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Class 11
Chemistry

Chapter 9

Acid Base Chemistry

Handwritten Notes

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CHAPTER-9 ACIDS- BASES CHEMISTRY

Q. Define acids. Bases and Salt and also write their reactions with water.

Acids

i. Acids produce H^+ ion when dissolved in H_2O

$HC\ell \rightleftharpoons H^+ + C\ell^-$

ii. Acids are sour in taste.

iii. They are classified into

a) Mineral acids (e.g HCl, H_2SO_4,HNO_3)

b) Organic acids

iv. Organic acids are weaker than mineral acids and are found in vegetables, fruits etc.

Organic Acids		Where it is found
(i)	Lactic acid	Sour milk
(ii)	Citric acid	citrus fruits
(iii)	Formic acid	insect bites
(iv)	Tartaric acid	grape juice
(v)	Maleic acid	apples and pears.

BASE

(i) Base produces OH^- ions when dissolved in water.

$$NaOH \overset{H_2O}{\rightleftharpoons} Na^+ + OH^-$$

(ii) It has a bitter taste.

(iii) It has a slippery feel.

(iv) When a base is dissolved in water, an alkali is produced.

$$CaO + H_2O \longrightarrow Ca(OH)_2$$

$Ca(OH)_2$ is an alkali and base

(v) Alkalis are used in making soap and detergents.

Do you know

Onions release a gas which turns into sulphuric acid when it reaches your eyes, making them burn

Salt:

The substance obtained as a result of neutralization of acids and bases are called salts.

Acid + Base $\longrightarrow$ Salt + H_2O

Types:

i) **Neutral salt**

It is obtained by neutralization of strong acid and strong base.

$HC\ell$ + $NaOH \longrightarrow NaC\ell + H_2O$

(strong) (strong) (neutral)

ii) **Acidic salt**

It is obtained by the neutralization of a strong acid with a weak base.

$2HC\ell$ + $Ca(OH)_2 \longrightarrow CaC\ell_2 + H_2O$

strong acid weak base acidic salt

iii) **Basic Salt**

It is obtained by the reaction of a strong base with a weak acid.

$2NaOH + H_2CO_3 \longrightarrow Na_2CO_3 + H_2O$

strong	weak	basic
base	acid	salt

ACIDIC SUBSTANCE:

(i) A substance which gives acidic solution when dissolved in water is acidic

(ii) "Oxides of non-metals" e.g.; CO_2, SO_2 and N_2O_5 produce H_2CO_3, H_2SO_3 and HNO_3 when dissolved in water.

(iii) $CO_2 + H_2O \longrightarrow H_2CO_3$

$SO_2 + H_2O \longrightarrow H_2SO_3$

$N_2O_5 + H_2O \longrightarrow 2HNO_3$

BASIC SUBSTANCE:

(i) A substance which gives basic solution when dissolved in water is basic.

(ii) "Oxides of metals" like CaO and Na_2O when dissolved in water produce NaOH and $Ca(OH)_2$

(iii) $Na_2O + H_2O \longrightarrow 2NaOH$

$CaO + H_2O \longrightarrow Ca(OH)_2$

AMPHOTERIC SUBSTANCE:

OR

Q. Elaborate any two amphoteric compound by using chemical equations.

Ans.

(i) Some substances when dissolved in water produce such solutions which are of amphoteric nature i.e. behave as acid as well as base. Such substance is called amphoteric substance.

(ii) **Al_2O_3**

$Al_2O_3 + 6H\,C\ell \longrightarrow 2Al\,C\ell_3 + 3H_2O$ (as base)

$Al_2O_3 + 2NaOH \longrightarrow 2NaAlO_2 + H_2O$ (as acid)

(iii) **Cr_2O_3**

$Cr_2O_3 + 6HNO_3 \longrightarrow 2Cr(NO_3)_3 + 3H_2O$ (as base)

$Cr_2O_3 + 6KOH \longrightarrow 2K_3CrO_3 + 3H_2O$ (as acid)

Q. What is Bronsted-Lowery concept of acids and Bases, explain with suitable examples.

Ans. **Bronsted-Lowery Acid:**

Bronsted Acid is anything that donates proton or hydrogen ion (H^+).

Example:

$HC\ell + H_2O \longrightarrow H_3O^+ + C\ell^-$

In this reaction, the $HC\ell$ donate H+ to water and becomes $C\ell^-$ so it is an acid here while water would be base.

Interesting Information

The wild lupine plant takes nitrogen from the atmosphere and produces ammonia. It was that ammonia to fertilize the soil for itself and surrounding plants.

Bronsted-Lowery Base:

Bronsted Base is anything that accepts proton or hydrogen ion (H^+).

Example:

(i) $H_2SO_4 + H_2O \longrightarrow H_3O^+ + HSO^{-4}$

In this reaction, water accepts H^+ and becomes H_2O+, so it acts as a base.

(ii) $NH_3 + H_2O \longrightarrow NH_4^+ + OH^-$

In this reaction, water donates H^+ which is accepted by NH_3 to become NH_4^+, so here NH_3 acts as a base.

Note: Amphoteric Substance:

There are substances like H_2O and CH_3COOH which act as both acid and base. Such

substances are called as amphoteric.

H_2O

$H_2SO_4 + H_2O \longrightarrow H_2O^+ + HSO_4^-$ (as base)

$NH_3 + H_2O \longrightarrow NH_4^+ + OH^-$ (as acid)

CH_3COOH:

$CH_3COOH + H_2O \longrightarrow CH_3COO^- + H_3O^+$ (as acid)

$CH3\ COOH + H\ C\ell \longrightarrow CH_3OOH_2^+ + C\ell^-$ (as base)

Q. What are conjugate acids and conjugate bases?

Conjugate Acid and conjugate Base:

Consider the reaction between $H\ C\ell$ and H_2O

$H\ C\ell + H_2O \longrightarrow H_3O^+ + C\ell^-$

(acid) (base) (conjugate acid) (conjugate base)

(i) A positively charged specie formed as a result of acceptance of proton by a base is conjugate acid (H_3O^+)

(ii) A negatively charged specie formed as a result of donation of proton by a an acid is called conjugate base ($C\ell^-$).

Q. What are conjugate acid-base pairs and how can they identified?

When an acid reacts with base, a conjugate acid and conjugate base is formed.

The pair of base and conjugate acid is called as conjugate acid-base pair.

Weak acids have strong conjugate base.

Strong acids have weak conjugate base.

Identification of conjugate acid-Base pairs.

$A_1 + B_2 \rightarrow B_1 + A_2$

acid base conjugate base conjugate acid

A_1,B_1 and A_2,B_2 are conjugate acid base pairs.

Typical examples of conjugate acid and base pairs.

1. $H\ C\ell + H_2O \rightleftharpoons H_3O^+ + C\ell^-$
2. $H\ C\ell + NH_3 \rightleftharpoons C\ell^- + NH_4^{+1}$
3. $HNO_3 + H_2O \rightleftharpoons NO_3^{-1} + H_3O^{+1}$
4. $NH_4^{+1} + H_2O \rightleftharpoons NH_3 + H$
5. $H_2O + H_2O \rightleftharpoons H_3O^+ + OH^-$
6. $H_2O + CO_3^{-2} \rightleftharpoons OH^{-1} + HCO_3^{-1}$
7. $HC\ell O_4 + H_2O \rightleftharpoons C\ell O_4^{-1} + H_3O^{+1}$

All the pairs are conjugate acid-base pairs.

LEWIS DEFINITIONS OF ACID AND BASE:

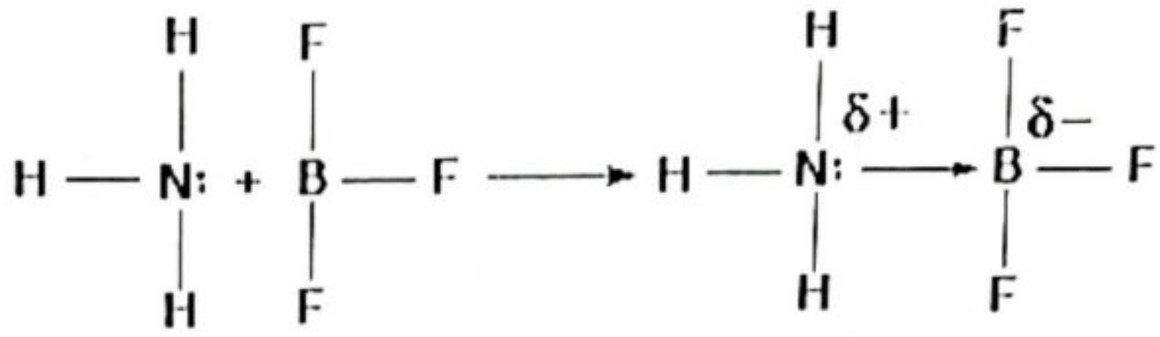
Lewis Acid:

An acid is a substance which can accept a pair of electrons. So, Lewis acid has no unpaired electrons on central atom.

Lewis Base:

A base is a substance, which can donate a pair of electrons. So, Lewis base has a pair of electrons on central atom.

Example:



In this reaction, ammonia donates a pair of electrons to BF_3. So, NH_3 is base and BF_3 is acid.
The arrow → shows a coordinate covalent bond between Lewis acid and Lewis base.

Importance of Lewis Concept:

Lewis's concept of acid and base more general and easily explains those acids and base which are not explained by Bronsted base acid concept.

Strength of Acids and Bases

An acid which can donate proton to a higher degree than other acid is called to be relatively strong acid e.g HCl is stronger acid than CH_3COOH

IONIZATION CONSTANT OF WATER:

Q. Write a detail note on Self-Ionization of Water?

Q. Prove that
$K_w = [H_3O^+]\,[OH^-]$

Water is amphoteric substance which can act as both acid or base.

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

This reaction is called as auto ionization of water.

(i) One molecule of H_2O acts as an acid and donates H^+ ion. This molecule then becomes OH^-.

(ii) Other molecule of H_2O accepts that H^+ and becomes a base. This, then converts to H_3O^+.

(iii) H_3O^+ is called as conjugate acid. OH^- is called as conjugate base.

When there is an equilibrium between H_2O, H_3O^+ and OH^-, K_C is written as:

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

Since water is a solvent, so its concentration remains constant

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

Where $K_c[H_2O]^2 = Kw$

$$Kw = [H_3O^+][OH^-]$$

At 25°C, the K_w for pure water is

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

Where Kw = ionic product constant of water

OR

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

This is the simple ionization of water at equilibrium,

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

As water is solvent

$$K_c[H_2O] = [H^+][OH^-]$$

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

Q. Prove that Kw is temperature dependent.

(3) (i) At 25°C, the K_w for pure water is

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

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

$$[H^+] = [OH^-] = 1 \times 10^{-7} \text{ each}$$

(ii) At 40°C, K_w fr pure water is

$$K_w = [H^+][OH^-] = 3.8 \times 10^{-14}$$

where $[H^+] = [OH^-] = 1.9 \times 10^{-7}$ each

So, K_w is temperature dependent.

pH

pH is defined as negative log of concentration of H^+.

$$pH = \frac{1}{\log[H^+]} = -\log[H^+]$$

if pH = 7, neutral solution

if pH < 7, acidic solution

if pH > 7, basic solution

Lesser is the pH, stronger is the acid.

pOH

pOH is defined as negative log of concentration of OH^- ions.

$$pOH = \frac{1}{\log[OH^-]} = -\log[OH]$$

Q. Prove that PH^+ pOH = 14

i) **pH + pOH = 14**

ionic product of water at 25°C

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

applying log

$\log[H^+][OH^-] = \log 1 \times 10^{-14}$ $\quad \log(a.b) = \log a + \log b$

$\log[H^+] + \log[OH^-] = -14 \log 10$ $\quad \log x^n = n \log x$

multiplying by (−1)

$$(-\log[H^+] + (-\log[OH^-])) = 14 \log 10$$

$$pH + pOH = 14$$

Example 9.2:

The concentration of $[OH^-]$ ion in a household ammonia solution is 0.005M. Calculate the concentration of $[H^+]$ in it.

Solution:

$$[OH^-] = 0.005 \text{ M}$$

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

$$1.0 \times 10^{-14} = [H^+] \times 0.005$$

$$[H^+] = \frac{1.0 \times 10^{-14}}{0.005}$$

$= 2.0 \times 10^{-12}$ M

Example 9.3:

Calculate the pH of 0.001 M aqueous hydrochloric acid solution.

Solution:

Hydrochloric acid ionizes in water completely therefore,

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

0.001 M 0.001 M

$[H_3O^+]$ is in fact the same as $[H^+]$

Therefore $[H^+] = 0.001$ M

$pH = -\log (0.001)$

$= -\log 10^{-3}$

$= 3.00$

Example 9.4:

Calculate the pH of 0.062 M NaOH solution.

Solution:

$$NaOH_{(aq)} \longrightarrow Na^+_{(aq)} + OH^-_{(aq)}$$

0.062M 0.062M 0.062M

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

$= -\log [0.062]$

$= 1.21$

Now

$pH + pOH = 14$

$pH = 14 - pOH$

$= 14 - 1.21$

$= 12.79$

Concept Assessment Exercise 9.2

What is the pH of a solution containing 1.95g pure H_2SO_4 per dm^3 of solution?

Mass of H_2SO_4 = 1.95g

Molecular Mass of H_2SO_4 = 98gmole^{-1}

$$\text{No of moles of } H_2SO_4 = \frac{\text{Mass of } H_2SO_4}{\text{Molecular Mass of } H_2SO_4}$$

$$= \frac{1.95}{98}$$

= 0.02 moles

H_2SO_4 ionizes in water completely therefore,

$$H_2SO_4 + 2H_2O \rightleftharpoons 2H_3O^+ + SO_4^-$$

0.02 2×0.02

$[2H_3O^+]$ Contains $2[H^+]$

Therefore $[H^+] = 2\times0.2$

$= 0.04$m

So

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

$= -\log (0.04)$

$= 1.4$

Titration

Method to find the volume of standard solution required to react completely with known volume of another solution under analysis is called titration

Titration Process:

i. Prepare standard solution and titrant.
ii. Fill burette with titrant and record initial volume.
iii. Measure sample volume and add indicator.
iv. Titrate slowly until endpoint is reached.
v. Record final volume and calculate concentration.

$$M_1V_1/n_1 = M_2V_2/n_2$$

Variables:

V_1= Volume of base taken in flask
V_2= Volume of acid used from the burette
M_1= Molarity of base
M_2= Molarity of acid
n_1= No. of mole of base
n_2 = No. of mole of acid

Selection of Indicators in Acid-Base Titration

To choose the right indicator for an acid-base titration, consider the pH range of the reaction. The indicator should change color within the pH range of the equivalence point.

1- Strong Acid-Strong Base: pH at equivalence point is 7, so indicators like bromothymol blue (pH 6.0-7.6) or phenolphthalein (pH 8.3-10.0) work well.
2-WeakAcid-StrongBase: pH at equivalence point is basic (>7), so indicators like phenolphthalein are suitable.
3-Strong Acid-Weak Base: pH at equivalence point is acidic (<7), so indicators like methyl orange (pH 3.1-4.4) is suitable.

Q. STRONG AND WEAK ACIDS.

According to Bronsted, acid is a substance that donates proton. The acid which has high tendency of donating proton is a strong acid and vice versa.

Ionization Constant of Acid (K_a)

K_a is the ionization constant of acid. Suppose reaction with general acid "HX".

$$HX + H_2O \rightleftharpoons H_3O^+ + X^-$$

K_c at equilibrium is

$$K_c = \frac{[H_3O^+][X^-]}{[HX][H_2O]}$$

as water is solvent and has constant concentration so,

$$K_c[H_2O] = \frac{[H_3O^+][X^-]}{[HX]}$$

The product of two constants K_c and $[H_2O]$ is equal to ionization constant of acid, K_a.

$$\text{So} \quad K_a = \frac{[H_3O^+][X^-]}{[HX]}$$

K_a is a temperature dependent quantity.

"Higher is value of K_a, greater is the strength of acid."

if $K_a > 1$ strong acid

if K_a = 1 to 10^{-3} moderate acid

if $K_a < 10^{-3}$ weak acid

pK_a

pK_a is defined as negative log of K_a.

$$pK_a = -\log K_a$$

Smaller is the value of pK_a, large will be K_a and stronger would be the acid.

Example 9.5:

Calculate concentration of H^+ ions of a solution that contains 1.0M HF ($K_a = 7.2 \times 10^4$)

Solution:

$$HF_{(aq)} \rightleftharpoons H^+_{(aq)} + F^-_{(aq)}$$

Initial conc.	1.0M	0	0
Eq. conc. (Mole dm^{-3})	$1.0 - x$	x	x

$$K_a = \frac{[H^+][F^-]}{[HF]}$$

$$7.2 \times 10^{-4} = \frac{x \cdot x}{1.0 - x}$$

Since x is very small as compared to 1.0, the term in the denominator can be approximated as follows:

$$1.0 - x = 1.0$$

$$7.2 \times 10^{-4} = \frac{x^2}{1}$$

$$x = 0.268M$$

$$[H^+] = 0.268M$$

STRONG AND WEAK BASES.

According to Bronsted, base is a proton acceptor. The base which has high tendency to accept proton is stronger and vice versa.

Ionization Constant of Base (K_b)

K_b is defined as ionization constant of base.

Suppose a reaction with general base B

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

K_c at equilibrium is

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

As water is a solvent, so its concentration is constant

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

her K_c $[H_2O]$ is product of two constants equal to ionization constant of base K_b.

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

Greater is the value of K_b, stronger is the base.

pK_b

pK_b is defined as negative log of K_b.

$pK_b = -\log K_b$

Smaller is the value of pK_b, larger will be the K_b and stronger is the base.

Relationship of K_a and K_b.

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

$$HA \rightleftharpoons H^+ + A^-$$

$$K_a = \frac{[H^+][A^-]}{[HA]} \quad \ldots\ldots\ldots (i)$$

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

$$K_c = \frac{[HA][OH^-]}{[A^-][H_2O]}$$

$$K_c[H_2O] = \frac{[HA][OH^-]}{[A^-]}$$

$$K_b = \frac{[HA][OH^-]}{[A^-]} \quad \ldots\ldots\ldots (ii)$$

multiplying I and ii

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

$$K_a \times K_b = [H^+][OH^-]$$

$$K_a \times K_b = K_w \quad \ldots\ldots\ldots (iii)$$

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

Taking K_w as constant.

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

Q. **Prove that $pK_a + pK_b = pK_w$.**

$pK_a + pK_b = pK_w$

as mentioned above in (iii)

$K_a \times K_b = K_w$

applying log

$\log K_a \times \log K_b = \log K_w$

$\log (K_a + K_b) = \log K_w$

multiplying both sides by (–1)

$-\log K_a - \log K_b = \log K_w$

or $pK_a + pK_b = pK_w$

Q. **Prove that $pK_a + pK_b = 14$**

$pK_a + pK_b = 14$

as $K_a \times K_b = K_w$

$$K_w = 1 \times 10^{-14}$$

applying log on both sides

$$\log K_a \times K_b = \log 10^{-14}$$

$$\log K_a + \log K_b = 14 \log 10$$

multiplying (–1) with both sides

$$-\log Ka - \log Kb = (-14)(-1)$$

$$pK_a + pK_b = 14$$

BUFFER SOLUTION AND THEIR APPLICATIONS

Q. Write a detail note on Buffer solution and its applications in daily life.

Solution whose pH do not change by addition of small amount of acid, base, or dilution keeping it for a long time, is called buffer solution.

Types:

There are two types of buffer solution:

(i) Acidic buffer solution:

It is the mixture of weak acid and its salt with strong base is called acidic buffer solutic Its pH is less than 7.

$CH_3COOH + CH_3COONa$

$H_3PO_4 + Na_3PO_4$

(ii) Basic Buffer Solution:

It is the mixture of weak base and its salt with strong acid is called basic buffer solution. I pH is greater than 7.

$NH_4OH + NH_4C\ell$

$Al(OH)_3 + AlC\ell_3$

Buffer Action:

Buffer solution has a constant value of pH. The method by which a buffer can maintain it pH is called buffer action.

Acidic Buffer solution is ionized as,

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

$$CH_3COONa \rightleftharpoons CH_3COO^- + Na^+$$

On adding acid

When acid is added in buffer solution, H^+ of acid combine with acetate ion (CH_3COO^- and form acetic acid which is already present in solution so there is no change in pH.

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

On adding base

When acidic buffer solution gets some base, the OH^- of base combine with H^+ of buffer tc form water which is neutral, so there is no change in pH

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

Basic Buffer solution is ionized as,

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

$$NH_4C\ell \rightleftharpoons NH_4^+ + C\ell^-$$

On adding acid:

When acid is added in basic buffer solution, H+ of acid combine with OH^- ion and form water which is neutral. So pH do not change.

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

On adding base

When base is added in buffer solution, OH^- of base combine with NH_4^+ to form NH_4OH which is already present in solution, so there is no pH change.

$$NH_4^+ + OH^- \longrightarrow NH_4OH$$

Facts about buffer solution:

- When a weak acid and its conjugate base is added in water OR when a weak base and its conjugate acid is added in water, a buffer solution is formed.
- Buffer solution always resists change in pH and hence its pH is constant.
- Whenever the acid or base is added in buffer, the H^+ or OH^- ions are produced which react with negative and positive ions of buffer and hence pH is made constant.
- When a strong base or strong acid is used, the buffer solution is not formed.

Examples of buffer are:

$CH_3COOH + CH_3COONa$
$H_3PO_4 + Na_3PO_4$
Formic acid + sodium format

Applications of buffer:

(i) **Industrial Application:**
They are used in industries to prepare photographic material, leather and dyes.

(ii) **Use in Processes:**
They are used in process of electro-plating and analytical procedures like titration.

(iii) **PH measurement:**
The buffer solutions are used as a standard for calibration of pH meters.

(iv) **Preparation of culture:**
Buffer is added in culture of bacteria so that pH of culture remains constant.

Natural Buffer Solutions:

(i)	human blood	(7.35 – 7.45)
(ii)	tears	(7.40)
(iii)	milk	(1.65 – 1.75)
(iv)	egg white	(8.0 – 8.1)

Example 9.6:

What is the pH of buffer if concentration of CH_3COOH is 0.1 M dm^{-3} and CH_3COONa is 1.0 M dm^{-3} pK_a for CH_3COOH is 4.76? Solution:

Concentration of CH_3COOH = 0.1 M dm^{-3}
Concentration of CH_3COONa = 1.0 M dm^{-3}
The formula to determine pH value is

$$pH = pK_a + \log \frac{\text{salt (i.e. } CH_3COONa)}{\text{acid (i.e. } CH_3COOH)}$$

$$pH = 4.76 + \log\left(\frac{1.0}{0.1}\right)$$

$$pH = 4.76 + \log 10 \quad \left(\frac{1.0}{0.1} = \frac{10}{1}\right)$$

$pH = 4.76 + 1.00$ ($\log 10 = 1$)
$pH = 5.76$

Concept Assessment Exercise 9.4

Calculate the pH of a buffer solution in which 0.11 Molar CH_3COONa and 0.09 Molar CH_3COOH solutions are present. K_a for CH_3COOH is 1.8×10^{-5}.

(i) Concentration of CH_3COOH = 0.09m
(ii) Concentration of CH_3COONa = 0.11M

K_a for $CH_3COOH^+ = 1.8 \times 10^{-5}$

Since $pK_a = -\log Ka$

$= -\log 1.8\times10^{-5}$

$= -(-5) \log 1.8\times10$

$= 5[\log 1.8 + \log10]$

$= 4.745$

$$pH = Pk_a + \log \frac{[salt]}{[acid]}$$

$$= 4.745 + \log \frac{[0.1]}{[0.09]}$$

$$= 4.745 + \log \left(\frac{0.11}{0.9}\right)$$

$= 4.745 + 0.087$

$pH = 4.83$

SALT HYDROLYSIS

Q. What is salt hydrolysis? How salts are characterized as acidic or basic in nature?

i) Hydrolysis is the break down of water into H^+ and OH^- ions.

ii) Salt hydrolysis is defined as a reaction between cations and anions of salts with H^+ and OH^- of water.

Generalized salt Hydrolysis:

Consider a salt MX dissolved in water. It produces M^+ and X^- ions.

"$MX \longrightarrow M^+ + X^-$"

The M^- reacts with OH^- of water as follows:

"$M^+ + H_2O \longrightarrow MOH + H^+$"

The X- reacts with H+ of water as follows:

"$X^- + H_2O \longrightarrow HX + OH^-$"

Four types of Salt Hydrolysis:

(i) When the salts of weak acids or strong bases hydrolyze, basic solution is produced. pH of solution is greater that 7.

e.g. $CH_3COONa + H_2O \longrightarrow CH_3COOH + NaOH$

So overall solution is basic.

(ii) When the salts of strong acids or weak bases hydrolyse in water, acidic solution is produced with pH < 7.

e.g. $CuSO_4 + 2H_2O \longrightarrow Cu(OH)_2 + H_2SO_4$

So overall, solution is acidic.

(iii) When salt of strong acid and strong base are hydrolyzed, they give strong acids and bases. So, the solution becomes neutral with pH = 7.

e.g $Na\,C\ell + H_2O \longrightarrow NaOH + H\,C\ell$

So, the strong base and strong acid neutralize each other and pH of solution becomes 7.

(iv) Salts of weak acids and weak bases hydrolyze but the resulting solution can be neutral, acidic or basic as follows:

if, $K_a = K_b$ neutral

$K_a > K_b$ acidic

$K_a < K_b$ basic

Salt Type		Common Example	Ions which impart Hydrolysis	Solution pH (Nature)
Acid	Base			

Strong	Strong	Na Cl, KBr	None	= 7.0 (Neutral)
Strong	Weak	NH_4NO_3, NH_4Cl	Cations	< 7.0 (Acidic)
Weak	Strong	NaCN, K_2CO_3	Anions	> 7.0 (Basic)
Weak	Weak	NH_4CN, NH_4NO_2	Anions & Cations	May be equal, smaller or greater than 7.0

SOLUBILITY PRODUCT AND PRECIPITATION REACTIONS

"Product of molar concentration of apposite charge ion in equilibrium with solid state i. called solubility product."

OR

Product of equilibrium concentrations of ions each raised toa power which is the coefficient of ion in the balance chemical equation

It is denoted by K_{sp} (stability product constant).

Its value depends upon temperature.

It can be written only for slightly soluble salts

e.g.

$$CaF_{2(s)} \rightleftharpoons Ca^{2+} + 2F^-$$

$$K_c = \frac{[Ca^{+2}][F^-]^2}{[CaF_2]}$$

$$K_c[CaF_2] = [Ca^{+2}][F^-]^2$$

$$K_{sp} = [Ca^{+2}][F^-]^2$$

For general reaction:

$$A_m + B_n \rightleftharpoons mA^{+n} + nB^{-m}$$

$$K_{sp} = [A^{+n}]^m [B^{-m}]^n$$

Applications:

When K_{sp} of the reaction is given, solubility of the reaction can be calculated and vice versa

PRECIPITATION REACTION:

"Reaction in which two or more than two solutions are mixed together. As a result, insoluble substance is precipitated out. Such reactions are called precipitated reactions."

$$CaF_{2(s)} \rightleftharpoons Ca^{2+} + 2F^-$$

ion product (Q′) =[Initial conc. of Ca^{2+}] [Initial conc. of F^-]^2

There are three cases:

Case I: If ion product (Q′) = solubility product constant (K_{sp}), solution is sat urated but precipitation will not occur.

Case II: It Q′>K_{sp}, solution is super saturated and precipitation will occur until Q′ = K_{sp}.

Case III: If Q′ < K_{sp}, solution will be unsaturated and precipitation will not occur.

Q′ is noted before equilibrium
K_{sp} is calculated at equilibrium

Example 9.7

The solubility of Ag Br is 7.1×10^{-7}M at 25°C. Calculate its K_{SP}.

Solution

7.1×10^{-7}M of dissolved Ag Br produces equal moles of Ag^+ and Br ions.

$$AgBr_{(s)} \rightleftharpoons Ag^{+}_{(aq)} + Br^{-}_{(aq)}$$

7.1×10^{-7}M $\quad 7.1 \times 10^{-7}$M $\quad 7.1 \times 10^{-7}$ M

$K_{sp} = [Ag^+][Br^-]$

$K_{SP} = (7.1 \times 10^{-7})\ (7.1 \times 10^{-7})$

$K_{SP} = 5.041 \times 10^{-13}$

COMMON ION EFFECT:

"Decrease in solubility or ionization of an electrolyte by the addition of another electrolyte having a common ion is called common ion effect."

Examples:

(1) $NaCl_{(s)} \rightleftharpoons Na^+ + Cl^-$

$HCl_{(g)} \rightleftharpoons H^{\oplus}_{(aq)} + Cl$

H Cℓ is strong electrolyte than Na Cℓ.

When H Cℓ is passed through saturated solution of Na Cℓ, Na Cℓ will precipitin out due to common ion effect of $C\ell^-$.

(2) $HF_{(\ell)} \rightleftharpoons H^{\oplus}_{(aq)} + F^{-}_{(aq)}$

$NaF_{(s)} \rightleftharpoons Na^{\oplus}_{(aq)} + F^{-}_{(aq)}$

NaF is strong electrolyte than HF, so when it passes through saturated solution of HF, the solubility of HF is suppressed due to common ion effect of F^-.

Applications:

(1) Na Cℓ is detained from sea water due to common ion effect.

(2) Buffer solution is prepared on the basis of common ion effect.

(3) Quantitative analysis (salt analysis) is based on common ion effect.

Examples 9.8

Potassium per chlorate $KC\ell O_4$ is moderately soluble in water. When highly soluble KCℓ is added to the saturated solution of $KC\ell O_4$. It causes to increase the concentration of K^+ ion.

$$KC\ell O_4 \rightleftharpoons K^{+}_{(aq} + C\ell O_4^{-}{}_{(aq)}$$

$$KC\ell(s) \rightleftharpoons K^{+}_{(aq} + C\ell^{-}_{(aq)}$$

According to the Le Chatelier's principle K^+ ions will react with $C\ell O_4^-$ ions to form $KC\ell_{(s)}$. This will suppress, the ionization of $KC\ell O_4$. Thus it will precipitate out. ii. When HCℓ gas is passed through the saturated solution of NaCℓ (Brine), it causes to increases the concentration of $C\ell^{-1}$ ion.

Examples 9.9

$$NaC\ell_{(S)} \rightleftharpoons Na^{+}_{(aq)} + C\ell^{-}_{(aq)}$$

$$NaC\ell_{(aq)} \rightleftharpoons H^{+}_{(aq)} + (aq) + C\ell^{-1}_{(aq)}$$

EXERCISE

Q.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 0.01 M H Cℓ 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) Amphotric

(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) Always 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 (CO_2) acidic

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

(ix) **If 25cm³ of 1 mol.dm³ nitric acid is added to 50cm³ of 0.5 mol.dm⁻³ 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 neutralise an add. 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) Na_2O (c) SO_2 (d) SiO_2

(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_2.
(c) Reaction with CO_2 in the air.
(d)Loss of water, sulphur dioxide and oxygen.

ANSWERS

1. (a)	2. (c)	3. (b)	4. (d)	5. (b)	6. (d)	7. (b)	8. (d)	9. (b)
10. (d)	11. (c)	12. (b)	13. (c)	14. (a)				

SHORT QUESTIONS/ANSWERS

Q.2 Short Questions

(i) Conjugate acid – base pairs:

See in Notes

(ii) Define Lewis acid and base:

See in notes also iii, iv, v, vi See in notes

(vii) Relationship of K_a and K_b:

$$HA \rightleftharpoons H^+ + A^-$$

$$K_a = \frac{[H^+][A^-]}{[HA]} \quad \ldots\ldots\ldots (i)$$

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

$$K_c = \frac{[HA][OH^-]}{[A^-][H_2O]}$$

$$K_c[H_2O] = \frac{[HA][OH^-]}{[A^-]}$$

$$K_b = \frac{[HA][OH^-]}{[A^-]} \quad \ldots\ldots\ldots (ii)$$

multiplying I and ii

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

$$K_a \times K_b = [H^+][OH^-]$$

$$K_a \times K_b = K_w \quad \ldots\ldots\ldots (iii)$$

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

Taking K_w as constant.

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

(viii) Buffer Solution:

"Solution whose pH do not change by addition of small amount of acid, base, or dilution or keeping it for a long time, is called buffer solution."

Examples:

- $CH_3COOH + CH_3COONa$

- $H_3PO_4+Na_3PO_4$

Q.3 See in Notes

Q.4 See in Notes

Q.5 See in Notes

Q.6 **Hydrolysis**

"Hydrolysis is break down of water into H^+ and OH^-." OR

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

Explanation:

Hydrolysis is different from hydration. In hydrolysis H – OH bond is broken whereas in hydration water molecule adds up to a substance-without bond breakage.

Generalized salt Hydrolysis:

Consider a salt MX dissolved in water. It produces M^+ and X^- ions.

$$\text{"MX} \longrightarrow M^+ + X^-\text{"}$$

The M^- reacts with OH^- of water as follows:

$$\text{"}M^+ + H_2O \longrightarrow MOH + H^+\text{"}$$

The X- reacts with H+ of water as follows:

$$\text{"X- + } H_2O \longrightarrow HX + OH^-\text{"}$$

(a) K_2CO_3

$K_2CO_3 + H_2O \longrightarrow KOH + H_2CO_3$

(b) $NH_4C\ell$

$NH_4\,C\ell + H_2O \longrightarrow NH_4OH + H\,C\ell$

(c) $NaNO_3$

$NaNO_3 + H_2O \longrightarrow NaOH + HNO_3$

Q.7 [HCOOH] = 1M

[HCOONa] = 1M

pH= ?

$K_a = 1.8 \times 10^{-4}$

$pK_a = -\log K_a$

$pK_a = -\log 1.8 \times 10^{-4}$

$pK_a = 3.74$

$$pH = pK_a + \log \frac{[HCOONa]}{[HCOOH]}$$

$$pH = 3.74 + \log\left(\frac{1}{1}\right)$$

$$\boxed{pH = 3.74}$$

Q.8 When solid $PbC\ell_2$ is added to pure water at 25°C, the salt dissolves until the concentration of Pb^{+2} reaches 1.6×10^{-2}M. After this concentration is reached, excess solid remains un-dissolved. What is K_{SP} for this salt.

$Pb\,C\ell_2$ ionizes in water as

$$Pb\,C\ell_{2(g)} \rightleftharpoons \underset{1.6\times10^{-2}M}{Pb^{+2}_{(aq)}} + \underset{2\times 1.6\times10^{-2}}{2\,C\ell^-_{aq}}$$

$$Ksp = [Pb^{+2}]\,[C\ell^{-}]^2$$
$$= (1.6\times10^{-2})\,(1.6\times10^{-2})^2$$
$$Ksp = 1.6384\times10^{-5}$$

Q.9 (a) **Given data:**

pH = 10.6, H^+ ion =?

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

Taking antilog on both sides

$$H^+ = \text{anti log} - pH$$
$$H^+ = \text{anti log} -10.6$$
$$H^+ = 2.5 \times 10^{-11}\ \text{moles/dm}^3$$

(b) **Given Data:**

$OH^- = 1.0\times10^{-9}$ **moles/dm³, pOH =?**

Solution:

$$pOH = -\log OH^{-1}.$$
$$pOH = -\log (1.0\times10^{-9})$$
$$pOH = 9M$$

Q.10 **Hydrolysis:**

Hydrolysis defined as "break down of water (H_2O) molecule into H^+ and OH^- ions."

(i) K^+

$$K^+ + H_2O \rightleftharpoons KOH + H^+$$

(ii) NH^+_4

$$NH_4^+ + H_2O \longrightarrow NH_4OH + H^+$$

(iii) CN^-

$$CN^- + H_2O \longrightarrow NCN + OH^-$$

Q.11 See in Notes
Q.12 See in Notes
Q.13 See in Notes
Q.14 See in Notes
Q.15 See in Notes
Q.5 See in Notes
Q.5 See in Notes

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