CLASS : CC -AD

Revision Worksheet Ionic Equilibrium

Single correct

Q.1 Three solution of strong electrolytes 25 ml of 0.1 M HX, 25 ml of 0.1 M H_2Y and 50 ml of 0.1 $NZ(OH)_2$ are mixed. What is the pOH of the solution ? [Given log 5 = 0.7]
(A) 1.6 (B) 7 (C) 11.6 (D) 12.6

Q.2 0.1 M NH_3 (aq) solution is titrated against 0.1 M HNO_3 (aq) solution. What would be the difference in pOH between $\frac{1}{4}$ and $\frac{3}{4}$ stages of neutralization of base?
(A) 2 log 3/4 (B) 2 log 1/4 (C) log 1/3 (D) 2 log 3

Q.3 What is the pH of 0.1M $NaHCO_3$? $K_1 = 5.0 \times 10^{-7}$, $K_2 = 7.0 \times 10^{-11}$ for carbonic acids.
(A) pH = 7 (B) pH much less than 7
(C) pH slightly less than 7 (D) pH greater than 7

Q.4 The degree of hydrolysis of 0.1 M solution of conjugate base of HA is 0.01. Find the $\mathbf{H^+}$ concentration in 0.4 M solution of $\mathbf{A^-}$.
(A) 5×10^{-11} M (B) 5×10^{-12} M (C) 2×10^{-3} M (D) 2×10^{-4} M

Q.5 What is the aq. ammonia concentration of a solution prepared by dissolving 0.15 mole of $NH_4^+CH_3COO^-$ in 1L H_2O.[$K_{a(CH_3COOH)} = 1.8\times10^{-5}$; $K_{b(NH_4OH)} = 1.8\times10^{-5}$]
(A) 8.3×10^{-4} (B) 0.15 (C) 6.4×10^{-4} (D) 3.8×10^{-4}

Q.6 An aqueous solution contains 0.01 M RNH_2 ($K_b = 2 \times 10^{-6}$) & 10^{-4} M NaOH.
The concentration of OH^- is nearly :
(A) 2.414×10^{-4} (B) 10^{-4} M (C) 1.414×10^{-4} (D) 2×10^{-4}

Q.7 At 90°C, pure water has $[H^+]=10^{-6}$M, if 100 ml of 0.2 M HNO_3 is added to 20 ml of 1 M NaOH at 90°C then pH of the resulting solution will be
(A) 5 (B) 6 (C) 7 (D) None of these

Q.8 Calcium lactate is a salt of weak organic acid and strong base represented as $Ca(LaC)_2$. A saturated solution of $Ca(LaC)_2$ contains 0.6 mole in 2 litre solution. pOH of solution is 5.60. If 90% dissociation of the salt takes place then what is pK_a of lactic acid
(A) 2.8 – log(0.54) (B) 2.8 + log(0.54) (C) 2.8 + log (0.27) (D) none

Q.9 What is the degree of dissociation of weak acid HA (C = 0.1 M) in presence of strong acid HB (C = 0.1M). Given : K_a (weak acid) = 10^{-6}
(A) 10^{-5} (B) 10^{-6} (C) 10^{-4} (D) 10^{-3}

Q.10 2.5 mL of $\frac{2}{5}$ M weak monoacidic base ($K_b = 1 \times 10^{-12}$ at 25°C) is titrated with $\frac{2}{15}$ M HCl in water at 25°C. The concentration of H^+ at equivalence point is
($K_w = 1 \times 10^{-14}$ at 25°C)
(A) 3.7×10^{-13} M (B) 3.2×10^{-7} M (C) 3.2×10^{-2} M (D) 2.7×10^{-2} M

Q.11 An aqueous solution of CH_3COOH has a pH = 3 and acid dissociation constant of CH_3COOH is 10^{-5} . What will be the concentration of acid taken initially?
(A) 0.1M (B) 0.11 M (C) 0.09 M (D) 0.101 M

Q.12 pH of a saturated solution of silver salt of monobasic acid HA is found to be 9.
Find the K_{sp} of sparingly soluble salt AgA(s). **Given :** K_a (HA) = 10^{-10}
(A) 1.1×10^{-11} (B) 1.1×10^{-10} (C) 10^{-12} (D) None of these

Q.13 Which of the following expression regarding % ionisation of a monobasic weak acid in aqueous solution at appreciable concentration is not correct?

(A) $100\sqrt{\frac{K_a}{C}}$ (B) $\left\{\frac{K_a}{K_a+[H^+]}\right\}\times 100$ (C) $\left\{\frac{[H^+]}{K_a+[H^+]}\right\}\times 100$ (D) $\frac{100}{1+10}(pK_a - pH)$

Q.14 The pH of a saturated solution of $X(OH)_3$ will be if its solubility product = 2.7×10^{-43}.
[**Given :** log 3 = 0.5 , log 2 = 0.3]
(A) 10.5 (B) 7 (C) 3.5 (D) 3.8

Q.15 Let solubility of a sparingly soluble salt of strong base and weak acid in pure water is equal to $\mathbf{s_1}$, in a buffer solution with pH < 7 is equal to $\mathbf{s_2}$ and in another buffer solution with pH > 7 is equal to $\mathbf{s_3}$. If hydrolysis is considered in all the solutions then the **correct** relation between $\mathbf{s_1}$, $\mathbf{s_2}$ and $\mathbf{s_3}$ is :
(A) $s_1 = s_2 = s_3$ (B) $s_1 > s_2 > s_3$ (C) $s_1 < s_2 < s_3$ (D) $s_2 > s_1 > s_3$

Q.16 The salt $Zn(OH)_2$ is involved in the following two equilibria,
$Zn(OH)_2\,(s) \rightleftharpoons Zn^{2+}\,(aq) + 2OH^-\,(aq)$; $K_{sp} = 10^{-17}$
$Zn(OH)_2\,(s) + 2OH^-\,(aq) \rightleftharpoons [Zn(OH)_4]^{2-}\,(aq.)$; $K_c = 0.1$
The pH of solution at which solubility is minimum is
(A) 10 (B) 9 (C) 8 (D) 11

Q.17 From separate solutions of four sodium salts NaW, NaX, NaY and NaZ had pH 7.0, 9.0, 10.0 and 11.0 respectively. When each solution was 0.1 M, the strongest acid is:
(A) HW (B) HX (C) HY (D) HZ

Q.18 How many moles NH_3 must be added to 2.0 litre of 0.80 M $AgNO_3$ in order to reduce the Ag^+ concentration to 5×10^{-8} M. K_f of $[Ag(NH_3)_2^+] = 10^8$
(A) 0.4 (B) 2 (C) 3.52 (D) 4

Q.19 A_3B_2 is a sparingly soluble salt of molar mass M (g mol^{-1}) and solubility x g lit^{-1}. The ratio of the molar concentration of B^{3-} to the solubility product of the salt is

(A) $108\frac{x^5}{M^5}$ (B) $\frac{1}{108}\frac{M^4}{x^4}$ (C) $\frac{1}{54}\frac{M^4}{x^4}$ (D) None

Assertion and Reason

Q.20 **Statement-1**: pH of acidic buffer solution always increases on increasing dilution.
Statement-2: pH of any acidic solution always increases on increasing dilution.
(A) Statement-1 is true, statement-2 is true and statement-2 is correct explanation for statement-1.
(B) Statement-1 is true, statement-2 is true and statement-2 is NOT the correct explanation for statement-1.
(C) Statement-1 is true, statement-2 is false. (D) Statement-1 is false, statement-2 is true.

Comprehension :

Paragraph for question nos. 21 to 23

Solubility is defined as the maximum number of moles of solute which can be dissolved in 1 litre of solvent. It is affected by presence of common ion, hydrolysis of ions. In a container equal volume of 0.02 M $AgNO_3$ and 0.02 HCN solution is mixed.
K_{sp} of AgCN = 4×10^{-16} M^2 ,
K_a (HCN) = 4×10^{-10} M

Q.21 Concentration of Ag^+ ions in the final solution would be
(A) 10^{-6} (B) 2×10^{-5} (C) 2×10^{-8} (D) 10^{-4}

Q.22 pH of final solution would be
(A) 2 (B) 3 (C) 4 (D) Can not find

Q.23 Concentration of CN^- in the final solution would be
(A) 4×10^{-10} (B) 2×10^{-11} (C) 4×10^{-12} (D) 2×10^{-8}

Paragraph for question nos. 24 to 25

Selective precipitation of ions from a mixture in the form of salts can be done by adding common ion gradually. Let us consider selective precipitation of Cl^- & CrO_4^{2-} ions in the form of AgCl & Ag_2CrO_4 from a mixture having 0.01 M Cl^- and 0.02 M CrO_4^{2-}. For this, salt having Ag^+ [like $AgNO_3(s)$] is added gradually.

Given : K_{sp} AgCl = 10^{-10} and K_{sp} $Ag_2CrO_4 = 4 \times 10^{-14}$.

Q.24 Minimum concentration of Ag^+ ion at which precipitation of at least one ion starts.
(A) 10^{-8} (B) 1.41×10^{-6} (C) 10^{-10} (D) 2×10^{-8}

Q.25 The percentage of one of the ion precipitated when another ion starts precipitation is

(**Given :** $\frac{1}{\sqrt{2}} = 0.7$)

(A) 98.7 % (B) 99.3 % (C) 97.6 % (D) 92.7 %

More than one correct :

Q.26 At 25°C , the dissociation constant of a weak monoprotic acid, HA **(K_a)** is **numerically equal** to the dissociation constant of its conjugate base, A^- **(K_b)**. Which of the following statement(s) is/are **correct**?
(A) The dissociation constant of the acid, HA, is 10^{-7}.
(B) The pH of 0.1 M aqueous solution of the acid, HA, is 4.0.
(C) The pH of 0.1 M aqueous solution of the conjugate base (A^-) is 10.0.
(D) The pH of an aqueous solution containing 0.1 M - HA and 0.01 M - HCl is 2.0.

Match the column

Q.27 Match column I having components with column II having pH or pOH of solution.

	Column I		Column II
(A)	Saturated solution of $A(OH)_2$ having $k_{sp} = 5 \times 10^{-31}$	(P)	7
(B)	Mixing equal volumes of 0.1 M CH_3COONa & 0.1 M XNO_3	(Q)	10
(C)	Solution obtained at $\frac{10}{11}$ of the equivalence point when 0.1 M XOH is titrated with HCl 0.1 M	(R)	5
		(S)	4

[Given data : K_h $X^+ = 10^{-5}$, $K_{a,CH_3COOH} = 10^{-5}$]

Subjective :

Q.28 A buffer of pH 9.26 is made by dissolving x moles of ammonium sulphate and 0.1 mole of ammonia into 100 mL solution. If pK_b of ammonia is 4.74, calculate value of x.

Q.29 50ml of a weak monoprotic acid was titrated with 0.1M solution of NaOH. If pH of the solution after adding 50 ml and 75 ml of base is 4.699 and 5 then calculate a four digit number **'abcd'** where :
ab = pK_a of the weak acid; **cd** = pH of the solution when 7.5 millimoles of HCl are added in the solution at the equivalence point without changing volume.[**Given :** log 2 = 0.301]

Q.30 A solution contains 0.01M Mg^{+2} and 0.1M Sr^{+2} and H_2CO_3 maintained at 0.05M. At what minimum & maximum pH will $SrCO_3$ will precipitate without any precipitation of $MgCO_3$.
[**Given:** $K_{sp}MgCO_3 = 2.5 \times 10^{-7}$, $K_{sp}SrCO_3 = 9 \times 10^{-9}$, $K_{a(overall)}$ $H_2CO_3 = 5\times10^{-16}$, log5 = 0.7, log 3 = 0.5] ***Express your asnwer as abcd where*** **ab** = Ten times minimum pH for the above condition;

REVISION FLOW CHART

Ionic Equilibrium

Theoretical Concept

- Different theories of Acid-base, Arrhenius, Bronsted lowry & Lewis — T
- Arrhenius Theory of dissociation K_w, K_a & K_b relation between K_a & K_b of conjugate acid-base — T + N
- Ostwald dilution law & factor affecting α — N + G
- Factor affecting α — T + N
- Dissociation of water, pH definition & variation of pH with T — T + N

pH Calculation

- Solution of only acid or bases (i) strong (ii) weak — N
  - mono
  - polyprotic
- Mixtures of only Acids or Only bases — N
  - SA + SA
  - SA + WA
  - WA + WA
- Salt solutions — N
  - Salt of SA/SB
  - Salt of SA/WB
  - Salt of WA/WB
  - Amphiprotic salt
  - Polyvalent hydrolysis
- Mixtures of salt & Acid or salt & base — N
  - Buffer solution
  - Mixture of only acids/ or only base

Buffer

- Type of Buffer — T
- pH calculation — N
- pH variation of dilution & on adding acid or base — N
- Buffer capacity — N + G

Indicator / Titration

- Theory of indicator — T
- Titration of SA+SB, WA+SB, SA+WB, WA+WB — useful indicator — N + G

Solubility

- Definition — N
- K_{sp} & solubility — N
- K_{SP} & K_{IP} — T + N
- Factor affecting solubility — T
- Common ion effect — N
- Complex formation — N
- Hydrolysis of cation, anion — N
- Precipitation/Condition for precipitation — T + N

(G) → represents topic of graphical importance

(T) → represents topic of theoretical importance

(N) → represents topic of numerical importance

587, Nitikhand-1, Indirapuram, Gzb.
7292077839 / 7292047839 | smartachievers.online

LIST OF IMPORTANT FORMULAS

* **Basic Relationships**

(a) $K_w = [H^+][OH^-] = 10^{-4}$ (at 25°C) , K_a (for water) $= \dfrac{K_w \times 18}{1000}$

(b) $\ln\dfrac{K_{w_2}}{K_{w_1}} = \dfrac{\Delta H}{R}\left[\dfrac{1}{T_1} - \dfrac{1}{T_2}\right]$; $\Delta H = \Delta H$ neutralise of SA and SB

(c) $pH + pOH = pK_w = 14$ at 25°C

(d) For weak conjugate acid / base pair

$pK_a + pK_b = pK_w$

(e) Ostwald dil law for an electrolyte AB, $\dfrac{C\alpha^2}{1-\alpha} = K$

pH CALCULATUION

1. pH of a solution containing strong monoprotic acid.

(i) at moderate conc. ($C \geq 10^{-6}$ M)

$pH = -\log C$

(ii) at low conc. ($C < 10^{-6}$ M)

$K_\omega = (C + x)(x)$

$[H^+] = C + x$ $x \rightarrow$ contribution of water

2. pH of a solution containing weak monoprotic acid.

Case-I : *If $C \geq 10^{-6}$ M*

$K_a = \dfrac{x^2}{C - x}$ If $C >> x$ $pH = \dfrac{1}{2}(pK_b - \log C)$

Case-II : If $C < 10^{-6}$ M

Weak acid can be treated as a strong acid but contribution due to water cannot be neglected.

$[H^+] = c + x$

$K_\omega = (c + x)x$

3. pH of a solution containing more than one acid.

Case I : **Mixture of SA + SA**

$[HA] = c_1$

$[HB] = c_2$

$[H^+] = c_1 + c_2$

Depending upon the final [H^+] condition, Contribution from water may be neglected or taken.

Case II : **Mixture of SA + WA**

[HCl] = c_1 [HA] = c_2

$$HA \rightleftharpoons H^+ + A^-$$
$$c_2 - x \quad c_1 + x \quad x$$

$$K_a = \frac{(c_1 + x)x}{(c_2 - x)}$$

Case-III : **Mixture of WA + WA**

$$HA \rightleftharpoons H^+ + A^- \qquad K_1 = \frac{(x+y)x}{c_1 - x}$$
$$c_1 - x \quad x + y \quad x$$

$$HB \rightleftharpoons H^+ + B^- \qquad K_2 = \frac{(x+y)y}{c_2 - y}$$
$$c_2 - y \quad x + y \quad y$$

* For polyprotic acid $K_1 >> K_2 >> K_3$

All the formulaes derived above are applicable for bases also. Only [H^+] will be replaced with [OH^-]

SALT HYDROLYSIS

Case-I : ***Salts of strong acids and strong bases*** do not undergo hydrolysis.

Neutral $\Rightarrow pH = 7$

Case-II : ***Salts of strong base and weak acid*** give a basic solution (pH>7) when dissolved in water.

e.g. : CH_3COONa, NaCN

$$K_b = \frac{K_\omega}{K_a} = K_h$$

$$K_h = K_h = \frac{K_\omega}{K_a} = \frac{(ch)^2}{c(1-h)}$$

if $\frac{K_h}{c} \le 10^{-3}$ then $h << 1$

$$pH = \frac{1}{2}\left(pK_\omega + pK_a + \log C\right)$$

In case of substantial hydrolysis the above expression will not be applicable

Case-III : ***Salts of a strong acids and weak bases*** give an acidic solution.

$$K_a = \frac{K_\omega}{K_b} = K_2$$

$$\frac{K_\omega}{K_b} = \frac{(ch)^2}{c(1-h)} = K_h$$

if $\frac{K_h}{c} \le 10^{-3}$ $h << 1$

$$= \frac{1}{2}\left(-\log K_\omega + \log K_b - \log C\right) = \frac{1}{2}\left(pK_\omega - pK_b - \log C\right)$$

In case of substantial hydrolysis the above expression will not be applicable

(iv) ***Salts of weak base and weak acid***

Assuming degree of hydrolysis to be same for the both the ions,

$K_h = K_w / (K_a.K_b)$, $[H^+] = [K_a K_w/K_b]^{1/2}$

$$pH = \frac{1}{2}(pk_w + pk_a - pk_b)$$

(v) ***Amphiprotic salts***

e.g. $NaHCO_3$

$$pH = \frac{1}{2}(pK_1 + pK_2)$$

pH of such salt solution is independent of concentration od salt.

* ***Buffer solution***

Acidic Buffer

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

$$\left[\text{Actually } pH = pK_a + \log\frac{[\text{conjugate base}]}{[\text{acid}]}\right]$$

* Basic buffer

$$pOH = pK_b + \log \frac{[salt]}{[base]}$$

$$\left[\text{Actually } pOH = pK_b + \log \frac{[\text{conjugate acid}]}{[base]} \right]$$

* Simple buffer (Salt of Weak acid & weak base)

$$pH = \frac{1}{2}(pK_w + pK_a - pK_b)$$

* ***Solubility***

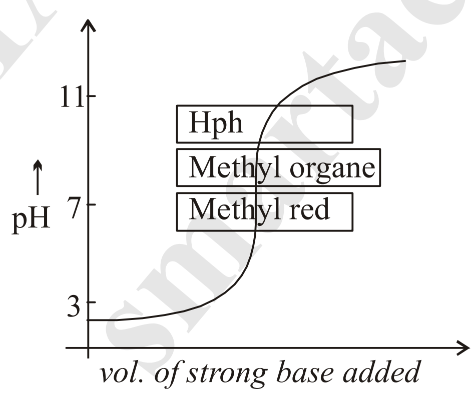
For A_xB_y $K_{sp} = (xS)^x (yS)^y$

Effect of hydrolysis on solubility $S = \sqrt{K_{sp}\left[1 + \frac{[H^+]}{K_a}\right]}$

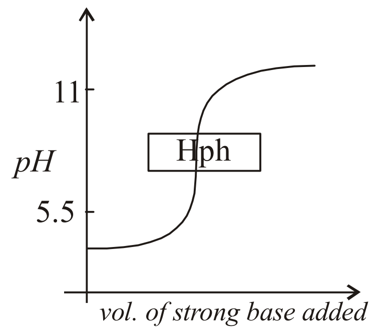
* **Indicators**

Indicators	**pH range**		**Colour**
		acid medium	basic medium
Methyl Orange	3.1-4.4	pink	yellow
Phenolphathlene	8.3-10	colourless	pink

Case I : **Titration of SA + SB**

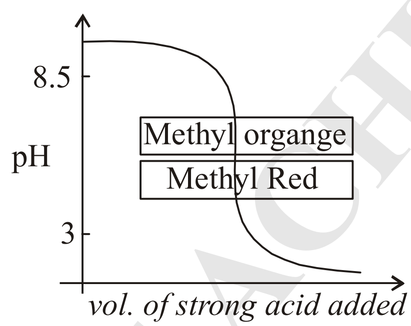


Case II : **Titration of WA + SB**



Methyl orange can't be used. Nothing definite can be said about the use of methyl red.

Case-III : **Titration of WB + SA**



Case-IV : **Titration of WB + WA :** No sharp change in pH. No suitable indicator.

* For complex formation

$$Ag^+ + NH_3 \rightleftharpoons Ag(NH_3)^+ \qquad K_1$$

$$Ag(NH_3)^+ + NH_3 \rightleftharpoons Ag(NH_3)_2^+ \qquad K_2$$

$$Ag^+ + 2NH_3 \rightleftharpoons Ag(NH_3)_2^+ \qquad K_{form} = K_1K_2$$

$$K_d \text{(dissociation or instability constant)} = \frac{1}{K_{form}\text{(formation or stability constant)}}$$

LAST MOMENT REVIEW

IONIC EQUILIBRIUM

Theory : ....

....

....

....

Exercise - 1 : Question ....

Exercise - 2 : Question ....

Exercise - 3 : Question ....

Exercise - 4 : Question ....

DPPs : ....

....

....

....

Other Sources : ....

....

....

CLASS : CC -AD

Revision Worksheet Ionic Equilibrium

Single correct

Q.1 Three solution of strong electrolytes 25 ml of 0.1 M HX, 25 ml of 0.1 M H_2Y and 50 ml of 0.1 $NZ(OH)_2$ are mixed. What is the pOH of the solution ? [Given log 5 = 0.7]
(A) 1.6 (B) 7 (C) 11.6 (D*) 12.6

Q.2 0.1 M NH_3 (aq) solution is titrated against 0.1 M HNO_3 (aq) solution. What would be the difference in pOH between $\frac{1}{4}$ and $\frac{3}{4}$ stages of neutralization of base?
(A) 2 log 3/4 (B) 2 log 1/4 (C) log 1/3 (D*) 2 log 3

Q.3 What is the pH of 0.1M $NaHCO_3$? $K_1 = 5.0 \times 10^{-7}$, $K_2 = 7.0 \times 10^{-11}$ for carbonic acids.
(A) pH = 7 (B) pH much less than 7
(C) pH slightly less than 7 (D*) pH greater than 7

Q.4 The degree of hydrolysis of 0.1 M solution of conjugate base of HA is 0.01. Find the $\mathbf{H}^+$ concentration in 0.4 M solution of $\mathbf{A}^-$.
(A) 5×10^{-11} M (B*) 5×10^{-12} M (C) 2×10^{-3} M (D) 2×10^{-4} M

Q.5 What is the aq. ammonia concentration of a solution prepared by dissolving 0.15 mole of $NH_4^+CH_3COO^-$ in 1L H_2O.[$K_{a(CH_3COOH)} = 1.8\times10^{-5}$; $K_{b(NH_4OH)} = 1.8\times10^{-5}$]
(A*) 8.3×10^{-4} (B) 0.15 (C) 6.4×10^{-4} (D) 3.8×10^{-4}

Q.6 An aqueous solution contains 0.01 M RNH_2 ($K_b = 2 \times 10^{-6}$) & 10^{-4} M NaOH.
The concentration of OH^- is nearly :
(A) 2.414×10^{-4} (B) 10^{-4} M (C) 1.414×10^{-4} (D*) 2×10^{-4}

Q.7 At 90°C, pure water has $[H^+]=10^{-6}$M, if 100 ml of 0.2 M HNO_3 is added to 20 ml of 1 M NaOH at 90°C then pH of the resulting solution will be
(A) 5 (B*) 6 (C) 7 (D) None of these

Q.8 Calcium lactate is a salt of weak organic acid and strong base represented as $Ca(LaC)_2$. A saturated solution of $Ca(LaC)_2$ contains 0.6 mole in 2 litre solution. pOH of solution is 5.60. If 90% dissociation of the salt takes place then what is pK_a of lactic acid
(A*) 2.8 – log(0.54) (B) 2.8 + log(0.54) (C) 2.8 + log (0.27) (D) none

Q.9 What is the degree of dissociation of weak acid HA (C = 0.1 M) in presence of strong acid HB (C = 0.1M). Given : K_a (weak acid) = 10^{-6}
(A*) 10^{-5} (B) 10^{-6} (C) 10^{-4} (D) 10^{-3}

Q.10 2.5 mL of $\frac{2}{5}$ M weak monoacidic base ($K_b = 1 \times 10^{-12}$ at 25°C) is titrated with $\frac{2}{15}$ M HCl in water at 25°C. The concentration of H^+ at equivalence point is
($K_w = 1 \times 10^{-14}$ at 25°C)
(A) 3.7×10^{-13} M (B) 3.2×10^{-7} M (C) 3.2×10^{-2} M (D*) 2.7×10^{-2} M

Q.11 An aqueous solution of CH_3COOH has a pH = 3 and acid dissociation constant of CH_3COOH is 10^{-5} . What will be the concentration of acid taken initially?
(A) 0.1M (B) 0.11 M (C) 0.09 M (D*) 0.101 M

Q.12 pH of a saturated solution of silver salt of monobasic acid HA is found to be 9.
Find the K_{sp} of sparingly soluble salt AgA(s). **Given :** K_a (HA) = 10^{-10}
(A*) 1.1×10^{-11} (B) 1.1×10^{-10} (C) 10^{-12} (D) None of these

Q.13 Which of the following expression regarding % ionisation of a monobasic weak acid in aqueous solution at appreciable concentration is not correct?

(A) $100\sqrt{\frac{K_a}{C}}$ (B) $\left\{\frac{K_a}{K_a+[H^+]}\right\}\times 100$ (C*) $\left\{\frac{[H^+]}{K_a+[H^+]}\right\}\times 100$ (D) $\frac{100}{1+10}(pK_a - pH)$

Q.14 The pH of a saturated solution of $X(OH)_3$ will be if its solubility product = 2.7×10^{-43}.
[**Given :** log 3 = 0.5 , log 2 = 0.3]
(A) 10.5 (B*) 7 (C) 3.5 (D) 3.8

Q.15 Let solubility of a sparingly soluble salt of strong base and weak acid in pure water is equal to $\mathbf{s_1}$, in a buffer solution with pH < 7 is equal to $\mathbf{s_2}$ and in another buffer solution with pH > 7 is equal to $\mathbf{s_3}$. If hydrolysis is considered in all the solutions then the **correct** relation between $\mathbf{s_1}$, $\mathbf{s_2}$ and $\mathbf{s_3}$ is :
(A) $s_1 = s_2 = s_3$ (B) $s_1 > s_2 > s_3$ (C) $s_1 < s_2 < s_3$ (D*) $s_2 > s_1 > s_3$

Q.16 The salt $Zn(OH)_2$ is involved in the following two equilibria,
$Zn(OH)_2\,(s) \rightleftharpoons Zn^{2+}\,(aq) + 2OH^-\,(aq)$; $K_{sp} = 10^{-17}$
$Zn(OH)_2\,(s) + 2OH^-\,(aq) \rightleftharpoons [Zn(OH)_4]^{2-}\,(aq.)$; $K_c = 0.1$
The pH of solution at which solubility is minimum is
(A*) 10 (B) 9 (C) 8 (D) 11

Q.17 From separate solutions of four sodium salts NaW, NaX, NaY and NaZ had pH 7.0, 9.0, 10.0 and 11.0 respectively. When each solution was 0.1 M, the strongest acid is:
(A*) HW (B) HX (C) HY (D) HZ

Q.18 How many moles NH_3 must be added to 2.0 litre of 0.80 M $AgNO_3$ in order to reduce the Ag^+ concentration to 5×10^{-8} M. K_f of $[Ag(NH_3)_2{}^+] = 10^8$
(A) 0.4 (B) 2 (C) 3.52 (D*) 4

Q.19 A_3B_2 is a sparingly soluble salt of molar mass M (g mol^{-1}) and solubility x g lit^{-1}. The ratio of the molar concentration of B^{3-} to the solubility product of the salt is

(A) $108\dfrac{x^5}{M^5}$ (B) $\dfrac{1}{108}\dfrac{M^4}{x^4}$ (C*) $\dfrac{1}{54}\dfrac{M^4}{x^4}$ (D) None

Assertion and Reason

Q.20 **Statement-1**: pH of acidic buffer solution always increases on increasing dilution.

Statement-2: pH of any acidic solution always increases on increasing dilution.

(A) Statement-1 is true, statement-2 is true and statement-2 is correct explanation for statement-1.

(B) Statement-1 is true, statement-2 is true and statement-2 is NOT the correct explanation for statement-1.

(C) Statement-1 is true, statement-2 is false. (D*) Statement-1 is false, statement-2 is true.

Comprehension :

Paragraph for question nos. 21 to 23

Solubility is defined as the maximum number of moles of solute which can be dissolved in 1 litre of solvent. It is affected by presence of common ion, hydrolysis of ions. In a container equal volume of 0.02 M $AgNO_3$ and 0.02 HCN solution is mixed.

K_{sp} of AgCN = 4×10^{-16} M^2 ,
K_a (HCN) = 4×10^{-10} M

Q.21 Concentration of Ag^+ ions in the final solution would be
(A) 10^{-6} (B) 2×10^{-5} (C) 2×10^{-8} (D*) 10^{-4}

Q.22 pH of final solution would be
(A*) 2 (B) 3 (C) 4 (D) Can not find

Q.23 Concentration of CN^- in the final solution would be
(A) 4×10^{-10} (B) 2×10^{-11} (C*) 4×10^{-12} (D) 2×10^{-8}

Paragraph for question nos. 24 to 25

Selective precipitation of ions from a mixture in the form of salts can be done by adding common ion gradually. Let us consider selective precipitation of Cl^- & CrO_4^{2-} ions in the form of AgCl & Ag_2CrO_4 from a mixture having 0.01 M Cl^- and 0.02 M CrO_4^{2-}. For this, salt having Ag^+ [like $AgNO_3(s)$] is added gradually.

Given : K_{sp} AgCl = 10^{-10} and K_{sp} Ag_2CrO_4 = 4×10^{-14}.

Q.24 Minimum concentration of Ag^+ ion at which precipitation of at least one ion starts.
(A*) 10^{-8} (B) 1.41×10^{-6} (C) 10^{-10} (D) 2×10^{-8}

Q.25 The percentage of one of the ion precipitated when another ion starts precipitation is

(**Given :** $\dfrac{1}{\sqrt{2}} = 0.7$)

(A) 98.7 % (B*) 99.3 % (C) 97.6 % (D) 92.7 %

More than one correct :

Q.26 At 25°C , the dissociation constant of a weak monoprotic acid, HA **(K_a)** is **numerically equal** to the dissociation constant of its conjugate base, A^- **(K_b)**. Which of the following statement(s) is/are **correct**?
(A*) The dissociation constant of the acid, HA, is 10^{-7}.
(B*) The pH of 0.1 M aqueous solution of the acid, HA, is 4.0.
(C*) The pH of 0.1 M aqueous solution of the conjugate base (A^-) is 10.0.
(D*) The pH of an aqueous solution containing 0.1 M - HA and 0.01 M - HCl is 2.0.

Match the column

Q.27 Match column I having components with column II having pH or pOH of solution.

	Column I	Column II	
(A)	Saturated solution of $A(OH)_2$ having $k_{sp} = 5 \times 10^{-31}$	(P)	7
(B)	Mixing equal volumes of 0.1 M CH_3COONa & 0.1 M XNO_3	(Q)	10
(C)	Solution obtained at $\frac{10}{11}$ of the equivalence point when 0.1 M	(R)	5
	XOH is titrated with HCl 0.1 M	(S)	4

Given data : $K_h\ X^+ = 10^{-5}$, $K_{a,CH_3COOH} = 10^{-5}$

[Ans. (A) P (B) R (C) QS]

Subjective :

Q.28 A buffer of pH 9.26 is made by dissolving x moles of ammonium sulphate and 0.1 mole of ammonia into 100 mL solution. If pK_b of ammonia is 4.74, calculate value of x. [Ans.0.05 mol]

Q.29 50ml of a weak monoprotic acid was titrated with 0.1M solution of NaOH. If pH of the solution after adding 50 ml and 75 ml of base is 4.699 and 5 then calculate a four digit number **'abcd'** where :
ab = pK_a of the weak acid; **cd** = pH of the solution when 7.5 millimoles of HCl are added in the solution at the equivalence point without changing volume.[**Given :** log 2 = 0.301] [Ans. 0505]

Q.30 A solution contains 0.01M Mg^{+2} and 0.1M Sr^{+2} and H_2CO_3 maintained at 0.05M. At what minimum & maximum pH will $SrCO_3$ will precipitate without any precipitation of $MgCO_3$.
[**Given:** $K_{sp}MgCO_3 = 2.5 \times 10^{-7}$, $K_{sp}SrCO_3 = 9 \times 10^{-9}$, $K_{a(overall)}\ H_2CO_3 = 5\times10^{-16}$, log5 = 0.7, log 3 = 0.5] ***Express your asnwer as abcd where*** **ab** = Ten times minimum pH for the above condition; **cd** = Ten times maximum pH for the above condition. [Ans. 4860]

REVISION FLOW CHART

Ionic Equilibrium

Theoretical Concept
- Different theories of Acid-base, Arrhenius, Bronsted lowry & Lewis — T
- Arrhenius Theory of dissociation K_w, K_a & K_b relation between K_a & K_b of conjugate acid-base — T + N
- Ostwald dilution law & factor affecting α — N + G
- Factor affecting α — T + N
- Dissociation of water, pH definition & variation of pH with T — T + N

pH Calculation
- Solution of only acid or bases — N
  - (i) strong
  - (ii) weak
    - mono
    - polyprotic
- Mixtures of only Acids or Only bases — N
  - SA + SA
  - SA + WA
  - WA + WA
- Salt solutions — N
  - Salt of SA/SB
  - Salt of SA/WB
  - Salt of WA/WB
  - Amphiprotic salt
  - Polyvalent hydrolysis
- Mixtures of salt & Acid or salt & base — N
  - Buffer solution
  - Mixture of only acids/ or only base

Buffer
- Type of Buffer — T
- pH calculation — N
- pH variation of dilution & on adding acid or base — N
- Buffer capacity — N + G

Indicator / Titration
- Theory of indicator — T
- Titration of — N + G
  - SA+SB
  - WA+SB
  - SA+WB
  - WA+WB
  - useful indicator

Solubility
- Definition — N
- K_{sp} & solubility — N
- K_{SP} & K_{IP} — T + N
- Factor affecting solubility — T
- Common ion effect — N
- Complex formation — N
- Hydrolysis of cation, anion — N
- Precipitation/Condition for precipitation — T + N

(G) → represents topic of graphical importance

(T) → represents topic of theoretical importance

(N) → represents topic of numerical importance

LIST OF IMPORTANT FORMULAS

* **Basic Relationships**

(a) $K_w = [H^+][OH^-] = 10^{-4}$ (at 25°C) , K_a (for water) $= \frac{K_w \times 18}{1000}$

(b) $\ln\frac{K_{w_2}}{K_{w_1}} = \frac{\Delta H}{R}\left[\frac{1}{T_1} - \frac{1}{T_2}\right]$; $\Delta H = \Delta H$ neutralise of SA and SB

(c) $pH + pOH = pK_w = 14$ at 25°C

(d) For weak conjugate acid / base pair

$pK_a + pK_b = pK_w$

(e) Ostwald dil law for an electrolyte AB, $\frac{C\alpha^2}{1-\alpha} = K$

pH CALCULATUION

1. pH of a solution containing strong monoprotic acid.

(i) at moderate conc. ($C \geq 10^{-6}$ M)

$pH = -\log C$

(ii) at low conc. ($C < 10^{-6}$ M)

$K_\omega = (C + x)(x)$

$[H^+] = C + x$ $x \rightarrow$ contribution of water

2. pH of a solution containing weak monoprotic acid.

Case-I : *If $C \geq 10^{-6}$ M*

$K_a = \frac{x^2}{C - x}$ If $C >> x$ $pH = \frac{1}{2}(pK_b - \log C)$

Case-II : If $C < 10^{-6}$ M

Weak acid can be treated as a strong acid but contribution due to water cannot be neglected.

$[H^+] = c + x$

$K_\omega = (c + x)\,x$

3. pH of a solution containing more than one acid.

Case I : **Mixture of SA + SA**

$[HA] = c_1$

$[HB] = c_2$

$[H^+] = c_1 + c_2$

Depending upon the final $[H^+]$ condition, Contribution from water may be neglected or taken.

Case II : **Mixture of SA + WA**

$[HCl] = c_1$ $\quad$ $[HA] = c_2$

$$HA \rightleftharpoons H^+ + A^-$$
$$c_2 - x \quad c_1 + x \quad x$$

$$K_a = \frac{(c_1 + x)x}{(c_2 - x)}$$

Case-III : **Mixture of WA + WA**

$$HA \rightleftharpoons H^+ + A^-$$
$$c_1 - x \quad x + y \quad x$$

$$K_1 = \frac{(x+y)x}{c_1 - x}$$

$$HB \rightleftharpoons H^+ + B^-$$
$$c_2 - y \quad x + y \quad y$$

$$K_2 = \frac{(x+y)y}{c_2 - y}$$

* For polyprotic acid $K_1 >> K_2 >> K_3$

All the formulaes derived above are applicable for bases also. Only $[H^+]$ will be replaced with $[OH^-]$

SALT HYDROLYSIS

Case-I : ***Salts of strong acids and strong bases*** do not undergo hydrolysis.

Neutral $\Rightarrow pH = 7$

Case-II : ***Salts of strong base and weak acid*** give a basic solution (pH>7) when dissolved in water.

e.g. : CH_3COONa, NaCN

$$K_b = \frac{K_\omega}{K_a} = K_h$$

$$K_h = K_h = \frac{K_\omega}{K_a} = \frac{(ch)^2}{c(1-h)}$$

if $\frac{K_h}{c} \le 10^{-3}$ then $h << 1$

$$pH = \frac{1}{2}\left(pK_{\omega} + pK_a + \log C\right)$$

In case of substantial hydrolysis the above expression will not be applicable

Case-III : ***Salts of a strong acids and weak bases*** give an acidic solution.

$$K_a = \frac{K_{\omega}}{K_b} = K_2$$

$$\frac{K_{\omega}}{K_b} = \frac{(ch)^2}{c(1-h)} = K_h$$

if $\frac{K_h}{c} \le 10^{-3}$ $h << 1$

$$= \frac{1}{2}\left(-\log K_{\omega} + \log K_b - \log C\right) = \frac{1}{2}\left(pK_{\omega} - pK_b - \log C\right)$$

In case of substantial hydrolysis the above expression will not be applicable

(iv) ***Salts of weak base and weak acid***

Assuming degree of hydrolysis to be same for the both the ions,

$K_h = K_w / (K_a.K_b)$, $[H^+] = [K_a K_w/K_b]^{1/2}$

$$pH = \frac{1}{2}\ (pk_w + pk_a - pk_b)$$

(v) ***Amphiprotic salts***

e.g. $NaHCO_3$

$$pH = \frac{1}{2}\ (pK_1 + pK_2)$$

pH of such salt solution is independent of concentration od salt.

* ***Buffer solution***

Acidic Buffer

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

$$\left[\text{Actually } pH = pK_a + \log \frac{[\text{conjugate base}]}{[acid]} \right]$$

* Basic buffer

$$pOH = pK_b + \log \frac{[salt]}{[base]}$$

$$\left[\text{Actually } pOH = pK_b + \log \frac{[\text{conjugate acid}]}{[base]} \right]$$

* Simple buffer (Salt of Weak acid & weak base)

$$pH = \frac{1}{2}(pK_w + pK_a - pK_b)$$

* ***Solubility***

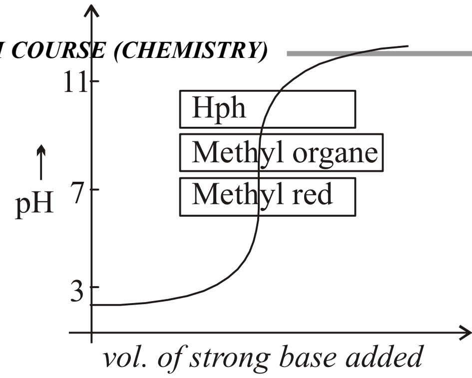
For A_xB_y $K_{sp} = (xS)^x (yS)^y$

Effect of hydrolysis on solubility $S = \sqrt{K_{sp}\left[1 + \frac{[H^+]}{K_a}\right]}$

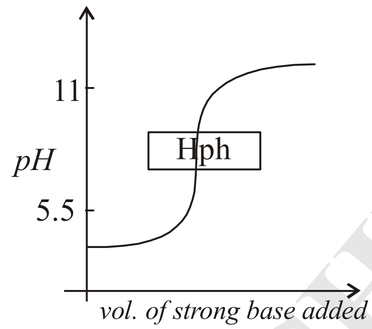
* **Indicators**

Indicators	**pH range**		**Colour**
		acid medium	basic medium
Methyl Orange	3.1-4.4	pink	yellow
Phenolphathlene	8.3-10	colourless	pink

Case I : **Titration of SA + SB**

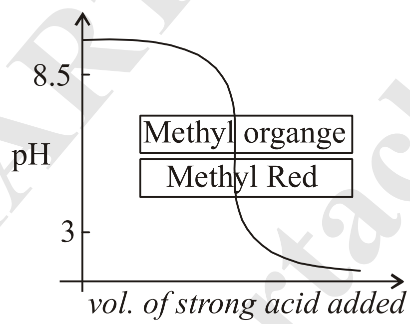


Case II : **Titration of WA + SB**



Methyl orange can't be used. Nothing definite can be said about the use of methyl red.

Case-III : **Titration of WB + SA**



Case-IV : **Titration of WB + WA :** No sharp change in pH. No suitable indicator.

* For complex formation

$$Ag^+ + NH_3 \rightleftharpoons Ag(NH_3)^+ \qquad K_1$$

$$Ag(NH_3)^+ + NH_3 \rightleftharpoons Ag(NH_3)_2{}^+ \qquad K_2$$

$$Ag^+ + 2NH_3 \rightleftharpoons Ag(NH_3)_2{}^+ \qquad K_{form} = K_1K_2$$

$$K_d \text{ (dissociation or instability constant)} = \frac{1}{K_{form}\text{ (formation or stability constant)}}$$

587, Nitikhand-1, Indirapuram, Gzb.
7292077839 / 7292047839 | smartachievers.online

LAST MOMENT REVIEW

IONIC EQUILIBRIUM

Theory : ....

....

....

....

Exercise - 1 : Question ....

Exercise - 2 : Question ....

Exercise - 3 : Question ....

Exercise - 4 : Question ....

DPPs : ....

....

....

....

Other Sources : ....

....

....