EXERCISE

MULTIPLE CHOICE QUESTIONS

1. Encircle the correct answer in each case:

i. Water cannot act as:

(a) Lewis acid (b) Lewis base

(c) Bronsted acid (d) Bronsted base

ii. pH of 0.01M HCl solution is:

(a) 10^{2} (b) 10^{-2}

(c) 2.0 (d) 1.0

iii. An aqueous solution of ammonium chloride is:

(a) Basic (b) Acidic

(c) Neutral (d) Amphoteric

iv. An aqueous solution of which compound is basic:

(a) Ammonium nitrate (b) Calcium chloride

(c) Ammonium acetate (d) Potassium carbonate

v. Which statement about acids is not correct? An acid:

(a) Contains hydrogen ions in solution.

(b) Contains oxygen.

(c) Has a pH of less than 7.

(d) Gives off carbon dioxide from a carbonate

vi. If a liquid has a pH of 7.

(a) It must be colourless (b) It has boiling point of 100°C.

(c) It must be a solution (d) It must be neutral.

vii. When air is bubbled through pure water, the pH is lowered from 7.0 to 5.6, which gas in the air is responsible for this change.

(a) Argon (b) Carbon dioxide

(c) Nitrogen (d) Oxygen

viii. Which of the following oxides is classified incorrectly!

(a) Zinc oxide (ZnO)amphoteric (b) Carbon dioxide (CO2) acidic

(c) Carbon monoxide neutral (d) Lead oxide (PbO) basic

ix. If 25 cm^3 of 1 mol. dm^{-3} nitric acid is added to 50 cm^3 of 0.5 potassium hydroxide solution, what would be the pH of the resulting solution

(a) 5 (b) 7

(c) 9 (d) 14

x. If dry citric acid crystals are placed on dry litmus paper they will

(a) turn yellow (b) turn green

(c) turn red (d) remains unchanged

xi. A base is a substance which will neutralize an acid which of these substances is not a base:

(a) Aqueous ammonia (b) Copper oxide

(c) Potassium chloride (d) Sodium carbonate

xii. A strong acid:

(a) Is always partially ionized when in solution

(b) is always fully ionized when in solution

(c) Always decomposes carbonates.

(d) Always contains oxygen

xiii. Which one of the following oxides dissolves in water to form acidic solution:

(a) MgO (b) NaO

(c) SO (d) SiO

xiv. When crystals of copper sulphate are heated, the colour changes from blue to white. This is caused by:

(a) Loss of water only

(b) Loss of water and SO

(c) Reaction with co, in the air.

(d) Loss of water, sulphur dioxide and oxygen

xv. The oxide of a metal was found to react with HCl along aqueous NaOH solution. Which of the following is the best description of the oxide?

(a) Acidic (b) Amphoteric

(c) Basic (d) Neutral

xvi. Which one of the following statements must be true of an acid salt?

(a) It can only be formed from a weak acid

(b) It is the only salt formed from a dibasic acid.

(c) It is the salt of a non-metal

(d) It contains hydrogen that is replaceable by a metal

xvii. One mole of each of the following compounds was dissolved in water to make one dm' of solution. Which of the solution would have the lowest pH value?

(a) NH_3 (b) CH_3COOH

(c) NaCl (d) NaOH

(e) H_2SO_4

xviii. Which salt could be obtained as the insoluble product of a reaction between a dilute acid and an aqueous salt?

(a) BaSO4 (b) $CuSO_4$

(c) $MgSO_4$ (d) $AgNO_3$

(e) $ZnCl_2$

xix. Which one of the following oxides react with aqueous NaOH to give a salt.

(a) Calcium oxide (b) Copper oxide

(c) Iron oxide (d) Magnesium oxide

(e) zinc oxide

xx. Which one of the following salts cannot be prepared by a reaction between a dilute acid and a metal?

a) $CaCl_2$ (b) $CuCl_2$

(c) $FeCl_2$ (d) $MgSO_4$

(e) $ZnSO_4$

Answers

		b	d	b	d	b
d	b	d	c	b	c	a
b	d	. e	i. a	c	b	

Q.2 (i) What are acidic, basic and amphoteric substances? Give one example of each substance.

Ans: J Bronsted and T. Lowry, independently elaborate acid base concept by defining them in a different manner.

According to these definitions,

A Bronsted acid is a proton donor

A Bronsted base is a proton acceptor

Acidic substance:

A substance which can donate a proton is called acidic substance. According to the following equation hydrochloric acid is a Bronsted acid since it donates a proton to water

$$HCl + H_2O \rightarrow H_3O^+ + Cl^-$$

Acid base

Since water accepts a proton, it is a Bronsted base. After accepting a proton, water is converted to hydroxonium ion (H_3O^+). The hydroxonium ion is in fact a hydrated proton.

Basic substances:

Those substance which can accept a proton are called basic substances. Thus, in the following equation, water is a Bronsted base since it accepts a proton from sulphuric acid,

$H_2SO_4 + H_2O \rightarrow H_3O^+ + HSO_4$

Acid Base

Also, in the equation below, ammonia accepts a proton from water. Therefore, ammonia is a base and a water is an acid here

$NH_3 + H_2O \rightarrow NH_4^+ + OH^-$

Base Acid

Amphoteric substance:

There are some substances which are on the border line of acids and bases these substances are called amphoteric substances. They show properties of both acids and bases.

Example:

Aluminium oxide (Al_2O_3) is amphoteric and it reacts with both acidic and basic solutions:

(basic) Al_2O_3 (s)+6HCl(aq) $\rightarrow 2AlCl_{3(aq)} + 3H_2O$

(acidic) Al_2O_3 (s)+2NaOH(aq) $\rightarrow 2NaAlO_{3(aq)} + H_2O$

Also chromium trioxide is amphoteric since it reacts with both acids and bases

(basic) $Cr_2O_{3(s)} + 6HNO_{(aq)} \longrightarrow 2Cr(NO_3)_{3(aq)} + 3H_2O_{(l)}$

(acidic) $Cr_2O_{3(s)} + 6KOH_{(aq)} \longrightarrow 2K_3CrO_{3(aq)} + 3H_2O_{(l)}$

(ii) Elaborate with equations two compounds which are amphoteric in nature.

Ans. Amphoteric substance:

There are some substances which are on the border line of acids and bases. These substances are called amphoteric substances. They show properties of both acids and bases.

Example:

Aluminium oxide is amphoteric and it reacts with both acidic and basic solutions:

$$\text{(basic) } Al_2O_{3(s)} + 6HCl_{(aq)} \longrightarrow 2AlCl_{3(aq)} + 3H_2O_{(l)}$$
$$\text{(acidic) } Al_2O_{3(s)} + 2NaOH_{(aq)} \longrightarrow 2NaAlO_{2(aq)} + H_2O_{(l)}$$

Also, chromium trioxide is amphoteric since it reacts with both acids and bases

$$\text{(basic) } Cr_2O_{3(s)} + 6HNO_{3(aq)} \longrightarrow 2Cr(NO_3)_{3(aq)} + 3H_2O_{(l)}$$
$$\text{(acidic) } Cr_2O_{3(s)} + 6KOH_{(aq)} \longrightarrow 2K_3CrO_{3(aq)} + 3H_2O_{(l)}$$

(iii) What is Bronsted - Lowry acid-base theory? Give examples.

Ans: Bronsted Lowery Concepts for Acids and Bases:

J Bronsted and T. Lowry, independently elaborate acid base concept by defining them in a different manner.

According to these definitions,

A Bronsted acids a proton donor

A Bronsted base is a proton acceptor

Definition of proton:

In this respect, a proton is defined as the nucleus of a hydrogen atom (H) and it has nothing to do with the protons in a carbon atom, or a sodium atom or any other atom.

Bronsted acid:

Hydrochloric acid is a Bronsted acid since it donates a proton to water according to the following equation

$HCl + H_2O \rightarrow H_3O^+ + Cl^-$

Acid Base

Since water accepts a proton, it is a Bronsted base After accepting a proton, water is converted to hydroxonium ion (H_3O^+). The hydroxonium ion is in fact a hydrated proton

Example:

Some more common Bronsted acids are sulphuric acid (H_2SO_4), phosphoric acid (H_3PO_4) and acetic acid (CH_3COOH) etc.

Bronsted bases:

The Bronsted bases have been defined as those species which accept a proton Thus, in the following equation water is a Bronsted base since it accepts a proton from sulphuric acid

$H_2SO_4 + H_2O \rightarrow H_3O^+ + HSO_4^{-2}$

Acid Base

Also in the equation below, ammonia accepts a proton from water

Therefore, ammonia is a Bronsted base and water is a Bronsted acid here

$NH_3 + H_2O \rightarrow NH_4^+ + OH^-$

Base Acid

(iv) What are conjugate acid-base pairs? Explain with examples.

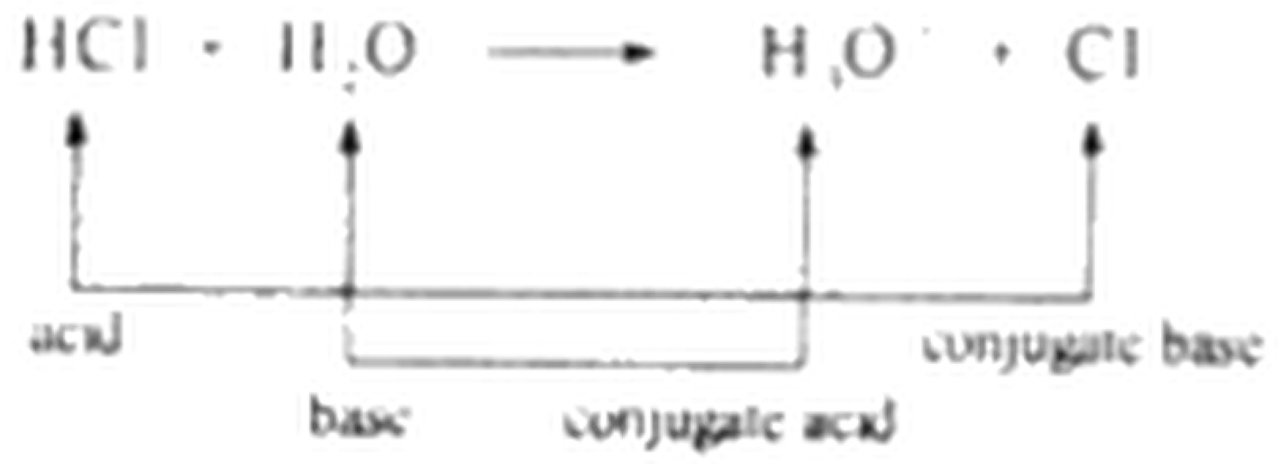
Ans: Conjugate acids:

Conjugate acid is a species which is formed as a result of acceptance of proton by a base Every Bronsted acid has a conjugate base

Conjugate bases:

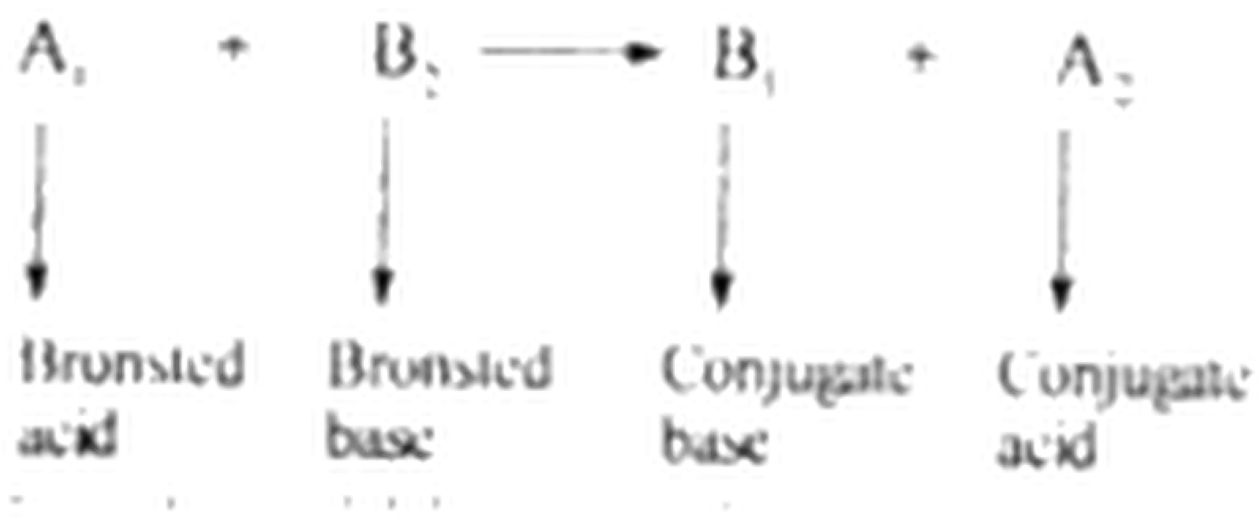
Conjugate base is a species which is left behind after donation of a proton from the acid

e.g. Ionization of HCl in water



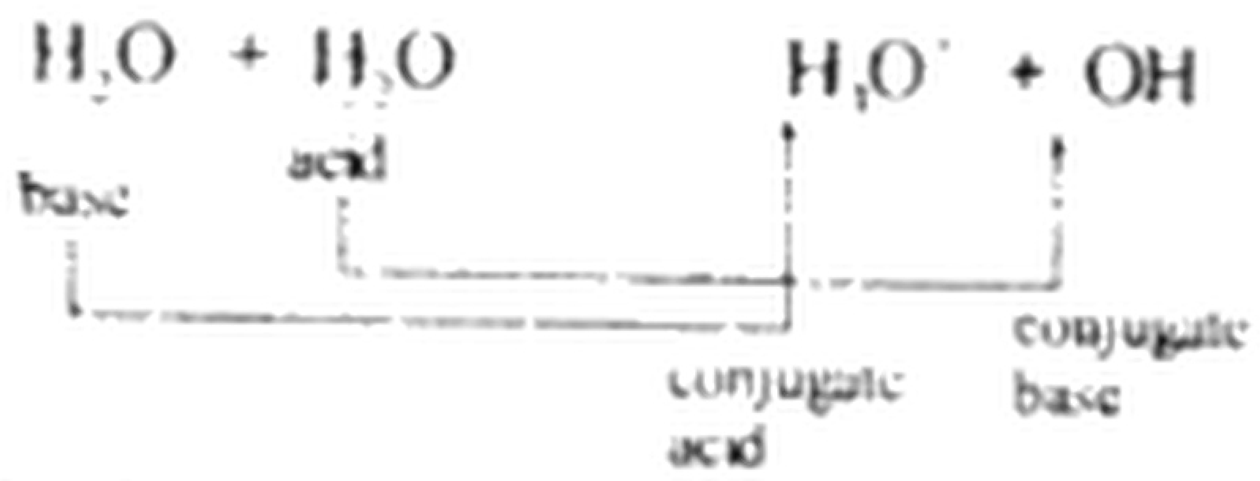
Conjugate Acid-Base Pairs (Major Concepts):

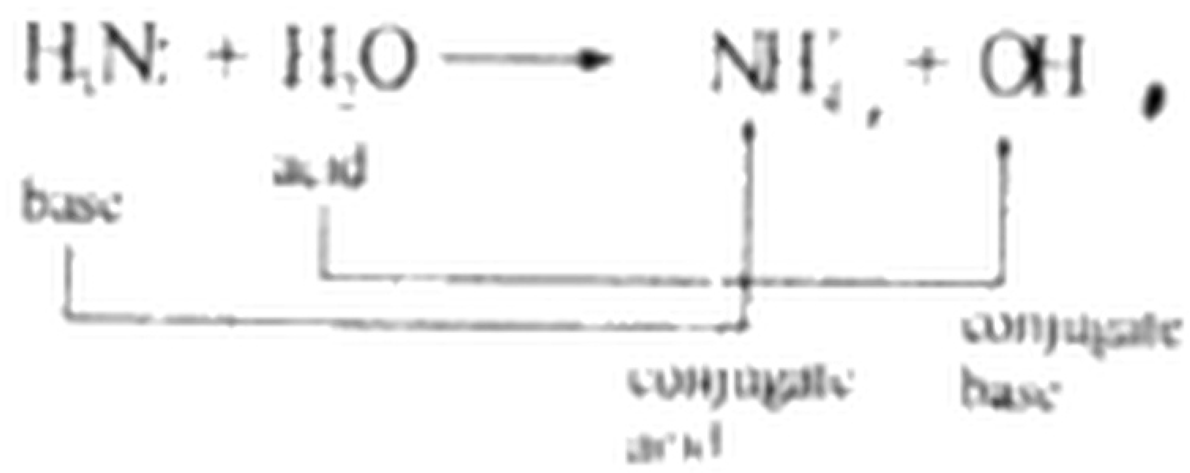
In an acid-base reaction, an acid yields a base (conjugate) and base after accepting proton yields a conjugate acid. The acid-base reaction is represented as



The conjugate acid-base pairs are species on opposite sides of an equation that differ by a proton

The weaker acids have strong conjugate base pairs and stronger acids have weaker conjugate bases e.g.

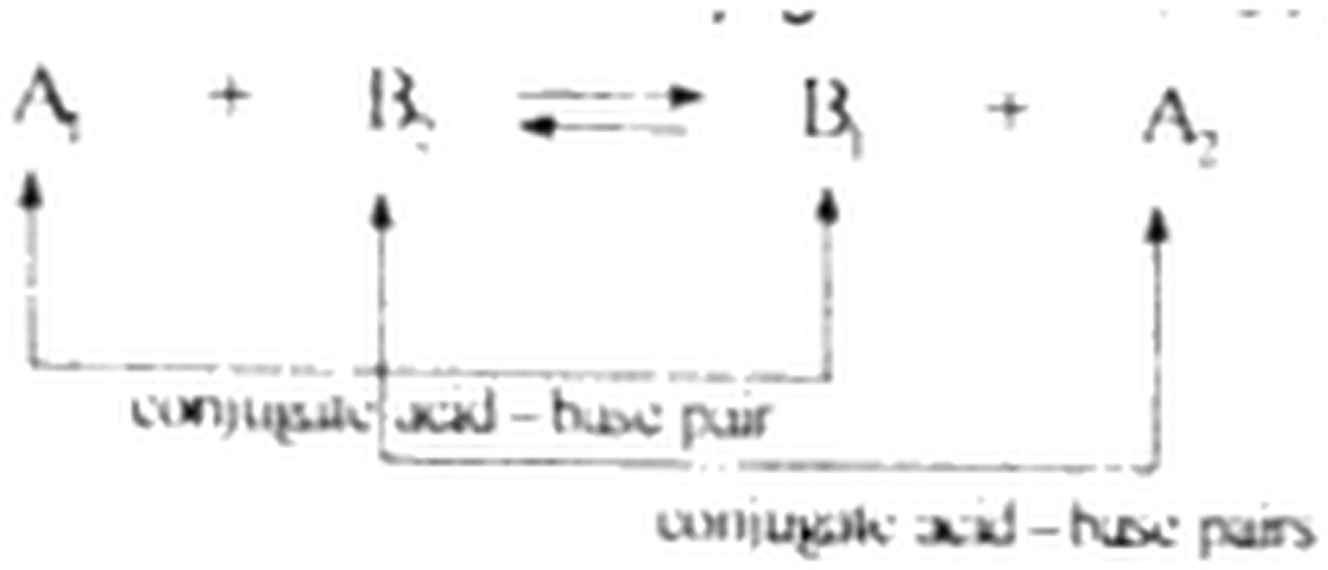




The water is amphoteric in nature i.e it is acidic as well as basic in nature

General notation:

The general notation of conjugate acid-base pair is



Each acid has a conjugate base and each base has a conjugate acid

Particular Examples of Conjugate Acid-Base Pairs:

A1 + B2 B1 + A2

$$HCl + H_2O \rightleftharpoons Cl^- + H_3O^+$$

$$HCl + NH_3 \rightleftharpoons Cl + NH_4^{-1}$$

$$HNO_3 + H_2O \rightleftharpoons NO_3^{-1} + H_3O^+$$

$$NH_4^{-1} + H_2O \rightleftharpoons NH_3 + H_3O^+$$

$$H_2O + H_2O \rightleftharpoons OH^- + H_3O^+$$

$$H_2O + CO_3^{-2} \rightleftharpoons OH^- + HCO_3^{-1}$$

$$HClO_4 + H_2O \rightleftharpoons ClO_4^{-1} + HCO_3^{-1}$$

A, B, and B, A, are known as conjugate acid-base pairs

(v) How preservatives are used in food products and allergic reactions

In people

Ans: Preservatives:

Preservatives are used to preserve food from bacteria They are chemicals that inhibit the growth of bacteria and fungi. It prevents decomposition of food from microbial growth or by chemical changes

Most commonly used preservatives are

i. Benzoates in bakery products

ii. Nitrates in meat products

iii. Sulphites in drinks.

iv. Potassium bromated.

v. Ammonium sulfate

vi. Benzoic acid and Parabens.

vii. Nitrates and Nitrites.

Allergic reactions in people:

The preservatives can cause severe allergic reactions in people. They increase the rate of cancer, childhood illness and auto-immune diseases sulphites can cause wheezing, chest tightness and coughing, Benzoates may cause urticaria, asthma etc

(vi) Define Lewis acid and Lewis bases. Give one example in each case.

Ans: Lewis Definitions of Acid and Base (Major Concept):

GN Lewis in 1932, put forward his acid-base theory on the basis of electron pair.

Base:

According to Lewis definition, a base is a substance that donates a pair of electrons

Acid:

According to Lewis definition, an acid is a substance which can accept a pair of electrons.

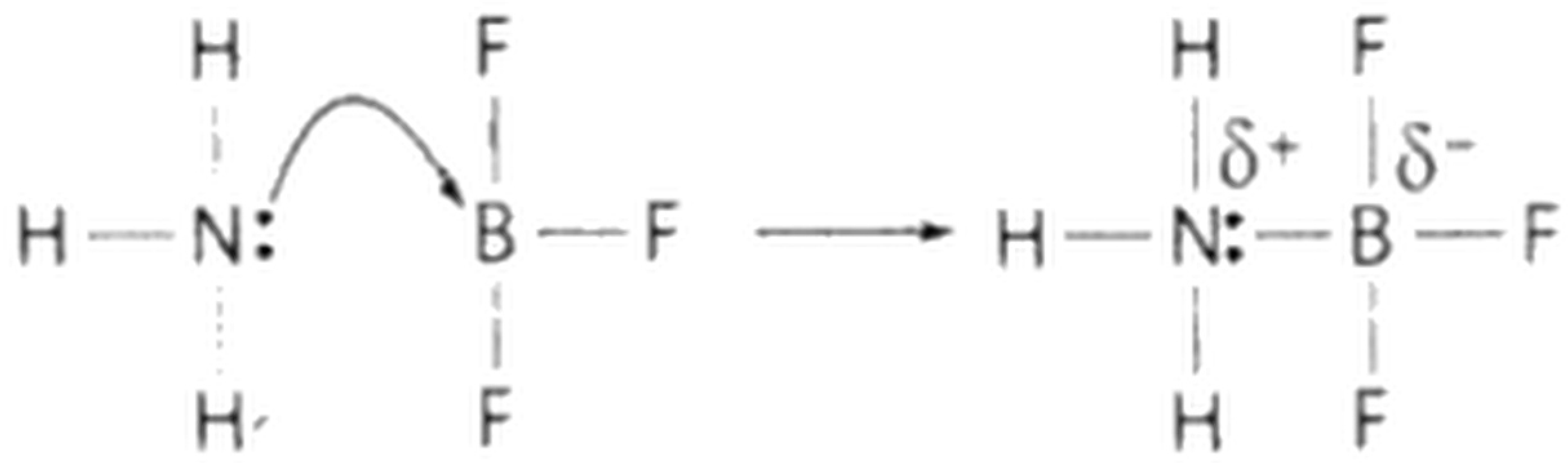
Example:

Tre hydroxide ion (OH) is a Lewis base because it donates a pair of electron and the proton (H) is a Lewis acid-since it accepts a pair of electrons.

$$H^+ + :\ddot{\underset{\cdot\cdot}{O}}^- — H \longrightarrow H — \ddot{\underset{\cdot\cdot}{O}} — H$$

Significance:

The significance of the Lewis acid-base concept is that it is much more general than other concepts. It may include such acid. base reactions which are not covered by the Bronsted-Lowery theory. One such example is the reaction between ammonia and boron trifluoride (BF3) illustrated by the following equation:



The vacant, unhybridized 2p orbital of boron atom in boron trifluoride accepts the electron pair from ammonia. Thus in this reaction ammonia is a Lewis-base and boron trifluoride is a Lewis acid although no proton transfer is observed here. Some common Lewis acids and bases are listed in the table.

vis Acids		**vis Bases**	
me	**mula**	**me**	**mula**
ton		droxide ion	
on trifluoride		nmonia	NH_3
minium chloride	$AlCl_3$	rbon monoxide	
er cation	*	ter	
rbon dioxide	C=O	anide ion	N

Conclusion:

According to this table Lewis acids have the ability to accept electron pair whereas the Lewis bases are capable of donating electron pair

(vii) Classify each of the following as Bronsted acid or Bronsted base,

(1) HCO_3^{-1} **(ii) HBr (iii)** CH_3COOH^{-}

Ans:

nsted acid	nsted base
A Bronsted acid is a proton donor	A Bronsted base is a proton acceptor
: a Bronsted acid as it donates a proton water $HBr + H_2O \rightarrow H_3O^{+} + Br^{-1}$	HCO_3^{-1} a Bronsted base because it accepts n water $HCO_3^{-1} + H_2O \rightarrow +H_2CO_3 + OH^{-}$ CH_3COO^{-} a Bronsted base as it accepts a ton from water $CH_3COO^{-} + H_2O \rightarrow CH_3COOH + OH^{-}$

(viii) Explain gastric acidity and use of anti-acid drug.

Ans: Human stomach produces HCl. It helps us to digest food. Production of excess HCl by stomach is called gastric acidity It is generally caused by eating spicy foods.

Anti-acid drug:

This problem can be deduced by taking arti-acidic drugs uke Milk of Magnesia (Mg(OH)2) or the medicines in which sodium bicarbonate ($NaHCO_3$) and Al(OH), are present These bases react with acid to neutralize it in this way acidity can be remove by taking anti-acidic drugs. These are generally present in a single medicinal preparation

(ix) Write briefly about the ionization of water.

Ans: Auto Ionization of Water:

Water is a unique compound due to its ability to accept or donate proton under different environments. It has been mentioned earlier that water acts as a Bronsted acid in presence of ammonia, in fact water itself undergoes ionization to a small extent.

$$H_2O + H_2O \rightleftharpoons OH^- + H_3O^+$$

This reaction is regarded as auto ionization of water.

Explanation:

Here a molecule of water donates a proton to another molecule of water and two species a hydronium ion and a hydroxide ion, are formed.

In this equation there are two acid-base conjugate pairs, the first one is water-hydroxide ion (acid-conjugate base) and the second water hydronium ion (base conjugate acid)

In fact there is an equilibrium between water molecules (on the left side of the equation) and the hydronium ions and hydroxide ions on the right side of the equation)

Equilibrium constant (K):

The equilibrium constant (K) for this equilibrium can be expressed by the following equation

$$K = \frac{[H_3O^+][OH^-]}{[H_2O]^2}$$

In this equation $[H_3O^+][OH^-]$ and $[H_2O]$ arc the concentration of hydronium ion, hydroxide ion and water at equilibrium stage

Large excess of water:

Since water is a solvent the concentration of water, $[H_2O]$ is in large excess, therefore it remains constant. The above equation may be re-written as.

K$[H_2O]^2$ = $[H_3O^+][OH^-]$

Ion-product constant of water:

The term KTH,' is the product of two constants, K and $[H_2O]$. Therefore this term is another constant represented by Ky in the following equation where

K$[H_2O]^2$ = K_w = $[H_3O^+][OH^-]$

K_w is termed as the ion-product constant of water. Since 1,0 is the concentration of hydrated protons at equilibrium, the above equation corresponds

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

Applications:

in pure water at 25°C, the concentration of H' and OH are equal and found to be 1.0 x 10^{-7} M each. Thus:

K_w = $[1.0 \times 10^{-7}]$-$[1.0 \times 10^{-7}]$= 1.0×10^{-14}

It has been noted that whether in pure water or in a solution of dissolved species, the following relationship always holds

K_w = [H] $[OH^-]$ = 1.0×10^{-14}

(x) Define pH what are the values of pH for acidic, basic and neutral solutions.

Ans: pH:

Since the concentration of [H] $[OH^-]$ are usually very small numbers and inconvenient to work with a more practical measure called pH was proposed and defined as

It means that pH of a solution is given by the negative logarithm of the [H] $\left[OH^{-}\right]$

concentration (in mol/dm).

pH= -log (H⁻]

No units:

However it must be kept in mind that pH being a logarithmic value, does not have any units,

Conditions:

The pH concept implies that at 25°C, the different types of solutions will show The following behaviour:

Acid solution: [H]>1.0 x 10^{-7} M, p < 7.00

Basic solution: [H] <1.0 x 10^{-7} m, pH> 7.00

Neutral solution [H] <1.0 x 10^{-7} m, pH=7.00

(xi) Define K_a and pK_a and their applications,

Ans: K_a

K_a is termed as the acid dissociation constant it is a measure of the extent to which acid is ionized or dissociated at the equilibrium state

$$K\left[H_2O\right]=K_a=\frac{\left[H_3O^{+}\right]\left[X^{-}\right]}{\left[HX\right]}$$

pK_a :

The negative Logarithm of K_a value is called pK_a The relationship between K_a and pK_a is as follows:

pK_a = -log K_a

Applications:

i. pK_a value is also used to determine the strength of acid

ii. Smaller the value pK_a value stronger will be the acid and vice versa. It is because smaller pK_a .value means greater K_a value.

iii. The K_a values are also used to determine the percentage dissociation of an acid

iv. The greater the value of K_a the stronger is the acid and vice versa.

$$PK_a \, \alpha \, \frac{1}{K_a}$$

(xii) Explain curdling of milk with lemon juice,

Ans:

i. Milk is actually made up of a lot of different components, the main ones being protein, fat, and water.

ii. When it comes to curding, we're mainly concerned with one specific milk protein called casern.

iii. Normally, little groupings of casein float around in the milk without bonding to anything

iv. These groupings (technically cared mice les) have a negative charge, which makes them repel other groupings of casein and keeps the casein evenly dispersed in the milk.

v. When milk becomes too acidic, like when we acid lemon juice or when goes scur, the negative charge on the casein groupings becomes neutralized.

vi. Now instead of pushing each other apart, the casein starts to clump together.

vii. Eventually large enough clumps are formed that we can actually see the separation, and then we have curdled milk

(xiii) What are K_b and pK_a and their applications.

Ans: K_b

Base dissociation constant is known as K_b. The extent of ionization and the base dissociation constant K_b is use to distinguish between strong and weak bases

Derivation of base dissociation constant K_b:

Consider a base B, an equilibrium reaction with water can be represented by the following equation

$$B + H_2O \rightleftharpoons BH^+(aq) + OH^-(aq)$$

Since water is a solvent, it is present in excess and therefore its. concentration may be regarded as constant. Thus, K $[H_2O]$ is another constant and is designated as K_b thus,

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

K_b value will be large if degree of ionization of the base B is high ie if the base B is strong The Value of K_b will be small for a weak base B.

pK_b:

pK_b is defined as the negative logarithm of K_b

pK_b = -log K_b

Applications:

i. PK_b also measures the strength of a base.

ii. Smaller the value pK_b value, stronger will be the base and vice versa. It is because smaller pK_b value means greater K_b value.

$$PK_b \; \alpha \frac{1}{K_b}$$

iii. K_b values are used to determine the strength of different bases.

iv. The greater the value of K_b the stronger is the base.

(xiv) What is an iodized salt? What is its use in practical life.

Ans: iodized Salt:

A mixture of sodium chloride (NaCl) that contain some amount of NaI (Sodium iodide) or KI (Potassium iodide) is called iodized salt.

Use in practical life:

i. Improved Thyroid Function:

Body needs iodine for the thyroid to produce a couple of essential hormones called thyroxine and triiodothyronine These hormones help to regulate the metabolism and the growth and development of the body Moreover the deficiency of iodized salt can cause a disease called goiter in this disease throat swells up due to enlargement of thyroid gland

ii. Improved Brain Function:

Iodized salt can improve the functions of the brain, such as memory concentration and the ability to learn. An iodine deficiency can lower the IQ by as much as 15 points

iii. Fights Depression:

Depress on and feelings of anxiety and frustration may be the result of an od ne deficiency, iodized salt can help to ensure that you are getting enough iodine to prevent these emotional feelings from occurring.

iv. Weight Control:

iodine is important for regulating the metabolism. When it is too high, you may not be able to gain a healthy weight, and if it is too low, you may gain weight or be unable to lose it Additionally, iodized salt can increase your energy levels so they you will exercise more

v. Improved Appearance:

It can help to cure dry and flaky skin, and improve the growth of your hair and nails. It also plays a role in developing and maintaining healthy teeth.

vi. Removes Toxins:

iodized salt can help to remove harmful metals such as lead and mercury as well as other harmful toxins from our body. This can help to restore a healthy pH level in your body

vii. Heart Health:

Iodized salt can help to create hormones which regulate the heart rate and blood pressure. It can also help the body to burn extra fat deposits that contribute to heart disease.

(xv) What is the relationship between K_a and K_b

Ans: As we know that

K_a x K_b. = Kw

and

$$K_a = \frac{K_w}{K_b}$$

And

$$K_b = \frac{K_w}{K_a}$$

Conclusion

Since Kw is constant at a given temperature, it may be deduced that K_a is inversely proportional to K_b

$$K_a \ \alpha \frac{1}{K_b}$$

Thus, stronger the acid, weaker is its conjugate base. It can also be said that stronger a base, weaker is its conjugate acid.

(xvi) Calculate concentrations of ions of slightly soluble salts.

Ans: Slightly soluble salts are partially ionized on dissolving in water.
product of the equilibrium concentrations of ions, each raised to a power which is the coefficient of the ion in the balanced equation is called solubility

product.

For example:

When AgCl is mixed with water. Following equilibrium is established

$$AgCl(s) \rightleftharpoons Ag^+ + Cl^-$$

Mathematically:

K_c for this equilibrium can be written as

$$K_c = \frac{[Ag^+][Cl^-]}{[AgCl]}$$

Since CaF_2 is slightly soluble salt ils concentration almost remains constant.

Therefore,

$$K_c = \begin{matrix}[AgCl] = [Ag^+][Cl^-] \\ Ksp = [Ag^+][Cl^-]\end{matrix}$$

Where Ksp is a constant known as the solubility product constant.

Therefore the value of concentrations of ions of slightly soluble salts can be determine by using the value of Ksp.

(xvii) Give two examples of a buffer solution.

Ans: A buffer solution contains a weak dissociating acid and the salt of that acid with a strong base.

Examples of a buffer solution:

i. H_2CO_3 + Na HCO

H_3PO_4 + Na H_2PO_4

$CH_3COOH + CH_3COONa$

ii. A solution containing a weak base e.g. NH_4 OH and its salt with a strong acid like NH_4 Cl can also act as a buffer.

(xviii) Perform acid-base titrations to calculate molarity and strength of a

given sample solution.

Ans: Consider a titration process in which 25 cm^3 of 0.12M NaOH is neutralized with 30 cm^3 of HCl of unknown concentration.

Calculate concentration and strength of HCl solution.

Solution: HCl. + NaOH. $\rightarrow$ NaCl + H_2O

HCl NaOH.

$$\frac{M_1V_1}{n_1} = \frac{M_2V_2}{n_1}$$

$$\frac{M_1 \times 30}{1} = \frac{0.12 \times 25}{1}$$

= 0.1M

Thus Molarity of HCl solution is 0.1M.

Strength of solution = Molarity x Molar mass

Strength of HCl solution = 0.1 x 36.5 = 3.65g dm^3

(xix) Write two acidic, two basic and two neutral salts.

Ans: Acidic Salts:

Salt formed by the reaction between strong acid and weak base is called "Acidic salt" these salts have pH is more than 7.

Examples: $CuSO_4$, NH_4Cl, NH_4NO_3 etc.

Basic Salts:

Salt formed by the reaction between strong base and weak acid is called "basic salt" these salts have pH less than 7.

Examples: CH_3COONa, NaCN, Na2S etc.

Neutral Salts:

A salt which is formed by the neutralization of strong acid and strong base is called neutral salts. They have pH = 7.

Examples: NaCl, Na_2SO_4, KNO_3 etc.

(xx) Use the concept of hydrolysis to explain why aqueous solutions of some salts are acidic or basic.

Ans: Hydrolysis:

The reaction of cations and anions of salts with water is called hydrolysis

Acidic Salts:

Those salts which produce strong acid and weak base on hydrolysis are called acidic salts. The solution becomes acidic due to the presence of strong acid and the hydrolysis is called acidic hydrolysis

Examples: $CuSO_4$, NH_4Cl, NH_4NO_3 etc.

$$CuSO_4 + H_2O \rightarrow Cu(OH)_2 + H_2SO_4$$

(Weak base) (Strong acid)

$$NH_4Cl + H_2O \rightarrow HCl + NH_4OH$$

(Strong acid) (Weak base)

$$NH_4NO_3 + H_2O \rightarrow HNO_3 + NH_4OH$$

(Strong acid) (Weak base)

Basic Salts:

Those salts which produce weak acid and strong base on hydrolysis are called basic salts. The solution becomes basic due to the presence of strong base and the hydrolysis is called basic hydrolysis

Examples:

$$CH_3COONa + H_2O \rightarrow CH_3COOH + NaOH$$

Weak acid) (Strong bane)

$$NaHCO_3 + H_2O \rightarrow NaOH + H_2CO_3$$

(Weak acid) (Strong bane)

Q3. Elaborate the ionization equation of water. How does it lead to the ion-product, constant of water? Describe an important application

of the ion-product of water,

Ans: Auto Ionization of Water:

Water is a unique compound due to its ability to accept or donate proton under different environments It has beer: mentioned earlier that water acts as a Bronsted acid in presence of ammonia, and as a Bronsted base in presence of hydrochloric acid. In fact. Water itself undergoes ionization to a small extent as shown in the following equation

$$H_2O + H_2O \rightleftharpoons OH^- + H_3O^+$$

This reaction is regarded as auto ionization of water

Explanation:

Here, a molecule of water (which is acting as an acid) donates a proton to another molecule of water (which is acting as a-base) and two species, a hydronium ion and a hydroxide ion, are formed.

Acid-base conjugate pairs:

In this equation there are two acid-base conjugate pairs the first one is water-hydroxide ion (acid-conjugate base) and the second water hydronium ion (base conjugate acid).

In fact, there is equilibrium between water molecules (on the left side of the equation) and the hydronium ions and hydroxide ions on the right side of the equation)

Equilibrium constant (K):

The equilibrium constant (K) for this equilibrium can be expressed by the following equation

$$K = \frac{[H_3O^+][OH^-]}{[H_2O]^2}$$

In this equation $[H_3O^+][OH^-]$ and $[H_2O]$ arc the concentration of hydronium ion. hydroxide ion and water at equilibrium stage

Large excess of water:

Since water is a solvent the concentration of water, $[H_2O]$ is in large excess, therefore it remains constant. The above equation may be re-written as.

$K[H_2O]^2 = [H_3O^+][OH^-]$

Ion-product constant of water:

The term KTH,' is the product of two constants, K and $[H_2O]$. Therefore this term is another constant represented by Ky in the following equation where

$K[H_2O]^2 = K_w = [H_3O^+][OH^-]$

K_w is termed as the ion-product constant of water. Since 1,0 is the concentration of hydrated protons at equilibrium, the above equation corresponds

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

in pure water at 25°C, the concentration of H' and OH are equal and found to be 1.0 x 10^{-7} M each. Thus:

K_w = [1.0 x 10^{-7}]-[1.0 x 10^{-7}]= 1.0 x10^{-14}

It has been noted that whether in pure water or in a solution of dissolved species, the following relationship always holds

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

K_w is temperature dependent:

It is to be noted that, because K is an equilibrium constant, it is temperature dependent thus, at 40 C K = 3.8 x 10^{-14} which corresponds to

$[H^+]$- 1.9 x 10^{-7}' M and $[OH^-]$ = 1.9 x 10^{-7}M as$[H^+] = [OH^-]$

Application:

i. pH scale is based on the ionic product of water.

il. It tells us about the behavior of water.

pH:

Since the concentration of $[H^+]$ and $[OH^-]$ are usually very small numbers and inconvenient to work with, a more practical measure called PH was proposed and defined as

It means that pH of a solution is given by the negative logarithm of the II concentration (in mol/dm).

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

No units: However it must be kept in mind that pH being a logarithmic value, does not have any units

Conditions:

The pH concept implies that at 25°C, the different types of solutions will show the following behaviour.

Acid solution: [H]>1.0 x 10^{-7} M, p < 7.00

Basic solution: [H] <1.0 x 10^{-7} m, pH> 7.00

Neutral solution [H] <1.0 x 10^{-7} m, pH=7.00

A scale analogous to the pH can be devised using the negative logarithm of the H^+ concentration.

Q.4 What are buffer solutions? Elaborate with suitable examples, their significance in acid-base reactions. Write three common applications of buffer solutions

Ans: Definition:

Those solutions which resist the change in their pH when a small amount of an acid or a base is added them, are called buffer solutions. They have a specific constant value of PH.

Constitution of buffer:

A buffer solution contains a weak dissociating acid and the salt of that acid with a strong base

i. H_2CO_3 + Na HCO

H_3PO_4 + Na H_2PO_4

$CH_3COOH + CH_3COONa$

ii. A solution containing a weak base e.g. NH_4 OH and its salt with a strong acid like NH_4 Cl can also act as a buffer.

Significance:

It resists the change in the pH of a solution even when a small amount of a strong acid and base is added in the solution

Buffer action:

Let us take a buffer solution of CH_3COOH and CH_3COONa Common ion effect helps us to understand how will buffer work. CH_3COOH being a weak electrolyte undergoes very little dissociation. When CH_3COONa a strong electrolyte is added to CH_3COOH solution, the dissociation of CH_3COOH is suppressed due to common effect of CH_3COO

$$CH_3COOH + H_2O \rightleftharpoons CH_3COO^- + H_3O^+$$
$$CH_3COONa \rightleftharpoons CH_3COO^- + Na^+$$

Addition of strong acid:

Suppose we add a few drops of HCl to it. Its H⁺ ions are used up by CH_3COO and the equilibrium is shifted towards loft.

Thus the addition of HCl will not change the pH of the buffer solution

Addition of strong base:

in the same buffer solution if a strong base is added it is neutralized by the acid

$$CH_3COOH \rightleftharpoons CH_3COO^- + H^+$$
$$NaOH \rightleftharpoons Na^+ + OH^-$$
$$H^+ + OH^- \rightarrow H_2O$$

Thus the addition of NaOH will not change value of pH.

Applications of buffer solutions:

(i) Buffer solutions play an important role in several industrial processes For example, they are used in the manufacture of photographic materials, leather and dyes

(ii) They are also used in the process of electroplating and analytical procedures.

(iii)The buffer solutions are also used for calibration of pH meters

Q.5 Write detailed notes on each of the followings:

(a) Conjugate acid base pairs (b) p K_a

(c) P K_b

(d) Relative strength of acids.

Ans: (a) Conjugate acid base pairs

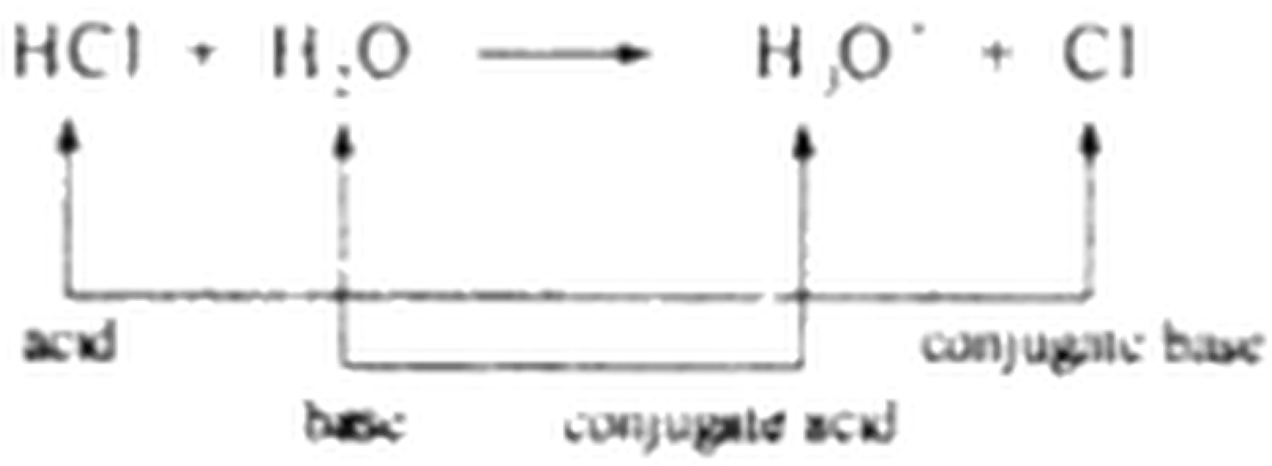
Conjugate acids:

Conjugate acid is a species which is formed as a result of acceptance of proton by a base Every Bronsted acid has a conjugate base

Conjugate bases:

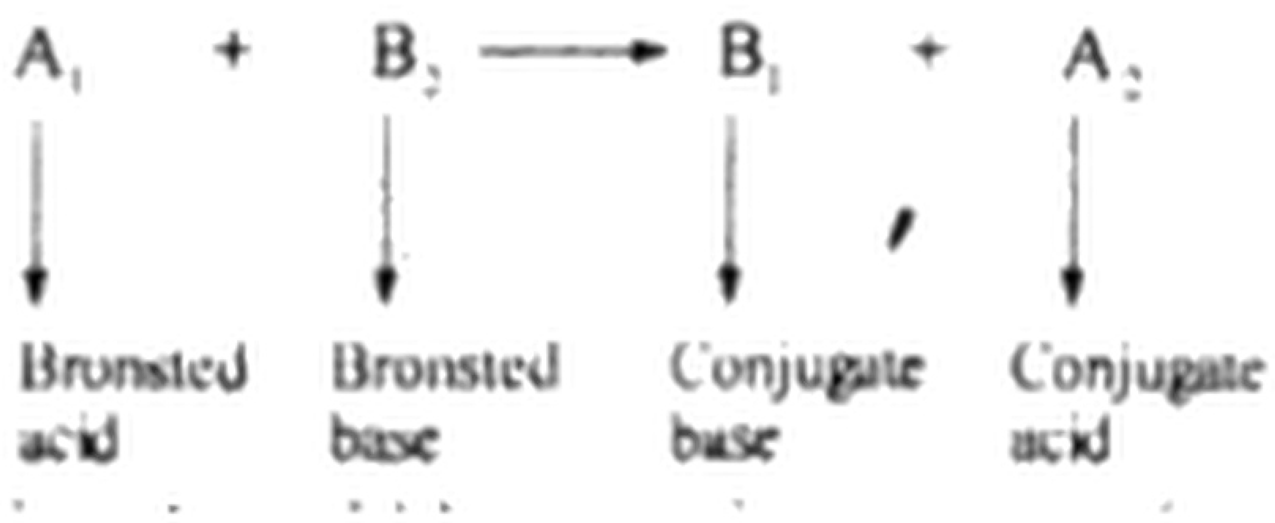
Conjugate base is a species which is left behind after donation of a proton from the acid

e.g. Iionization of HCl in water



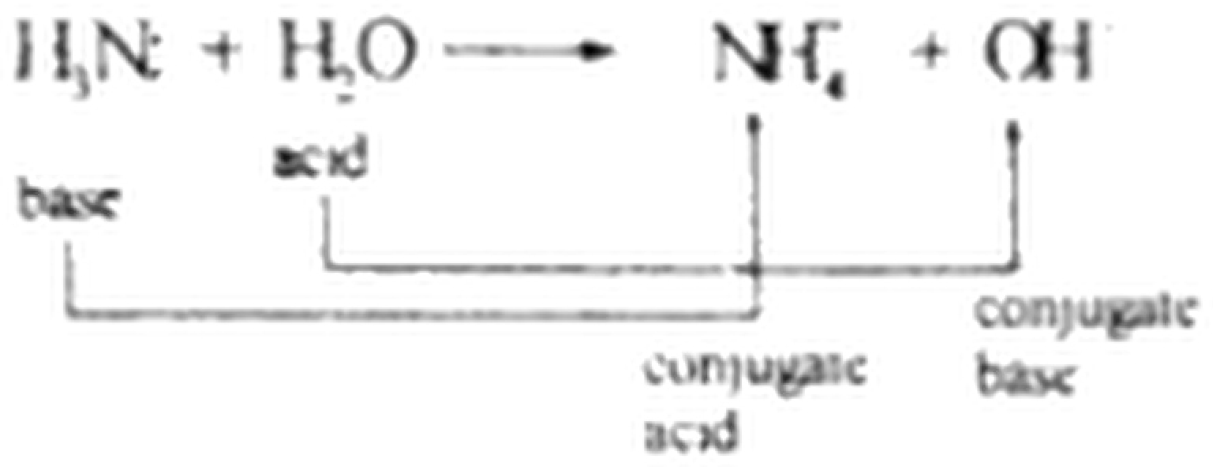
Conjugate Acid-Base Pairs (Major Concepts):

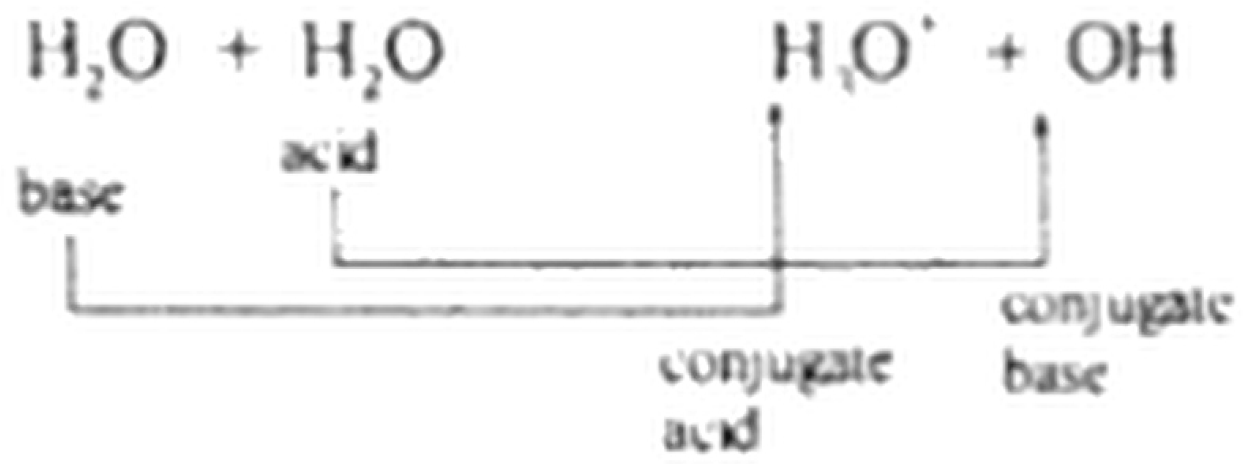
In an acid-base reaction, an acid yields a base (conjugate) and base after accepting proton yields a conjugate acid. The acid-base reaction is represented as



The conjugate acid-base pairs are species on opposite sides of an equation that differ by a proton.

The weaker acids have strong conjugate base pairs and stronger acids huge weaker conjugate bases e.g.

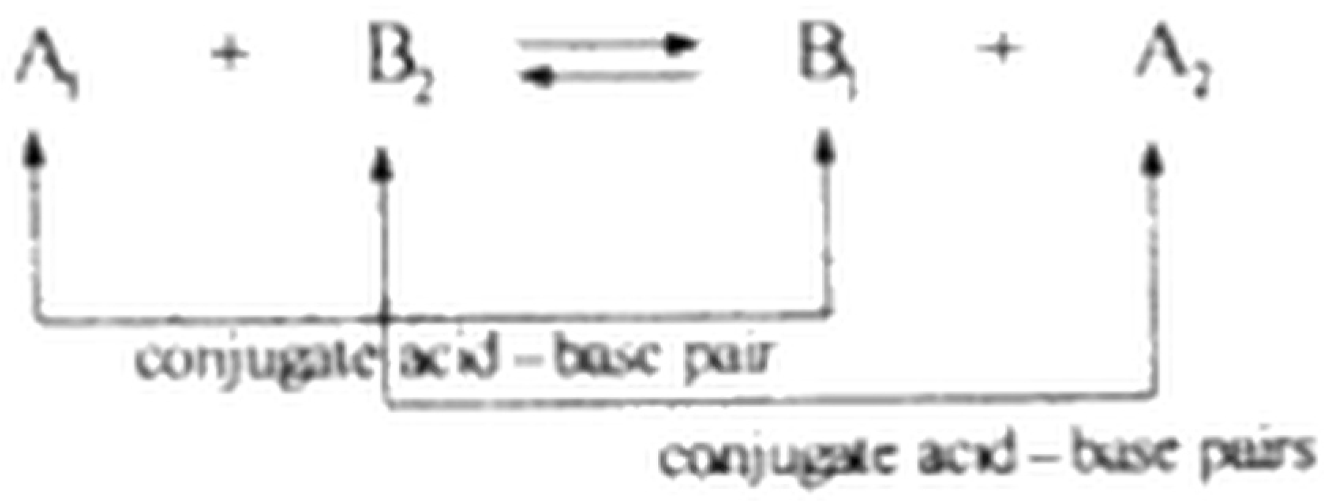




The water is amphoteric in nature i.e it is acidic as well as basic in nature.

General notation:

The general notation of conjugate acid-base pair is



Each acid has a conjugate base and each base has a conjugate acid

Particular Examples of Conjugate Add-Base Pairs:

A1 + B2 B1 + A2

$HCl + H_2O \rightleftarrows Cl^- + H_3O^+$

$HCl + NH_3 \rightleftarrows Cl^- + NH_4^{-1}$

$HNO_3 + H_2O \rightleftarrows NO_3^{-1} + H_3O^+$

$NH_4^{-1} + H_2O \rightleftarrows NH_3 + H_3O^+$

$H_2O + H_2O \rightleftarrows OH^- + H_3O^+$

$H_2O + CO_3^{-2} \rightleftarrows OH^- + HCO_3^{-1}$

$HClO_4 + H_2O \rightleftarrows ClO_4^{-1} + HCO_3^{-1}$

(b) pK_a

The value of K_a are usually inconvenient numbers, therefore, for convenience these values are converted to pk, values. The relationship between K_a and pK_a is as follows:

pK_a = -log K_a

Conclusion:

Since pK_a refers to the negative logarithm of K_a Thus smaller the value pK_a stronger shall be the acid because smaller pK_a value corresponds to a greater K_a value

(c) PK_b

For converlience, a parameter pK_b has been devised to express K_b value in convenient numbers. Thus, pK_b is defined as the negative logarithm of K_b

PK_b= -log K_b

Table: K_b and pK_b Values of Some Common Bases:

Name of Base	Formula	K_b	pK_b
Diethlyamine	$(C_2H_5)_2NH$	9.6×10^{-4}	3.02
Ethylamine	$C_2H_5NH_2$	5.6×10^{-4}	3.25
Methylamine	CH_3NH_2	4.5×10^{-4}	3.34
Ammonia	NH_3	1.7×10^{-5}	4.76
Pyridine	C_5H_5N	5.6×10^{-9}	8.25
Aniline	$C_6H_5NH_2$	4.3×10^{-10}	9.37

Conclusion:

According to these values ammonia is a stronger base than pyridine and aniline but weaker than methylamine and ethylamine Also, diethyl amine is a strongest base among all those listed in the table.

(d) Relative strength of acids

Different Bronsted acids donate proton to different extents

The ability of an acid to donate proton is called 'strength of acid' or the 'acid strength'.

An acid which can donate proton to a higher degree than another acid is said to be relatively strong acid.

Example:

Hydrochloric acid is a relatively stronger acid than acetic acid Also, acetic acid is relatively stronger than water

Q.6 What is hydrolysis? Discuss in detail, the behaviour of each of the following salts in their aqueous solutions.

(a) K_2CO_3 (b) NH_4Cl (c) NaNO_3

Ans: Hydrolysis:

"The reaction of cations and anions of salts with water is called hydrolysis."

Example:

When a salt (MX) is dissolved in water, it splits up into its M and Xions.

These ions may react with water and give following reactions:

M +H-OH ...MOH (aq)+ H⁺

X +H-OH = HX (aq)+ OH

Since H' and OH' ions are produced in these reactions, the solution of the salt may be acid c or basic in salts anions are derived from acids and cations from bases. The anions of weak acids are strong conjugate bases. Such anions react with water producing basic solutions.

Hydrolysis of some salts.

(a) K_2CO_3

$$K_2CO_3 + 2H_2O \rightleftharpoons 2KOH + H_2CO_3$$

Strong base Weak acid

As strong base and weak ac d is produce therefore the solution become basic.

(b) NH_4Cl

$$NH_4Cl + H_2O \rightleftharpoons HCl + NH_4OH$$

Weak base Strong acid

As strong acid and weak bases produce therefore the solution become acidic

(c) $NaNO_3$

$$NaNO_3 + H_2O \rightarrow HNO_3 + NaOH$$

Q.7 (a) Calculate the pH of formic acid-sodium format buffer solution containing 1.0 mole of each component. (Ans: 3.7447)

(b) What will be the pH of the solution after addition of 0.10 mole of hydrochloric acid gas to 1.01dm volume of the buffer solution in part (a) Assume that the volume of solution remains unchanged on addition of hydrochloric acid (K_a for formic acid is 1.8 x 10^{-4}) (Ans: 3.6997)

Solution: Let the volume of buffer solution is 1 dm

[HCOONa] = 1 0 mole dm^{-3}

(HCOOH) = 10 mole dm^{-3}

K_a for formic acid = 1.8 x 10+

P K_a for formic acid = -log(1.8 x 10-4)

p K_a = 3.745

Using Henderson's equation

PH=P K_a + log salt

$$pH = \log\frac{(CH_3COONa)}{(CHCOOH)}$$

$$pH = 3.745 + \log\left(\frac{1.0}{0.1}\right)$$

pH = 3.745 + 0

PH = 3.745

(b) What will be the pH of the solution after addition of 0.10 mole of hydrochloric acid gas to 1.01dm volume of the buffer solution in part (a) Assume that the volume of solution remains unchanged on addition of hydrochloric acid (K_a for formic acid is 1.8 x 10)

Solution:

As HCl is a strong acid it ionizes completely. Thus, it means that 0.1 moles of H^+ produces! These H^+ ions react with 0.1 moles of $HCOO^-$ ions. Hence, out o 1.0 moles of salt (HCOONa), 0.9 roles (1.0.0.1 = 0.9) of salt are left behind

As the equilibrium is shifted towards the left therefore the concentrate on of acid (HCOOH) is increased from 1.0 moles to 1.1 moles (1.0 + 0.1 = 11)

New concentrations of salt and acid:

Number of moles of HCOONa = 0.9 moles

Number of moles of HCOOH = 1.1 moles

Since volume of the buffer solution is 1 01 dm, therefore the molar concentrations will be

[HCOONa) = 0 = 0.891 mole dm .

IHCOOH) =* = 1089 mole dm-3

K_a of formic acid = 1.84 x 10-4

P K_a =-log K_a

= -log (1.84 x 10-) = 3.745

Using Henderson's eq. pH of buffer will be

PH = P K_a + toga

pH = p K_a + log HCOON

pH = 3.745+ log [0.891

pH = 3.745 - 0.0872

OH = 3658

Q.8 (a) Calculate the H^+ ion concentration of an aqueous solution having pH 10.6. (Ans: 2.5×10^{-11} moles/dm^3)

(b) An aqueous solution contains $1.0*10^{-9}$ moles/hm of hydronium ions. Calculate the pOH of these solutions. (Ans: 5.0)

Solution: (a)

pH = 10.6

$[H^+]$ = ?

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

10.6= -log$[H^+]$

Taking antilog on both sides

$[H^+]$ = antilog (-10.6)

$[H^+] = 2.51 \times 10^{-11}$

(b) Hydronium ion (H_3O^+) is actually the hydrated proton,

Therefore

$[H] = 1.0 \times 10^{-9}$

POH = ?

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

PH = -log[1.0 x 10-9

pH = -log(10-9)

pH = -log (10)

PH = 9 x 1 = 9

As, pH + DOH = 14

9 + POH = 14

pOH = 14.9

pOH = 5

Q.9 (a) What is acid dissociation constant? How is it related to pk.

Write equation to elaborate.

(b) Define and briefly describe the levelling effect of water in

acid-base reactions.

Solution:

(a) What is acid dissociation constant? How is it related to pk

Write equation to elaborate

Ans: Acid dissociation constant:

Consider the case of ionization of a general acid HX in water in this aqueous solution the established equilibrium may be represented as follows

$$HX + H_2O \rightleftharpoons H_3o^+ + \bar{X}$$

The equilibrium constant K for this ionization process may be written as follows.

$$K = \frac{\left[H_3O^+\right]\left[X^-\right]}{\left[H_2O\right]\left[HX\right]}$$

$$K\left[H_2O\right] = \frac{\left[H_3O^+\right]\left[X^-\right]}{\left[HX\right]}$$

Since water is a solvent. It is present in excess and therefore its concentration may be regarded as constant. Thus, K (HO) is another constant and is designated as K, thus,

$$K\left[H_2O\right] = K_a = \frac{\left[H_3O^+\right]\left[X^-\right]}{\left[HX\right]}$$

K_a is termed as the acid dissociation constant. It is a measure of the extent to which an acid is ionized or dissociated at the equilibrium state

"The greater the value of K, the stronger is the acid".

Relationship between pK_a and K_a

The value of K, are usually inconvenient numbers, therefore, for convenience these values are converted to pk, values. The relationship between K_a and K_a is as follows:

pK_a = -log K_a

Conclusion:

Since pK_a refers to the negative logarithm of K_a Thus smaller the value pK_a stronger shall be the acid because smaller pK_a value corresponds to a greater value.

(b) Define and briefly describe the levelling effect of water in acid-base reactions.

The levelling Effect:

Whenever acid react with water hydroxonium ion (H_3O^+) will produce which is acidic in nature and is form in each case. The equal behavior of water, levels the strength of the acids and this leveling of strength is called leveling effect

Explanation:

The reaction between any stronger acid and water produce hydroxonium ion (H_3O^+) and water goes completely to the right. Strong acids like perchloric acid, hydrochloric acid, nitric acid and Sulphuric acid in water appear to be of equal strength. Thus, water as a base, is unable to differentiate among the relative acid strength of acids stronger than hydroxonium ion. Such inability of any solvent to differentiate among the relative strength of all acids is termed as the levelling effect. Because the solvent s said to level the strength of all the acids, making them appear identical

effect of water compensated:

The leveling effect of water can be compensated for, if a more weak basic solvent like acetic acid is employed in place of water. Since acetic acid is a much weaker base than water it is not easily protonated Thus, appreciable differences in proton donation of acids are observed in acetic acid (solvent)

The relative acid strength is

HCl > HI > HBr > H_2SO_4 > H NO_3

It may be pointed out that all these acids are of identical strength in water (solvent) due to the levelling effect

e.g.

$$HCl + H_2O \longrightarrow H_3O^+ + Cl^-$$
$$HI + H_2O \longrightarrow H_3O^+ + I^-$$
$$HBr + H_2O \longrightarrow H_3O^+ + Br^-$$
$$H_2SO_4 + H_2O \longrightarrow H_3O^+ + HSO^-_4$$
$$HNO_3 + H_2O \longrightarrow H_3O^+ + NO^-_3$$

Hydroxonium ion (H_3O^+) which is acidic in nature, is formed in each case

The equal behaviour of such ions formed levels the strength of the acid

Q.10 (a) What is hydrolysis? Write the equations of hydrolysis equilibrium for each of the followings:

(i) Li* (ii) NH. (iii) CN

(b) Write a unique property of buffer solution. What types of acids are required to prepare buffer solutions? Name two such acids.

Ans: (a)

"The reaction of cations and anions of salts with water is called hydrolysis."

Explanation:

These observations can be explained on the base of Bronsted-Lowry acid-
The theory. When a salt (MX) is dissolved in water, splits up into its M and X
on These ions may react with water and give following reactions

M + H-OH MOH(-4) + H)

X +H-OHHX) + OH 4)

Since Hand OH' ions are produced in these reactions, the solution of the
on may be acidic or basic. In salts anions are derived from acids and cations from bases. The an ions of weak acids are strong conjugate bases. Such anions react with water producing basic solutions

Examples:

CH_3COO^{-1}, CN^{-1}, HCO^{-1}_3, CO^{-2}_3 etc.

[redacted]tions:

(i)	**Li^+**	$Li^+ + H_2O \rightleftharpoons$	$LiOH + H^+$
(ii)	**NH^+_4**	$NH_4^+ + H_2O \rightleftharpoons$	$NH_4OH + H^+$
(iii)	**CN^-**	$CN^- + H_2O \rightleftharpoons$	$HCN + OH^-$
([redacted])			

Unique property of a buffer solution:

They resist the change in their pH when a small amount of an acid or a base
Is added them. They have a specific constant value of pH.

Constitution of buffer:

A buffer solution contains a weak dissociating acid and the salt of that acid with a strong base

$H_2CO_3 + NaHCO_3$

$H_3PO_4 + NaH_2PO_4$

$CH_3COOH + CH_3COONa$