Chapter # 08

Acids, Bases and Salts

Major Concepts

8.1 Acidic, Basic and Amphoteric Substances.
8.2 Bronsted-lowry Definitions of Acids and Bases
8.3 Relative Strength of Acids and Bases
8.4 Expressing the Strength of Acids and Bases
8.5 Lewis Definitions of Acids and Bases
8.6 Buffers Solutions and their Applications
8.7 Salt Hydrolysis

Learning Outcomes

The students will be able to:

- Define Bronsted and Lowery concepts for acids and bases(Remembering)
- Define salts, conjugate acids and conjugate bases. (Remembering)
- Identify conjugate acid-base pairs of Bronsted-Lowery acid and base(Analyzing)
- Explain ionization constant of water and calculate pH and pOH in aqueous medium using given K_w values. (Applying)
- Use the extent of ionization and the acid dissociation constant, K_a, to distinguish between strong and weak acids. (Applying)
- Use the extent of ionization and the base dissociation constant, K_b, to distinguish between strong and weak bases. (Applying)
- Define a buffer, and show with equations how a buffer system works. (Applying)
- Make a buffered solution and explain how such a solution maintains a constant pH even with the addition of small amounts of strong acid or strong base. (Understanding)
- Use the concept of hydrolysis to explain why aqueous solutions of some salts are acidic or basic. (Applying)
- Use concept of hydrolysis to explain why the solution of a salt is not necessarily neutral. (Understanding)
- Define and explain leveling effect. (Understanding)

Introduction

Acids, bases and salts are very important and common chemical compounds. They are used in the home and in the chemical industries. For example, vinegar contains acetic acid that is common in many foods and cleaning products. Citrus fruits such as lemons, oranges and tomatoes juices contain citric and ascorbic acids. Apples contain malic acid, and grape juice contains tartaric acid. Dairy products such as sour milk, yogurt and cheese contain lactic acid. Carbonated beverages (soft drinks) contain carbonic acid and may also contain benzoic acid, citric acid and phosphoric acid. Gastric juice contains hydrochloric acid which is essential to digestion and is secreted by the lining of our stomachs in a significant quantity of 1.2-1.5 L/day. The sulphuric acid is used in the lead storage batteries and is also used in the manufacture of fertilizers, explosives, dyes and glues. Bacteria in our mouths yield acids that can dissolve tooth enamel to produce cavities. On the other hand, the sodium hydroxide is a base and is used in the manufacture of soaps, detergents and in the manufacture of other compounds. The ammonia is also a base and is present in many household cleaners. The suspension of magnesium hydroxide also known as milk of magnesia is used as an antacid. Aluminum hydroxide and sodium bicarbonate are also bases and are also used as antacids. The sodium carbonate also known as washing soda is a common example of base and is used for washing purposes.

Keep in mind

Never check the taste of acids in the laboratory because some acids are very corrosive and some are poisonous.

Salts are formed when acids react with bases. Some well-known examples of salts are sodium chloride, sodium nitrate, and potassium iodide. Sodium chloride is the only one of many hundreds of salts that is used for seasoning and preserving foods.

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Preservatives in Food Products and Allergic Reactions in People:

A preservative is a naturally occurring or synthetically produced substance that is added to food or other organic materials to prevent decomposition or fermentation. Preservatives may either be antimicrobial, antioxidants or anti-enzymatic. Antimicrobial preservatives inhibit the growth of bacteria, yeast and molds. Antioxidant preservatives inhibit the oxidation of food constituents. They keep foods from becoming rancid, browning, or developing black spots. Anti-enzymatic preservatives block the enzymatic processes such as ripening of fruits and vegetables occurring in foodstuffs even after harvest. Salt, sugar and vinegar and natural spices are also considered as food preservatives. However, the main concerns of using food preservatives are mostly related to chemical substances and artificial ingredients.

Some people are sensitive to particular food preservatives and may ha reactions like hives, diarrhoea, digestive disorders and respiratory problems such a asthma. Most of the food preservatives are safe.

8.1 Acidic, Basic and Amphoteric Substance

Acids and bases were first recognized by their characteristic properties. The w acid has been taken from Latin word "acidus", which means "sour". Acids a compounds whose water solutions taste sour, turn blue litmus red, and react w bases to form salts. The bases are the compounds that taste bitter, feel slippery, turn red litmus blue and react with acids to form salts. Many theories have been proposed to define an acid and a base. One of the earliest, most significant of these theories was Arrhenius theory of acids and bases. In 1887, Svante Arrhenius (1859-1927), a Swedish chemist, stated that ***acids*** are substances that produce hydrogen ions (H^+ ions) and anions when dissolved in water while ***bases*** are substances that produce hydroxide ions (OH^- ions) and cations when dissolved in water. For example, HCl is an Arrhenius acid becaus produces hydrogen ions and anions (Cl^- ions) when dissolved in water.

Svante Arrheni
(1859-1927)

$$HCl_{(g)} \xrightarrow{H_2O} H^+_{(aq)} + Cl^-_{(aq)}$$

The aqueous solution of HCl is called hydrochloric acid.

Keep in Mind

The hydrogen ion is very reactive and cannot exist freely in the aqueous solutions it combines with solvent water molecule to produce hydronium ion.

$$H^+_{(aq)} + H_2O_{(l)} \longrightarrow H_3O^+_{(aq)}$$

The NaOH is an Arrhenius base because it produces hydroxide ions cations (Na^+ ions) when dissolved in water.

$$NaOH_{(s)} \xrightarrow{H_2O} Na^+_{(aq)} + OH^-_{(aq)}$$

The neutralization of an acid by a base produces water and salt. Salt i ionic compound that is composed of the cation from the base and the anion fr acid.

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

Arrhenius observed that process of neutralization occurs when the from the acid and OH^- ion from the base reacts to produce water.

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

The substances which have both the properties of acids and bases are called amphoteric substances. An amphoteric substance acts as an acid when combined with strong base and as a base when combined with strong acid. For example, water shows basic properties by reacting with HCl.

$$\underset{\text{Acid}}{HCl_{(aq)}} + \underset{\text{Base}}{H_2O_{(l)}} \rightleftharpoons H_3O^+_{(aq)} + Cl^-_{(aq)}$$

Here, H_2O accepts a proton from the HCl molecule in the forward reaction. It also shows acidic properties by reacting with NH_3.

$$\underset{\text{Base}}{NH_{3(aq)}} + \underset{\text{Acid}}{H_2O_{(l)}} \rightleftharpoons NH^+_{4(aq)} + OH^-_{(aq)}$$

Here, H_2O donates a proton to NH_3 molecule in the forward reaction. The amphoteric substances may be either molecules or ions. For example, bicarbonate ion of baking soda is amphoteric.

$$\underset{\text{Acid}}{HCO^-_{3(aq)}} + \underset{\text{Base}}{OH^-_{(aq)}} \rightleftharpoons CO^-_{3(aq)} + H_2O_{(l)}$$

Here, the HCO_3^- ion donates a proton to a base and acts as an acid.

$$\underset{\text{Base}}{HCO^-_{3(aq)}} + \underset{\text{Acid}}{H_3O^+_{(aq)}} \rightleftharpoons H_2CO_{3(aq)} + H_2O_{(l)}$$

Keep in mind

Soluble bases are known as alkalis. All the alkalis are bases but all the bases are not alkalis. Most bases are water insoluble.

Here, the HCO_3^- ion accepts a proton from an acid and acts as a base.

The acidic and basic character of amphoteric substance depends on the other substance to which it reacts.

8.2 Bronsted-Lowry Definitions of Acids and Bases

8.2.1 Proton Donors and Acceptors

Although Arrhenius theory is widely used but it failed to answer the question why compounds such as ammonia form a basic solution when dissolved in water although it does not contain hydroxide ion. The Arrhenius definition also does not explain those acid-base reactions which occur in non-aqueous solvents and those which occur in gas phase where no solvent is present. A more general definition of acids and bases was proposed

Bronsted (1879-1947) Thomas Lowry (1874-1936) by a Danish chemist Johannes Bronsted and an English chemist Thomas Lowry in 1923. According to the Bronsted-Lowry theory, an acid is a substance that donates a proton (H^+ ion) to another substance, and a base is a substance that accepts the proton from another substance. In short, an acid is proton donor and a base is proton acceptor and acid-base reactions are proton transfer reactions. For example, HCl is a Bronsted-Lowry acid because its molecule donates a proton to water molecule when dissolved in water.

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

On the other hand, water acts as Bronsted-Lowry base, because its molecule accepts a proton from HCl molecule. Consider the reaction between gaseous HCl and ammonia.

$$HCl_{(g)} + NH_{3(g)} \longrightarrow NH_4^+ + Cl^- \longrightarrow NH_4Cl_{(s)}$$

In the above equation, the HCl donates proton to NH_3 and acts as Bronsted-Lowry acid while NH_3 accepts proton from HCl and acts as the Bronsted-Lowry base. This is gas phase reaction and no solvent is required.

8.2.2 Conjugate Acid-Base Pairs

The dissociation of an acid can best be represented by the general reaction given below:

Conjugate acid-base pair

$$\underset{\text{Acid}}{HA_{(aq)}} + \underset{\text{Base}}{H_2O_{(l)}} \rightleftharpoons \underset{\text{Conjugate acid}}{H_3O^+_{(aq)}} + \underset{\text{Conjugate base}}{A^-_{(aq)}}$$

Conjugate acid-base pair

In the forward reaction, the HA donates protons to water and acts as a Bronsted-Lowry acid and water accepts proton form HA and acts as a Bronsted-Lowry base. In the reverse reaction, the H_3O^+ acts as an acid and A^- acts as a base. An acid and a base such as HA and A^- that differ only by a single proton are called conjugate acid-base pair. Every acid has a conjugate base that is formed by loss of proton from the acid. For example, OH^- is a conjugate base of H_2O and A^- is a conjugate base of HA. Every base has a conjugate acid that is formed by adding the proton to the base. Thus, H_3O^+ is the conjugate acid of H_2O, and HA is the conjugate acid of A^-. We may also say that a conjugate acid is formed by ionization of a base and a conjugate base is formed by the ionization of an acid.

Consider the dissociation of HCl in water:

$$\underset{\text{Acid}}{HCl_{(aq)}} + \underset{\text{Base}}{H_2O_{(l)}} \longrightarrow \underset{\substack{\text{Conjugate}\\\text{Acid}}}{H_3O^+_{(aq)}} + \underset{\substack{\text{Conjugate}\\\text{Base}}}{Cl^-_{(aq)}}$$

In this reaction HCl acts as an acid because it has donated proton and H_2O acts as a base because it has accepted proton. For any given acid-base pair, the acid always has one more hydrogen atom and one unit fewer negative charges than the base. On the other hand, the base has one fewer hydrogen atom and one unit more negative charge than the acid.

8.3 Relative Strengths of Acids and Bases

The tendency of an acid to give hydrogen ion is called strength of an acid. The tendency of a base to accept hydrogen ion is called strength of a base. Strong acids and bases are strong electrolytes which are assumed to ionize completely in water. Weak acids and bases are weak electrolytes which ionize only to a limited extent in water. For example, HCl is a strong acid and CH_3COOH is a weak acid because HCl is almost 100% ionized and CH_3COOH is about 1% ionized in water. The strength of acids and bases can be determined by passing electricity through their solutions. The aqueous solutions of strong acids and strong bases are good conductors of electricity and the aqueous solutions of weak acids and bases are not the good conductors of electricity.

The strength of conjugate acids and bases are related to the parent acids and bases as:

(i) The stronger the acid, the weaker is its conjugate base and vice versa.

(ii) The weaker the base, the stronger is its conjugate acid and vice versa.

Consider the ionization of HCl in water:

$$\underset{\text{Stronger acid}}{HCl_{(aq)}} + \underset{\text{Weaker base}}{H_2O_{(l)}} \longrightarrow \underset{\text{Stonger acid}}{H_3O^+_{(aq)}} + \underset{\text{Weaker base}}{Cl^-_{(aq)}}$$

HCl is a strong acid that completely donates its protons to water. The chloride ion (the conjugate base of HCl) is an extremely weak base and has very little tendency to gain proton from water. Thus the strongest acids have the weakest conjugate bases.

Consider another example of the ionization of ammonia in water:

$$\underset{\text{Weaker base}}{NH_{3(aq)}} + \underset{\text{Weaker acid}}{H_2O_{(l)}} \longrightarrow \underset{\text{Stronger acid}}{NH^+_{4(aq)}} + \underset{\text{Stronger base}}{OH^-_{(aq)}}$$

Ammonia is a weak base and its conjugate base (NH_4^+ ion) is a strong acid.

The NH_4^+ ion (the conjugate acid of NH_3) is a strong acid and has greater tendenc donate proton to OH^- ion. Thus the weakest bases have the strongest conjug acids.

Table 8.1: Relative Strengths of Acids and Bases

	Acid	Name	Base	Name	
Strongest acids	$HClO_4$	Per chloric acid	ClO_4^-	Perchlorate ion	Weakest bases
	HCl	Hydrochloric acid	Cl^-	Chloride ion	
	H_2SO_4	Sulphuric acid	HSO_4^-	Hydrogen sulphate ion	
	HNO_3	Nitric acid	NO_3^-	Nitrate ion	
	H_3O^+	Hydronium ion	H_2O	Water	
	HSO_4^-	Hydrogen suphate ion	SO_4^{2-}	Sulphate ion	
	H_3PO_4	Phosphoric acid	$H_2PO_4^-$	Dihydrogen phosphate ion	
	HNO_2	Nitrous acid	NO_2^-	Nitrite ion	
	HF	Hydrofluoric acid	F^-	Fluoride ion	
	CH_3COOH	Acetic acid	CH_3COO^-	Acetate ion	
	H_2CO_3	Carbonic acid	HCO_3^-	Bicarbonate ion	
	NH_4^+	Ammonium ion	NH_3	Ammonia	
	HCN	Hydrocyanic acid	CN^-	Cyanide ion	
	HCO_3^-	Bicarbonate ion	CO_3^{2-}	Carbonate ion	
	H_2O	Water	OH^-	Hydroxide ion	
	NH_3	Ammonia	NH_4^+	Ammonium ion	
Weakest acids	OH^-	Hydroxide ion	H_2O	water	Strongest bases

8.4 Expressing the Strength of Acids and Bases

The strength of an acid or a base is expressed in terms of hydrogen and hydroxide ions concentrations, pH and pOH scales, K_a, K_b, pK_a, and pK_b values.

8.4.1 Ionic Product of Water (K_w)

Even highly purified water is observed to conduct very small electricity which shows the presence of very small concentration of ions. These ions are produced by the auto-ionization or self-ionization of water molecules.

$$\underset{\text{Acid}}{H_2O_{(l)}} + \underset{\text{Base}}{H_2O_{(l)}} \rightleftharpoons \underset{\text{Acid}}{H_3O^+_{(aq)}} + \underset{\text{Base}}{OH^-_{(aq)}}$$

As water is amphoteric, hence it acts both as a proton donor and proton acceptor. The ionization reaction of water can also be written as:

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

Here pure water ionizes into H^+ and OH^-. According to law of mass action, the equilibrium constant for the reaction is:

$$K_c = \frac{[H^+][OH^-]}{[H_2O]}$$

Since water is in excess, and very few of its molecules undergo ionization, so its concentration remains constant. Therefore:

$$K_c.[H_2O] = [H^+][OH^-] \quad \text{or}$$

$$K_w = [H^+][OH^-]$$

The equilibrium constant K_w is called ionic product of water and its value at 25°C is always $1.0 \times 10^{-14} mol^{-2}dm^{-6}$ obtained by multiplying together the molar concentrations of H^+ ions and OH^- ions present in pure water at room temperature.

$$\text{Ionic product of water } (K_w) = [H^+][OH^-]$$

$$= (10^{-7} mol\ dm^{-3})(10^{-7} mol\ dm^{-3})$$

$$= 1.0 \times 10^{-14} mol^2 dm^{-6}$$

K_w is temperature dependent. Its value increases with increase in temperature and decreases with decrease in temperature and has a constant value at constant temperature.

Table 8.2: Value of K_w at different temperatures

Temperature (°C)	K_w value (mol^2dm^{-6})
0	1.2×10^{-15}
10	3.0×10^{-15}
25	1.0×10^{-14}
50	5.3×10^{-14}

The relative concentrations of H^+ and OH^- ions indicates the acidic, neutral, or basic character of the aqueous solution. There are three possibilities for all aqueous solutions.

a) An aqueous solution in which the concentration of H^+ ion is equal to OH^- ion is called neutral solution. For neutral solution, the $[H^+]=[OH^-]=10^{-7} mol\ dm^{-3}$.

b) An aqueous solution in which the concentration of H^+ ion is greater than OH^- ions is called acidic solution. For acidic solution, the $[H^+] > [OH^-]$. In this solution $[H^+] >$ 10^{-7} mol dm^{-3} and $[OH^-] < 10^{-7}$ mol dm^{-3}.

c) An aqueous solution in which the concentration of H^+ ion is less than that of OH^- ion is called basic solution. For basic solution, the $[H^+] < [OH^-]$. In this solution $[H^+]$ $< 10^{-7}$ mol dm^{-3} and $[OH^-] > 10^{-7}$ mol dm^{-3}.

When acid is added to water then $[H^+] > [OH^-]$ and when a base is added to water then $[H^+] < [OH^-]$.

Example 8.1

The concentration of H^+ ions in an orange juice is 5.5×10^{-4}M. Calculate the concentration of OH^- ions; classify the solution as acidic, basic or neutral.

Solution:

$[H^+] = 5.5 \times 10^{-4}$M

$[OH^-] = ?$

$K_w = [H^+][OH^-]$

or $$[OH^-] = \frac{K_w}{[H^+]} = \frac{1.0 \times 10^{-14}}{5.5 \times 10^{-4}} = 1.82 \times 10^{-11}\ M$$

The solution is acidic because the $[OH^-] < 10^{-7}$M.

Practice Exercise 1:

The concentration of OH^- ions in gastric juice is 1.0×10^{-13}M. Calculate the concentration of H^+ ions; classify the solution as acidic, basic or neutral.

8.4.2 pH, pOH and pK_w

The molar concentrations of hydrogen ions and hydroxide ions are very small in aqueous solutions. Their concentrations generally range from 10 mol dm^{-3} to 10^{-15} mol dm^{-3}. It is, therefore, more convenient to express these concentrations on a logarithm scale, known as the pH scale.

The concept of pH scale was first introduced by the Danish chemist Soren Sorenson (1868 - 1939) in the year 1909 as easy and convenient way for expressing hydrogen ion concentration (acidity and basicity).

The pH of solution is defined as:

The negative logarithm of hydrogen ion concentration in an aqueous solution is called pH.

$pH = -\log[H^+]$

For example, the pH of pure water is 7 and that of apples is 3.1

The pOH of solution may be defined as:

The negative logarithm of hydroxide ion concentration in an aqueous is called pOH.

$pOH = -\log[OH^-]$

For example, the pOH of pure water is 7 and that of apples is 10.9

The pH values range from 0 to 14. However, solutions of negative pH and having pH > 14 are also known.

For Acidic solution the pH value is less than 7(pH < 7).

For Neutral solution the pH value is equal to 7(pH = 7).

For Basic solution the pH value is greater than 7(pH >7).

The pK_w is the negative logarithm of K_w in the solution.

> **Keep in mind**
>
> The pH scale is basically the reverse of pOH scale.

$pK_w = -\log K_w$

The numerical value of K_w is 1.0×10^{-14} and its pK_w value can be calculated as:

$$pK_w = -(\log 1.0\times10^{-14})$$
$$= -(-14)$$
$$= 14$$

The pK_w value can also be calculated as:

$$pK_w = pH + pOH = 7+7 = 14$$

Table 8.3: The pH and pOH Values for a Range of $[H^+]$ and $[OH^-]$ Concentrations

$[H^+]$	pH		$[OH^-]$	pOH	
10^{-14}	14		10^{0}	0	
10^{-13}	13	Decreasing acidity	10^{-1}	1	Increasing basicity
10^{-12}	12		10^{-2}	2	
10^{-11}	11		10^{-3}	3	
10^{-10}	10		10^{-4}	4	
10^{-9}	9		10^{-5}	5	
10^{-8}	8		10^{-6}	6	
10^{-7}	7	Neutral	10^{-7}	7	Neutral
10^{-6}	6		10^{-8}	8	
10^{-5}	5		10^{-9}	9	
10^{-4}	4		10^{-10}	10	
10^{-3}	3	Increasing acidity	10^{-11}	11	Decreasing Basicity
10^{-2}	2		10^{-12}	12	
10^{-1}	1		10^{-13}	13	
10^{0}	0		10^{-14}	14	

The pH values are the inverse of $[H^+]$. Greater the value of pH, smaller will be the value of $[H^+]$ and weaker will be the acid.

Keep in Mind

- The acids which have pH less than 4 are strong acids (pH < 4).
- The acids which have pH equal to or greater than 4 but less than 6 are weak acids (pH $\geq$ 4<6).
- The solutions which have pH equal to or greater than 6 but less than 8 are neutral or nearly neutral (pH $\geq$ 6<8).
- The bases which have pH equal to or greater than 8 but less than 10 are weak bases (pH $\geq$ 8<10).
- The bases which have pH equal to or greater than 10 are strong bases (pH $\geq$10).

Table 8.4: The pH of Some Common Substances

Substance	pH	Substance	pH	Substance	pH
Gastric juice (stomach acid)	1.0-3.0	Tomatoes	4.0-4.5	Milk (cows)	6.3-6.6
Lemon juice	1.8-2.4	Bear	4.0-5.0	Pure water	7.0
Soft drinks	2.0-4.0	Banana	4.5-5.7	Blood (human)	7.3-7.5
Vinegar	2.4-3.4	Urine (human)	4.8-8.4	Tears	7.4
Plums	2.8-3.0	Carrots	4.9-5.3	Egg white	7.6-8.0
Apples	2.9-3.3	Peas	5.8-6.4	Sea water	8.0-8.5
Cherries	3.2-4.0	Butter	6.1-6.4	Milk of magnesia	10.6
Peaches	3.4-3.6	Rain water	6.2	Household ammonia	11.5
Pears	3.6-4.0	Potatoes	5.6-6.0	Drain cleaner	13.0
Orange juice	3.5-4.0	Saliva (human)	6.4-6.9	1 M NaOH	14.0

Most of the fruits are acidic in nature. The sour taste is the indication of the acidic character of fruits. When the difference of the pH values of two acids is one then one acid is ten times stronger than other (or it has ten times more hydrogen ion concentration). For example, a lemon with a pH of 2.0 is ten times more acidic than an apple with a pH of 3.0 and is hundred times more acidic than cherry with a pH of 4.0. An increase of one in pH corresponds to a tenfold decrease in $[H^+]$.

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Curdling of Milk with Lemon Juice:

Curdling means to change into curds. Curds are a dairy product obtained by curdling of milk by acid or rennet, used in making cheese or eaten as food.

Milk is actually made up of water, suspended fat, and protein. It contains a specific protein called casein. Casein floats freely in groups (called micelles) throughout the milk and does not usually bond to anything. These grouping (micelles) have negative charge. They repel each other, which stops them from clumping and keeps the casein groupings (micelles) evenly dispersed in the milk. When lemon juice that contains citric acid is added to milk, the pH of milk drops and it becomes more acidic. The positive hydrogen atoms get attracted to the negative micelles, making them neutral. Now instead of pushing one another part, the micelles start to attract one another and stick together. Thus the milk gets curdled. The process of curdling happens much faster when lemon juice is added to hot milk. But it is better to add a lemon juice to milk when both the liquids are cold and lemon juice is added slowly. The more lemon juice you add, the larger your curds will be and the faster they will form.

Example 8.2

Calculate the pH of the carrots juice which has the hydrogen ion concentration 5.5×10^{-6}M.

Solution:

The pH of carrots juice = ?

$[H^+] = 5.5\times10^{-6}$M

pH $= -\log[H^+]$

pH of carrots juice $= -(\log 5.5\times10^{-6})$

$= -[(\log 5.5)+(-6\log 10)]$

$= -[(0.740)+(-6\times1)]$

$= -[0.740-6]$

$= -[-5.26]$

$= 5.26$

Practice Exercise 2:

Calculate the pH and pOH of the egg white having hydrogen ion concentration 2.51×10^{-8}M. Is the solution acidic or basic?

Example 8.3

(a) Calculate the hydrogen ion concentration of banana, which has a pH value of 5.0.

(b) What are the hydrogen and hydroxide ion concentrations of human [blood] that has a pH of 7.4?

Solution (a):

If we may be given the pH value of a solution and asked to calculate hydrogen ion concentration, then we need to take the antilog of the equation

$[H^+] = 10^{-pH}$

$[H^+] = ?$

pH of banana $= 5.0$

As we know, $[H^+] = 10^{-pH}$

By putting the value of pH we have,

$[H^+] = 10^{-5}$

Solution (b):

$[H^+] = ?$

$[OH^-] = ?$

pH of human blood $= 7.4$

As we know, $[H^+] = 10^{-pH}$

By putting the value of pH in the above equation, we obtain

$[H^+] = 10^{-7.4}$

$[H^+] = 3.98 \times 10^{-8}$

We also know that,

$$[H^+][OH^-] = 1.0 \times 10^{-14} \quad \text{or}$$

$$[OH^-] = \frac{1.0 \times 10^{-14}}{[H^+]}$$

$$\text{Then, } [OH^-] = \frac{1.0 \times 10^{-14}}{3.98 \times 10^{-8}}$$

$$[OH^-] = 0.251 \times 10^{-14} \times 10^{8}$$

$$[OH^-] = 0.251 \times 10^{-6}$$

$$[OH^-] = 2.51 \times 10^{-5}$$

Practice Exercise 3:

Calculate the H^+ ions concentrations of the following solutions:

(a) Milk which has pH 6.40.

(b) Sea water which has pH 8.30.

8.4.3 Acid Ionization Constant, K_a and pK_a

The equilibrium constant which shows the strength of an acid is called ionization constant of an acid (K_a). Let us consider the ionization of an acid, HA in water:

$$HA + H_2O \rightleftharpoons H_3O^+ + A^-$$

According to law of mass action, the equilibrium constant for the reaction is given as:

$$K_c = \frac{[H_3O^+][A^-]}{[HA][H_2O]}$$ (The bracket shows molar concentration)

In dilute solutions, the concentration of water is almost constant because it is always in excess:

So, $$K_c . [H_2O] = \frac{[H_3O^+][A^-]}{[HA]}$$

or $$K_a = \frac{[H_3O^+][A^-]}{[HA]}$$

The K_a is called ionization constant of an acid. Greater the value of K_a, the stronger is the acid and the greater is the concentration of H^+ ions at equilibrium because of its ionization.

Strength of an acid $\propto K_a$

- For weak acids, the K_a value is less than 10^{-3}. ($K_a < 10^{-3}$)
- For moderately weak acids, the K_a values range from 1 to 10^{-3} (K_a = 1 to 10^{-3})
- For strong acids, the K_a value is more than 1. ($K_a > 1$)

The negative logarithm of K_a is called pK_a.

$$pK_a = -\log K_a$$

The pK_a is the inverse of K_a.

$$pK_a \propto \frac{1}{K_a}$$

Greater the value of K_a, smaller will be the value of pK_a and stronger will be the acid. When the difference of pK_a values of two acids is equal to one, then one acid is ten times stronger than the other. When difference is two, then one acid is 100 times stronger than the other. When difference is three, then one acid is 1000 times stronger than the other.

Table 8.5: K_a and pK_a Values for Some Acids at Room Temperature

Acids	Name	K_a	pK_a
HI	Hydroiodic acid	1.0×10^{10}	−10
$HClO_4$	Perchloric acid	1.0×10^{10}	−10
HBr	Hydrobromic acid	1.0×10^{9}	−9
HCl	Hydrochloric acid	1.0×10^{7}	−7
H_2SO_4	Sulphuric acid	1.0×10^{3}	−3
HNO_3	Nitric acid	25	−1.4
H_3O^+	Hydronium ion	1.0	Zero
$(COOH)_2$	Oxalic acid	5.6×10^{-2}	1.25
H_2SO_3	Suphorus acid	1.7×10^{-2}	1.77
$HClO_2$	Chlorus acid	1.2×10^{-2}	1.92
HF	Hydrofluoric acid	6.8×10^{-4}	3.17
HNO_2	Nitrous acid	4.5×10^{-4}	3.35
$HCOOH$	Formic acid	3.0×10^{-4}	3.52
CH_3COOH	Acetic acid	1.8×10^{-4}	3.74
C_6H_5COOH	Benzoic acid	6.3×10^{-5}	4.20
H_2CO_3	Carbonic acid	4.3×10^{-7}	6.37
H_2S	Hydrogen sulphide	1.0×10^{-7}	7.00
$HClO$	Hypochlorus acid	6.8×10^{-8}	7.17
H_3BO_3	Boric acid	5.9×10^{-10}	9.23
HCN	Hydrocyanic acid	4.9×10^{-10}	9.31
C_6H_5OH	Phenol	1.3×10^{-10}	9.89
H_2O	Water	1.0×10^{-14}	14

Keep in Mind

You can determine the pH value or H^+ ion concentration of a weak acid solution, if you know the initial concentration of acid and its K_a value. On the other hand, you can determine the K_a value of an acid solution, if you know the pH value and H^+ ion concentration of its solution.

Example 8.4

Calculate the pH of 0.05 M mono-protic acid whose K_a value is 2.5×10^{-5}.

Solution:

Initial concentration of mono-protic acid = 0.05M

The K_a value of mono-protic acid = 2.5×10^{-5}

The pH of mono-protic acid = ?

The equation for the ionization of an acid is:

$$HA_{(aq)} \rightleftharpoons H^+_{(aq)} + A^-_{(aq)}$$

	$HA_{(aq)}$	$H^+_{(aq)}$	$A^-_{(aq)}$
Initial conc. (M):	0.05	0	0
Equilibrium conc. (M):	$(0.05-x)$	x	x

or

$$K_a = \frac{(x)(x)}{0.05-x} = 2.5 \times 10^{-5}$$

$$2.5 \times 10^{-5} = \frac{(x)(x)}{(0.05-x)} = \frac{(x^2)}{(0.05-x)}$$

The value of K_a is small, hence, the ionization of HA is assumed to be very small and the value of x is considered as zero.

$$2.5 \times 10^{-5} = \frac{x^2}{0.05}$$

$$(2.5 \times 10^{-5})(0.05) = x^2 \quad \text{or}$$

$$x^2 = (2.5 \times 10^{-5})(0.05)$$

$$x^2 = 0.125 \times 10^{-5}$$

$$x^2 = 1.25 \times 10^{-6}$$

by taking under root,

$$x = \sqrt{1.25 \times 10^{-6}}$$

$$x = 1.12 \times 10^{-3} M$$

$$pH = -\log[H^+] = -\log(1.12 \times 10^{-3})$$

$$= -(-2.95)$$

$$= 2.95$$

Practice Exercise 4:

Calculate the pH of 0.10 M formic acid whose K_a value is 1.8×10^{-14}.

Example 8.5

Calculate the K_a value of mono-protic acid which has $[HA] = 2.5 \times 10^{-1} M$, $[H_3O^+] = [A^-] = 4.2 \times 10^{-2} M$ at room temperature.

Solution:

$[HA] = 2.5 \times 10^{-1} M$

$[H_3O^+] = [A^-] = 4.2 \times 10^{-2} M$

$[A^-] = 4.2\times10^{-2}M$

K_a value of mono-protic acid = ?

$$HA + H_2O \rightleftharpoons H_3O^+ + A^-$$

$$K_a = \frac{[H_3O^+][A^-]}{[HA]}$$

$$K_a = \frac{(4.2\times10^{-2})(4.2\times10^{-2})}{(2.5\times10^{-1})}$$

$$K_a = \frac{(0.042)(0.042)}{(0.25)}$$

$$K_a = 7.06\times10^{-3}$$

Practice Exercise 5:

A dilute solution of acetic acid has $[H_3O^+] = 4.2\times10^{-4}M$ and $[CH_3COO^-] = 4.2\times10^{-4}M$. The concentration of unionized acetic acid is $2.1\times10^{-2}M$. What is the K_a value for acetic acid at 25°C?

Percentage Ionization of Weak Acids

The strength of an acid can be measured by its %age ionization. We can determine %age ionization of an acid by the formula:

$$\%\text{age ionization} = \frac{\text{Concentration of ionized acid at equilibrium}}{\text{Initial concentration of acid}}\times100$$

The greater the %age ionization of an acid, the stronger the acid and vice versa.

The %age ionization of weak acids depends upon the extent of dilution of solution. The extent of dilution increases with decrease in the number of moles of an acid.

$$\uparrow\text{Extent of dilution} \propto \frac{1}{\text{No. of moles}\downarrow}$$

When the solution is infinitely diluted, the weak acid gets 100% ionization.

The % ionization of 0.1M acetic acid solution is 1.3% and that of 0.005M acetic acid solution is 5.86%.

8.4.4 Levelling Effect

The strength of acids is not only depending upon the ability of a substance to donate a proton, it also depends upon the ability of solvent to accept protons. The acid strengths are determined by comparing their relative ionizations in water. The strongest acid that exists in water is the hydronium ion (H_3O^+). Any acid that is stronger than hydronium ion will ionize completely in water and produce H_3O^+ and so appear to be of equal strength and have nearly same pK_a values. Hence, it becomes difficult to tell which acid is stronger, that is, they are levelled out. This phenomenon is called leveling effect. Although strong acids have close pK_a values in water, but they are not equally strong. To compare the strengths of strong acids, a solvent that is weaker proton acceptor than water such as liquid HF or CH_3COOH has to be used. Strong acids show different pK_a values in a solvent that is weaker base than water. For example, the pK_a values of $HClO_4$, H_2SO_4 and HCl in the acetic acid medium are 5.3, 6.8, and 8.8 respectively. The relative strengths of some strong acids in acetic acid (known as glacial acetic acid) medium are:

$$HClO_4 > HBr > H_2SO_4 > HCl > HNO_3$$

Water has no leveling effect on the strength of weak acids such as HCOOH, CH_3COOH, H_2CO_3 and H_3PO_4 etc. because they are not ionized 100% in water and have different pK_a values.

Similarly, the strongest base that is present in water is OH^- ion. Any base that is stronger than OH^- hydrolyzes in water to produce OH^- ions. For example, the amide (NH_2^-) and oxide (O^{2-}) ions are stronger bases than OH^-, and hydrolyze completely in water producing OH^- ions. Water has no leveling effect on the strength of those bases which are weaker than OH^- ions.

Table 8.6: The Strong Bases and Acids that Ionize Completely in the Solution

Strong Bases	Name	Strong Acids	Name
LiOH	Lithium hydroxide	$HClO_4$	Perchloric acid
NaOH	Sodium hydroxide	$HClO_3$	Chloric acid
KOH	Potassium hydroxide	H_2SO_4	Suphuric acid
RbOH	Rubidium hydroxide	HI	Hydroiodic acid
CsOH	Cesium hydroxide	HBr	Hydrobromic acid
$Ca(OH)_2$	Calcium hydroxide	HCl	Hydrochloric acid
$Sr(OH)_2$	Strontium hydroxide	HNO_3	Nitric acid
$Ba(OH)_2$	Barium hydroxide		

Base Ionization Constant, K_b and pK_b

Equilibrium constant which shows the strength of a base is called dissociation constant of a base (K_b). Consider the ionization of a base in water:

$$B + H_2O \rightleftharpoons BH^+ + OH^-$$

The equilibrium constant is given as:

$$K_c = \frac{[BH^+][OH^-]}{[B][H_2O]}$$

The water is in excess, so its concentration is constant:

$$K_c.[H_2O] = \frac{[BH^+][OH^-]}{[B]} \qquad \text{or}$$

$$K_b = \frac{[BH^+][OH^-]}{[B]}$$

K_b is called ionization or dissociation constant of a base. Greater the value of K_b, stronger is the base and vice versa.

Strength of a base $\propto K_b$

The negative logarithm of K_b is called pK_b.

$pK_b = -\log K_b$

The greater the value of K_b, smaller will be the value of pK_b and stronger will be the base.

Table 8.7: K_b and pK_b Values for Some Acids at Room Temperature

Bases	Name	K_b	pK_b
NaOH	Sodium hydroxide	Very large	Very small
KOH	Potassium hydroxide	Very large	Very small
CH_3NH_2	Methyl amine	4.4×10^{-14}	3.36
NH_3	Ammonia	1.8×10^{-5}	4.74
N_2H_4	Hydrazine	1.7×10^{-6}	5.77
NH_2OH	Hydroxylamine	9.1×10^{-9}	8.04
C_5H_5N	Pyridine	1.4×10^{-9}	8.85
C_6H_5NH	Aniline	4.3×10^{-10}	9.37
$CO(NH_2$	Urea	1.3×10^{-14}	13.89

8.4.6 Relation between K_a and K_b

The strength of an acid can be measured by its acid ionization constant (K_a) and the strength of a base can be measured by its base ionization constant (K_b). An important relationship between the acid ionization constant (K_a) and the ionization constant of its conjugate base (K_b) can be derived as follows:

Consider the ionization of a weak acid, HA in water:

$$HA + H_2O \rightleftharpoons H_3O^+ + A^-$$

The equilibrium constant for the reaction is given as:

$$K_c = \frac{[H_3O^+][A^-]}{[HA][H_2O]} \quad \text{or} \quad K_a = \frac{[H_3O^+][A^-]}{[HA]} \quad \text{.............} \quad (i)$$

Consider the ionization of a conjugate base in water:

$$A^- + H_2O \rightleftharpoons HA + OH^-$$

The equilibrium constant is given as:

$$K_c = \frac{[HA][OH^-]}{[A^-][H_2O]} \quad \text{or} \quad K_b = \frac{[HA][OH^-]}{[A^-]} \quad \text{...............} \quad (ii)$$

By multiplying equation (i) and (ii), we get

$$K_a.K_b = \frac{[H_3O^+][A^-]}{[HA]} \times \frac{[HA][OH^-]}{[A^-]}$$

$$K_a.K_b = [H_3O^+][OH^-]$$

As we know, $K_w = [H^+][OH^-]$, hence,

$$K_a.K_b = K_w \quad \text{...............} \quad (iii)$$

The equation (iii) can be expressed as

$$K_a = \frac{K_w}{K_b} \quad \text{or} \quad K_b = \frac{K_w}{K_a}$$

From the above expressions it is clear that the larger the value of K_a, the smaller the value of K_b and vice versa.

If we take negative logarithm of both sides of the equation (iii), then pK values of the conjugate acid and base pairs are related to each other as:

$$pK_a + pK_b = pK_w = 14 \quad (\text{at } 25°C)$$

If we know the pK_a value of an acid we can calculate the pK_b value of its conjugate base as:

$$pK_b = 14 - pK_a$$

If we know the pK_b value of a base we can calculate the pK_a value of its conjugate acid as:

$$pK_a = 14 - pK_b$$

Table 8.8: Relationship of K_a, pK_a, K_b and pK_b Values

Acids	K_a	pK_a	Conjugate bases	K_b	pK_b
H_2SO_3	1.7×10^{-2}	1.7	SO_3^{2-}	5.9×19^{-13}	12.3
HF	6.8×10^{-4}	3.2	F^-	1.5×10^{-11}	10.8
HNO_2	4.5×10^{-4}	3.3	NO_2^-	2.2×10^{-11}	10.7

Acids	K_a	pK_a
CH_3COOH	1.8×10^{-4}	3.7
H_2CO_3	4.3×10^{-7}	6.3
HCN	4.9×10^{-10}	9.3
C_6H_5OH	1.3×10^{-10}	9.9
H_2O	1.0×10^{-14}	14

Conjugate bases	K_b	pK_b
CH_3COO^-	5.5×10^{-11}	10.3
HCO_3^-	2.3×10^{-8}	7.6
CN^-	2.0×10^{-5}	4.7
$C_6H_5O^-$	7.7×10^{-5}	4.1
OH^-	1.0×10^{0}	0

8.5 Lewis Definitions of Acids and Bases

The Bronsted and Lowry concept failed to explain those acids and bases which do not have hydrogen ions such as $AlCl_3$, and $FeCl_3$. It also failed to explain those acid-base reactions which occur in non-protonic solvents such as liquid SO_2, and liquid BF_3. In 1923, Gilbert N. Lewis, an American chemist focused on the role of electron pair and proposed an even more general concept of acids and bases. This concept is not restricted to any particular element or solvent. According to this concept, an acid is a substance (atom, molecule or ion) that can accept an electron pair to form a coordinate covalent bond and a base is substance that can donate an electron pair and to form a coordinate covalent bond. In short, acids are electron pair acceptors and bases are electron pair donors.

i) All simple cations such as H^+, Li^+, Ni^+, Sn^{2+} and Sn^{4+} act as Lewis acids. Smaller cations such as H^+, Li^+ and Be^{2+} act as stronger Lewis acids because of higher tendency to accept electrons.

ii) The electron deficient molecules which have incomplete octet such as $BeCl_2$, $AlCl_3$, and BF_3 act as Lewis acids.

iii) All anions such as Cl^-, O^{2-}, OH^-, X^- (Halide ion) and NH_2^- (amide ion) act as Lewis bases because they have tendency to donate electron pairs.

iv) The molecules which have lone pair of electrons such as $H_2\ddot{O}$ and $\ddot{N}H_3$ act as Lewis bases.

According to Lewis concept, an acid-base reaction occurs when a base donates an electron pair to an acid with the formation of a coordinate covalent bond between the two substances.

$$\underset{\text{Base}}{Cl^-} + \underset{\text{Acid}}{AlCl_3} \longrightarrow Cl \rightarrow \bar{A}lCl_3 \quad \text{or} \quad AlCl_4^-$$

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Gastric Acidity and use of Anti-acid Drugs

Gastric juice (stomach acid) is essential for the digestion of foods and to kill the bacteria that is ingested. The gastric juice has 0.5% HCl that is secreted by cells in the stomach walls and pH value of the gastric juice is normally 1-3. When too much food is eaten or when the stomach is irritated by very spicy foods, fatty foods, fried foods, acidic foods (tomatoes, oranges), acidic juices (grape fruits, oranges), and alcoholic beverages, or when a person becomes emotional, can cause the stomach to produce too much HCl. Some of acid contents can flow back into the esophagus (the tube that carries food to your stomach) and cause "heart burn". The condition is medically known as "pyrosis" or "acid indigestion".

The symptoms of heart burn include a burning sensation in the chest (just above the stomach) or behind your breast bone (sternum) and some patients experience sour taste in the mouth from gastric juice. Some patients may experience the burning sensation in the throat. Heart burn gets worse when bending over or lying down.

The heart burn is reduced:

- By taking an antacid (anti-acid) to neutralize the stomach acid. The antacid has one or more basic substances which are used to neutralize the excess but not all of the stomach acid.
- By taking acid-inhibitors to reduce the secretion of stomach acid. The acid-inhibitors include famotidine, ranitidine and cimetidine.

Bases (neutralizing agents) used as antacids
$NaHCO_3$ (Baking Soda)
$Al(OH)_3$ and $NaHCO_3$
$Mg(OH)_2$ (Milk of Magnesia)
$Al(OH)_3$ and $Mg(OH)_2$

- By taking proton pump inhibitors to block the production of acid. The proton pump inhibitors include omeprazole, rabeprazole and esomeprazole.

The heart burn can be controlled by some lifestyle change such as:

- To avoid the spicy foods, fatty foods, alcohol, aspirin, ibuprofen, caffeine, beverages, overeating.
- To avoid eating before bedtime, stop smoking, elevate the head of the bed and remain cool in the crisis.

8.6 Buffer Solution

A solution which resists any change in pH when a small amount of an acid or a base is added to it is called buffer solution.

If a solution comes in contact with air, it will absorb CO_2 and become more acidic. If a solution is stored in a glass bottle for long time, the silicates of glass make it more basic. In both cases the pH of solution changes. Buffer solution is used to maintain the pH of solution and prevent these changes.

Composition of Buffer Solution

A buffer solution usually contains:

i) A weak acid and its salt with strong base.

For example, CH_3COOH and CH_3COONa

ii) A weak base and its salt with strong acid.

For example, NH_4OH and NH_4Cl

Action of Buffer

Consider the buffer solution of acetic acid and sodium acetate:

i) When we add few drops of a strong base (NaOH) to the solution, the OH^- ion of NaOH reacts with acetic acid to give its conjugate base.

$$CH_3COOH + OH^- \rightleftharpoons CH_3COO^- + H_2O$$

ii) When we add few drops of strong acid (HCl) to the solution, the H^+ ion of HCl reacts with acetate ion of acetic acid to give back CH_3COOH.

$$CH_3COO^- + H^+ \rightleftharpoons CH_3COOH$$

Both OH^- and H^+ ions are consumed completely in the above reactions. Hence, pH of the solution remains constant.

Applications of Buffer Solutions

i) They are used on industrial scale in tanning of leather, electroplating and manufacture of sugar where constant pH is essential.

ii) Normal human blood has a pH of about 7.4. It is maintained by use of bicarbonates and phosphates. If it decreases up to 7 or goes up to 8, death may occur.

iii) They are used to study the growth of bacteria at constant pH.

iv) For proper production of crops, the pH control of the soil is essential.

v) They are used in beverage industries.

Buffer Capacity

The number of moles of an acid or a base required by one dm^3 of a buffer solution for changing its pH by one unit is called buffer capacity of a solution. It may also be defined as: the capability of a buffer to resist the change of pH is called buffer capacity. The buffer capacity depends upon the concentration (amount) of acids and salts present in the buffer solution.

Consider a buffer solution having 1.78M CH_3COONa and 1.0M acetic acid. The pH of this solution is 5. A solution having 0.178M CH_3COONa and 0.10M CH_3COOH shall also have pH equal to 5. Since anion/acid ratio is same. But the dilute solution has lower buffer capacity and would more easily be destroyed by addition of acid or base. However, the solution with 1M CH_3COOH and 1.78M CH_3COONa has larger buffer capacity and is able to show small pH changes when strong acids or bases are added and is not easily destroyed by such additions.

8.7 Salt Hydrolysis

The neutralization reaction of an acid and a base produces salts and water. The reaction of cations or anions or both cations and anions of salt with water to form acidic or basic or neutral solution is called hydrolysis. Salt hydrolysis usually affects the pH of a solution. For example:

When sodium acetate (CH_3COONa), a salt of weak acid and a strong base, is dissolved in water, its anion (CH_3COO^-) reacts with water as:

$$CH_3COO^-_{(aq)} + H_2O_{(l)} \rightleftharpoons CH_3COOH_{(aq)} + OH^-_{(aq)}$$

Similarly, NH_4Cl, a salt of weak base and a strong acid when dissolved in water, its cation (NH_4^+) reacts with water as:

$$NH^+_{4(aq)} + H_2O_{(l)} \rightleftharpoons NH_4OH_{(aq)} + H^+_{(aq)}$$

There are *four types of salts* on the basis of reactivity with water.

(i) The salts of weak acids and strong bases react with water and give basic solutions. For example, consider the ionization and hydrolysis reactions of CH_3COONa in water.

$$CH_3COONa_{(s)} \xrightleftharpoons{H_2O} CH_3COO^-_{(aq)} + Na^+_{(aq)}$$

$$CH_3COO^-_{(aq)} + H_2O_{(l)} \rightleftharpoons CH_3COOH_{(aq)} + OH^-_{(aq)}$$

The sodium acetate solution will be basic due to the formation of OH^- ions. The pH of this solution is more than 7 (pH > 7). The cations of strong bases such as alkali metal and alkaline earth metal ions except Be, have hardly any acidic character; that is to say, these ions do not hydrolyze and has no effect on the pH of the solution.

$$Na^+_{(aq)} + H_2O_{(l)} \rightleftharpoons \text{No reaction}$$

The extent of hydrolysis is determined by the value of K_b for acetate ion. The greater the value of K_b of a base, the more basic the solution it produces.

(ii) The salts of strong acids and weak bases react with water to give acidic solution. For example, consider the ionization of NH_4Cl in water.

$$NH_4Cl_{(s)} \xrightleftharpoons{H_2O} NH^+_{4(aq)} + Cl^-_{(aq)}$$

$$NH^+_{4(aq)} + H_2O_{(l)} \rightleftharpoons NH_4OH_{(aq)} + H^+_{(aq)}$$

The ammonium ion solution is acidic because of the formation of hydrogen ions. The pH of this solution is less than 7 (pH < 7). The anions of strong acids such as Cl^- ions have hardly any basic character; that is to say, these ions do not hydrolyze and has no effect on the pH of the solution.

$$Cl^-_{(aq)} + H_2O_{(l)} \rightleftharpoons \text{No reaction}$$

The extent of hydrolysis is determined by the value of K_a for ammonium ion. The greater the value of K_a of the acid, the more acidic the solution it produces.

(iii) The salts of strong acids and strong bases give neutral solutions. Because these salts are not hydrolyzed in water. For example consider the ionization of NaCl in water.

$$NaCl_{(s)} + H_2O_{(l)} \rightleftharpoons Na^+_{(aq)} + Cl^-_{(aq)}$$

The sodium ion is a very weak Bronsted base and has no effect on the pH of the solution. On the other hand, the chloride ion is a very weak Bronsted acid and it also has no effect on the pH of the solution. The pH of this solution is equal to 7 (pH=7).

(iv) The salts of weak acids and weak bases may not give neutral solutions, but whether the resulting solution is basic, acidic or neutral depends upon the relative strength of the weak acid and weak base.

To calculate the pH of such solution, you are required to compare the K_a value of the cation with the K_b value of the anion. There are three possibilities:

(a) If the value of K_a of the cation is greater than the value of K_b of the anion ($K_a > K_b$), the solution is acidic. This solution has pH value less than 7 (pH<7).

(b) If the value of K_a of the cation is less than the value of K_b of the anion ($K_a < K_b$), the solution is basic. This solution has pH value greater than 7 (pH> 7).

(c) If the value of K_a of the cation is close to the value of K_b of the anion ($K_a \approx K_b$), the solution is neutral or nearly neutral. This solution has pH value equal to or nearly equal to 7 (pH≈ 7).

To determine whether CH_3COONH_4 solution is acidic, basic, or neutral, you are required to compare the values of K_a of the ammonium ions and K_b of the acetate ions. Consider the ionization and hydrolysis reactions of CH_3COONH_4 in water.

$$CH_3COONH_4 \xrightleftharpoons{H_2O} CH_3COO^-_{(aq)} + NH^+_{4(aq)}$$

$$CH_3COO^-_{(aq)} + H_2O_{(l)} \rightleftharpoons CH_3COOH_{(aq)} + OH^-_{(aq)} \qquad K_a = 5.6 \times 10^{-10}$$

$$NH^+_{4(aq)} + H_2O_{(l)} \rightleftharpoons NH_4OH_{(aq)} + H^+_{(aq)} \qquad K_b = 5.6 \times 10^{-10}$$

This solution is neutral because the value of K_a is equal to the value of K_b. It means that the number of hydrogen ions produced by ammonium ions is equal to the number of hydroxide ions produced by acetate ions.

Table: 8.9: Types of Salts and the pH of Solutions

Type of Salt	Ions that are Hydrolyzed	pH of the Solution	Examples
Cation from strong base and anion from weak acid	Anion	>7	CH_3COONa, Na_2CO_3
Cation from weak base and anion from strong acid	Cation	<7	NH_4Cl, NH_4NO_3
Cation from strong base and anion from strong acid	None	≈ 7	$NaCl$, KI, KNO_3, Na_2SO_4
Cation from weak base and anion from weak acid	Anion and cation	Depends upon the relative strength of the weak acid and weak base.	NH_4F, CH_3COONH_4, NH_4CN

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Why Essential Elements like Iodine are added to Table Salt?

Iodine is a trace mineral that is naturally present in small amounts in marine organisms, eggs, and dairy products such as yogurt, milk, and cheese. It is essential for the formation of iodine containing hormones (thyroxine and triiodothyronine) secreted by the thyroid gland. Hormones are molecules that are carried in the bloodstream from one part of the body to another part and are responsible for growth, development, and maintenance of all body tissues. Iodine containing hormones regulate the heart rate and blood pressure and can also help the body to burn extra fat deposits. Daily requirement of iodine for adults is placed at 150 micrograms per day or 5 gram (1 teaspoonful) of iodized salt per day. If you do not get enough iodine, your thyroid enlarges and works overtime to produces more hormones, and as a consequence, the person suffers from low energy, low blood pressure, and weight gain. Extreme cases can lead to *goiter,* in which the gland is grossly enlarged. On the other hand, excess iodine in the diet can lead to an overactive metabolism. The addition of small amount of iodine to table salt in the form of NaI or KI at very little cost can prevent the deficiency of iodine. Good quality iodized salt provides the right quantity of iodine, therefore regular use of iodized salt cannot lead to excess iodine in the body. Daily consumption of iodized salt can protect entire generations of both humans and animals from mental and physical disabilities caused by Iodine deficiency.

Summary of Facts and Concepts

- Acids have sour taste, turn blue litmus paper red, react with active metals, conduct electricity in aqueous solutions and react with bases to form salts and water.
- Bases have bitter taste, soapy touch, turn red litmus paper blue, conduct electricity in aqueous solutions and react with acids to form salts and water.
- The sour taste of many foods such as vinegar, tomatoes, yogurt, lemons, and oranges is because of the presence of acids in them. The bitter taste of beer and coffee is because of the presence of bases in them. The bases are less common in foods. Both the acids and bases are used on large scale in industries and laboratories.
- According to Arrhenius concept, an acid is a substance that gives H^+ ions in water and a base is substance that gives OH^- ions in water.
- The strength of Arrhenius acid or base depends on the extent of ionization of acid or base in water.
- According to Bronsted-Lowry concept, an acid is a proton donor substance while a base is proton acceptor substance.
- According to Lewis concept, acid is electron pair acceptor substance while a base is electron pair donor substance.
- A strong acid or base is almost completely ionized in the solution while a weak acid or base is ionized to a slight extent in the solution.
- An acid increases the concentration of H^+ ions when dissolved in water while a base increases the concentration of OH^- ions when dissolved in water.
- The pH scale is used to measure acidity and pOH scale is used to measure basicity of solution. One-fold change in pH is equal to ten-fold change in the concentration of hydrogen ion.
- The K_a is the equilibrium constant that describes the ionization of an acid and the K_b is the equilibrium constant that describes the ionization of a base. Acid strength is denoted by the value of K_a and base strength is denoted by the value of K_b. Greater the value of K_a, stronger is the acid and vice versa. Greater

value of K_b, stronger is the base and vice versa.

- Greater the value of K_a, smaller the value of pK_a, and stronger will be the acid.
- Greater the value of K_b, smaller the value of pK_b, and stronger will be the base.
- A solution that resists pH change when small amounts of acid or base is added to it is called buffer solution. Buffer solutions contain either a mixture of weak acid and a salt of its conjugate base (such as CH_3COOH and CH_3COONa) or a mixture of weak base and a salt of its conjugate acid (such as NH_4OH and NH_4Cl).
- The number of moles of an acid or a base required by one dm^3 of a buffer solution for changing its pH by one unit is called buffer capacity of a solution.
- Salts that have cations of strong bases and anions of strong acids give neutral solutions. Salts that have cations of strong bases and anions of weak acids give basic solutions. Salts that have cations of weak bases and anions of strong acids give acidic solutions.

Questions and Problems

Q.1. Four answers are given for each question. Select the correct one:

i) According to ______ concept, acid is an electron pair acceptor substance.

(a) Arrhenius (b) Bronsted

(c) Lowry (d) Lewis

ii) Ionic product of water is represented by:

(a) K_w (b) K_c

(c) pK_a (d) pK_b

iii) The value of ionic product of water at 25°C is:

(a) $1.0\times10^{-14} mol^2dm^{-6}$ (b) $1.0\times10^{14} mol^2dm^{-6}$

(c) $1.5\times10^{14} mol^2dm^{-6}$ (d) $1.5\times10^{-14} mol^2dm^{-6}$

iv) When an acid is added to water, then

(a) $[H^+]>[OH^-]$ (b) $[H^+]<[OH^-]$

(c) $[H^+]=[OH^-]$ (d) $[H^+]\approx[OH^-]$

v) The pH of the aqueous solution of 10^{-2}M HCl is:

(a) 1.0 (b) 1.5

(c) 2.0 (d) 2.5

vi) The pH value of tears is 7.4, and pOH value is

(a) 4.6 (b) 6.6

(c) 7.0 (d) 7.6

vii) Which one of the following is **NOT** the Arrhenius acid:

(a) HCl (b) HNO_3

(c) $AlCl_3$ (d) H_2SO_4

viii) When the difference of pK_a values of two acids is equal to 2 then one acid is stronger than the other.

(a) 10 times (b) 100 times

(c) 1000 times (d) 10000 times

ix) Which one of the following statements is **NOT** correct for bases:

(a) Have bitter taste (b) turn blue litmus red

(c) have high pH values (d) reacts with acids to form salts

x) Which salt in water has pH value less than 7:

(a) NaCl (b) NH_4Cl

(c) CH_3COONa (d) CH_3COONH_4

Q.2. Fill in the blanks with suitable words given in the brackets:

i) A substance is termed as amphoteric when it acts ______. (as proton acceptor/as proton donor/both as proton acceptor and proton donor)

ii) Pure water is ______ conductor of electricity. (poor/good)

iii) The basic solution has pH ______ than seven. (less/greater)

iv) The sum of pH and pOH is ______ 14. (equal to/less than)

v) The ionic product of water increases by ______ temperature. (decrease/increase)

vi) The stronger acids have ______ value of pK_a. (smaller/greater).

vii) The value of pH is _______ proportional to $[H^+]$. (directly/ inversely)

viii) Weak acids are _______ electrolytes. (weak/strong)

ix) The electron deficient molecules act as Lewis _______. (acids/bases)

x) Aqueous solution of Na_2CO_3 is _______ while that of NH_4NO_3 is _______. (acidic/basic)

Q.3. Label the following statements as True or False:

i) The pH of the normal human blood is 7.4.

ii) The conjugate base of HCl is Cl^- *ion*.

iii) A strong base has a pH value of 14.

iv) The strength of an acid depends on the conjugate base: the stronger the conjugate base, the weaker the acid.

v) Orange is more acidic than apple.

vi) The antacids are used to reduce the secretion of stomach acid.

vii) Milk of magnesia has sour taste.

viii) For the neutralization of the stomach acid most of the people take NaOH.

ix) The larger the value of K_a, the smaller the value of K_b and vice versa.

x) The acids which have pH less than 4 are strong acids.

Q.4: What are the properties of acids and bases? Give some examples of acid containing foods and base containing substances.

Q.5: Define Arrhenius acids and bases and make it clear with examples.

Q.6: What is amphoteric substance? Give examples of substances that have amphoteric characteristics.

Q.7: Define Bronsted-Lowry acids and bases. Give a demonstration of each by chemical equations.

Q.8: What is a conjugate acid-base pair? Give relationship between the strength of an acid and its conjugate base.

Q.9: Write the conjugate base for each of the following acids:

(a) HNO_2 (b) H_2SO_3 (c) HSO_4^-

(d) H_2CO_3 (e) HF (f) HCO_3^-

Q.10: Write the conjugate acids for each of the following bases:

(a) NH_3 (b) Br^- (c) $CH_3CH_2NH_2$

(d) NO_3^- (e) NH_2^- (f) H_2O

Q.11: What is the difference between Arrhenius and Bronsted-Lowry definition of acids and bases?

Q.12: What do you know about Lewis acid-base concept? Explain it.

Q.13: What is the difference between strong and weak acids?

Q.14: What do you know about the strongest acid and strongest base which can exist in water?

Q.15: How do you measure the strength of an acid or base?

Q.16: What is the ionic product constant of water? Why the ionic product of water increases by increasing temperature? Write an expression for K_w what is its value at room temperature?

Q.17: Calculate the H^+ ions and OH^- ions concentrations of each of the following solutions:

(a) 0.1M HNO_3 solution ($K_a = 25$)

(b) 0.05 M NaOH solution. ($K_b = 10^{20}$)

Q.18: Calculate the concentration of H^+ ions in ammonia solution used for floor and glass cleaning whose hydroxide ion concentration is 0.025M.

Q.19: What do you mean by the pH and pOH? How are they related? Write an equation relating pH and pOH with pK_w.

Q.20: What is pOH of the plum having hydrogen ion concentration 1.58×10^{-3}M Is the solution acidic or basic?

Q.21: What is the pH of 2.0×10^{-4}M NaOH solution?

Q.22: What is H^+ and OH^- concentrations of a sample of tomatoes that has a pH value of 4.50?

Q.23: Calculate the pH of each of the following solution and classify each solution as acidic or basic:

(a) $[H^+] = 1.0\times10^{-10}$M (b) $[H^+] = 3.2\times10^{-8}$M

(c) $[H^+] = 4.3\times10^{-5}M$ (d) $[H^+] = 6.7\times10^{-3}M$

(e) $[H^+] = 9.5\times10^{-1}M$

Q.24: Calculate the pOH of each of the following solution and classify each solution as acidic, basic or neutral:

(a) $[OH^-] = 1.0\times10^{0}M$ (b) $[OH^-] = 2.4\times10^{-3}M$

(c) $[OH^-] = 1.0\times10^{-7}M$ (d) $[OH^-] = 7.8\times10^{-11}M$

(e) $[OH^-] = 1.0\times10^{-14}M$

Q.25: Calculate the pH of each of the following solution and classify each solution as acidic, basic or neutral:

(a) pOH = 0.00 (b) pOH = 1.62

(c) pOH = 3.33 (d) pOH = 7.00

(e) pOH = 13.92

Q.26: Calculate the pOH of each of the following solution and classify each solution as acidic, basic or neutral:

(a) pH = 14.02 (b) pH = 12.14

(c) pH = 7.61 (d) pH = 7.00

(e) pH = 11.44

Q.27: What is meant by K_a and K_b? What is their relation to the strength of acids and bases?

Q.28: Calculate the pH of 0.01M acetic acid whose K_a value is 1.7×10^{-5}.

Q.29: Formic acid has 2.45×10^{-6}M hydrogen ion concentration and 2.45×10^{-6}M formate ion concentration. The concentration of unionized formic acid is 0.025M at equilibrium. Find out the value of K_a for acid at room temperature.

Q.30: What are buffer solutions? Why buffers resist changes in pH? What are the applications of buffer solution?

Q.31: What is salt hydrolysis? Classify the salts on the basis of reactivity with water.

Q.32: Define and explain leveling effect.

Q.33: Answer the following questions:

(a) Why water is considered as a base in the Bronsted-Lowry concept?

(b) Why the value of K_b of a conjugate base decreases by increasing the value of K_a of a substance?

(c) Why a strong acid has a weak conjugate base and a weak acid is relatively a strong conjugate base?

(d) The aqueous solution of NH_4Cl is acidic and that of Na_2CO_3 is basic, why?

(e) $AlCl_3$ is said to be Lewis acid but not Bronsted-Lowry acid, why?

(f) Why the sum of pH and pOH of the aqueous solution at 25°C is equal to 14?

(g) Why the HCl has different pH values in water and acetic acid?

(h) Which has the larger value of the pK_a, a strong acid or the weak acid?

(i) Why chemists prefer to show the acid strength in terms of pH rather than $[H^+]$?