DPP - Daily Practice Problems

Name :

Date :

Start Time :

End Time :

CHEMISTRY

14

SYLLABUS : Equilibrium 3 : Buffer solution, solubility & common ion effect

Max. Marks : 120

Time : 60 min.

GENERAL INSTRUCTIONS

- The Daily Practice Problem Sheet contains 30 MCQ's. For each question only one option is correct. Darken the correct circle/ bubble in the Response Grid provided on each page.
- You have to evaluate your Response Grids yourself with the help of solution booklet.
- Each correct answer will get you 4 marks and 1 mark shall be deduced for each incorrect answer. No mark will be given/ deducted if no bubble is filled. Keep a timer in front of you and stop immediately at the end of 60 min.
- The sheet follows a particular syllabus. Do not attempt the sheet before you have completed your preparation for that syllabus. Refer syllabus sheet in the starting of the book for the syllabus of all the DPP sheets.
- After completing the sheet check your answers with the solution booklet and complete the Result Grid. Finally spend time to analyse your performance and revise the areas which emerge out as weak in your evaluation.

DIRECTIONS (Q.1-Q.21) : There are 21 multiple choice questions. Each question has 4 choices (a), (b), (c) and (d), out of which ONLY ONE choice is correct.

Q.1 The solubility product of chalk is 9.3×10^{-8}. Calculate its solubility in gram per litre–

(a) 0.3040 gram / litre (b) 0.0304 gram / litre

(c) 2.0304 gram / litre (d) 4.0304 gram / litre

Q.2 Maximum conductivity would be of–

(a) $K_3[Fe(CN)_6]$ [0.1 M solution]

(b) $K_2[Ni(CN)_4]$ [0.1M solution]

(c) $FeSO_4.Al_2(SO_4)_3.24H_2O$ [0.1 M solution]

(d) $Na[Ag(S_2O_2)_3]$ [0.1 M solution]

Q.3 Calculate pH of the buffer solution containing 0.15 moles of NH_4OH and 0.25 moles of NH_4Cl, K_b for NH_4OH is 1.8×10^{-5}.

(a) 7.034 (b) 9.04

(c) 8.043 (d) None of these

Q.4 0.15 mole of pyridinium chloride has been added into 500 cm^3 of 0.2 M pyridine solution. Calculate pH of the resulting solution assuming no change in volume.
(K_b for pyridine = 1.5×10^{-9} M)

(a) 4 (b) 9 (c) 5 (d) 8

Q.5 The solubility of CaF_2 in water at 20°C is 15.6 mg per dm^3 solution. What will be the solubility product of CaF_2 ?

(a) 4.0×10^{-4} (b) 8.0×10^{-8}

(c) 32.0×10^{-12} (d) None

RESPONSE GRID 1. ⓐⓑⓒⓓ 2. ⓐⓑⓒⓓ 3. ⓐⓑⓒⓓ 4. ⓐⓑⓒⓓ 5. ⓐⓑⓒⓓ

Space for Rough Work

Q.6 Given the solubility product of A_3B_2 is 2×10^{-30}. What will be the solubility in moles/litre?

(a) $(1.85 \times 10^{-32})^{1/5}$ (b) $\left(\frac{2 \times 10^{-30}}{108}\right)^{1/5}$

(c) $\left(\frac{10-28}{5400}\right)^{1/5}$ (d) All

Q.7 Which of the following pairs constitutes a buffer ?

(a) HNO_3 and NH_4NO_3 (b) HCl and KCl

(c) HNO_2 and $NaNO_2$ (d) NaOH and NaCl

Q.8 The solubility of AgCl is 0.0014 g per litre at 18°C. What will be its solubility product at 18°C? (Molecular weight of AgCl = 143.5)

(a) 3.952×10^{-10} (b) 0.0952×10^{-10}

(c) 1.952×10^{-10} (d) 0.952×10^{-10}

Q.9 A salt M_2X_3 dissolves in water such that its solubility is x g. mole/litre. Its K_{sp} is–

(a) x^5 (b) $6x^2$ (c) $108x^5$ (d) $6x^5$

Q.10 The solubilities of AgCl in water, in 0.02 M $CaCl_2$, in 0.01M NaCl and in 0.05 M$AgNO_3$ are S_0, S_1, S_2 and S_3 respectively. Which of the following relationships between these quantities is correct ?

(a) $S_0 > S_1 > S_2 > S_3$ (b) $S_0 > S_2 > S_1 > S_3$

(c) $S_0 > S_1 = S_2 > S_3$ (d) $S_0 > S_2 > S_3 > S_1$

Q.11 Solubility product of AgCl is 2.8×10^{-10} at 25°C. Calculate solubility of the salt in 0.1 M $AgNO_3$ solution

(a) 2.8×10^{-9} mole/litre (b) 2.8×10^{-10} mole/litre

(c) 3.2×10^{-9} mole/litre (d) 3.2×10^{-12} mole/litre

Q.12 What amount of sodium propanoate should be added to one litre of an aqueous solution containing 0.02 mole of propanoic acid ($K_a = 1.0 \times 10^{-5}$ at 25°C) to obtain a buffer solution of pH 6 ?

(a) 0.1 M (b) 0.2 M (c) 0.3 M (d) 1.3 M

Q.13 What will be the pH of the solution, if 0.01 mole of HCl is dissolved in a buffer solution containing 0.03 mole of propanoic acid ($K_a = 1.0 \times 10^{-5}$) and 0.02 moles of salt, at 25°C ?

(a) 3.4 (b) 4.4 (c) 5.4 (d) 6.4

Q.14 20 ml of 0.2 M NaOH is added to 50 ml of 0.2 M CH_3COOH to give 70 ml of the solution. What is the pH of the solution? (The ionization constant of acetic acid is 2×10^{-5})

(a) 4.522 (b) 5.568 (c) 6.522 (d) 7.568

Q.15 The concentration of H^+ ions in a 0.2 M solution of HCOOH is 6.4×10^{-3} mol L^{-1}. To this solution, HCOONa is added so as to adjust the concentration of HCOONa to one mole per litre. What will be the pH of this solution? (K_a for HCOOH is 2.4×10^{-4} and the degree of dissociation of HCOONa is 0.75)

(a) 3.19 (b) 4.19 (c) 5.19 (d) 6.19

Q.16 What amount of HCl will be required to prepare one litre of a buffer solution (containing NaCN and HCN) of pH 10.4 using 0.01 mole of NaCN?

Given $K_{ion}(HCN) = 4.1 \times 10^{-10}$

(a) 8.55×10^{-3} mole (b) 8.65×10^{-3} mole

(c) 8.75×10^{-3} mole (d) 9.9×10^{-4} mole

Q.17 Calculate the pH of a solution which contains 10 ml of 1 M HCl and 10 ml of 2 M NaOH

(a) 11.7 (b) 12.7 (c) 13.7 (d) 10.7

Q.18 Calculate pH of a solution of given mixture (0.1 mole CH_3COOH + 0.2 mole CH_3COONa) in 100 ml of mixture. ($K_a = 2 \times 10^{-5}$)

(a) 5 (b) 5.6 (c) 6.6 (d) 7.6

Q.19 Calculate the pH of a buffer solution prepared by dissolving 10.6 g of Na_2CO_3 in 500 ml of an aqueous solution containing 80 ml of 1 M HCl. (K_a for $HCO_3^- = 6 \times 10^{-11}$)

(a) 8.6 (b) 9.6

(c) 11.6 (d) 12.6

RESPONSE GRID

6. ⓐⓑⓒⓓ 7. ⓐⓑⓒⓓ 8. ⓐⓑⓒⓓ 9. ⓐⓑⓒⓓ 10. ⓐⓑⓒⓓ
11. ⓐⓑⓒⓓ 12. ⓐⓑⓒⓓ 13. ⓐⓑⓒⓓ 14. ⓐⓑⓒⓓ 15. ⓐⓑⓒⓓ
16. ⓐⓑⓒⓓ 17. ⓐⓑⓒⓓ 18. ⓐⓑⓒⓓ 19. ⓐⓑⓒⓓ

Space for Rough Work

Q.20 What volume of 0.1 M HCOONa solution should be added to 50 ml of 0.05 M formic acid to produce a buffer solution of pH = 4.0? (pK_a of formic acid = 3.7)

(a) 50ml (b) 40ml (c) 30ml (d) 60ml

Q.21 Saccharin ($K_a = 2 \times 10^{-12}$) is a weak acid of formula HSaC. A amount 4×10^{4} mole of saccharin is dissolved in 200cc water of pH = 3. Assuming no change in volume, calculate the concentration of Sac. ions in the resulting solution at equilibrium.

(a) 2×10^{-12} M (b) 3×10^{-12} M

(c) 4×10^{-12} M (d) 5×10^{-12} M

DIRECTIONS (Q.22-Q.24) : In the following questions, more than one of the answers given are correct. Select the correct answers and mark it according to the following codes:

Codes :

(a) 1, 2 and 3 are correct **(b)** 1 and 2 are correct

(c) 2 and 4 are correct **(d)** 1 and 3 are correct

Q.22 Which of the following mixtures forms a basic buffer?

(1) $NH_4OH + NH_4NO_3$

(2) $NH_4OH + NH_4Cl$

(3) $H_2CO_3 + (NH_4)_2CO_3$

(4) $CH_3COOH + CH_3COONa$

Q.23 The solubility of a sparingly soluble salt A_xB_y in water at 25°C is 1.4×10^{-4} M. The solubility product is 1.1×10^{-11}. The possible values of x and y are :

(1) x = 1, y = 2 (2) x = 2, y = 1

(3) x = 0, y = 3 (4) x = 3, y = 0

Q.24 The solubility of silver acetate in pure water at 25°C is 8.35 g L^{-1} and 61.8 g L^{-1} in an acid solution of pH = 3. Choose the correct options-

(1) K_{sp} of silver acetate is 2.5×10^{-3}

(2) Dissociation constant of acetic acid is 1.8×10^{-5}

(3) K_{sp} of silver acetate is 2.5×10^{-2}

(4) Dissociation constant of acetic acid is 1.8×10^{-3}

DIRECTIONS (Q.25-Q.27) : Read the passage given below and answer the questions that follows :

pH of an acidic or a basic buffer can be calculated by Henderson-Hasselbalch equation.

For acidic buffers, $pH = pK_a + \log\frac{[salt]}{[acid]}$

When $\frac{[salt]}{[acid]} = 10$, then, $pH = 1 + pK_a$ and

When $\frac{[salt]}{[acid]} = \frac{1}{10}$, then $pH = pK_a - 1$

Q.25 Calculate the ratio of pH of a solution containing 1 mole of CH_3COONa + 1 mole of HCl per litre and of other solution containing 1 mole CH_3COONa + 1 mole of acetic acid per litre.

(a) 1 : 1 (b) 2 : 1

(c) 1 : 2 (d) 2 : 3

Q.26 Calculate the $pOH - pK_b$ for the buffer, 0.20 M NH_3, 0.40M NH_4Cl, K_b for ammonia = 10 -

(a) 0.50 (b) 0.60

(c) 0.30 (d) 0.75

Q.27 pH of 0.01 M HS^- will be -

(a) $pH = 7 + \frac{pK_a}{2} + \frac{\log C}{2}$

(b) $pH = 7 - \frac{pK_a}{2} - \frac{\log C}{2}$

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

(d) $pH = 7 + \left(\frac{pK_a - pK_b}{2}\right)$

RESPONSE GRID	20. ⓐⓑⓒⓓ	21. ⓐⓑⓒⓓ	22. ⓐⓑⓒⓓ	23. ⓐⓑⓒⓓ	24. ⓐⓑⓒⓓ
	25. ⓐⓑⓒⓓ	26. ⓐⓑⓒⓓ	27. ⓐⓑⓒⓓ		

Space for Rough Work

DIRECTIONS (Q. 28-Q.30) : Each of these questions contains two statements: Statement-1 (Assertion) and Statement-2 (Reason). Each of these questions has four alternative choices, only one of which is the correct answer. You have to select the correct choice.

(a) Statement-1 is True, Statement-2 is True; Statement-2 is a correct explanation for Statement-1.

(b) Statement-1 is True, Statement-2 is True; Statement-2 is NOT a correct explanation for Statement-1.

(c) Statement -1 is False, Statement-2 is True.

(d) Statement -1 is True, Statement-2 is False.

Q.28 Statement 1 : Sb (III) is not precipitated as sulphide when H_2S is passed in its alkaline solution.

Statement 2 : The concentration of S^{2-} ion in alkaline medium is inadequate for precipitation.

Q.29 Statement 1 : An ionic product is used for any types of electrolytes whereas solubility product is applicable only to sparingly soluble salts.

Statement 2 : Ionic product is defined at any stage of the reaction whereas solubility product is only applicable to the saturation stage.

Q.30 Statement 1 : $BaCO_3$ is more soluble in HNO_3 than in plain water.

Statement 2 : Carbonate is a weak base and reacts with the H^+ from the strong acid, causing the barium salt to dissociate.

RESPONSE GRID 28. ⓐⓑⓒⓓ 29. ⓐⓑⓒⓓ 30. ⓐⓑⓒⓓ

DAILY PRACTICE PROBLEM SHEET 14 - CHEMISTRY

Total Questions	30	Total Marks	120
Attempted		Correct	
Incorrect		Net Score	
Cut-off Score	36	Qualifying Score	60
Success Gap = Net Score – Qualifying Score			
Net Score = (Correct × 4) – (Incorrect × 1)			

Space for Rough Work

DAILY PRACTICE PROBLEMS

CHEMISTRY SOLUTIONS

14

(1) **(b)** $CaCO_3 \rightleftharpoons Ca^{2+} + CO_3^{2-}$

Let the solubility of $CaCO_3$ be 's' mole per litre

$\therefore K_{sp} = [Ca^{2+}][CO_3^{2-}] = s.s$

$\therefore s = \sqrt{K_{sp}} = \sqrt{9.3\times10^{-8}} = 0.000304$ mole/litre

Solubility in g/l = mole/litre × Molecular wt.

$= 0.000304 \times 100 = 0.0304$ gram / litre

(2) **(c)** For same concentration, the conductivity will depend on the number of ions produced in solution. Now,

$K_3[Fe(CN)_6] \xrightarrow{H_2O} 3K^+ + [Fe(CN)_6]^{3-} \Rightarrow$ 4 ions

$K_2[Ni(CN)_4] \xrightarrow{H_2O} 2K^+ + [Ni(CN)_4]^{2-} \Rightarrow$ 3 ions

$FeSO_4 \cdot Al_2(SO_4)_3 \cdot 24H_2O \xrightarrow{H_2O} Fe^{2+} + 4SO_4^{2-} + 2Al^{3+} \Rightarrow$ 7 ions

$Na[Ag(S_2O_3)_3] \xrightarrow{H_2O} Na^+ + [Ag(S_2O_3)_3]^-$

$\Rightarrow$ 2 ions

$\therefore FeSO_4 \cdot Al_2(SO_4)_3 \cdot 24H_2O$ gives maximum ions and hence shows maximum conductivity

(3) **(b)** Applying the equation,

$$pOH = \log\frac{[Salt]}{[Base]} - \log K_b$$

$$= \log\frac{0.25}{0.15} - \log 1.8\times10^{-5}$$

$$= \log 5 - \log 3 - \log 1.8\times10^{-5} = 4.96$$

$pH = (14 - 4.96) = 9.04$

(4) **(c)** Concentration of pyridinium chloride

$= 0.15\times2 = 0.3$ M.

$$pOH = \log\frac{[Salt]}{[Base]} - \log K_b$$

$$= \log\frac{0.3}{0.2} - \log 1.5\times10^{-9} = 9$$

$[OH^-] = 10^{-pOH} = 10^{-9}$

$pH = (14\text{-}pOH) = (14\text{-}9) = 5$

(5) **(c)** Solubility in moles per dm^3

$$= \frac{15.6\times10^{-3}}{78\,g/mole} = 2.0\times10^{-4}$$

$\because CaF_2 \rightleftharpoons Ca^{2+} + 2F^-$

$\therefore [Ca^{2+}] = 2.0\times10^{-4}$

and $[F^-] = 2\times2.0\times10^{-4}$

Hence, solubility product K_{sp}

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

$= [2.0\times10^{-4}][4.0\times10^{-4}]$

$= 32\times10^{-12}$

(6) **(b)** K_{sp} of $A_3B_2 = 2\times10^{-30}$.

$A_3B_2 \rightleftharpoons 3A^{2+} + 2B^{3-}$

Assuming 's' is the solubility of A_3B_2

then $K_{sp} = (3s)^3\,(2s)^2 = 108\,s^5$

$$s = \sqrt[5]{\frac{K_{sp}}{108}} \quad \therefore \quad s = \left(\frac{2\times10^{-30}}{108}\right)^{1/5}$$

(7) **(c)** HNO_2 is a weak acid and $NaNO_2$ is its salt with a strong base.

(8) **(d)** Solubility in moles $= \dfrac{0.0014}{143.5} = 0.9757\times10^{-5}$

Solubility product $K_{sp} = [Ag^+][Cl^-]$

$= s\times s = s^2 = (0.9757\times10^{-5})^2$

$= 0.952\times10^{-10}$

(9) **(c)** Solubility of M_2X_3 = x gm mole/litre

$M_2X_3 \rightleftharpoons 2M^{+3} + 3X^{-2}$

$\therefore [M^{+3}] = 2x$

$[X^{-2}] = 3x$

Solubility product $K_{sp} = (2x)^2.(3x)^3 = 108\,x^5$

(10) **(b)** $$\text{Solubility} = \frac{\text{Solubility product}}{\text{Concentration of common ion}}$$

$$\therefore S_1 = \frac{K_{sp}}{0.02} = 50\,K_{sp}$$

$$S_2 = \frac{K_{sp}}{0.01} = 100\,K_{sp}$$

$$S_3 = \frac{K_{sp}}{0.05} = 20\,K_{sp}$$

So, $S_2 > S_1 > S_3$

Again solubility will be greatest in water.

So, $S_0 > S_2 > S_1 > S_3$

(11) **(a)** In 0.1 M $AgNO_3$

$AgNO_3 \rightleftharpoons Ag^+ + NO_3^-$

$AgCl \rightleftharpoons Ag^+ + Cl^-$

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

Now $[Ag^+]$ can be taken as $[AgNO_3]$ while $[Cl^-]$ is the solubility of AgCl

$$\therefore Cl^- = \frac{K_{sp}}{[Ag^+]} = \frac{2.8\times10^{-10}}{0.1}$$

$\therefore$ Solubility of AgCl $= 2.8\times10^{-9}$ mole/litre

(12) **(b)** Using the expression

$$pH = pK_a + \log\frac{[Salt]}{[Acid]}$$

We get, $6 = -\log(1.0 \times 10^{-5}) + \log \dfrac{[\text{Salt}]}{[0.02\ \text{M}]}$

Which gives $6 = 5 + \log \dfrac{[\text{Salt}]}{[0.02\ \text{M}]}$

or $\dfrac{[\text{Salt}]}{[0.02\ \text{M}]} = 10$ or $[\text{Salt}] = 0.2\ \text{M}$

(13) **(b)** $\text{pH} = \text{pK}_a + \log \dfrac{[\text{Salt}]}{[\text{Acid}]}$

$= -\log(1.0 \times 10^{-5}) + \log \dfrac{(0.02 - 0.01)}{(0.03 + 0.01)}$

$= 5 + \log\left(\dfrac{1}{4}\right) = 5 - 0.6 = 4.4$

(14) **(a)** The addition of NaOH converts equivalent amount of acetic acid into sodium acetate. Hence,

Concentration of acetic acid after the addition of

$\text{NaOH} = \dfrac{30}{70} \times 0.2\text{M}$

Concen. of CH_3COONa after the addition of NaOH

$= \dfrac{20}{70} \times 0.2\ \text{M}$

Hence, using the expression

$\text{pH} = \text{pK}_a + \log \dfrac{[\text{Salt}]}{[\text{Acid}]}$

$= -\log(2 \times 10^{-5}) + \log\left(\dfrac{20}{30}\right)$

$= 4.699 - 0.177 = 4.522$

(15) **(b)** Assuming that the addition of HCOONa suppresses the ionization of HCOOH, we can use the expression

$\text{pH} = \text{pK}_a + \log \dfrac{[\text{Salt}]}{[\text{Acid}]}$

Since salt is 75% dissociated, we will get

$\text{pH} = -\log(2.4 \times 10^{-4}) + \log \dfrac{0.75}{0.2}$

$= 3.62 + 0.57 = 4.19$

(16) **(d)** The addition of HCl converts NaCN into HCN. Let x be the amount of HCl added. We will have.

$[\text{NaCN}] = (0.01 - x)$

$[\text{HCN}] = x$

Substituting these values along with pH and K_a in the expression.

$\text{pH} = -\log K_a + \log \dfrac{[\text{Salt}]}{[\text{Acid}]}$

We get $10.4 = -\log[4 \times 10^{-10}] + \log \dfrac{0.01 - x}{x}$

or $10.4 = 9.4 + \log \dfrac{0.01 - x}{x}$

or $\log \dfrac{0.01 - x}{x} = 1$

or $\dfrac{0.01 - x}{x} = 10$

$\Rightarrow 11x = 10^{-2}$ or $x = 9.9 \times 10^{-4}$ mole

(17) **(c)**

	HCl	+ NaOH →	NaCl	+ H_2O
$M_{eq.}$ before reaction	10×1 $= 10$	10×2 $= 20$	0	0
$M_{eq.}$ After reaction	0	10	10	10

$[OH^-]$ left from NaOH $= \dfrac{10}{20} = 0.5\text{M}$

$\text{pOH} = -\log[OH]^- = -\log 0.5$

$\text{pOH} = 0.3$

$\text{pH} = 13.7$

(18) **(a)** We have

$\text{pH} = -\log K_a + \log \dfrac{[\text{Salt}]}{[\text{Acid}]}$

$\therefore \text{pH} = -\log 2 \times 10^{-5} + \log \dfrac{\dfrac{0.2 \times 1000}{100}}{\dfrac{0.1 \times 1000}{100}} = 5$

(19) **(b)**

	Na_2CO_3	+ HCl →	NaCl	+ $NaHCO_3$
$M_{eq.}$ before reaction	$\dfrac{10.6}{106} \times 1000$ $= 100$	80×1 $= 80$	0	0
$M_{eq.}$ After reaction	20	0	80	80

The solution contains Na_2CO_3 and HCO_3^- and thus acts as buffer

$\text{pH} = -\log K_a + \log \dfrac{[CO_3^{2-}]}{[HCO_3^-]}$

$= -\log 6 \times 10^{-11} + \log \dfrac{20}{80} = 9.6$

(20) **(a)** Let V ml of 0.1 M HCOONa be mixed with 50 ml of 0.05 M HCOOH in mixture

$[\text{HCOONa}] = \dfrac{0.1 \times V}{(V + 50)}$

$[\text{HCOOH}] = \dfrac{50 \times 0.05}{(V + 50)}$

$\therefore \text{pH} = -\log K_a + \log \dfrac{[\text{Salt}]}{[\text{Acid}]}$

$4.0 = 3.7 + \log \dfrac{(0.1 \times V)/(V + 50)}{2.5/(V + 50)}$

V = 50 ml

(21) **(c)** $[\text{HSaC}] = \dfrac{4 \times 10^{-4}}{200/1000} = 2 \times 10^{-3}\text{M}$

	HSaC ⇌	H^+ +	SaC^-
Conc. before dissociation	2×10^{-3}	10^{-3}	0

In presence of H^+ the dissociation of HSaC is almost negligible because of common ion effect. Thus at equilibrium

$[HSaC] = 2 \times 10^{-3}$, $H^+ = 10^{-3}$

$$K_a = \frac{[H^+][SaC^-]}{[HSaC]}$$

$$\therefore 2 \times 10^{-12} = \frac{[10^{-3}][SaC^-]}{[2 \times 10^{-3}]}$$

$\therefore [SaC^-] = 4 \times 10^{-12}\,M$

(22) **(b)** Basic buffer is a solution of a mixture of a weak base and a salt of this weak base with a strong acid. Thus, $(NH_4NO_3 + NH_4OH)$ and $(NH_4OH + NH_4Cl)$ form a basic buffer.

(23) **(b)** $K_{sp} = 1.1 \times 10^{-11} = (1.4 \times 10^{-4})^{x+y} x^x.y^y$

So, we have put $x + y = 3$

$$\text{so, } x^x.y^y = \frac{1.1 \times 10^{-11}}{1.4 \times 1.4 \times 1.4 \times 10^{-12}} = \frac{110}{1.96 \times 1.4} = 4$$

Hence, $x = 1, y = 2$ or $y = 1, x = 2$

(24) **(b)** CH_3COOAg (mol. wt) = 167 and

$$S = \frac{8.35}{167} \text{ mole/litre} = 0.05\,M$$

$K_{sp} = S^2 = (0.05)^2 = 2.5 \times 10^{-3}$.

Use expression of salt hydrolysis to calculate K_a

$$pH = \frac{1}{2}[14 + pK_a + \log C] \text{ or } K_a = 1.8 \times 10^{-5}$$

(25) **(c)** **Case I** - pH when 1 mole CH_3COONa and 1 mole HCl are present

$$CH_3COONa + HCl \rightarrow CH_3COOH + NaCl$$

Before reaction	1	1	0	0
After reaction	0	0	1	1

$[CH_3COOH] = 1M$

$$\therefore [H^+]\,C\alpha = C.\sqrt{\frac{K_a}{C}} = \sqrt{C.K_a} = \sqrt{K_a}$$

$\because C = 1$

$$\therefore pH_1 = -\frac{1}{2}\log K_a$$

Case II - pH when 1 mole CH_3COONa and 1 mole of CH_3COOH, a buffer solution

$$\therefore pH_2 = -\log K_a + \log\frac{[Salt]}{[Acid]} = -\log K_a$$

$\because [Salt] = [Acid] = 1M$

$$\therefore \frac{pH_1}{pH_2} = \frac{1}{2} \text{ or } pH_1 : pH_2 = 1 : 2$$

(26) **(c)** $pOH = pK_b + \log\frac{[Salt]}{[Base]}$;

$$pOH - pK_b = \log\frac{0.4}{0.2} = 0.30$$

(27) **(a)** $HS^- + H_2O \rightleftharpoons H_2S + OH^-$

$$\therefore [OH^-] = Ch = \sqrt{\frac{K_w.C}{K_a}}$$

$$\therefore [H^+] = \frac{K_w}{\sqrt{\frac{K_w.C}{K_a}}} = \sqrt{\frac{K_w.K_a}{C}}$$

$$\text{or } pH = \frac{1}{2}[pK_w + pK_a + \log C] = \frac{14}{2} + \frac{pK_a}{2} + \frac{\log C}{2}$$

$$= 7 + \frac{pK_a}{2} + \frac{\log C}{2}$$

(28) **(d)** Sb (III) is precipitated in (group II B) by H_2S in acidic medium (dil HCl). In acidic medium, S^{2-} concentration decreases so that radicals of group IV are not precipitated. In alkaline medium, S^{2-} ion concentration increases and not decreases.

(29) **(b)** For sparingly soluble salts, reason is not a correct explanation. Hence both statement 1 and statement 2 are true but statement 2 is not a correct explanation of statement 1.

(30) **(a)** Barium carbonate is more soluble in HNO_3 than in water because carbonate is a weak base and reacts with the H^+ ion of HNO_3 causing the barium salt to dissociate.

$$BaCO_3 + 2HNO_3 \rightarrow Ba(NO_3)_2 + CO_2 + H_2O$$