



Ranjan Singh
Chemistry Classes

JEE (Main & Advanced) • NEET • XI • XII

Study Package



Ranjan Singh

M.Sc. Bio-Chemistry(P.U)

Ex-Faculty : Narayana & Goal

PHYSICAL CHEMISTRY

IONIC EQUILIBRIUM



Director's Message



Ranjan Singh
M.Sc.(Biochemistry), P.U.

Chemistry plays a central and important role in all competitive examinations as well as in day to day life. For last so many years, I have constantly been in touch with students, guiding them in Chemistry and looking into their difficulties for them to succeed in their board as well as competitive examinations JEE(Mains & Advance) | NEET.

I have felt a need for a good coaching centre to fulfil the requirements of students. Students need a highly experienced and qualified faculty in chemistry, who can guide them well, clear their doubts, provide them the effective & tricky notes, and make them do much needed practice. More importantly they should also be provided Classroom Monitoring, Periodical & Surprise Tests to guide them in the proper direction. I realize that, it is very important to diagnose the basic weaknesses and problems of students not succeeding in JEE(Mains & Advance) | NEET and Board exams. In fact, as question patterns are changing, now they need to have a different approach for these Examinations.

At RANJAN SINGH CHEMISTRY CLASSES, we have our own way to prepare students for Competitive Examinations as well as Board Examination at a time so they can crack the entrance exam like JEE(Mains & Advance) and NEET as well as 12th Board simultaneously. We act as a medium to provide the simplest, easiest and a comfortable way to make students achieve their target. At RANJAN SINGH CHEMISTRY CLASSES(RSCC), we guide our students with the best motivational classes so weak students are also able to believe that, They can do it.

When you join RANJAN SINGH CHEMISTRY CLASSES you become a part of the powerful force which propels you towards your goal and if you get a position among the rankers with my excellent guidance, I will think that our efforts have borne fruits.

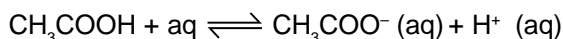
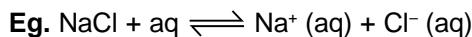
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Introduction

Degree of dissociation

When an electrolyte is dissolved in a solvent (H_2O), it spontaneously dissociates into ions. It may dissociate partially ($\alpha \ll 1$) or sometimes completely ($\alpha \cong 1$)



The degree of dissociation of an electrolyte (α) is the fraction of one mole of the electrolyte that has dissociated under the given conditions.

$$\alpha = \frac{\text{No. of moles dissociated}}{\text{No. of moles taken initially}}$$

Ostwald's Dilution Law (for weak electrolyte's)

For a weak electrolyte A^+B^- dissolved in water, if α is the degree of dissociation then

	$AB \rightleftharpoons A^+ + B^-$
initial conc	C 0 0
conc-at eq.	$C(1-\alpha)$ $C\alpha$ $C\alpha$

Then according to law of mass action,

$$K_{eq} = \frac{[A^+][B^-]}{[AB]} = \frac{C\alpha \cdot C\alpha}{C(1-\alpha)} = \frac{C\alpha^2}{(1-\alpha)} = \text{dissociation constant of the weak electrolyte}$$

If α is negligible in comparison to unity $1 - \alpha \approx 1$, so $k_{eq} = \alpha^2 C \Rightarrow \alpha = \sqrt{\frac{k_{eq}}{C}} =$

$$\Rightarrow \alpha \propto \frac{1}{\text{concentration}}$$

as concentration increases $\Rightarrow \alpha$ decreases

Acids Bases and Salts :

◆ **Arrhenius concept :**

Arrhenius Acid : Substance which gives H^+ ion on dissolving in water (H^+ donor)
 eg. HNO_3 , $HClO_4$, HCl , HI , HBr , H_2SO_4 , H_3PO_4 etc.

◆ **Arrhenius base :** Any substance which releases OH^- (hydroxyl) ion in water (OH^- ion donor)

Bronsted - Lowery concept : (Conjugate acid - base concept) (Protonic concept)

Acid : substances which donate H^+ are Bronsted Lowery acids (H^+ donor)

Base : substances which accept H^+ are Bronsted Lowery bases (H^+ acceptor)

◆ **Lewis concept (electronic concept) :**

An acid is a molecule/ion which can accept an electron pair with the formation of a coordinate bond.

Acid $\rightarrow e^-$ pair acceptor

e.g. Electron deficient molecules : BF_3 , $AlCl_3$
 Cations : H^+ , Fe^{2+} , Na^+
 Molecules with vacant orbitals : SF_4 , PF_3

A base is any molecule/ion which has a lone pair of electrons which can be donated.

Base \rightarrow (One electron pair donate)

e.g. Molecules with lone pairs : NH_3 , PH_3 , H_2O , CH_3OH



Properties of Water :

◆ Amphoteric (amphiprotic) Acid/base nature:

Water - an acid as well as base according to Arrhenius and Bronsted-Lowry theory but according to Lewis concept it can only be taken as base only.

In pure water $[H^+] = [OH^-]$ so it is Neutral.

◆ Molar concentration / Molarity of water :

Molarity = No. of moles/litre = $\frac{1000\text{gm/litre}}{18\text{gm/mole}} = 55.55 \text{ mole /litre} = 55.55 \text{ M}$ (density = 1 gm/cc)

◆ Ionic product of water :

According to Arrhenius concept

$H_2O \rightleftharpoons H^+ + OH^-$ so, ionic product of water, $K_w = [H^+][OH^-] = 10^{-14}$ at 25° (experimental)

○ dissociation of water, is endothermic, so on increasing temperature K_{eq} increases

K_w increases with increase in temperature.

Now $pH = -\log[H^+] = 7$ and $pOH = \log[OH^-] = 7$ for water at 25° (experimental)

$pH = 7 = pOH \Rightarrow$ neutral

$pH < 7$ or $pOH > 7 \Rightarrow$ acidic at $25^\circ C$

$pH > 7$ or $pOH < 7 \Rightarrow$ Basic

Ionic product of water is always a constant whatever has been dissolved in water since its an equilibrium constant so will be dependent only on temperature.

◆ Degree of dissociation of water :

$H_2O \rightleftharpoons H^+ + OH^- \Rightarrow \alpha = \frac{\text{no. of moles dissociated}}{\text{Total No. of moles initially taken}} = \frac{10^{-7}}{55.55} = 18 \times 10^{-10}$ or $1.8 \times 10^{-7}\%$

Acidity and pH scale :

Acidic strength means the tendency of an acid to give H_3O^+ or H^+ ions in water.

So greater the tendency to give H^+ , more will be the acidic strength of the substance.

Basic strength means the tendency of a base to give OH^- ions in water.

So greater the tendency to give OH^- ions, more will be basic strength of the substance.

$pH = -\log a_{H^+}$ (where a_{H^+} is the activity of H^+ ions)

The pH scale was marked from 0 to 14 with central point at 7 at $25^\circ C$ taking water as solvent.

If the temperature and the solvent are changed, the pH range of the scale will also change. For example

0 – 14	at $25^\circ C$	Neutral point, $pH = 7$
0 – 13	at $80^\circ C$ ($K_w = 10^{-13}$)	Neutral point, $pH = 6.5$

pH can also be negative or > 14

pH Calculations of Different Types of Solutions :

◆ Strong acid solution :

(i) If concentration is greater than 10^{-6} M

In this case H^+ ions coming from water can be neglected,

so $[H^+] =$ normality of strong acid solution

(ii) If concentration is less than 10^{-6} M

In this case H^+ ions coming from water cannot be neglected.

So $[H^+] =$ normality of strong acid + H^+ ions coming from water in presence of this strong acid

◆ Strong base solution :

Calculate the $[OH^-]$ which will be equal to molarity of the strong base solution and then use

$K_w = [H^+] \times [OH^-] = 10^{-14}$, to calculate $[H^+]$.

◆ **pH of mixture of two strong acids :**

If V_1 volume of a strong acid solution of normality N_1 is mixed with V_2 volume of another strong acid solution of normality N_2 , then

$$[H^+] = N = \frac{N_1V_1 + N_2V_2}{V_1 + V_2}$$

◆ **pH of mixture of two strong bases :**

similar to above calculation

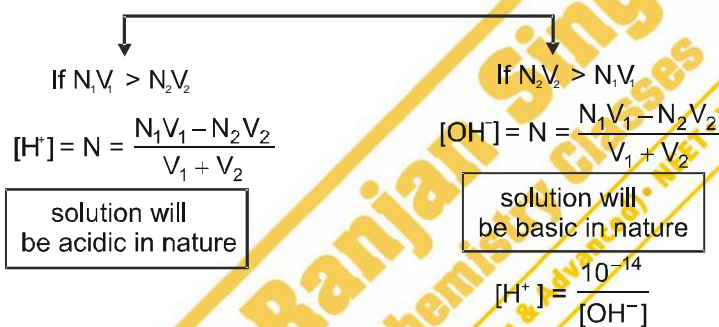
$$[OH^-] = N = \frac{N_1V_1 + N_2V_2}{V_1 + V_2} \quad [H^+] = \frac{10^{-14}}{[OH^-]}$$

◆ **pH of mixture of a strong acid and a strong base :**

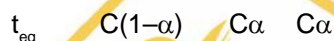
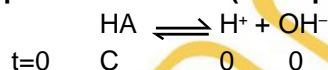
Acid Base neutralisation reaction will take place.

The solution will be acidic or basic depending on which component has been taken in excess.

If V_1 volume of a strong acid solution of normality N_1 is mixed with V_2 volume of a strong base solution of normality N_2 , then



◆ **pH of a weak acid(monoprotic) solution :**



$$K_a = \frac{[H^+][OH^-]}{[HA]} = \frac{C\alpha^2}{1-\alpha}$$

if $\alpha \ll 1 \Rightarrow (1-\alpha) \approx 1$

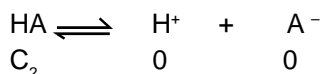
$$\Rightarrow K_a \approx C\alpha^2 \quad \Rightarrow \alpha = \sqrt{\frac{K_a}{C}} \quad (\text{is valid if } \alpha < 0.1 \text{ or } 10\%)$$

$$[H^+] = C\alpha = C\sqrt{\frac{K_a}{C}} = \sqrt{K_a \times C} \quad \text{So } \text{pH} = (\log \sqrt{K_a \times C})$$

on increasing the dilution $\Rightarrow C \downarrow \Rightarrow \alpha \uparrow$ and $[H^+] \downarrow \Rightarrow \text{pH} \uparrow$

◆ **pH of a mixture of weak acid(monoprotic) and a strong acid solution :**

$[SA] = C_1$ and $[WA] = C_2$, then $[H^+]$ from $SA = C_1$
the weak acid will dissociate as follows.



$$C_2(1-\alpha) \quad C_2\alpha + C_1 \quad C_2\alpha \quad K_a = \frac{(C_2\alpha + C_1)C_2\alpha}{C_2(1-\alpha)} \quad (\alpha \ll 1)$$

(The weak acids dissociation will be further suppressed because of presence of strong acid, common ion effect)

$$K_a = (C_2\alpha + C_1)\alpha$$

$$\text{Total } H^+ \text{ ion concentration} = C_1 + C_1\alpha$$

If the total $[H^+]$ from the acid is greater than 10^{-6} M, then contribution from the water can be neglected, if not then we have to take $[H^+]$ from the water also.

◆ **pH of a mixture of two weak acid(both monoprotic) solutions :**

Let the acid are HA_1 & HA_2 and their **final** concentrations are C_1 & C_2 respectively, then

$$[H^+] = C_1\alpha_1 + C_2\alpha_2 = \frac{C_1K_{a1}}{\sqrt{C_1K_{a1} + C_2K_{a2}}} + \frac{C_2K_{a2}}{\sqrt{C_1K_{a1} + C_2K_{a2}}} \Rightarrow [H^+] = \sqrt{C_1K_{a1} + C_2K_{a2}}$$

If the dissociation constant of one of the acid is very much greater than that of the second acid then contribution from the second acid can be neglected.

$$\text{So, } [H^+] = C_1\alpha_1 + C_2\alpha_2 \approx C_1\alpha_1$$

◆ **pH of a mixture of a weak acid/ weak base with weak/strong base/acid respectively.**

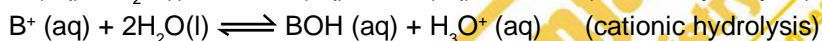
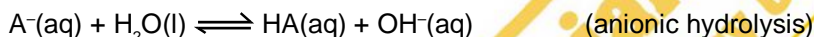
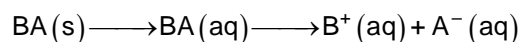
For this type of mixtures there can be two cases in general,

- (i) if the acids and bases are mixed in equal amounts (equivalents)
- (ii) if the acids and bases are mixed in different amounts (equivalents)

First case will lead to phenomenon of **SALT HYDROLYSIS** and second case will lead to formation of **BUFFER SOLUTIONS**.

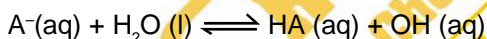
Salt Hydrolysis (Reverse of neutralisation)

Salt + Water \rightleftharpoons acid +base



ANIONIC HYDROLYSIS

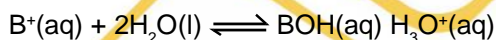
Anions can function as a base on reaction with water and hydrolyse as follows :



The extent of hydrolysis of a given anion depends on its basic strength

CATIONIC HYDROLYSIS

Cations can function as acid on reaction with water and hydrolyse as follows.



The extent of hydrolysis of a given cation depends on its acidic strength.

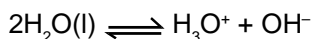
There are four types of salt.

- ◆ Salt of strong acid and strong base
- ◆ Salt of strong acid and weak base
- ◆ Salt of weak acid and strong base
- ◆ Salt of weak acid and weak base

Salts of first type does not undergo hydrolysis and rest three types undergo hydrolysis.

◆ **Salt of strong acid and strong base**

Neither of the ions will undergo hydrolysis so the solution contain only the equilibrium of ionization of water.

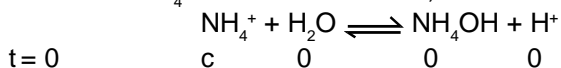


Thus, **the pH of solution will be 7(neutral solution).**

◆ **Salt of strong acid and weak base**

The examples can be $NH_4Cl, (NH_4)_2SO_4, AlCl_3$

Only the cation will undergo hydrolysis and the solution will be acidic in nature. for example in the solution of NH_4Cl of concentration c , we will have

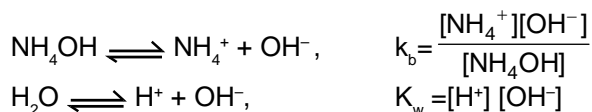


t = 0	c	0	0	0
at eq.	$c(1-h)$	ch	ch	

(h - degree of hydrolysis)



$$k_h = \frac{[\text{NH}_4\text{OH}][\text{H}^+]}{[\text{NH}_4^+]} = \text{called hydrolysis constant of the salt}$$



From above equations we can get

$$k_h \times k_b = k_w$$

$$k_h = \frac{ch \cdot ch}{c(1-h)} = \frac{ch^2}{(1-h)} \quad (\text{generally } 1-h \approx 1) \text{ so we get, } \Rightarrow h = \sqrt{\frac{k_h}{c}}$$

$$\Rightarrow [\text{H}^+] = ch = \sqrt{k_h \times c} =$$

$$\Rightarrow \text{pH} = -\log [\text{H}^+] = -\frac{1}{2} [\log k_w - \log k_b + \log c] = \frac{1}{2} [\text{pk}_w - \text{pk}_b - \log c]$$

◆ **Salt of weak acid and strong base**

The examples can be CH_3COONa , KCN , $\text{Na}_2\text{C}_2\text{O}_4$, K_3PO_4
Similar to above analysis we will get

$$k_h = \frac{ch \cdot ch}{c(1-h)} = \frac{ch^2}{(1-h)} \quad (\text{generally } 1-h \approx 1) \text{ so we get, } \Rightarrow h = \sqrt{\frac{k_h}{c}}$$

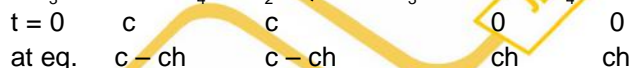
$$k_h \times k_a = k_w$$

$$[\text{OH}^-] = ch = \sqrt{k_h \times c}$$

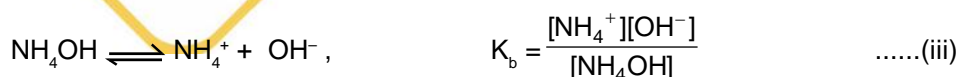
$$\text{pH} = -\log [\text{H}^+] = -\frac{1}{2} [\log k_w + \log k_a - \log c] = \frac{1}{2} [\text{pk}_w + \text{pk}_a + \log c]$$

◆ **Salt of weak acid and weak base**

Examples can include $\text{CH}_3\text{COONH}_4$, NH_4CN , $\text{Ca}(\text{CH}_3\text{COO})_2$, MgC_2O_4
 $\text{CH}_3\text{COO}^- + \text{NH}_4^+ + \text{H}_2\text{O} \rightleftharpoons \text{CH}_3\text{COOH} + \text{NH}_4\text{OH}$



$$K_h = \frac{[\text{CH}_3\text{COOH}][\text{NH}_4\text{OH}]}{[\text{CH}_3\text{COO}^-][\text{NH}_4^+]} \quad \dots(i)$$



$$\text{So, } k_h \times k_a \times k_b = k_w,$$

$$\Rightarrow k_h = \frac{ch \cdot ch}{c(1-h) \cdot c(1-h)} = \left(\frac{h}{1-h}\right)^2 \quad \Rightarrow \left(\frac{h}{1-h}\right) = \sqrt{k_h}$$

from (ii) equation

$$[\text{H}^+] = k_a \frac{[\text{CH}_3\text{COOH}]}{[\text{CH}_3\text{COO}^-]} = k_a \frac{ch}{c(1-h)} = k_a \times \frac{h}{1-h} = k_a \times \sqrt{k_h} = k_a \times \sqrt{\frac{k_w}{k_a \times k_b}} = \sqrt{\frac{k_w \times k_a}{k_b}}$$

$$\text{pH} = -\log [\text{H}^+] = \frac{1}{2} [\text{pk}_w + \text{pk}_a - \text{pk}_b]$$



Buffer Solution

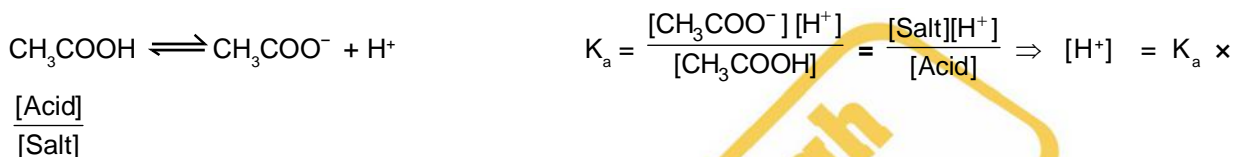
(if the acids and bases are mixed in different amounts (equivalents))

Buffer solutions are those, which resist a change in pH upon addition of small amount of small amount of acid or base. this does not mean that the pH will not change, and all it means is that the pH change would be less than the change that would have occurred had it not been a buffer.

There are various types of buffers :

- (i) Buffer of weak acid and its salt with a strong base ;
- (ii) Buffer of a weak base and its salt with a strong acid.
- (iii) The solution of the salt of a weak acid and a weak base.

To calculate the pH of a buffer solution made up of a weak acid and its salt with a strong base. We have



Taking log of both sides. $\log [\text{H}^+] = \log K_a + \log \frac{[\text{Acid}]}{[\text{Salt}]}$

$$\therefore \text{pH} = \text{p}K_a + \log \frac{[\text{Salt}]}{[\text{Acid}]}$$

This is known as the **Henderson's equation of a buffer.**

For a buffer made up of weak base and its salt with a strong acid the Henderson's equation looks like this

$$\text{pOH} = \text{p}K_b + \log \frac{[\text{Salt}]}{[\text{Base}]}$$

Buffer Capacity :

It is defined as the amount of a strong acid or strong base required to change the pH of a buffer by one unit.

$$\frac{dx}{d\text{pH}} = 2.303 \frac{(a+x)(b-x)}{a+b}$$

This is defined as buffer capacity. where a = conc of acid; b = conc of salt

and x = conc of H⁺ added. It is the ratio of the small amount of acid or base added to the change in pH caused in the buffer.

Maximum buffer capacity :

It can be proved that the maximum buffer is achieved when the salt and acid or base concentration is equal.

$$\therefore b = a$$

The buffer shows maximum buffer capacity when the amount of acid or base and the salt are same.

Solubility product (K_{sp}) is a type of equilibrium constant, so will be dependent only on temperature for a particular salt.

Following examples will illustrate the different type of solubilities and the effects of different factors or situations on solubility of a salt.

◆ Simple Solubility

Let the salt is A_xB_y , in solution in water, let the solubility in H₂O = 's' M, then





◆ **Effect of common ions on solubility :**

Because of the presence of common ion the solubility of the salt decreases.

Que. Calculate solubility of silver oxalate in 10^{-2} M potassium oxalate solution. Given that K_{sp} of silver oxalate = 10^{-10} .

Sol. $[C_2O_4^{2-}] = 10^{-2} + x$, where x is the solubility of silver oxalate, this can be neglected in comparison to 10^{-2} .

$$\text{So, } K_{sp} = 10^{-10} = 10^{-2} \times (2x)^2 \quad \Rightarrow \quad \frac{10^{-8}}{2 \times 2} = x^2 \quad \Rightarrow \quad x = 5 \times 10^{-5}$$

◆ **Calculation of simultaneous Solubility**

Que. Calculate simultaneous solubility of silverthiocyanate and silver bromide in water given that K_{sp} of silver thiocyanate = 10^{-12} and K_{sp} of silver bromide = 5×10^{-13} respectively.

Sol. Let the solubility of AgSCN be x and that of AgBr is y, then



$$10^{-12} = x(x + y) \qquad \text{-----(i)}$$

$$5 \times 10^{-13} = y(x + y) \qquad \text{-----(ii)}$$

On solving we get, $x = 2y$

$$\text{So, } y = 4.08 \times 10^{-7} \text{ and } x = 8.16 \times 10^{-7}$$

◆ **Condition of precipitation**

For precipitation ionic product should be greater than K_{sp}

Que. You are given 10^{-5} M NaCl solution and 10^{-8} M $AgNO_3$ solution, they are mixed in 1 : 1 volume ratio, predict whether AgCl will be precipitated or not, if solubility product of AgCl is = 10^{-10} mole per litre.

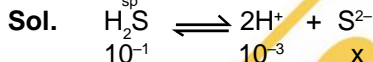
Sol. Ionic product = $\frac{10^{-5}}{2} \times \frac{10^{-8}}{2} = 25 \times 10^{-15} < K_{sp}$

Hence no precipitation will take place.

◆ **Selective Precipitation**

Que. 2×10^{-4} moles of Mn^{2+} and Cu^{2+} each is present in one litre solution of 10^{-3} M $HClO_4$, which is saturated with H_2S . Whether or not each of the ions will be precipitated? Given that concentration of H_2S in its saturated solution is = 0.1 M, net dissociation constant of H_2S = 10^{-21} ,

K_{sp} for MnS = 2.5×10^{-10} and K_{sp} for (CuS) = 8.5×10^{-36}



$$K = \frac{x \times (10^{-3})^2}{[H_2S]} = 10^{-21} \quad \Rightarrow \quad [S^{2-}] = 10^{-16} \text{ M}$$

Ionic Product of CuS = $[Cu^{2+}][S^{2-}] = 2 \times 10^{-4} \times 10^{-16} = 2 \times 10^{-20} > K_{sp}$ of CuS

Ionic Product of MnS = $2 \times 10^{-20} < K_{sp}$ of MnS

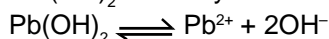
CuS will be precipitated

◆ **Solubility in appropriate buffer solutions.**

Appropriate buffer means that the components of buffer should not interfere with the salt or only H^+ or OH^- ions should be interacting with the ions of the salt.

Que. The Solubility of $Pb(OH)_2$ in water is 6.7×10^{-6} M. Calculate its solubility in buffer solution of pH = 8

Sol. $Pb(OH)_2$ solubility in water = 6.7×10^{-6} , $K_{sp} = 4s^3 = 4 \times 300 \times 10^{-18} = 1.2 \times 10^{-15}$



Let $[Pb^{2+}] = x$, then $K_{sp} = x [OH^-]^2$

$$1.2 \times 10^{-15} = x \times 10^{-12} \quad \Rightarrow \quad x = 1.2 \times 10^{-3} \text{ M}$$



TOPIC WISE MCQS

Arrhenius Ionization Theory, Ostwald dilution Law, Ionization Constant

- Q.1** For a binary weak electrolyte, the degree of dissociation is proportional to the-
- (A) Dilution
(B) Square root of dilution
(C) Concentration
(D) Square root of concentration
- Q.2** The ionisation constant of a tribasic acid is K_a . If its first, second and third ionisation constants are K_{a1} , K_{a2} and K_{a3} respectively then -
- (A) $K_a = K_{a1} \times K_{a2} \times K_{a3}$
(B) $K_a = \frac{K_{a1}}{K_{a2} \times K_{a3}}$
(C) $K_{a2} = \frac{K_{a1} \times K_a}{K_{a3}}$
(D) None of these
- Q.3** The dissociation constant of two weak acids are K_{a1} and K_{a2} respectively. Their relative strength is -
- (A) $\frac{K_{a2}}{K_{a1}}$ (B) $\left\{ \frac{K_{a1}}{K_{a2}} \right\}^{\frac{1}{2}}$
(C) $\frac{K_{a1}}{K_{a2}}$ (D) $\left\{ \frac{K_{a1} \times K_{a2}}{K_{a1}} \right\}^{\frac{1}{2}}$
- Q.4** K_{a1} , K_{a2} and K_{a3} are the three dissociation constants of H_3PO_4 . Which of the following is a correct order -
- (A) $K_{a1} > K_{a2} < K_{a3}$ (B) $K_{a1} > K_{a2} > K_{a3}$
(C) $K_{a3} > K_{a2} > K_{a1}$ (D) None of these
- Q.5** K_1 and K_2 are the first and second dissociation constants of oxalic acid. Choose the correct statement -
- (A) K_2 is equal to K_1
(B) K_2 is equal to 1000 times K_1
(C) K_2 is equal to $\frac{1}{1000}$ times K_1
(D) All are correct
- Q.6** Ostwald's dilution law gives satisfactory results with the solution of the electrolyte -
- (A) HCl (B) HNO_3
(C) CH_3COOH (D) NaOH

- Q.7** A monoprotic acid in a 0.1 M solution ionizes to 0.001%. Its ionization constant is -
- (A) 1.0×10^{-3} (B) 1.0×10^{-6}
(C) 1.0×10^{-8} (D) 1.0×10^{-11}
- Q.8** For two weak acids A and B, the ratio of their percent ionization is 4 : 9. The ratio of their K_a would be -
- (A) 4 : 9 (B) 2 : 3
(C) 16 : 81 (D) 3 : 2
- Q.9** Ostwald's dilution law for a weak acid HA may be given as -
- (A) $K_a = \frac{\alpha \cdot c}{(1-\alpha)c}$ (B) $K_a = \frac{\alpha^2 \cdot c}{(1-\alpha)}$
(C) $K_a = \left(\frac{\alpha^2}{(1-\alpha)} \cdot c \right)^2$ (D) $K_a = \frac{\alpha^2 \cdot c}{1-\alpha^2}$
- Q.10** In the aqueous solution of $MgCl_2$ -
- (A) The number of Mg^{2+} and Cl^- ions are equal
(B) The number of Mg^{2+} ions is half that of Cl^- ions
(C) The number of Mg^{2+} ions is double that of Cl^- ions
(D) None of these

Ionic Product of water, pH & pOH

- Q.11** The concentration of a monoprotic acid is C moles L^{-1} and ionisation constant K_a . The pH of the solution is -
- (A) $\frac{1}{2} pK_a - \frac{1}{2} \log C$
(B) $(K_a \times C)^{\frac{1}{2}}$
(C) $\frac{1}{2} pK_a + \frac{1}{2} \log C$
(D) $\alpha \cdot C$
- Q.12** Select the correct statement -
- (A) If $[H^+] = y \times 10^{-x}$ then $pH = x - \log y$
(B) If $[H^+] = \frac{1}{y} \times 10^{-x}$ then $pH = x + \log y$
(C) pH of a solution = $14 + \log [OH^-]$
(D) All of the above
- Q.13** The pK_w of water at $50^\circ C$ is 13.40. An aqueous solution at $50^\circ C$ has $pH = 7$. This solution is -
- (A) Acidic (B) alkaline
(C) Neutral (D) Amphoteric



- Q.14** The value of pK_w at 25°C is -
(A) 7 (B) -14
(C) 14 (D) 1×10^{-14}
- Q.15** K_w for water at 25°C is equal to 10^{-14} . What is its value at 90°C -
(A) 10^{-15} (B) 10^{-17}
(C) 10^{-14} (D) 10^{-12}
- Q.16** Pure water ionises as
 $2\text{H}_2\text{O} (\ell) \rightleftharpoons \text{H}_3\text{O}^+ (\text{aq.}) + \text{OH}^- (\text{aq.})$
At 25°C the pH of pure water is approximately 7.0. At 37°C its pH is -
(A) More than 7.0 (B) Less than 7.0
(C) Equal to 7.0 (D) None of these
- Q.17** The aqueous solution whose pH = 0 is -
(A) Acidic (B) Alkaline
(C) Amphoteric (D) Neutral
- Q.18** The formula to calculate pK_a for acetic acid is -
(A) $\text{pH} + \frac{1}{2} \log C$ (B) $2\text{pH} + \log C$
(C) $\text{pH} - \frac{1}{2} \log C$ (D) $2 [\text{pH} + \log C]$
- Q.19** For an acid solution the $[\text{OH}^-]$ is -
(A) $> 10^{-7}$ (B) $< 10^{-7}$
(C) 10^{-14} (D) 10^{-7}
- Q.20** The pH of two equimolar weak acids are 3.0 and 5.0 respectively. Their relative strength is -
(A) 3 : 5 (B) 5 : 3
(C) 100 : 1 (D) 1 : 100
- Q.21** 4.0 g of NaOH and 4.9 g of H_2SO_4 are dissolved in water and volume is made upto 250 ml. The pH of this solution is -
(A) 7.0 (B) 1.0
(C) 2.0 (D) 12.0
- Q.22** The pH of two solutions are 5 and 3 respectively. What will be the pH of the solution made by mixing the equal volumes of the above solutions -
(A) 3.5 (B) 4.5
(C) 3.3 (D) 4.0
- Salt Hydrolysis, Buffer Solution**
- Q.23** An aqueous solution of aluminium sulphate would show -
(A) An acidic reaction
(B) A neutral reaction
(C) A basic reaction
(D) Both acidic and basic reactions
- Q.24** Which of the following is an acidic salt -
(A) Na_2SO_4 (B) $\text{Ca}(\text{OH})\text{Cl}$
(C) $\text{Pb}(\text{OH})\text{Cl}$ (D) Na_2HPO_4
- Q.25** Select the correct combination -
(A) The aqueous solutions of each Na_3BO_3 and Na_3PO_4 - Acidic nature
(B) The aqueous solutions of each Na_3BO_3 and CH_3COONa - Basic nature
(C) The aqueous solutions of each CH_3COONa and NaCN - Acidic nature
(D) The aqueous solutions of each Na_3PO_4 and NH_4Cl - Acidic nature
- Q.26** Expression $\text{p}K_h = \text{p}K_w - \text{p}K_a - \text{p}K_b$ is not applicable to -
(A) Ammonium acetate
(B) Ammonium cyanide
(C) Aniline acetate
(D) Ammonium chloride
- Q.27** What is the pH of an aqueous solution of ammonium acetate ($K_a = K_b = 1.8 \times 10^{-5}$) -
(A) > 7 (B) 7.0 (C) < 7.0 (D) Zero
- Q.28** If $\text{p}K_b > \text{p}K_a$ then the solution of the salt of weak acid and weak base will be -
(A) Neutral (B) Acidic
(C) Basic (D) Amphoteric
- Q.29** $\text{pOH} = 7 - 0.5 \text{p}K_a + 0.5 \text{p}K_b$ is true for which pair of cation and anion.
(A) $\text{C}_6\text{H}_5\text{NH}_3^+$, CH_3COO^- (B) Na^+ , CN^-
(C) Al^{3+} , Cl^- (D) NH_4^+ , NO_3^-
- Q.30** The pH of buffer of $\text{NH}_4\text{OH} + \text{NH}_4\text{Cl}$ - type is given by -
(A) $\text{pH} = \text{p}K_b$
(B) $\text{pH} = \frac{1}{2}\text{p}K_b - \frac{1}{2} \log [\text{salt}]/[\text{base}]$
(C) $\text{pH} = 14 - \text{p}K_b - \log [\text{salt}]/[\text{base}]$
(D) $\text{pH} = \text{pOH} - \text{p}K_b + [\text{salt}]/[\text{base}]$
- Q.31** For a salt of weak acid and weak base $[\text{p}K_a - \text{p}K_b]$ would be equal to -
(A) $2 \text{pH} + \text{p}K_w$ (B) $2 \text{pH} - \log 10^{-14}$
(C) $2 \text{pH} - \text{p}K_w$ (D) None of these
- Q.32** A basic buffer solution can be prepared by mixing the solution of -
(A) Ammonium chloride and Ammonium acetate
(B) Ammonium acetate and acetic acid
(C) Ammonium chloride and ammonium hydroxide
(D) Ammonium cyanide and Ammonium hydroxide



- Q.33** Which of the following compound forms an aqueous solution which is acidic when compared with water –
 (A) NaOH (B) K_2CO_3
 (C) $BaCl_2$ (D) $Al_2(SO_4)_3$
- Q.34** The solution of blue vitriol in water is acidic because –
 (A) $CuSO_4$ reacts with water
 (B) Cu^{2+} reacts with water
 (C) SO_4^{2-} reacts with water
 (D) $CuSO_4$ removes OH^- ions from water
- Q.35** The salt of which of the following four weak acids will be most hydrolysed –
 (A) HA ; $K_a = 1 \times 10^{-8}$
 (B) HB ; $K_a = 2 \times 10^{-6}$
 (C) HC ; $K_a = 3 \times 10^{-8}$
 (D) HD ; $K_a = 4 \times 10^{-10}$
- Q.36** Formula for degree of hydrolysis 'h'; $h = [10^{-7} (K_a K_b)^{-1/2}]$ is applicable to the salt –
 (A) NH_4CN (B) $(NH_4)_2SO_4$
 (C) NH_4Cl (D) NH_4NO_3
- Q.37** $[H^+] = \sqrt{\frac{K_w K_a}{C}}$ is suitable for –
 (A) NaCl, NH_4Cl
 (B) CH_3COONa , NaCN
 (C) CH_3COONa , $(NH_4)_2SO_4$
 (D) CH_3COONH_4 , $(NH_4)_2CO_3$
- Q.38** Which one of the following mixture does not act as a buffer solution-
 (A) Boric acid and borax
 (B) Sodium phosphate & disodium hydrogen phosphate
 (C) Sodium propionate and propionic acid
 (D) Sod. acetate and sodium propionate
- Q.39** In the neutralization process of H_3PO_4 and NaOH, the number of buffers formed will be
 (A) 3 (B) 1 (C) 2 (D) 4
- Q.40** Addition of sodium acetate solution to acetic acid causes the following change–
 (A) pH increases (B) pH decreases
 (C) pH remains unchanged
 (D) pH becomes 7
- Q.41** H^+ ion concentration of water does not change by adding –
 (A) CH_3COONa (B) $NaNO_3$
 (C) NaCN (D) Na_2CO_3
- Q.42** $H_2CO_3 + NaHCO_3$ found in blood helps in maintaining pH of the blood close to 7.4 . An excess of acid entering the blood stream is removed by-
 (A) HCO_3^- (B) H_2CO_3 (C) H^+ ion (D) CO_3^{2-} ion

Solubility Product

- Q.43** If s is the molar solubility of Ag_2SO_4 , then –
 (A) $3 [Ag^+] = s$ (B) $[Ag^+] = s$
 (C) $[2Ag^+] = s$ (D) $[SO_4^{2-}] = s$
- Q.44** Which of the following would increase the solubility of $Pb(OH)_2$ –
 (A) Add hydrochloric acid
 (B) Add a solution of $Pb(NO_3)_2$
 (C) Add a solution of NaOH
 (D) None of the above—the solubility of a compound is constant at constant temperature
- Q.45** The aqueous solution of which of the following sulphides would contain maximum concentration of S^{2-} ions–
 (A) MnS ($K_{sp} = 1.1 \times 10^{-21}$)
 (B) ZnS ($K_{sp} = 1.1 \times 10^{-23}$)
 (C) PbS ($K_{sp} = 1.1 \times 10^{-35}$)
 (D) CuS ($K_{sp} = 1.1 \times 10^{-30}$)
- Q.46** Which of the following salts has maximum solubility –
 (A) HgS, $K_{sp} = 1.6 \times 10^{-54}$
 (B) $PbSO_4$, $K_{sp} = 1.3 \times 10^{-8}$
 (C) ZnS, $K_{sp} = 7.0 \times 10^{-26}$
 (D) AgCl, $K_{sp} = 1.7 \times 10^{-10}$
- Q.47** The necessary condition for saturated solution is –
 (A) Product of ionic concentrations = Solubility product
 (B) Product of ionic concentrations < solubility product
 (C) Product of ionic concentrations > solubility product
 (D) None of the above
- Q.48** At $30^\circ C$, the solubility of Ag_2CO_3 ($K_{sp} = 8 \times 10^{-2}$) will be maximum in –
 (A) 0.05 M Na_2CO_3 (B) 0.05 M $AgNO_3$
 (C) Pure water (D) 0.05 NH_3
- Q.49** Which of the following expressions shows the saturated solution of $PbSO_4$ –
 (A) $K_{sp} (PbSO_4) = [Pb^{2+}] [SO_4^{2-}]$
 (B) $K_{sp} (PbSO_4) > [Pb^{2+}] [SO_4^{2-}]$
 (C) $K_{sp} (PbSO_4) = [Pb^+] [SO_4^-]$
 (D) $K_{sp} (PbSO_4) < [Pb^{2+}] [SO_4^{2-}]$
- Q.50** The correct relation between K_{sp} and solubility for the salt $KAl(SO_4)_2$ is
 (A) $4s^3$ (B) $4s^4$ (C) $27s^4$ (D) None



MISCELLANEOUS QUESTIONS

- Q.1** Find the percentage ionisation of 0.2 M acetic acid solution, whose dissociation constant is 1.8×10^{-5}
(A) 0.198 (B) 0.290
(C) 0.950 (D) None of these
- Q.2** What will be the hydrogen ion concentration (moles L^{-1}) of 0.01 M HCN solution if it is 20% ionised
(A) 0.002 (B) 0.02 (C) 0.2 (D) 0.1
- Q.3** The dissociation constant of a weak acid is 1.0×10^{-4} . The equilibrium constant of its reaction with strong base is –
(A) 1.0×10^{-4} (B) 1.0×10^{-10}
(C) 1.0×10^{10} (D) 1.0×10^{-14}
- Q.4** The $[H^+]$ of a solution is 0.03 M. The pOH of this solution is –
(A) 12.48 (B) 12.52 (C) 12.54 (D) 12.58
- Q.5** The pH of a solution is 6.0. In this solution –
(A) $[H^+] = 100 [OH^-]$ (B) $[H^+] = 10 [OH^-]$
(C) $[H^+] = [OH^-]$ (D) $[H^+] = \frac{1}{10} [OH^-]$
- Q.6** At 298 K, the ratio of number of pure water molecules to number of hydroxyl ions is –
(A) 1.8×10^{-9} (B) 5.55×10^8
(C) 10^7 (D) 6.02×10^{23}
- Q.7** A sufficient quantity of acid is added to change its pH from 5 to 2. Its hydrogen ion concentration is increased by –
(A) 100 times (B) 1000 times
(C) 2.50 times (D) 5 times
- Q.8** A 0.01 M acetic acid solution is 1.0% ionised. An another acetic acid is 10% ionised. What will be the concentration of another acetic acid –
(A) 0.001 M (B) 0.0001 M
(C) 0.01 M (D) 0.1 M
- Q.9** For a 100 ml solution of 10^{-2} M NaOH the ratio pH: pOH would be –
(A) 6 : 1 (B) 1 : 6
(C) 2 : 1 (D) $10^{10} : 1$
- Q.10** How many moles of HCl must be removed from 1 litre of aqueous HCl solution to change its pH from 2 to 3 –
(A) 1 (B) 0.02 (C) 0.009 (D) 0.01
- Q.11** 0.01 M Acetic acid is 12.5 % dissociated its pH will be –
(A) 4.509 (B) 3.723
(C) 2.903 (D) 5.623
- Q.12** 10^{-2} mole of KOH is dissolved in 10 litres of water. The pH of the solution is –
(A) 12 (B) 2 (C) 3 (D) 11
- Q.13** % hydrolysis of 0.1M CH_3COONH_4 , when $K_a = K_b = 1.8 \times 10^{-5}$ is –
(A) 0.55 (B) 7.63
(C) 0.55×10^{-2} (D) 7.63×10^{-3}
- Q.14** Given the two concentration of HCN are 0.1 M & 0.001 M respectively. What will be the ratio of degree of dissociation –
(A) 1 (B) 0.1
(C) 0.003 (D) 0.01
- Q.15** On hydrolysis of sodium carbonate, the reaction takes place between –
(A) Na^+ and water (B) Na^+ and OH^-
(C) CO_3^{2-} and water (D) CO_3^{2-} and H^+
- Q.16** The pH of 0.001M sodium acetate solution is $[K_a(CH_3COOH) = 1.8 \times 10^{-5}]$ –
(A) ≈ 11 (B) ≈ 6.5
(C) ≈ 14 (D) ≈ 8.0
- Q.17** The pH of a buffer solution containing 0.1 mole of acetic acid and 0.15 mole of sodium acetate is (K_a for acetic acid = 1.75×10^{-5})–
(A) 4.9 (B) 3.0
(C) 4.2 (D) 5.4
- Q.18** A certain buffer solution contains equal concentration of X^- and HX . The K_b for X^- is 1×10^{-10} . The pH of the buffer is–
(A) 4 (B) 7 (C) 10 (D) 14
- Q.19** In a buffer solution of a weak acid and its salt, if the ratio of concentration of salt to acid is raised 10 times then pH of the solution will–
(A) Increase ten times
(B) Decrease by one unit
(C) Decrease ten times
(D) Increase by one unit
- Q.20** 500 ml of 0.2 M acetic acid are added to 500 ml of 0.30 M sodium acetate solution. If the dissociation constant of acetic acid is 1.5×10^{-5} then pH of the resulting solution is –
(A) 5.0 (B) 9.0 (C) 3.0 (D) 4.0
- Q.21** The pOH of a basic buffer (e.g. NH_4OH/NH_4Cl) is 5. If the concentration of the salt is tripled whereas that of base remains same. What is the new value of pOH (Given $\log 3 \approx 0.48$)
(A) 4.52 (B) 5.48
(C) 6.48 (D) 3.52



- Q.22** Let the solubility of AgCl in water, in 0.01 M CaCl₂, in 0.01 M NaCl and in 0.05 M AgNO₃ be s₁, s₂, s₃ and s₄ respectively. Which of the following relations between these quantities is correct –
(A) s₁ > s₂ > s₃ > s₄ (B) s₁ > s₂ = s₃ > s₄
(C) s₄ > s₂ > s₃ > s₁ (D) s₁ > s₃ > s₂ > s₄
- Q.23** K_{sp} of AgCl is 1 × 10⁻¹⁰. Its solubility in 0.1 M KNO₃ will be –
(A) 10⁻⁵ moles/litre (B) > 10⁻⁵ moles/litre
(C) < 10⁻⁵ moles/litre (D) None of these
- Q.24** At 298 K, how many milligrams of silver bromide can be dissolved in 20 litres of water – [K_{sp}(AgBr) = 5.0 × 10⁻¹³]
(Atomic wt. Ag = 108, Br = 80)
(A) 2.66 (B) 3.66
(C) 4.66 (D) None of these
- Q.25** At 25°C what will be the solubility of silver carbonate in 0.1 M Na₂CO₃ solution. At this temperature K_{sp} of silver carbonate is 4 × 10⁻¹³ –
(A) 2 × 10⁻⁷ (B) 2 × 10⁻⁶
(C) 10⁻⁶ (D) 10⁻⁷
- Q.26** When equal volumes of the following solutions are mixed, precipitation of CaF₂ (K_{sp} = 1.7 × 10⁻¹⁰) will occur only with –
(A) 10⁻⁴ M Ca²⁺ and 10⁻⁴ M F⁻
(B) 10⁻² M Ca²⁺ and 10⁻³ M F⁻
(C) 10⁻⁵ M Ca²⁺ and 10⁻³ M F⁻
(D) 10⁻³ M Ca²⁺ and 10⁻⁵ M F⁻
- Q.27** At 25°C, the solubility product of Ca(OH)₂ is 32 × 10⁻¹². What will be the pOH of its saturated solution at this temperature –
(A) 3.4990 (B) 3.3980
(C) 0.3010 (D) None of these
- Q.28** In the hydrolysis of sodium acetate –
(A) Anions of the salt are hydrolysed
(B) Cations of the salt are hydrolysed
(C) Both of the above ions are not hydrolysed
(D) None of these
- Q.29** When HCl gas is passed through a saturated solution of common salt, pure NaCl is precipitated because –
(A) The ionic product [Na⁺] and [Cl⁻] exceeds the solubility product of NaCl
(B) The impurities dissolve in HCl
(C) HCl is highly soluble in H₂O
(D) The solubility product of NaCl is lowered by the Cl⁻ ions from aqueous HCl
- Q.30** At 298K, the solubility of PbCl₂ is 6.3 × 10⁻³ moles L⁻¹. Its solubility product at this temperature is –
(A) (6.3 × 10⁻³) × (6.3 × 10⁻³)
(B) (6.3 × 10⁻³) × (12.6 × 10⁻³)
(C) (6.3 × 10⁻³) × (12.6 × 10⁻³)²
(D) (12.6 × 10⁻³) × (12.6 × 10⁻³)
- Q.31** Given : Heat of ionization of two acids :
ΔH⁰ (HCN) = 45.2 kJ mol⁻¹
ΔH⁰ (CH₃COOH) = 2.1 kJ mol⁻¹
which relationship for the two acids is true –
(A) pK_a (HCN) = pK_a (CH₃COOH)
(B) pK_a (HCN) > pK_a (CH₃COOH)
(C) pK_a (HCN) < pK_a (CH₃COOH)
(D) pK_a(HCN) = $\frac{45.2}{2.1}$ pK_a(CH₃COOH)
- Q.32** Which of the following is the ionisation constant of 0.01 M aniline (0.02% ionised)–
(A) 4.0 × 10⁻⁴ (B) 4.0 × 10⁻⁵
(C) 4.0 × 10⁻⁹ (D) 4.0 × 10⁻¹⁰
- Q.33** The dissociation constants of monobasic acids A, B, C and D are 6 × 10⁻⁴, 5 × 10⁻⁵, 3.6 × 10⁻⁶, and 7 × 10⁻¹⁰ respectively. The pH values of their 0.1 molar aqueous solutions are in the order –
(A) A < B < C < D (B) A > B > C > D
(C) A = B = C = D (D) A > B < C > D
- Q.34** In a saturated solution of AgCl (K_{sp} = 1.6 × 10⁻¹⁰ at 25° C), the [Ag⁺] = 1.3 × 10⁻⁵ mol/L. Then enough potassium chloride is added to this solution so that [Cl⁻] = 0.020 M. The solubility of AgCl in this solution of potassium chloride is –
(A) 3.2 × 10⁻⁸ mol/L (B) 8 × 10⁻¹¹ mol/L
(C) 8 × 10⁻⁸ mol/L (D) 8 × 10⁻⁹ mol/L
- Q.35** What will be the pH of 10⁻³ M monobasic weak acid solution if its dissociation constant is 1.8 × 10⁻⁵ –
(A) 4 (B) 3.872
(C) 2.52 (D) 2.00
- Q.36** Calculate the solubility of silver phosphate (Ag₃PO₄) in 0.10M AgNO₃ ?
K_{sp} = 1.1 × 10⁻¹⁶ –
(A) 0.10 M (B) 4.1 × 10⁻¹⁵ M
(C) 1.1 × 10⁻¹⁵ M (D) 1.1 × 10⁻¹³ M
- Q.37** K_{sp} for A₂B salt is 4 × 10⁻⁹. Evaluate its solubility –
(A) 4 × 10⁻² M (B) 2 × 10⁻⁴ M
(C) 1 × 10⁻⁴ M (D) 1 × 10⁻³ M



- Q.38** K_{sp} for $Pb(IO_3)_2$ is 3.2×10^{-14} at a certain temperature. Its solubility will be –
(A) 3.6×10^{-7} (B) 2×10^{-5}
(C) 4×10^{-5} (D) 8×10^{-5}
- Q.39** Equal volumes of 0.005 M and 0.005 N H_2SO_4 solutions are mixed together. The pH of resulting solution is –
(A) 5.0 (B) 3.125 (C) 2.10 (D) 2.125
- Q.40** Solubility of $AgCl$ [$K_{sp} = 1 \times 10^{-10}$] in 0.2 M $AgNO_3$ and 0.2 M $NaCl$ solution will be respectively –
(A) 5×10^{-10} M and 2×10^{-5} M
(B) 5×10^{-10} M and 5×10^{-10} M
(C) 2×10^{-5} M and 2×10^{-5} M
(D) 5×10^{-5} M and 5×10^{-5} M
- Q.41** Phenoxide ion ($C_6H_5O^-$) is a weak base, with $K_b = 7.7 \times 10^{-5}$. Calculate the pH of a 0.20 M solution of $C_6H_5O^-$ –
(A) 11.6 (B) 3.9×10^{-3} (C) 9.2 (D) 8.9
- Q.42** Half of the formic acid solution is neutralised on addition of a KOH solution to it. If $K_a(HCOOH) = 2 \times 10^{-4}$ then p^H of the solution is – ($\log 2 = 0.3010$)
(A) 3.6990 (B) 10.3010
(C) 3.85 (D) 4.3010
- Q.43** 100 ml of 1.0 N CH_3COOH are added to 100 ml of 1.0 N $NaOH$ solution. What will be the p^H of resulting solution –
(A) 7.0 (B) > 7.0 (C) < 7.0 (D) Zero
- Q.44** A solution contains 0.2 M NH_4OH and 0.2 M NH_4Cl . If 1.0 ml of 0.001 M HCl is added to it. What will be the $[OH^-]$ of the resulting solution ($K_b = 2 \times 10^{-5}$)
(A) 2×10^{-5} (B) 5×10^{-10}
(C) 2×10^{-3} (D) None of these
- Q.45** The concentration of a HCl solution is 10^{-2} N. If this solution is diluted ten times then its p^H will –
(A) Become ten times
(B) Become double
(C) Increase by one unit
(D) Decrease by one unit
- Q.46** What is the pH of the solution at half neutralization in the titration of 0.1N CH_3COOH and 0.1N KOH : ($K_a = 1.8 \times 10^{-5}$)–
(A) 4.75 (B) 1
(C) 13 (D) Zero
- Q.47** Acetic acid ($K_a = 2 \times 10^{-5}$) and propionic acid ($K_a = 2 \times 10^{-6}$) have same concentrations. What is the ratio of their degree of ionisation
(A) 100 (B) $\sqrt{10}$
(C) 10 (D) 0.1
- Q.48** The pH of 1 litre solution containing 0.5M NH_4OH and 0.5 M NH_4Cl ($pK_b = 5$) –
(A) 5 (B) 9
(C) 5 ± 1 (D) 9 ± 1
- Q.49** 0.05 M ammonium hydroxide solution is dissolved in 0.001 M ammonium chloride solution. What will be the OH^- ion concentration of this solution –
 $K_b(NH_4OH) = 1.8 \times 10^{-5}$
(A) 3.0×10^{-3} (B) 9.0×10^{-4}
(C) 9.0×10^{-3} (D) 3.0×10^{-4}
- Q.50** When 0.02 moles of $NaOH$ are added to a litre of buffer solution, its pH. changes from 5.75 to 5.80. What is its buffer capacity –
(A) 0.4 (B) 0.05
(C) -0.05 (D) 2.5



CHEMISTRY GURU
RANJAN SINGH
M.Sc.(Biochemistry), P.U.
— Ex-Faculty : Narayana IIT Academy & Goal Institute



ANSWER KEY

TOPIC WISE MCQS

Q.No.	1	2	3	4	5	6	7	8	9	10	11	12	13	14	15
Ans.	B	A	B	B	C	C	D	C	B	B	A	D	B	C	D
Q.No.	16	17	18	19	20	21	22	23	24	25	26	27	28	29	30
Ans.	B	A	B	B	C	A	C	A	D	B	D	B	B	A	C
Q.No.	31	32	33	34	35	36	37	38	39	40	41	42	43	44	45
Ans.	C	C	D	B	D	A	B	D	A	A	B	A	D	A	A
Q.No.	46	47	48	49	50										
Ans.	B	A	D	A	B										

MISCELLANEOUS QUESTIONS

Ques.	1	2	3	4	5	6	7	8	9	10	11	12	13	14	15
Ans.	C	A	C	A	A	B	B	B	A	C	C	D	A	B	C
Ques.	16	17	18	19	20	21	22	23	24	25	26	27	28	29	30
Ans.	D	A	A	D	A	B	D	A	A	C	B	B	A	A	C
Ques.	31	32	33	34	35	36	37	38	39	40	41	42	43	44	45
Ans.	B	D	A	D	B	D	D	B	D	B	A	A	B	A	C
Ques.	46	47	48	49	50										
Ans.	A	B	B	B	A										



**MEMORABLE YEARS
OF TOGETHERNESS**

Farewell Party

A warm hearted farewell ceremony is conducted every year to wish our students best of luck for their upcoming future and to glance over our memorable years of togetherness.

Our INFRASTRUCTURE

CLASSROOM



Individual interaction with Ranjan Singh Sir

DOUBT



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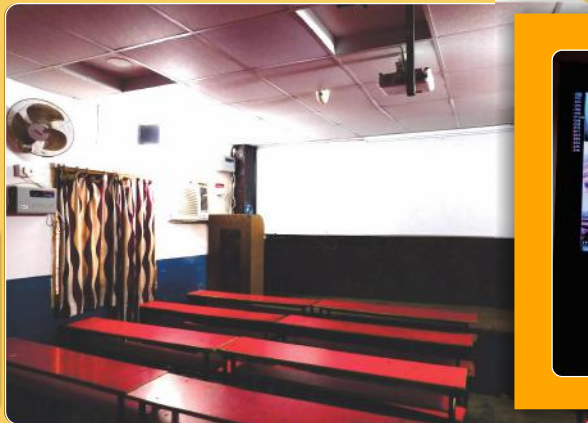
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PROJECTOR

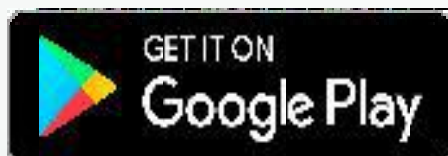


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
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