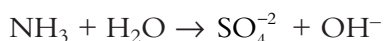
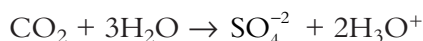
Therefore, the net reaction would be formation of water by the combination of H^+ and OH^- ions. The other ions that remain unchanged during the reaction are called spectator ions.

Examples:

The net reaction of neutralisation becomes $\text{H}^+ + \text{OH}^- \rightarrow \text{H}_2\text{O}$

After proposing his theory of electrolytic dissociation, Arrhenius himself gave a modified version of his theory.

According to this version, the acid is defined as a substance that increases the concentration of H^+ ions or hydronium ions in aqueous solutions. The base is defined as a substance that increases the concentrations of OH^- ions in their aqueous solutions.

Examples:**Limitations of Arrhenius Theory**

1. Arrhenius theory could not explain the acidic nature of substances, which do not contain hydrogen. For example, non-metal oxides are generally acidic, but they do not contain hydrogen.

Examples: CO_2 , SO_3 , etc.

2. Bases, which do not contain OH^- ions also could not be explained by Arrhenius concept. For example, metal oxides are generally basic, but they do not contain OH^- ions.

Examples: CaO , MgO , etc.

3. Arrhenius theory was confined only to aqueous solutions. It could not explain the nature of acidic and basic substances in gas phase or in non-aqueous solutions.

4. Arrhenius theory failed to explain the neutralisation reactions that do not involve formation of water.

Example: $\text{CaO} + \text{CO}_2 \rightarrow \text{CaCO}_3$

5. Later studies on solutions have revealed that H^+ ions have no independent existence in aqueous solutions. They exist as hydronium ions (H_3O^+). This is in contradiction to Arrhenius theory.

In order to account for these drawbacks, a more advanced theory was proposed, which introduced the protonic concept of acids and bases.

Brønsted-Lowry Theory of Acids and Bases

In 1923, J.N. Brønsted and J.M. Lowry independently proposed a broader concept of acids and bases in order to overcome some of the drawbacks of Arrhenius theory.

Postulates

1. The acid is a substance that donates protons to any other substance(s).

Example:

HCl acts as a proton donor and hence, is called an acid.



2. The base is a substance that accepts protons from any other proton donors.

Example:

NH₃ acts as a base as it is a proton acceptor.



3. An acid on losing a proton produces a species that has proton-accepting tendency that means the tendency of a base. The base so formed is called conjugate base of the acid.

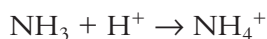
Example:



Acid Conjugate base

4. The base on accepting a proton produces a species that has proton-donating tendency that means the tendency of an acid. The acid so formed is called conjugate acid of the base.

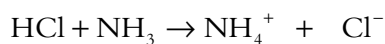
Example:



Base Conjugate acid

5. Neutralisation is defined as the transfer of a proton from the acid to the base to form a conjugate acid–base pair. As this theory mainly focuses on proton transfer, this is also known as protonic concept of acids and bases.

Example:



Acid Base Conjugate Conjugate
 acid base

Merits

1. The acidic and basic nature of substances in gas phase could be convincingly explained by this theory, which could overcome the major limitations of Arrhenius theory.
2. It involves a broader coverage of acids and bases in comparison to Arrhenius theory. This is because it is not confined to aqueous solutions alone.
3. The basic nature of substances that do not contain OH⁻ ions could also be explained on the basis of protonic concept.

Limitations

1. The acidic nature of substances that do not contain hydrogen cannot be explained.
Examples: CO₂, SO₂, etc.
2. This theory mainly involves proton transfer. Proton donation or acceptance being a relative tendency of substances, cannot give the absolute acidity or basicity of substances.

In order to explain the nature of acids and bases in cases where there is no role for protons, a more advanced theory was proposed that introduced the electronic concept of acids and bases.

Lewis Theory of Acids and Bases

In 1930, G.N. Lewis introduced another theory of acids and bases that could successfully explain the cases where there is no role for protons.

Postulates

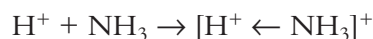
1. The acid is a substance that can act as an electron pair acceptor. The molecule or ion should be associated with shortage of electrons to act as the acid.

Examples: BF₃, H⁺, etc.

2. The base is a substance that can act as an electron pair donor. The molecule or ion should be associated with one or more lone pairs of electrons to act as the base.

Examples: NH₃, OH⁻, etc.

3. Acid–base neutralisation can be defined as the transfer of an electron pair from the base to the acid.

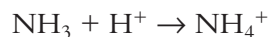


Formation of a hydronium ion and an ammonium ion can be explained by means of Lewis theory of acids.

Formation of hydronium ions (H₃O⁺)



Formation of ammonium ion



Merits

1. This theory could explain the acidic nature of those substances that do not contain hydrogen molecules.
2. It could explain the neutralisation reactions that do not involve any transfer of protons.

Limitations

1. This theory failed to explain the acidic and basic nature of general acids and bases.
2. This theory could not explain the general neutralisation reactions involving the formation of salt and water.
3. Generally, the acid–base reactions are fast reactions. However, the formation of a co-ordinate bond is a slow process.
4. This theory cannot explain the relative strengths of acids and bases.

The three different theories of acids and bases explain the nature of various acids and bases in different perceptions. However, none of the theories could convincingly explain all the properties of acids and bases and hence, cannot be universally applicable. Depending on the nature and conditions of a reaction, an appropriate theory is made applicable. Thus, all the three theories put together can cover the entire range of acids and bases.

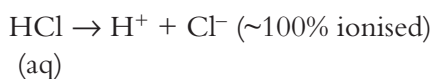
Acids and bases can be further classified on the basis of certain properties as discussed below.

Classification of Acids

1. On the basis of strength

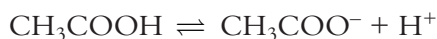
- (a) **Strong acids:** The acids that undergo nearly 100% ionisation in aqueous solutions are called strong acids. These produce a high concentration of H^+ ions in the solution.

Example:



- (b) **Weak acids:** Acids that undergo partial ionisation in aqueous solutions are called weak acids. These produce low concentration of H^+ ions in the solution.

Example:

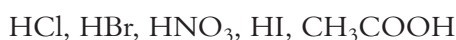


2. On the basis of basicity

Number of H^+ or H_3O^+ ions produced by ionisation of one molecule of an acid in aqueous solution is called its basicity.

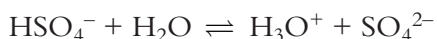
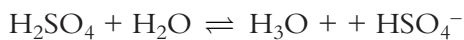
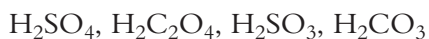
- (a) **Monobasic acids:** The acids that dissociate in aqueous solutions to produce one hydronium ion per molecule of the acid are called monobasic acids.

Examples:



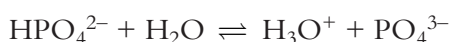
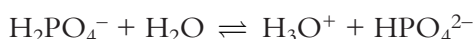
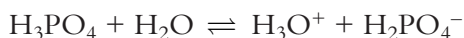
- (b) **Dibasic acids:** The acids that dissociate in aqueous solutions to produce two hydronium ions per molecule of the acid are called dibasic acids.

Examples:



- (c) **Tribasic acids:** The acids that dissociate in aqueous solution to produce three hydronium ions per molecule of the acid are called tribasic acids.

Examples:



3. On the basis of concentration

- (a) **Concentrated acid:** An aqueous solution that has high percentage of acid and low percentage of water is said to be a concentrated acid.
- (b) **Dilute acid:** An aqueous solution that has low percentage of acid and high percentage of water is said to be a dilute acid.

4. On the basis of volatility

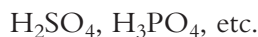
Acids that vaporise easily below 100°C are volatile acids.

Examples:



Acids that do not vaporise even on strong heating are non-volatile acids.

Examples:

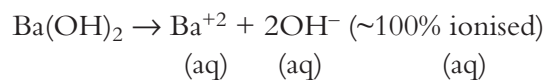


Classification of Bases

1. On the basis of strength

- (a) **Strong bases:** Bases that undergo nearly 100% (complete) ionisation in aqueous solutions. These produce high concentration of OH^- ions in the solution.

Examples:



- (b) **Weak bases:** Bases that undergo partial ionisation in aqueous solutions. It produces low concentration of OH^- ions in the solution.

Example:



2. On the basis of acidity

Number of OH^- ions produced by the ionisation of one molecule of the base in the aqueous solution is called its acidity.

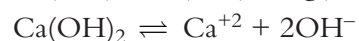
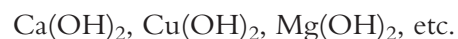
- (a) **Monoacidic bases:** The bases that dissociate in the aqueous solution to produce one hydroxyl ion per molecule of the base are called mono-acidic bases.

Examples:



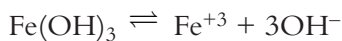
- (b) **Diacidic bases:** The bases that dissociate in aqueous solutions to produce two hydroxyl ions per molecule of the base are called di-acidic bases.

Examples:



- (c) **Triacidic bases:** Triacidic bases are those that dissociate in the aqueous solutions to produce three hydroxyl ions per molecule of the base.

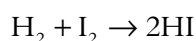
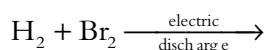
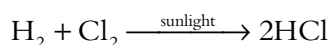
Examples:



General Preparation Methods of Acids

1. By synthesis/direct combination (for hydric acids and binary acids)

Hydrogen + Non-metal \rightarrow Acid



(Vapour)

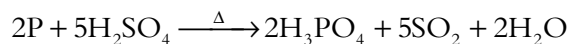
2. By the reaction of non-metallic (acidic) oxides with water.

Non-metallic oxide + Water \rightarrow Acid



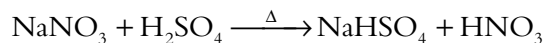
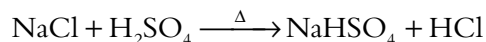
3. By the oxidation of non-metals with oxyacids

Non-metal + Acid \rightarrow Acid + Water + Oxide



4. By the displacement of salts of more volatile acids with less volatile acid.

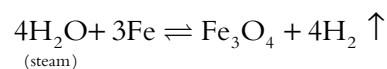
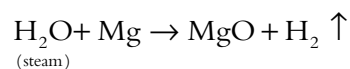
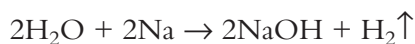
Salt + Less volatile acid \rightarrow Salt + More volatile acid



General Preparation Methods of Bases

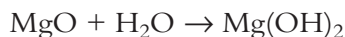
1. By the action of water on metals

**Water + Metal \rightarrow Base/Alkali + Hydrogen \uparrow
(cold)**



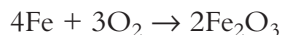
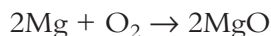
2. By the action of metallic oxides on water

Metallic oxide + Water → Base



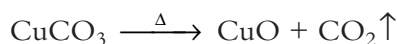
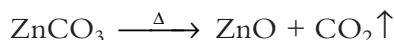
3. By the action of oxygen on metals

Metal + Oxygen → Metallic oxides



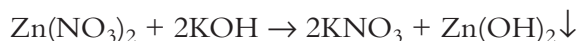
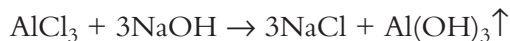
4. By the decomposition of metal carbonates

Metal carbonate $\xrightarrow{\Delta}$ Basic oxide + Carbon dioxide



5. Using soluble metal salts and NaOH/KOH

Soluble metal salt + NaOH/KOH → Salt + Insoluble metal hydroxide



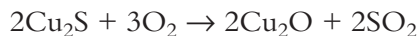
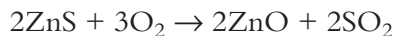
6. Strong heating of metal nitrates

Metal nitrate $\xrightarrow{\Delta}$ Metal oxide + NO₂ + O₂↑

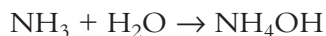


7. By the action of oxygen on metal sulphides

Metal sulphide + Oxygen → Metal oxide + Sulphur dioxide



8. Dissolution of NH₃ in H₂O

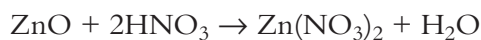
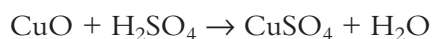


(weak alkali)

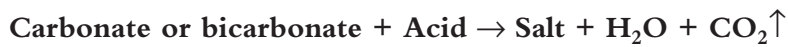
Chemical Properties of Acids

1. Action with metallic oxides

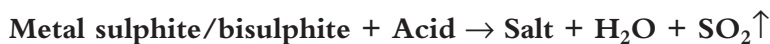
Metal oxide + dilute acid → Salt + Water



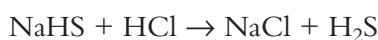
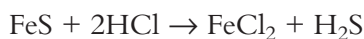
2. Action with metallic carbonates and bicarbonates



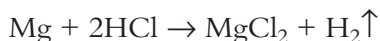
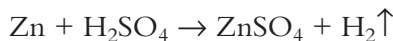
3. Action with metallic sulphite and bisulphite



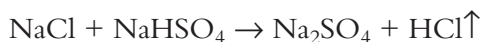
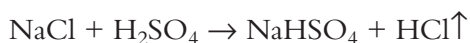
4. Action with metallic sulphide and bisulphide



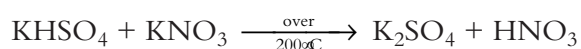
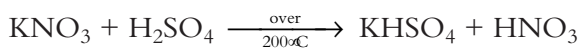
5. Action of active metals with dilute acids



6. Reaction with chlorides

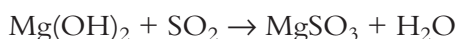
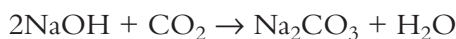


7. Reaction with nitrates

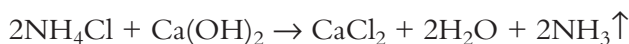
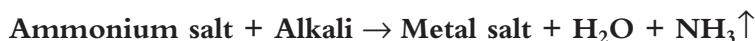


Chemical Properties of Bases

1. Action with non-metallic oxides



2. Action with ammonium salt



3. Action with soluble salts of metals other than that of Na, K and Ca

Soluble metallic salt + Base \rightarrow Soluble salt + Insoluble hydroxide

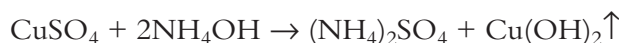


TABLE 5.1 Uses of acids and bases

Some important acids	Uses
(i) Sulphuric acid	Manufacture of fertilisers, dyes, drugs, explosives, paints, chemicals, artificial fabrics, etc.
(ii) Hydrochloric acid	Printing, tanning and tinning industries
(iii) Nitric acid	Manufacture of dyes, explosives, drugs, fertilisers and chemicals
(iv) Acetic acid	Preservation of food and as a flavouring agent
(v) Citric acid	Curdling of milk, flavouring of soft drinks and as a (source of vitamin C) medicine
(vi) Carbonic acid	Used in soft drink
(vii) Oxalic acid	Used as an ink stain remover
(viii) Boric acid	Preservation of grains, to wash eyes
(ix) Benzoic acid	As a food preservative
(x) Tartaric acid	Constituent of baking powder

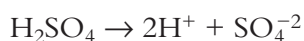
Some important bases	Uses
(i) NaOH	Manufacture of soap
(ii) KOH	Manufacture of soft soap and in alkaline batteries
(iii) NH_4OH	Used to remove grease stains from woollen clothes
(iv) $\text{Ca}(\text{OH})_2$	Manufacture of mortar, softening temporary hard water, neutralizing acidity of soil and as a general germicide
(v) $\text{Mg}(\text{OH})_2$	Used as an antacid for neutralizing gastric acidity
(vi) $\text{Al}(\text{OH})_3$	Used as a foaming agent in fire extinguishers

NEUTRALISATION

It is the reaction of acid and base to form salt and water and is associated with the liberation of heat. The relative quantities of acid and base undergoing neutralisation reaction and hence, the heat liberated depends upon the concentration of H^+ and OH^- produced by the dissociation of the acid or the base in a given reaction. Since the acids and bases differ in their basicities and acidities, the concentration of H^+ and OH^- ions furnished by them also differs. Therefore, the quantities of acids and bases are expressed in terms of equivalents that represent quantities of substances that have same combining capacities in chemical reactions.

In a neutralisation reaction, one equivalent of the acid is the quantity of the acid that furnished one mole of H^+ ions and for the base one mole of OH^- ions.

Example:



1 mole $\text{H}_2\text{SO}_4 \rightarrow 2$ moles of H^+ ions

1 mole of H^+ ions $\rightarrow \frac{1}{2}$ mole H_2SO_4

= 1 equivalent H_2SO_4

For all monobasic acids and monoacidic bases, the equivalent weight is equal to the molecular weight.

Example:

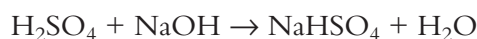
Molecular weight of HCl = Equivalent weight = 36.5

For the acids and bases that have their basicity or acidity greater than one, the equivalent weight must be calculated by taking the reaction into consideration.

$$\text{Equivalent weight of the acid} = \frac{\text{Molecular weight}}{\text{Number of replaceable hydrogen ions in the given react}}$$

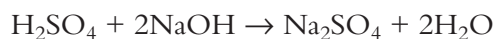
$$\text{Equivalent weight of the base} = \frac{\text{Molecular weight}}{\text{Number of replaceable OH}^- \text{ ions in the given reaction}}$$

H_2SO_4 .

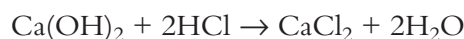


Example:

$$\text{Equivalent weight} = \frac{\text{Molecular weight}}{1} = 98$$



$$\text{Equivalent weight} = \frac{\text{Molecular weight}}{2} = 49.$$



$$\text{Equivalent weight} = \frac{\text{Molecular weight}}{2} = \frac{74}{2} = 37 \text{ g}$$

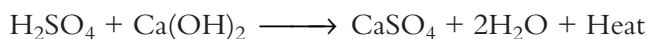
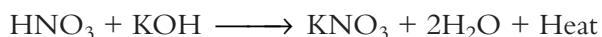
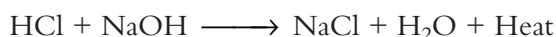
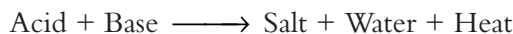
$$\text{Equivalent weight of salt} = \frac{\text{Molecular weight}}{\text{Net charge on ion}}$$



$$\text{Equivalent weight} = \frac{\text{Molecular weight}}{2} = \frac{142}{2} = 71 \text{ g}.$$

Neutralisation Reaction

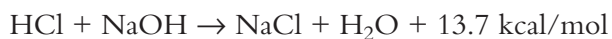
Reaction between acid and base to form salt and water liberating heat in the process.



Heat of Neutralisation

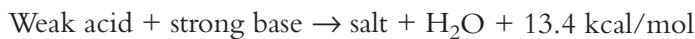
The amount of heat liberated when one equivalent of the acid reacts with one equivalent of the base. For any strong acid–strong base reaction, the heat of neutralisation has the same value, that is, 13.7 kcal/mol since the acid and base involved are completely ionised.

Example:



When the weak acid or the weak base is involved in the reaction, some amount of heat is utilised to ionise the acid or base to furnish ions. This is called heat of ionisation. Therefore, the net heat released is less than the heat of neutralisation for strong acid–strong base reaction.

Example:



$$\text{Heat of ionisation} = 13.7 - \text{net heat liberated}$$

$$= 13.7 - 13.4 = 0.3 \text{ kcal/mol}$$



Uses of Neutralisation

1. Slaked lime ($\text{Ca}(\text{OH})_2$) is added to reduce the acidity of soil.
2. Antacid tablets with $\text{Mg}(\text{OH})_2$ or cold milk are used to treat acidity by neutralizing HCl in stomach.
3. Sting of ants/bees containing HCOOH (formic acid) can be neutralised by rubbing soap (NaOH).
4. Sting of wasps that contains an alkali can be treated by neutralizing it with CH_3COOH .
5. Lithium hydroxide is used to neutralise CO_2 exhaled by astronauts in spaceships or submariners in submarines.

The concept of equivalent masses helps in determining the number of moles of the acid needed to neutralise the base and vice versa. This also plays a major role in determining the concentration of a solution in terms of normality.

NORMALITY

The number of equivalents of a solute present in one litre of the solution is called normality. It is denoted by 'N'.

$$\begin{aligned} \text{Normality} &= \frac{\text{Number of equivalents}}{\text{Volume of solution in litres}} \\ &= \text{Number of equivalents} \times \frac{1000}{V \text{ in mL}} \\ &= \frac{\text{weight}}{\text{Equivalent weight}} \times \frac{1}{V \text{ in L}} = \frac{\text{weight}}{\text{Equivalent weight}} \times \frac{1000}{V \text{ in mL}} \end{aligned}$$

The normality of the solution is helpful in expressing the concentration of the solution. However, in case of the acids and bases, the concentration can also be expressed in terms of pH. The pH of the solution represents the hydrogen ion concentration in it. This is a single factor that represents not only the concentration but also helps in determining whether a solution is acidic or basic. To calculate the pH of a solution, the concept of ionic product of water should be well understood.

IONIC PRODUCT OF WATER

Water molecules ionise partially to give H^+ and OH^- ions. Equilibrium exists between the ionised and unionised molecules:



Applying the law of mass action to the above equilibrium, $K = \frac{[\text{H}^+][\text{OH}^-]}{[\text{H}_2\text{O}]}$

Since $[\text{H}_2\text{O}]$ is constant (Molar concentration of water in its pure sample is $= \frac{1000}{18} = 55.55 \text{ M}$)

$$K[\text{H}_2\text{O}] = K_w = [\text{H}^+][\text{OH}^-],$$

where K_w is called the ionic product of water.

At normal temperature (25°C), the concentration of H^+ and OH^- ions remains equal, that is, 10^{-7} M

$$\therefore [\text{H}^+] = [\text{OH}^-] = 10^{-7} \text{ mol/L}$$

$$\begin{aligned} K_w &= [\text{H}^+][\text{OH}^-] = 10^{-7} \times 10^{-7} \\ &= 10^{-14} \text{ mol}^2/\text{L}^2 \end{aligned}$$

On adding a base, $[\text{OH}^-]$ increases and $[\text{H}^+]$ decreases but K_w remains constant at constant temperature.

Examples:

$[\text{H}^+]$ mol ion/L	$[\text{OH}^-]$ mol ion/L	K_w (mol. ion) $^2/\text{L}^2$
1×10^0	1×10^{-14}	1×10^{-14}
1×10^{-3}	1×10^{-11}	1×10^{-14}

At 25°C ,

Neutral solution $\rightarrow [\text{H}^+] = [\text{OH}^-] = 10^{-7} \text{ mol ion/L}$

Acidic solution $\rightarrow [\text{H}^+] > [\text{OH}^-] < 10^{-7} \text{ mol ion/L}$

Basic solution $\rightarrow [\text{H}^+] < [\text{OH}^-] > 10^{-7} \text{ mol ion/L}$

In all the cases, the product of $[\text{H}^+]$ and $[\text{OH}^-]$ remains constant and is equal to 10^{-14} at 25°C .

Thus, at a given temperature the value of ionic product of water remains constant. Addition of the acid or the base does not change the value of K_w . On the addition of acid, when $[\text{H}^+]$ increases, a corresponding decrease in $[\text{OH}^-]$ keeps the K_w value constant.

pH

The concept of pH was introduced by Sørensen in order to avoid the use of negative exponents in expressing the concentration of acids and bases.

In pH, p means *potenz* and 'H' stands for hydrogen.

pH is defined as the negative logarithm to base 10 of H^+ ion concentration (or) logarithmic reciprocal of H^+ ion concentration.

$$\text{pH} = -\log_{10}[\text{H}^+] \text{ or } \log_{10} \frac{1}{[\text{H}^+]}$$

Similarly,

$$\text{pOH} = -\log_{10}[\text{OH}^-] \text{ or } \log_{10} \frac{1}{[\text{OH}^-]}$$

At 25°C, for a neutral solution, $[H^+] = [OH^-] = 10^{-7}$ mol ion/L

$$\text{The pH} = -\log_{10} \frac{1}{[H^+]} = -\log_{10} \frac{1}{10^{-7}} = -\log_{10} 10^7 = 7$$

For the acidic solution: $[H^+] > [OH^-] < 10^{-7}$ mol ion/L

$$\text{pH} = -\log_{10} \frac{1}{> 10^{-7}} = < 7.$$

Lower the pH value, the greater will be the acidic strength of the solution.

For the basic solution: $[H^+] < [OH^-] > 10^{-7}$ mol ion/L

$$\text{pH} = -\log_{10} \frac{1}{> 10^{-7}} = > 7$$

Greater the pH value, greater will be the basic strength of the solution.

Relation between pH and concentration of H^+ ions

$[H^+]$	10^0	10^{-1}	10^{-2}	10^{-3}	10^{-4}	10^{-5}	10^{-6}	10^{-7}	10^{-8}	10^{-9}	10^{-10}	10^{-11}	10^{-12}	10^{-13}	10^{-14}
pH	0	1	2	3	4	5	6	7	8	9	10	11	12	13	14

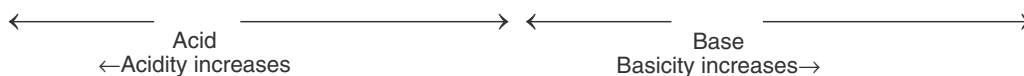


TABLE 5.2 pH of some common substances

Substances	pH values
HCl (gastric juice in stomach)	1.0
H_2SO_4	1.2
Citric acid	2.3
Lime juice	2.4
Acetic acid	2.9
Carbonic acid	3.0
Orange	3.5
Boric acid	5.2
Milk (lactic acid)	6.6
Pure water	7.0
Human blood	7.3
Albumin (egg)	7.9
$NaHCO_3$ solution	8.4
Sea water	8.5
Milk of magnesia	10.5
Ammonium hydroxide	11.1
Na_2CO_3 solution	11.6
NaOH	13

Uses of pH Value

1. In brewing, pH values are strictly controlled for wines of desired flavour.
2. pH value of soils is tested to enhance growth of crops in acidic/alkaline soils.
3. pH value of milk is controlled in dairies.
4. Depending on pH values of skin secretion, different soaps for oily/dry skin are prepared.
5. In medicine, for diagnosis of certain diseases, pH values of blood and urine are determined.

By mathematically calculating the pH, nature of a solution can be determined. In laboratory, however, certain chemicals are used as indicators to determine whether a solution is acidic or basic.

Indicators

Indicators are organic dyes that have one colour in acidic solutions and another colour in basic solutions. There are two types of indicators as listed below:

1. common indicator and
2. universal indicator.

Common Indicators

These are specific and work only in a specified pH range. Using these, it is possible to identify the acidic, basic and neutral solutions and the pH range of the solution. However, it is not possible to know the exact pH value of the solution.

TABLE 5.3 Common colour indicators

Indicators	Colours in acidic solution	Colours in basic solution
Litmus	Red	Blue
Methyl orange	Red	Yellow
Phenolphthalein	Colourless	Pink

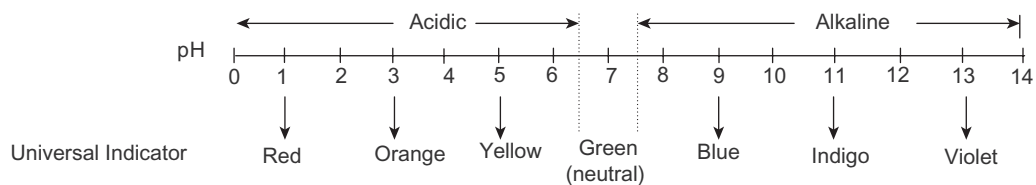
TABLE 5.4 Determination of pH using indicators

Name of the indicators	Acid colours	pH range of colour change	Base colours
Methyl violet	Yellow	0.0–1.6	Blue
Thymol blue	Red	1.2–2.8	Yellow
Methyl orange	Red	3.2–4.4	Yellow
Bromocresol green	Yellow	3.8–5.4	Blue
Methyl red	Red	4.8–6.0	Yellow
Litmus	Red	5.0–8.0	Blue
Bromothymol blue	Yellow	6.0–7.6	Blue
Thymol blue	Yellow	8.0–9.6	Blue
Phenolphthalein	Colourless	8.2–10.0	Pink
Thymolphthalein	Colourless	9.4–10.6	Blue
Alizarin yellow R	Yellow	10.1–12.0	Red

Universal Indicator

It is a mixture of organic dyes that gives different colours with solutions of different pH values. This is more sensitive than other common indicators.

Determination of pH using universal indicators

**EXAMPLE**

How does the concentration of hydronium ions change during dilution of an acid?

SOLUTION

Dilution of an acid decreases the concentration of hydronium ions per unit volume.

EXAMPLE

Give reasons for the following:

- (i) Metal oxides generally react with acids. However, zinc oxide reacts even with sodium hydroxide.
- (ii) Heat of neutralisation value for the reaction of HCl and NH_4OH is less than 13.7 kcal/mol.
- (ii) All indicators are not suitable for all titrations.

SOLUTION

- (i) Zinc oxide is amphoteric in nature. Hence, it can react both with acids as well as bases.
- (ii) NH_4OH is a weak base and hence, some heat is utilised to ionise the weak base.
- (iii) Indicator that undergoes colour change in the pH range of the titration is only suitable for a particular titration.

EXAMPLE

Explain the action of phenolphthalein indicator in the titration of acetic acid. sodium hydroxide.

SOLUTION

Acetic acid is a weak acid having a pH value of more than 6. Sodium hydroxide is a strong base and addition of OH^- ions neutralises the H^+ ions formed by the dissociation of acetic acid. Since H^+ ion concentration is very low, the addition of further OH^- ions increases the pH value abruptly. Therefore, the titration range of this reaction lies between 6.5 and 10.0.

The pH range of phenolphthalein is between 8.2 and 10.0 and hence, it undergoes colour change in this range. Hence, it acts as a suitable indicator for this titration.

EXAMPLE

Is a solution with pH 6 acidic, basic or neutral? Justify.

SOLUTION

A solution with pH 6 will be acidic when the ionic product of water is $1 \times 10^{-14} \text{ mol}^2/\text{L}^2$. In such a case, a solution with pH less than 7 is acidic in nature. On increasing the temperature, the ionic

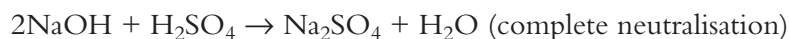
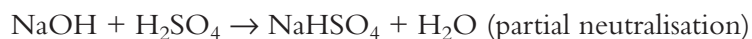
product of water increases due to increase in the ionisation of water. When the ionic product of water increases to $1 \times 10^{-12} \text{ mol}^2/\text{L}^2$, a solution with pH 6 will be neutral and below 6 will be acidic in nature.

SALTS

All salts are basically ionic compounds and are electrically neutral as the total positive charge on positive radicals is equal to the total negative charge on negative radicals. These oppositely charged ions are held together by strong electrostatic forces of attraction when the salts are in the solid state. In fused state or in aqueous solutions, these forces of attraction are broken and the salts are dissociated into ions. This accounts for the electrical conductivity of their solutions. Thus, all salts are electrolytes in their molten states as well as in their aqueous solutions.

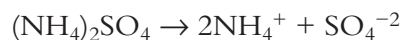
Formation of Salts

Salts are the compounds formed during the neutralisation reaction between the acid and the base. Therefore, salts can also be defined as the compounds formed by either partial or complete replacement of ionizable hydrogen atoms of the acid by one or more metal ions or ammonium ions of a base.



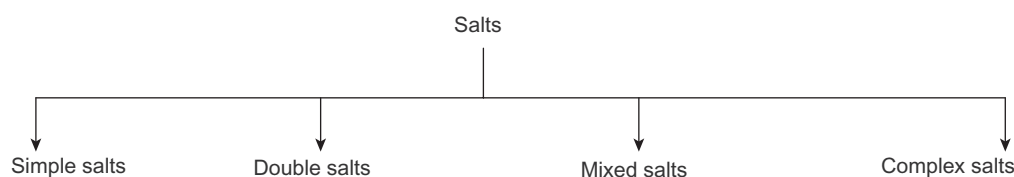
Ionic definition of salt can be given as the ionic compound that on dissociation in water yields one or more positive ions other than H^+ ions and one or more negative ions other than OH^- ions.

Examples:



There are a large number of salts present. To make the study of salts easier, they are classified into four different categories on the basis of their composition.

Classification of Salts



Simple Salts

The salts that contain only one type of positive ion and one type of negative ion are called simple salts. These are further classified into three types based on their chemical nature.

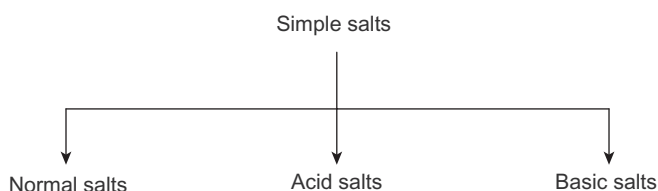


TABLE 5.5 Comparative study of normal, acid and basic salts

Parameters	Normal salts	Acid salts	Basic salts
Definition	The salts that do not contain any replaceable hydrogen or 'OH' groups	The salts that contain one or more replaceable hydrogen groups	The salts that contain one or more replaceable 'OH' groups
Formation	By complete neutralization of acids and bases. E.g. $2\text{NH}_4\text{OH} + \text{H}_2\text{SO}_4$ ↓ $(\text{NH}_4)_2\text{SO}_4 + \text{H}_2\text{O}$	By incomplete neutralisation of acids and bases E.g. $\text{NaOH} + \text{H}_2\text{SO}_4$ ↓ $\text{NaHSO}_4 + \text{H}_2\text{O}$	By incomplete neutralization of acids and bases. E.g. $\text{Cu}(\text{OH})_2 + \text{HNO}_3$ ↓ $\text{Cu}(\text{OH})\text{NO}_3 + \text{H}_2\text{O}$
Composition	All ionizable hydrogens are replaced by a metal or NH_4^+ ion	Partial replacement of ionizable hydrogen by a metal or NH_4^+ ion	Partial replacement of ionizable hydroxyl groups by non-metal ions or negative radicals
Properties	Neutral in nature. Do not react with acids or bases	Acidic in nature. React with bases.	Basic in nature. React with acids.
Examples	NaCl , $(\text{NH}_4)_2\text{SO}_4$, $\text{Al}_2(\text{SO}_4)_3$	NaHSO_4 , NH_4HSO_4	$\text{Cu}(\text{OH})\text{NO}_3$, $\text{Cu}(\text{OH})\text{Cl}$

Double Salts

The salts that contain more than one simple salt are called double salts. These salts are formed by the crystallisation and union of two simple salts dissolved in water. Double salts undergo complete dissociation in aqueous solutions. They give reactions of all ions present in both the simple salts. These salts are usually associated with water of crystallisation.

Examples:

Salt	Formulae
Potash alum	$\text{K}_2\text{SO}_4 \cdot \text{Al}_2(\text{SO}_4)_3 \cdot 24\text{H}_2\text{O}$
Mohr's salt	$\text{FeSO}_4 \cdot (\text{NH}_4)_2\text{SO}_4 \cdot 6\text{H}_2\text{O}$
Dolomite	$\text{CaCO}_3 \cdot \text{MgCO}_3$

Mixed Salts

The salts that contain more than one basic or acidic radicals are called mixed salts. These are formed by multi-step reactions involving corresponding acids and bases. Mixed salts undergo complete dissociation in aqueous solutions. They give the reactions of all the ions produced in aqueous solutions.

Examples:

Sodium potassium carbonate (NaKCO_3) (Na^+ , K^+ and CO_3^{-2} radicals)

Bleaching powder (CaOCl_2) (Ca^{+2} , OCl^- , Cl^- radicals)

Disodium potassium phosphate Na_2KPO_4 (Na^+ , K^+ , PO_4^{-3} radicals)

Complex Salts

The salts that contain one complex ion and one or more simple ions are called complex salts. These salts are formed by mixing saturated solutions of simple salts followed by crystallisation.

Examples:

Sodium argentocyanide $\text{Na}[\text{Ag}(\text{CN})_2]$

Potassium ferrocyanide $\text{K}_4[\text{Fe}(\text{CN})_6]$

A complex salt dissociates into complex ion and simple ions. Simple ions react individually, while the complex ions react as one unit.

Example:



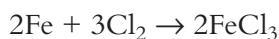
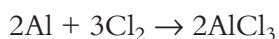
It gives reactions of potassium ions and ferrocyanide ions. However, it cannot give reactions of ferrous ions and cyanide ions separately.

Preparation of Salts

1. Preparation of Soluble Salts

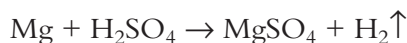
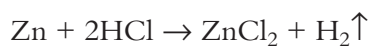
- (a) By direct combination (synthesis):** Heating a metal and a non-metal together. The non-metal is over the surface of the hot metal to get the salt.

Examples:



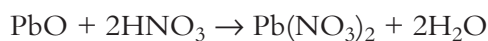
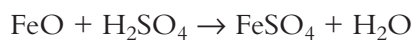
- (b) By the action of dilute acids on metals:** Active metals like sodium and potassium are treated with acids like HCl , H_2SO_4 and HNO_3 to give their respective salts.

Examples:



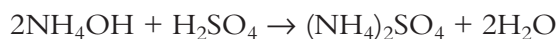
- (c) By the action of dilute mineral acids on oxides:** Metal oxides are generally insoluble and basic in nature. When these oxides are treated with dilute acids, displacement takes place. As, metal oxides are basic in nature, neutralisation takes place.

Examples:



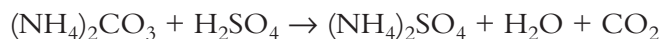
- (d) Action of dilute mineral acids on hydroxides:** Soluble salts are obtained by the titration of an acid with a soluble hydroxide.

Example:



- (e) Action of dilute acids on carbonates:** Soluble salts are obtained by the titration of a carbonate salt solution with an acid.

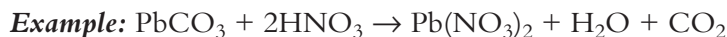
Example:



- (f) Action of dilute mineral acid on hydroxide:** Soluble salts are obtained by the neutralisation of an acid with its respective insoluble hydroxide.



- (g) Action of dilute acids on carbonates:** Soluble salts are obtained by the neutralisation of the acid with an insoluble carbonate.



2. Preparation of Insoluble Salts

- (a) By direct combination (synthesis):** Heating the powdered metal with sulphur.



- (b) Precipitation from soluble salts:** In this method, when two soluble salt solutions are mixed, double decomposition takes place and the insoluble salt is precipitated. This is the most common method for the preparation of insoluble salts.



General Characteristics of Salts

- 1. Physical state:** Salts are non-volatile crystalline solids.
- 2. Electrical conductivity:** Salts being electrovalent (ionic) compounds conduct current in molten as well as in aqueous state due to their dissociation into ions.
- 3. Solubility:** The salts are ionic compounds. However, all the salts are not soluble in water. Some salts remain insoluble.

TABLE 5.6 Table showing the solubility of salts

Solubility of salts in water	
Solubility of compounds	Exceptions
All NH_4^{+} , Na^{+} , K^{+} compounds are soluble	—
All nitrates and nitrites are soluble	—
All chlorides are soluble in water	Hg_2Cl_2 , AgCl , PbCl_2
	PbCl_2 —Soluble in hot water
All sulphates are soluble	CaSO_4 , PbSO_4 , BaSO_4
All oxides and hydroxides are insoluble	Oxides of Na^{+} , K^{+} , Mg^{+2}
All carbonates are insoluble	K_2CO_3 , Na_2CO_3 , $(\text{NH}_4)_2\text{CO}_3$
All sulphides are insoluble	K_2S , Na_2S , $(\text{NH}_4)_2\text{S}$

Salt Hydrolysis

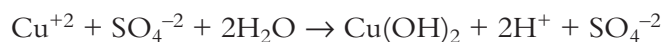
The phenomenon in which a simple salt on dissolution in water forms a parent acid and a parent alkali resulting in acidic, basic or neutral solution is called salt hydrolysis. It is the most important property of salts.

Depending on the nature of the salts, it is classified into the following four types.

- 1. Hydrolysis of salts of strong acids and weak bases:** This type of hydrolysis produces a strong acid that ionises and a weak base that does not ionise. This increases the concentration of H^+ ion making the solution acidic.

Example:

$(NH_4)_2SO_4$, $CuSO_4$, NH_4NO_3 , $FeCl_3$, etc.

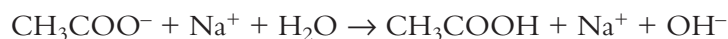


As only the cation of the salt reacts with water, it is also called cationic hydrolysis.

- 2. Hydrolysis of salts of weak acids and strong bases:** This produces a weak acid that does not ionise and a strong base that ionises and gives OH^- ions. This makes the solution basic and has a $pH > 7$.

Example:

Na_2CO_3 , CH_3OOK , CH_3COONa



As the anion of the salt reacts with water, it is also called anionic hydrolysis.

- 3. Hydrolysis of salts of strong acids and strong bases:** This type of hydrolysis produces a strong acid and a strong base that are completely ionised giving equal number of H^+ and OH^- ions. Thus, this solution is neutral with a $pH = 7$.

Example:

$NaCl$, K_2SO_4 , KNO_3



- 4. Hydrolysis of salts of weak acids and weak bases:** This produces a weak acid and weak base which exist in the molecular form. Thus, there are no free H^+ or OH^- ions and the solution is neutral with $pH = 7$.

Examples:

$(NH_4)_2CO_3$, CH_3COONH_4



Qualitative Analysis of Simple Salts

A simple salt contains an acidic (anion) and a basic radical (cation). The analysis of a simple salt involves identification of the constituent acidic and basic radicals of the salt.

TABLE 5.7 Identification of some important anions

Anions	Tests	Reactions involved	Observations and inference
Carbonate	Action of dilute HCl	$CaCO_3 + 2HCl$ ↓ $CaCl_2 + H_2O + CO_2$	Colourless gas with brisk effervescence that puts off burning splinter and turns lime water milky. CO_2 gas is evolved.

(Continued)

Anions	Tests	Reactions involved	Observations and inference
Chloride	Action of conc. H_2SO_4	$\text{BaCl}_2 + \text{H}_2\text{SO}_4$ \downarrow $\text{BaSO}_4 + 2\text{HCl}$ $\text{HCl} + \text{NH}_4\text{OH}$ \downarrow $\text{NH}_4\text{Cl} + \text{H}_2\text{O}$	A colourless and pungent smelling gas is evolved (HCl). The gas when exposed to a glass rod dipped in NH_4OH forms white dense fumes.
Nitrate	Action of conc. H_2SO_4 + in the presence of copper turnings	$4\text{NaNO}_3 + 2\text{H}_2\text{SO}_4 + \text{Cu}$ \downarrow $2\text{Na}_2\text{SO}_4 + 2\text{NO}_2 +$ $\text{Cu}(\text{NO}_3)_2 + 2\text{H}_2\text{O}$	A reddish brown pungent gas is evolved (NO_2) and the solution turns blue ($\text{Cu}(\text{NO}_3)_2$).

TABLE 5.8 Confirmatory tests for anions

Anions	Tests	Reactions involved	Observations
Carbonate	Reaction with BaCl_2 solution	$\text{CaCO}_3 + \text{BaCl}_2$ \downarrow $\text{BaCO}_3 + \text{CaCl}_2$	A white precipitate that is insoluble in conc. HCl (BaCO_3)
Chloride	Reaction with AgNO_3 solution	$2\text{BaCl}_2 + 2\text{AgNO}_3$ \downarrow $2\text{AgCl} + \text{Ba}(\text{NO}_3)_2$	White curdy precipitate (AgCl)
Nitrate	Brown ring test: Salt solution + Freshly prepared ferrous sulphate solution + conc. H_2SO_4		A reddish brown ring is formed at the junction of the two layers

Identification of Some Important Cations

For the identification of cations in the salts, NaOH and NH_4OH are the generally used reagents. They form insoluble precipitates with characteristic colours.

TABLE 5.9 Table showing different characteristic precipitations given by different metal ions

Cations of salts	Precipitates formed	
	NaOH	NH_4OH
Fe^{2+}	$\text{Fe}(\text{OH})_2$ Dirty green Insoluble in excess NaOH	$\text{Fe}(\text{OH})_2$ Dirty green Insoluble in excess NH_4OH
Fe^{3+}	$\text{Fe}(\text{OH})_3$ Reddish Brown	$\text{Fe}(\text{OH})_3$ Reddish Brown

(Continued)

Cations of salts	Precipitates formed	
	NaOH	NH ₄ OH
Ca ⁺²	Insoluble in excess NaOH	Insoluble in excess NH ₄ OH
	Ca(OH) ₂ Milky white	No reaction –
	Insoluble in excess NaOH	–
Zn ⁺²	Zn(OH) ₂ White	Zn(OH) ₂ White
	Soluble in excess NaOH (colourless solution)	Soluble in excess NH ₄ OH (colourless solution)
Pb ⁺²	Pb(OH) ₂ Chalky white	Pb(OH) ₂ Chalky white
	Soluble in excess NaOH	Insoluble in excess NH ₄ OH
Cu ⁺²	Cu(OH) ₂ Pale blue	Cu(OH) ₂ Pale blue
	Insoluble in excess NaOH	Soluble in excess NH ₄ OH (deep blue solution)
Al ⁺³	Al(OH) ₃ White gelatinous	Al(OH) ₃ White gelatinous
	Soluble in excess NaOH	Soluble in excess NH ₄ OH

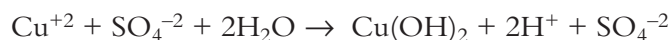
EXAMPLE

Define salt hydrolysis. Explain the nature of aqueous solution of CuSO₄ on the basis of the above phenomenon.

SOLUTION

The phenomenon in which a simple salt on dissolution in water forms a parent acid and a parent alkali resulting in acidic, basic or neutral solution is called salt hydrolysis. It is the most important property of salts.

CuSO₄ on hydrolysis produces a strong acid that ionises and a weak base that does not ionise. This increases the concentration of H⁺ ions making the solution acidic with pH < 7 (at 25 °C)



As cation of the salt reacts with water, it is called cationic hydrolysis.

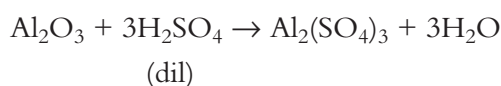
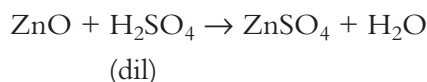
EXAMPLE

Explain the amphoteric nature of zinc oxide and aluminium oxide.

SOLUTION

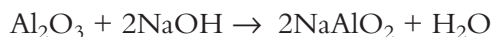
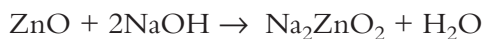
On warming zinc oxide with dilute sulphuric acid, the zinc oxide dissolves rapidly and forms clear solution of zinc sulphate.

Similarly, when aluminium oxide is warmed with dilute sulphuric acid forms aluminium sulphate.



On warming zinc oxide with concentrated sodium hydroxide or potassium hydroxide, the zinc oxide rapidly dissolves to form clear solution of sodium zincate or potassium zincate.

Similarly, aluminium oxide reacts with concentrated sodium hydroxide to form sodium metaluminate.



As the oxides of zinc and aluminium react with acids as well as alkali to form salt and water, they are amphoteric in nature.

EXAMPLE

How do you identify the following cations in their aqueous solutions by using caustic soda solution?

- (i) Fe^{+2} (ii) Zn^{+2} (iii) Al^{+3}

SOLUTION

- (i) Fe^{+2} gives light green precipitate of ferrous hydroxide ($\text{Fe}(\text{OH})_2$).
- (ii) Zn^{+2} gives white precipitate of zinc hydroxide ($\text{Zn}(\text{OH})_2$).
- (iii) Al^{+3} gives white gelatinous precipitate of aluminium hydroxide ($\text{Al}(\text{OH})_3$).

EXAMPLE

What is the use of an indicator in titration? Name the indicator used when sodium hydroxide is titrated against sulphuric acid.

SOLUTION

The neutralisation point can be determined by the use of indicator. The indicator used when sodium hydroxide is titrated against sulphuric acid is phenolphthalein.

TEST YOUR CONCEPTS

Very Short Answer Type Questions

- What is meant by ionic product of water?
- Give examples of some acidic substances without H^+ ions.
- Which acids/bases are used for the following?
 - to neutralise sting of wasps
 - in alkaline batteries
 - manufacturing of soap
- Define oxyacid. Give two examples.
- According to Bronsted–Lowry theory neutralisation reaction is the formation of _____.
- Give one example of acid–base reaction without formation of water.
- $X + H_2O \rightarrow Y$. Y is used to remove grease stains from woollen clothes. Identify X and Y.
- How do neutral, acidic and basic salt solutions respond to litmus test?
- Give examples of some basic substances without OH^- ions.
- How does a complex salt dissociate in its aqueous solution?
- What is the principle involved in the preparation of $FeCl_3$?
- In the equation $X + HCl \rightarrow NaCl + H_2S$, what is X?
- Mention any two methods of preparation of acids. Illustrate with equations.
- How are basic salts formed?
- How are soluble salts prepared by direct combination? Give an example.
- Illustrate any one method of preparation of bases with equations.
- What is the pH range in which phenolphthalein is colourless?
- Name the insoluble salts that can be prepared by direct combination?
- What are the pH values of the following solutions at $25^\circ C$?
 - 0.001 M of HCl
 - 0.05 M of H_2SO_4
 - 0.01 M of NaOH
- The formula of the basic salt formed by treating calcium hydroxide with HCl is _____.
- Which method is employed for the preparation of sulphate of Ba, Ca and Pb?
- Define pH and give its mathematical representation.
- What would be the nature of aqueous solutions of the following salts?
 - KNO_3
 - Na_2CO_3
 - CH_3COONa
- Define basicity of an acid. Give an example each for monobasic, dibasic and tribasic acids.
- Normality of 0.01M aluminium hydroxide is _____ (assuming complete ionisation).
- Write the equation for the action of a dilute acid on a metal oxide.
- The hydroxyl ion concentration of sodium hydroxide having a pH value of 12, at $25^\circ C$ is _____.
- Mention the steps involved in ionisation of phosphoric acid.
- Non-metallic oxides are either _____ or _____ in nature.
- The number of ions produced by one mole of $[Co(NH_3)_6]Cl_3$ in its aqueous solution is _____.

Short Answer Type Questions

- What is meant by a universal indicator? State the colours given by a universal indicator to the solutions of following pH values.

(i) pH = 1	(ii) pH = 8
(iii) pH = 5	(iv) pH = 12
- How is NH_4OH helpful as a reagent for the analysis of cations in the aqueous solutions of their salts? Give equations.



33. What is meant by acidity of a base? Give the acidities of the following bases.
 (i) caustic potash (ii) slaked lime
 (iii) aluminium hydroxide
34. Define indicator. Why are different indicators used in different pH ranges?
35. Define salt hydrolysis. Explain the hydrolysis of potassium carbonate by giving ionic equations.
36. Explain protonic concept of acids and bases with one example each.
37. How do bases react with the following? Write the equations.
 (i) non-metallic oxide (ii) ammonium salt
 (iii) soluble metallic salt
38. What is meant by neutralisation? Give some examples identifying spectator ions in the reactions.
39. What is the value of ionic product of water at 25°C? How does it change with temperature?
40. Give pH range for various solutions at 25°C.
41. Distinguish between acidic and basic salts giving examples.
42. Mention any four applications of pH.
43. Define basicity of an acid. Give an example each for monobasic, dibasic and tribasic acids.
44. What is the difference between mixed and complex salts with respect to chemical reactions? Explain with an example.

Essay Type Questions

46. State the postulates of Arrhenius theory. Explain with examples. What are the limitations of Arrhenius theory?
47. On the basis of ionisation explain classification of acids and bases with suitable examples. Classify acids/bases on the basis of amount of water present in the solution.
48. Explain the formation of hydronium ion and ammonium ion.
49. Define salt hydrolysis. Explain the nature of aqueous solution of the following salts on the basis of salt hydrolysis.
 (i) $(\text{NH}_4)_2\text{CO}_3$ (ii) CuSO_4
 (iii) NaNO_3 (iv) CaCl_2
50. Explain with examples the use of NaOH and NH_4OH as reagents for the identification of cations in salts.

For Answer key, Hints and Explanations, please visit: www.pearsoned.co.in/IITFoundationSeries

CONCEPT APPLICATION

Level 1

Direction for questions from 1 to 7:

State whether the following statements are true or false.

- Formation of co-ordinate covalent bond between NH_3 and BF_3 is an example of neutralisation reaction according to Bronsted theory.
- A di-acidic base has one replaceable hydroxide ion.
- For a dibasic acid, normality is twice the molarity for a given solution.
- S^{2-} is the conjugate base of HS^- .
- Methyl orange exhibits yellow colour in basic medium.

- Curd is sour in taste due to the presence of an acidic substance.
- $\text{K}_2\text{SO}_4 \cdot \text{Al}_2(\text{SO}_4)_3 \cdot 24\text{H}_2\text{O}$ gives reactions of both K^+ and Al^{3+} ions respectively.

Direction for questions from 8 to 14:

Fill in the blanks.

- Phenolphthalein is a/an _____ indicator.
- The salt obtained by partial neutralisation of sodium hydroxide and sulphuric acid is _____ in nature.
- Metallic oxides are either _____ or _____ in nature.



11. $[H^+]$ of nitric acid having a pH value of 6 is _____
12. The pH of aqueous solution of salt obtained by the reaction between NaOH and CH_3COOH is _____ at $25^\circ C$.
13. A salt, X is prepared by the complete neutralisation of $Cu(OH)_2$ with H_2SO_4 . The nature of X is _____
14. The units of K_w are _____.

Direction for question 15:

Match the entries given in Column A with appropriate ones from Column B.

Column A	Column B
A. Sodium chloride	() a. Baking powder
B. Tartaric acid	() b. Freezing mixture in ice creams
C. Potassium nitrate	() c. Antacid in medicine
D. Sodium bicarbonate	() d. Gun powder
E. Citric acid	() e. Bleaching powder
F. Calcium hydroxide	() f. Flavouring agent

Direction for questions from 16 to 45:

For each of the questions, four choices have been provided.

Select the correct alternative.

16. Which of the following is not a base according to any of the theories?
- (a) $Mg(OH)_2$ (b) NH_3
(c) $H_2PO_4^-$ (d) BF_3
17. According to Lewis acid–base concept, which is a correct statement?
- (a) Species in which central atom has completed octet cannot act as acid.
(b) All negatively charged ions are acids.
(c) Molecule in which the central atom has vacant 'd' orbitals acts as acid.
(d) All positively charged ions are bases.
18. At a certain temperature, K_w was found to be $1 \times 10^{-10} \text{ mol}^2/L^2$. At this temperature, a solution is taken and red litmus is added to this solution. The solution turned to blue. pH range of the solution could be
- (a) between 5 and 7 (b) between 7 and 10
(c) less than 5 (d) both 1 and 2

19. Which of the following indicators is the best indicator for the detection of end point in the titration of CH_3COOH and KOH?
- (a) methyl orange (b) phenolphthalein
(c) methyl red (d) none of these
20. According to Lewis theory, neutralisation is the
- (a) transfer of proton from acid to base.
(b) transfer of proton from base to acid.
(c) transfer of electron pair from acid to base.
(d) transfer of electron pair from base to acid.
21. The indicator that shows colour changes in the entire pH range is
- (a) phenolphthalein (b) methyl orange
(c) universal indicator (d) thymol blue
22. Which of the following has the highest value of pH?
- (a) HCl (b) NaOH
(c) $Ca(OH)_2$ (d) CH_3COOH
23. Heat of neutralisation is less than 13.7 kcal/mol for the reaction
- (a) $HCl + NaOH \rightarrow NaCl + H_2O$
(b) $H_2SO_4 + 2NaOH \rightarrow Na_2SO_4 + 2H_2O$
(c) $HNO_3 + NaOH \rightarrow NaNO_3 + H_2O$
(d) $CH_3COOH + NaOH \rightarrow CH_3COONa + H_2O$
24. Which among the following does not undergo hydrolysis?
- (a) Na_2CO_3 (b) CH_3COOH
(c) CH_3COONH_4 (d) K_2SO_4
25. Which of the following acts as both Bronsted acid and Lewis acid?
- (a) Cu^{+2} (b) SO_2
(c) Fe^{+3} (d) NH_4^+
26. Which among the following factors changes the value of ionic product of water?
- (a) change in temperature
(b) addition of acid
(c) addition of base
(d) addition of either acid and base
27. Which of the following solutions has the same concentration of H^+ ions as 0.1 N HCl?
- (a) 0.1N H_2SO_4 (b) 0.3 N H_3PO_4
(c) 0.2 N HNO_3 (d) All of these



28. Which of the following cannot show acidic nature?
 (a) H_2CO_3 (b) CaCO_3
 (c) HCl (d) HSO_4^-
29. The spectator ions in the hydrolysis of $(\text{NH}_4)_2\text{CO}_3$ are
 (a) NH_4^+ , H^+ , OH^-
 (b) NH_4^+ , CO_3^{2-} , H^+ , OH^-
 (c) CO_3^{2-} , H^+ , OH^-
 (d) no spectator ions
30. Disodium potassium phosphate is a/an _____ salt.
 (a) acid (b) mixed
 (c) complex (d) double
31. V_1 mL of x molar hydrochloric acid is mixed with V_2 mL of y molar sodium hydroxide. Arrange the following steps in correct sequence for calculation of pH of the mixture. ($y > x$ and $V_1 > V_2$)
 (1) Calculation of the number of equivalents of acid and base.
 (2) Calculation of the total volume of mixture.
 (3) Calculation of the normality of acid and base.
 (4) Calculation of the net OH^- ion concentration.
 (5) Taking the negative logarithm of H^+ ion concentration.
 (6) Dividing K_w by OH^- ion concentration.
 (a) 1 3 2 6 4 5 (b) 3 1 2 4 5 6
 (c) 3 1 2 6 4 5 (d) 1 2 3 4 5 6
32. A strongly acidic solution is taken and a base is slowly added to it drop by drop. Arrange the colours observed during the process of addition in proper sequence.
 (1) green (2) indigo
 (3) orange (4) red
 (5) blue (6) violet
 (7) yellow
 (a) 3 4 1 7 2 5 6 (b) 4 3 7 1 5 2 6
 (c) 4 3 1 7 5 2 6 (d) 3 4 7 1 6 2 5
33. A red litmus paper when dipped in an aqueous solution of potassium acetate turned blue. Arrange the following steps in proper sequence to explain the above observation.
 (1) identification of spectator ions.
 (2) identification of free ions that caused the colour change of the litmus paper.
 (3) identification of positive and negative radicals.
 (4) writing the net reaction taking place.
 (5) writing ionic equation for the reaction of positive and negative radicals with water.
 (a) 3 5 1 4 2 (b) 3 1 5 4 2
 (c) 5 1 4 3 2 (d) 5 3 1 4 2
34. Identify the species that can give both conjugate acid and conjugate base.
 (a) NH_4^+ (b) H_2PO_4^-
 (c) PO_4^{3-} (d) H_3O^+
35. The basicity of acetic acid is the same as the basicity of _____.
 (a) HNO_3 (b) H_2SO_4
 (c) H_3PO_4 (d) H_2CO_3
36. Which of the following species cannot have conjugate base?
 (a) HSO_4^- (b) HPO_4^{2-}
 (c) PO_4^{3-} (d) H_2PO_2^-
37. Which among the following pairs of acid and base are weak respectively?
 (a) HCOOH and NH_4OH
 (b) HCl and NaOH
 (c) CH_3COOH and KOH
 (d) HClO_4 and CsOH
38. The pH of 0.001 N sodium hydroxide solution at 25°C is
 (a) 3 (b) 4
 (c) 11 (d) 12
39. A solution turns methyl orange red. It can turn the universal indicator to _____.
 (a) violet (b) blue
 (c) orange (d) green
40. Which of the following pH values at 25°C corresponds to the strongest acid?
 (a) 2 (b) 4
 (c) 6 (d) 8
41. Which of the following during hydrolysis gives an acidic solution in water?
 (a) KNO_3 (b) NH_4Cl
 (c) $(\text{NH}_4)_2\text{CO}_3$ (d) KCl



42. X is formed by the partial replacement of hydroxyl groups of a diacidic base by an acidic radical. The number of ionizable hydroxyl groups in X is _____.
- (a) 0 (b) 1
(c) 2 (d) 3
43. In the reaction between silver nitrate and sodium chloride, the spectator ion is _____.
- (a) Ag^+ (b) Cl^-
(c) NO_3^- (d) All of these
44. Which among the following statements is false?
- (a) Every protonic acid has its conjugate acid.
- (b) Pair of Bronsted acid and base that differ by a proton is conjugate acid–base pair.
- (c) A substance that accepts an electron pair to form co-ordinate covalent bond is an acid.
- (d) Arrhenius theory is confined to aqueous solutions.
45. Which of the following salts can produce only two types of radicals?
- (a) NaKCO_3
(b) CaOCl_2
(c) NH_4HSO_4
(d) $\text{Na}(\text{NH}_4)\text{HPO}_4$

Level 2

- Calculate the resultant pH of a solution when 20 mL of 0.1N NaOH is mixed with 20 mL of 0.05 M $\text{Ca}(\text{OH})_2$ at 25°C.
- The ionic product of water is found to be $1 \times 10^{-12} \text{ mol}^2/\text{L}^2$ at 60°C. What would be the nature of solution with pH = 7? Explain giving reasons.
- Explain why phosphorous acid is not a tribasic acid like phosphoric acid though both acids contain three hydrogen in their molecules.
- What volume of 0.45 M H_3PO_4 is needed to neutralise 60 mL of 0.60 N NaOH solution? (Assume that there is a complete ionisation of acid and base.)
- Explain the nature of resultant solution when 20 mL of 0.05 M NaOH is mixed with 10 mL of 0.1N HCl solution at 25°C. When 10 mL of this solution is diluted 100 times what is the change in pH?
- Account for the variation of acidic character with increase in the oxidation state of central atom.
- Two salts are dissolved in water separately. When phenolphthalein is added to the two solutions, one solution turned pink and the other solution was colourless. How do you justify this?
- How many moles of NaOH must be added to 200 mL of 0.1 M solution of HNO_3 to get a solution having pH value 2.00?
- Explain the nature of the following salts:
(a) NaH_2PO_2 (b) Na_2HPO_3
(c) Na_2HPO_4
- An aqueous solution of Mohr's salt can give a precipitate with NaOH or NH_4OH . But the aqueous solution of potassium ferrocyanide cannot give a precipitate with both reagents. How do you account for this?
- Two coloured salt solutions of two metals A and B gave precipitates with NaOH and NH_4OH . On further addition of NH_4OH , the precipitate disappears, in the case of B. There is no such change in the case of 'A' either with NaOH or NH_4OH . Identify A and B. Give equations.
- Compare and contrast the nature of aqueous solutions of ammonium acetate and sodium sulphate.
- Find the pH of resultant solution when 100 mL of 0.005 M H_2SO_4 at 25°C is diluted to 1000 mL. What is the amount of NaOH required to be dissolved in 500 mL to exactly neutralise the above solution?
- A strong alkali forms an acidic salt with a dibasic acid. From the aqueous solution of the salt formed on the addition of dilute acid evolves a gas that turns lime water milky. Identify the salt and predict the reactions involved in it.
- Explain the variation in basic strengths of hydrides of VA group elements.

Direction for questions from 16 to 25:

Application-Based Questions

16. 14.3 g of $\text{Na}_2\text{CO}_3 \cdot x\text{H}_2\text{O}$ is dissolved in water and the volume is made up to 200 mL. 20 mL of this



solution required 40 mL of $\frac{N}{4}$ HNO_3 for complete neutralisation. Calculate x .

17. 10 mL of 1 N HCl, 25 mL of 2 N H_2SO_4 and 40 mL of X N HNO_3 are mixed and made up to 2000 mL. 100 mL of this solution required 30 mL of NaOH taken from a solution containing 4 g of NaOH in 250 mL of solution. What is the value of X ?
18. 8 g of mixture of NaNO_3 and Na_2CO_3 are dissolved and made into 500 mL of solution. 50 mL of this solution neutralises completely 25 mL of $N/5$ HNO_3 . Calculate the percentage composition of the mixture.
19. Do equimolar solutions of NaOH and $\text{Ca}(\text{OH})_2$ possess same pH value at the same temperature? Justify.
20. Explain why no indicator can be used for the titration between formic acid and ammonium hydroxide.
21. Sodium bicarbonate can be manufactured by Solvay process where addition of brine solution to ammonium bicarbonate solution gives precipitation of sodium bicarbonate. But, its analogue, potassium bicarbonate cannot be prepared by the above process. Explain with respect to the concept of solubility product.
22. Explain the acidic and basic nature of the following substances on the basis of protonic concept.

(a) H_2SO_4	(b) HPO_4^{-2}
(c) NH_3	(d) NH_4^+
23. 'All Lewis bases are Bronsted bases. But all Lewis acids are not Bronsted acids.' Comment on the statement.
24. Find the pH of resultant solution when 100 mL of 0.005M H_2SO_4 at 25°C is diluted to 1000 mL. What is the amount of NaOH required to be dissolved in 500 mL to exactly neutralise the above solution?
25. Explain the strengths of conjugate acid base pairs of the following species:

(a) HSO_4^-	(b) NH_3
(c) HS^-	(d) ClO_4^-

Level 3

1. Acetic acid contains four hydrogen atoms in each molecule. Explain why it is not a tetrabasic acid?
 2. How do you account for the acidic nature of 10^{-8} M nitric acid? Justify by mathematical calculation ($\log 1.1 = 0.041$).
 3. Though zinc and manganese get precipitated as their sulphides, they do not get precipitated along with lead or copper when H_2S is passed through acidified solutions of their salts. Justify and also explain how the precipitation of Zn^{+2} and Mn^{+2} ions takes place.
 4. How does the acidic character vary among the trihalides of boron? Give a reason in support of your answer.
- Direction for questions from 5 to 9:*
Application-Based Questions
5. Among H_3PO_2 , H_3PO_3 and H_3PO_4 which has the maximum acidic strength? Explain.
 6. Calculate the pH value of the 10^{-6} M HCl solution when diluted by 100 times. Justify ($\log 1.1 = 0.041$).
 7. Why do both acidic indicator and basic indicator show different colours in acidic and basic solutions?
 8. A complex, salt with formula $[\text{Co}(\text{NH}_3)_6]\text{Cl}_3$ can form a white curdy precipitate with silver nitrate solution. Another complex with same ligands cannot form precipitate with AgNO_3 solution. What would be the probable formula of that complex? Give reasons in support of your answer.
 9. Explain why the strengths of strong acids and strong bases cannot be determined in water. How is it possible to compare the acidic strengths in acetic acid?



CONCEPT APPLICATION

Level 1

True or false

- | | | | |
|----------|----------|---------|---------|
| 1. False | 2. False | 3. True | 4. True |
| 5. False | 6. True | 7. True | |

Fill in the blanks

- | | | | |
|-----------|------------------------|--------------------|--|
| 8. acidic | 10. basic, amphoteric | 12. greater than 7 | 14. mol.ion ² /L ² |
| 9. acidic | 11. 0.000001 mol ion/L | 13. normal salt | |

Match the following

- | | | |
|-----------|-------|-------|
| 15. A : b | C : d | E : f |
| B : a | D : c | F : e |

Multiple Choice Questions

- | | | | |
|-------|-------|-------|-------|
| 16. d | 20. d | 24. d | 28. b |
| 17. c | 21. c | 25. d | 29. d |
| 18. d | 22. b | 26. a | 30. b |
| 19. b | 23. a | 27. a | |

Solutions for questions from 31 to 45:

31. (i) calculation of the normality of acid and base
 (ii) calculation of the number of equivalents of acid and base
 (iii) calculation of the total volume of the mixture
 (iv) calculation of the net OH⁻ ion concentration
 (v) dividing K_w by OH⁻ ion concentration
 (vi) taking the negative logarithm of H⁺ ion concentration
32. (i) red (ii) orange (iii) yellow
 (iv) green (v) blue (vi) indigo
 (vii) violet
33. (i) identification of positive and negative radicals
 (ii) writing ionic equations for the reaction of positive and negative radicals with water
 (iii) identification of spectator ions
 (iv) writing the net reaction taking place
 (v) identification of free ions that caused the colour change of litmus paper
34. H₂PO₄⁻ can form conjugate base HPO₄²⁻ and conjugate acid H₃PO₄
35. Basicity of acetic acid = 1 since it ionises as CH₃COOH ⇌ CH₃COO⁻ + H⁺ and gives only one H⁺ ion. HNO₃ also possesses only one replaceable hydrogen and has basicity = 1.
36. PO₄³⁻ has only conjugate acid, that is, HPO₄²⁻. It has no conjugate base as it does not possess any replaceable H⁺ ion.
37. HCOOH is a weak acid and NH₄OH is a weak base.
38. [OH⁻] of sodium hydroxide = 0.001 = 10⁻³
 pOH = 3, pH = 14 - 3 = 11
39. Methyl orange shows red colour in acidic medium and universal indicator shows orange colour in the same pH range.
40. The lower is the pH value, the stronger is the acid. Therefore pH value of 2 corresponds to the strongest acid.



41. In the aqueous solution of NH_4Cl , NH_4^+ ion undergoes hydrolysis forming NH_4OH and HCl . Since HCl is a strong acid, the solution is acidic in nature.
42. Di-acidic base contains two $-\text{OH}^-$ groups in a molecule. If one $-\text{OH}^-$ group is replaced by an acidic radical, the number of ionizable hydroxyl radicals in the salt 'X' is one.
43. In the given reaction, Na^+ and NO_3^- ions are spectator ions.
44. Every protonic acid has a conjugate base.
45. $\text{NaKCO}_3 \rightarrow \text{Na}^+ + \text{K}^+ + \text{CO}_3^{2-}$
 $\text{CaOCl}_2 \rightarrow \text{Ca}^{+2} + \text{OCl}^- + \text{Cl}^-$
 $\text{NH}_4\text{HSO}_4 \rightarrow \text{NH}_4^+ + \text{HSO}_4^-$
 (two types of radicals)
 $\text{Na}(\text{NH}_4)\text{HPO}_4 \rightarrow \text{Na}^+ + \text{NH}_4^+ + \text{HPO}_4^-$

Level 2

Solutions for questions from 16 to 25:

- calculation of net H^+ ion concentration in the mixture
 - calculation of the total number of OH^- ions generated from the given volume of NaOH and $\text{Ca}(\text{OH})_2$
 - calculation of concentration of OH^- ions in the resultant solution
 - calculation of pOH
 - calculation of pH from the value of pOH at 25°C
 - 3
- concentration of H^+ and OH^-
 - $[\text{H}^+]/[\text{OH}^-]$ from the given ionic product of water
 - calculation of pH from $[\text{H}^+]$ (or) pOH from $[\text{OH}^-]$
 - relation between pH value and nature of the solution
 - relation between given pH and nature of the solution
- comparison of the structures of two acids
 - condition required for a substance to behave like an acid in its aqueous solution
 - comparison of positions of hydrogen atoms present in phosphorus acid to that present in phosphoric acid
 - relation between the position of hydrogen in the respective molecules and its replaceability
 - relation between the number of replaceable hydrogen and the basicity of the acid
- balanced equation and comparison of equivalents
 - basicity of acid
 - calculation of normality of H_3PO_4
 - calculation of number of equivalents of base
 - calculation of volume of acid required based on normality of acid and number of equivalents of base
 - 26.6 ml.
- calculation of net H^+ ions concentration in the mixture
 - calculation of the difference of number of H^+ and OH^- ions generated from the given amount of respective acid and base
 - determination of the pH of the given mixture at 25°C
 - relation between the pH of the solution and its nature
 - relation between the nature of the solution and the effect of dilution
- effect of electronegative atom on the $\text{O}-\text{H}$ bond strength
 - effect of oxidation state of the central atom on the position of electron cloud present between oxygen atom and the central atom
 - relation between the position of the above electron cloud and the charge separation between hydrogen and oxygen atom of the hydroxyl radical attached to the central atom
 - relation between the above charge separation and ease of formation of hydrogen ion
 - relation between ease of formation of hydrogen ion with the acidic character of the respective acids

7. (i) nature of solution where phenolphthalein is pink and colourless
 (ii) nature of the components produced on hydrolysis of salts
 (iii) relation between the components produced and the colour change of indicator
8. (i) pH of initial solution
 (ii) calculation of normality of the solution after neutralisation
 (iii) calculation of number of equivalents of HNO_3 neutralised
 (iv) calculation of number of moles of NaOH required to neutralise this HNO_3 in 200 ml
 (v) 0.018.
9. (i) identification of parent acid from which the respective salts are formed
 (ii) structures of the given salts
 (iii) number of hydroxyl radical present in the respective acid
 (iv) number of replaceable hydrogen in the respective acid
 (v) identification of the negative radical(s) that can be formed from the given acids
 (vi) determination of the number of replaceable hydrogen(s) present in the given salts based on the composition of the above negative radicals
 (vii) relation between the number of replaceable hydrogen(s) present in the salt and its nature
10. (i) characteristics of double salts
 (ii) type of salts
 (iii) comparison of ions present in their salt solutions
 (iv) requisite of a salt to give a precipitate
11. (i) metals giving coloured salts
 (ii) coloured salts that give precipitates with NaOH and NH_4OH
 (iii) comparison between the reactions of the precipitates obtained with excess NaOH and NH_4OH
 (iv) solubility of the product formed
12. (i) identification of species undergoing hydrolysis in the respective salts
 (ii) nature of constituent ions of the given salts
 (iii) comparison of reactions that take place in the respective solutions
 (iv) comparison of the concentration of hydrogen and hydroxyl ions in the respective solution
13. (i) calculation of H^+ ion concentration in the diluted solution
 (ii) calculation of the number of moles of hydrogen ions present in 100 mL of 0.005 M H_2SO_4
 (iii) calculation of the concentration of the hydrogen ions of H_2SO_4 solution after dilution
 (iv) calculation of the number of moles of NaOH required to neutralise the above solution
 (v) calculation of the amount of NaOH required
 (vi) 4×10^{-2} g.
14. (i) gas that turns lime water milky
 (ii) identification of gas
 (iii) dibasic acid responsible to produce this gas
 (iv) acidity of the alkali
 (v) requisite for acidic salt
15. (i) requisite for basic nature according to Lewis theory
 (ii) elements present in group VA
 (iii) number of lone pairs present in the valence shell of the elements of group VA
 (iv) concept of base according to Bronsted–Lowry theory
 (v) comparison of the size of the above elements
 (vi) relation between the size of the element and tendency of accepting the proton
 (vii) comparison of stability of the respective hydrides of the elements of group VA
16. $V_1N_1 = N_2V_2$
 $20(N_1) = 40 \times 0.25$
 $N_1 = 0.5 \text{ N}$
 Normality of $\text{Na}_2\text{CO}_3 \cdot x\text{H}_2\text{O}$ solution is 0.5
- $$0.5 = \frac{14.3}{\text{Eq.wt}} \times \frac{1000}{200}$$
- $$\text{Eq. Wt} = \frac{14.3 \times 5}{0.5}$$
- $$\text{Eq. Wt} = 143$$
- But equivalent weight $\text{Na}_2\text{CO}_3 \cdot x\text{H}_2\text{O}$ is
- $$\frac{106 + x(18)}{2} = 143.$$
- $$\therefore x = \frac{286 - 106}{18}$$
- $$x = 10$$
- $$\therefore \text{Na}_2\text{CO}_3 \cdot 10\text{H}_2\text{O}$$



17. Normality of NaOH

$$N_1 = \frac{4}{10} \times \frac{1000}{250}$$

$$N_1 = 0.4N$$

$$\therefore N_1 V_1 = N_{\text{mix}} V_{\text{mix}}$$

$$0.4 \times 30 = N_{\text{mix}} \times 100$$

$$N_{\text{mix}} = 0.12N$$

$$\frac{(10 \times 1) + (25 \times 2) + (40 \times x)}{2000} = 0.12N$$

$$240 = 60 + 40x$$

$$180 = 40x$$

$$x = \frac{180}{40} = 4.5$$

$$N_{\text{HNO}_3} = 4.5N$$

18. $\text{Na}_2\text{CO}_3 + 2\text{HNO}_3 \rightarrow 2\text{NaNO}_3 + \text{H}_2\text{CO}_3$

Only sodium carbonate reacts with acid (HNO_3)

$$N_1 V_1 = N_2 V_2$$

$$\text{Normality of } \text{Na}_2\text{CO}_3 \text{ is } \frac{0.2 \times 25}{50} = 0.1N$$

$$\therefore 0.1 = \frac{\text{wt of } \text{Na}_2\text{CO}_3}{53} \times \frac{1000}{500}$$

$$\therefore \text{wt of } \text{Na}_2\text{CO}_3 = \frac{0.1 \times 53 \times 500}{1000} = 2.65 \text{ g}$$

$$\therefore \text{wt of } \text{NaNO}_3 \text{ is } = 8 - 2.65 = 5.35 \text{ g}$$

$$\% \text{ of } \text{NaNO}_3 = \frac{5.35}{8} \times 100 = 66.8\%$$

$$\% \text{ of } \text{Na}_2\text{CO}_3 = \frac{2.65}{8} \times 100 = 33.2\%$$

19. Sodium hydroxide being a strong base than calcium hydroxide furnishes more number of OH^- as it undergoes complete dissociation. Greater the OH^- concentration greater will be the basic character. A stronger base will have pH value greater than a weak base. Hence, NaOH, which is a strong base has higher pH value than $\text{Ca}(\text{OH})_2$.

20. In acid–base titrations, the end point is detected by the use of a substance called acid–base indicator. Acid–base indicator exhibits a sharp change in its colour at the end point. The experimentally determined point of completion of an acid–base titration is called end point. For the titration between a weak

acid and weak base, there is no sharp change in pH. Hence, no indicator can be used.

21. Sodium bicarbonate has low value of solubility product. Addition of sodium chloride solution to ammonium carbonate solution results in the double displacement reaction that results in the precipitation of sodium bicarbonate, as ionic product easily exceeds the low value of solubility product. Potassium bicarbonate has high value of solubility product and is highly soluble. In this case, though double displacement takes place, ionic product cannot exceed the high value of solubility product of KHCO_3 . Therefore, precipitation cannot take place. Hence, this method cannot be employed for the preparation of potassium bicarbonate.

22. (a) H_2SO_4 is a proton donor and hence, can act as an acid according to protonic concept.

(b) HPO_4^{-2} can act as a proton donor and form PO_4^{-3} ion. Therefore, it acts as an acid. HPO_4^{-2} ion can also accept proton and form H_2PO_4^- . Therefore, it acts as a base.

(c) NH_3 can act as a proton donor to form NH_2^- and can act as an acid. NH_3 can act as a proton acceptor to form NH_4^+ ion and can act as a base.

(d) NH_4^+ can act as a proton donor to form NH_3 and it can act as an acid only.

23. A Lewis base is an electron pair donor. A molecule that has excess electron pair can only act as an electron pair donor. As it has an excess electron pair, it invariably accepts a proton. Therefore, it can act as a proton acceptor that is a Bronsted base. However, this is not the case with Lewis acids. A Lewis acid is an electron pair acceptor and hence, should be deficient in electrons. It need not necessarily be a proton donor as it may or may not possess hydrogen.

24. Normality of $\text{H}_2\text{SO}_4 = \text{Molarity} \times 2 = 0.005 \times 2 = 0.01$

According to dilution law

$$100 \times 0.01 = 1000 \times x$$

$$= \frac{100 \times 0.01}{1000} = 0.001 = 1 \times 10^{-3}$$

$$\text{pH} = -\log 1 \times 10^{-3} = 3$$

If this solution neutralised by 500 mL of NaOH solution



$$1000 \times 10^{-3} = 500 \times c$$

$$c = 2 \times 10^{-3}$$

Amount of NaOH required

$$2 \times 10^{-3} = \frac{x}{40} \times \frac{1000}{500}$$

$$2 \times 10^{-3} \times 40 / 2 = 4 \times 10^{-2} \text{ g.}$$



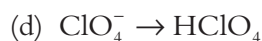
H_2SO_4 is a strong acid, HSO_4^- is a strong weak conjugate base. So, HSO_4^- is a strong acid; so, its conjugate base SO_4^{2-} is weak.



As ammonia is a weak base its conjugate acid is strong.



H_2S is a weak acid. HS^- is a strong conjugate base So, HS^- is weak acid so its conjugate base S^{2-} is strong.



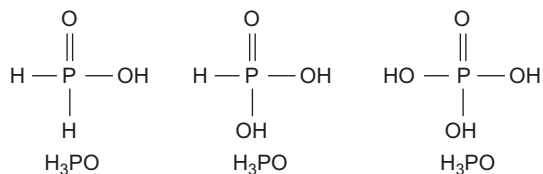
ClO_4^- is the conjugate base of strong acid HClO_4 . As ClO_4^- is the conjugate base of strong acid it has to be weak.

Level 3

- replaceable hydrogens in acetic acid
 - structure and bonding in acetic acid
 - comparison of the position of the hydrogen atoms present in the molecule of acetic acid
 - relation between the position of hydrogen atom and the effect of strong alkali on it
- pH range of acidic solutions
 - number of hydrogen ions furnished by nitric acid, present in the solution
 - determination of concentration of the hydrogen ions furnished by nitric acid
 - influence of the concentration of hydrogen ions furnished by the acid on the degree of dissociation of water
 - concentration of hydrogen ions furnished by water
 - total concentration of hydrogen ions present in the solution
 - calculation of pH
- requisite for precipitation of the respective sulphides
 - difference in the types of ions present in an acidified salt solution and normal salt solution
 - ions present in the solution when H_2S is passed through it
 - influence of extra ions present in the acidified solution on the degree of dissociation of H_2S
 - relation between the degree of dissociation of H_2S and the availability of sulphide ions in the solution
 - relation between the number of sulphide ions and the precipitation of the selective salts
- comparison of electronegativities of halogens attached to boron in various trihalides
 - reason for the trihalides of boron to act as an acid
 - comparison of valence electronic configuration of fluorine, chlorine, bromine, iodine with that of boron
 - orbitals in which the lone pairs of electrons of halogens are present
 - identification of vacant orbital of boron
 - comparison of energy of vacant orbital of boron with the orbitals containing lone pairs of electrons in halogens
 - comparison of tendency of donating the electron pair to the vacant orbital of boron among the halogens
 - relation between the tendency of formation of the above kind of bond and availability of vacant orbital of boron for accepting the lone pairs of electrons from other species
 - comparison of acidic characters of trihalides of boron according to Lewis concept
- Generally, the acidic strength of oxyacids of an element with different oxidation states increases with the increase in oxidation state of central atom. So, the expected order of acidic strength should be $\text{H}_3\text{PO}_4 > \text{H}_3\text{PO}_3 > \text{H}_3\text{PO}_2$. But actual order is $\text{H}_3\text{PO}_2 > \text{H}_3\text{PO}_3 > \text{H}_3\text{PO}_4$. This can be explained based on the number of OH^- groups attached to



central P-atom. With increasing number of OH groups on central P-atom the acidic strength decreases. Hence, H_3PO_4 is having minimum acidic strength and H_3PO_2 is having maximum acidic strength.



With the increases in number of $-\text{OH}$ groups, the strength of conjugate base formed increases due to greater proton accepting tendency. Therefore, the strengths of corresponding acids decrease accordingly.

6. 10^{-6} M HCl solution when diluted by 100 times, should become 10^{-8} M solution. Then the pH value should increase from 6 to 8. However, a solution having pH value of 6 is acidic and the solution having pH value of 8 indicates basic nature. An acidic solution can never become neutral or basic howsoever dilute it is. The pH value of the solution is slightly less than 7. This is because in very low concentrated solution, the dissociation of water also should be taken into consideration. H^+ ion concentration should be equal to the sum of H^+ ion concentration from HCl and H_2O .

$$[\text{H}^+] \text{ from HCl} = 10^{-8}$$

$$[\text{H}^+] \text{ from H}_2\text{O} = 10^{-7}$$

$$[\text{H}^+] = 10^{-8} + 10^{-7} = 10^{-7} (10^{-1} + 1) = 1.1 \times 10^{-7}$$

$$\text{pH} = -\log (1.1 \times 10^{-7})$$

$$= 7 - \log 1.1 = 7 - 0.041 = 6.96$$

7. Acid–base indicators are either weak organic acids or weak organic bases. These indicators show one colour in molecular form and another colour in ionic form. Acidic indicator is represented by HIn .



In acidic solution, the equilibrium shifts towards left.

Hence, an acidic indicator has the colour of HIn in acidic solutions. In basic solution, the equilibrium shifts towards right and shows colour of In^- in alkaline solutions.

Basic indicator is represented by InOH . $\text{InOH} \rightleftharpoons \text{In}^+ + \text{OH}^-$

In acidic solution, the equilibrium shifts towards right. In basic solution, the equilibrium shifts towards left. Hence, basic indicator is having colour of In^+ in acidic solutions, and shows colour of InOH in basic solutions.

8. In a complex salt, there are two parts namely coordination sphere and ionisation sphere. The part present in ionisation sphere gives reactions of the respective ions. But the part present in coordination sphere gives the reactions of complex ions only and cannot give reactions of individual ions. In the complex of the formula $[\text{Co}(\text{NH}_3)_3]\text{Cl}_3$, chloride ions are present in the ionisation sphere and give the reactions of Cl^- ions. The complex $[\text{Co}(\text{NH}_3)_3\text{Cl}_3]$ has no chloride ions in the ionisation sphere and does not give the reactions of Cl^- ions because the components present within the sphere do not further undergo any hydrolysis and hence, do not behave independently.
9. In presence of any strong acid, water acts as a proton acceptor. In presence of any strong base, water acts as a proton donor. Therefore, the ability of a strong acid to donate protons is determined by the proton accepting capacity of water and not on the acid. Hence, all strong acids show equal tendency to donate protons. Same thing happens with bases also. When the strong acids are taken in acetic acid, which is a very weak proton acceptor than water, the different acids donate protons to different extents. Therefore, the strengths of various strong acids can be compared in acetic acid.

