Ionic Equilibrium

SOLUTIONS

Exercise-I (Conceptual Questions)

1. Ans. (2)

$M = \frac{w \times 1000}{M_w \times V_{ml}}$

$0.1 = \frac{w \times 1000}{98 \times 400}$

$w = 3.92g$

2. Ans. (2)

$N = n_f \times M$

$N = 2 \times 2 = 4N$

3. Ans. (2)

$pH = 3.31$

$[H^+] = 10^{-pH} = 10^{-3.31}$

$[H^+] = 5 \times 10^{-4}$

4. Ans. (2)

$[OH^-] = 5 \times 10^{-5}M$

$pOH = -\log(OH^-)$

$= -\log(5 \times 10^{-5})$

$= 5 - \log 5$

$pH = 14 - pOH$

$= 14 - (5 - \log 5)$

$= 9 + \log 5$

5. Ans. (4)

Basicity is equal to number of replaceable H–atom attached to more electronegative atoms.

For H_3PO_3 Basicity is 2

H_3PO_2 Basicity is 1

6. Ans. (4)

$[OH^-] = \frac{n}{V_{(L)}} = \frac{2 \times 10^{-3}}{2} = 10^{-3}M$

$pOH = -\log[OH^-] = -\log(10^{-3})$

$= 3$

$pH = 14 - pOH$

$= 14 - 3 = 11$

7. Ans. (2)

$pH = 4.4$

$[H^+] = 10^{-pH} = 10^{-4.4}$

$[H^+] = 3.9 \times 10^{-5}$

8. Ans. (1)

$M = \frac{n}{V_{(L)}} = \frac{w}{M_w \times V_{(L)}}$

$= \frac{8}{40 \times 1} = 0.2M$

9. Ans. (3)

$M = \frac{w \times 1000}{M_w \times V_{(ml)}}$

$M = \frac{w \times 1000}{60 \times 400} = 0.1$

$w = 0.60g$

10. Ans. (1)

Mili equivalent $= N \times V_{(ml)}$

$= 0.5 \times 100$

$= 50$

11. Ans. (1)

$MgCl_2 \rightarrow Mg^{+2} + 2Cl^-$

C – –

– C 2C

$[Cl^-] = 2C = 2 \times 1.5 \times 10^{-3}$

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

12. Ans. (4)

$\alpha = \sqrt{\frac{k_a}{c}} = \sqrt{\frac{10^{-5}}{0.1}}$

$\alpha = 10^{-2}$

13. Ans. (1)

$[H^+] = C\alpha$

$[H^+] = 0.1$

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

$= -\log(0.1)$

$= 1$

14. Ans. (1)

Ostwald's dilution law is not applicable for strong electrolyte because strong electrolytes are completely ionised.

15. Ans. (2)

The degree of ionisation of a compound depends upon nature of solute molecules.

16. Ans. (4)

pOH = 10

pH = 14 – pOH = 4

$[H^+] = 10^{-4}M = \sqrt{K_a \times C}$

$[H^+] = \sqrt{K_a \times C}$

$10^{-4} = \sqrt{K_a \times 10^{-2}}$

$K_a = 10^{-6}$

17. Ans. (1)

On dilution, 1.0M to 0.01M, percentage ionisation will increase.

18. Ans. (3)

On adding water, extent of ionisation of weak electrolytes increases.

19. Ans. (4)

$[H^+] = \sqrt{K_a \times C}$

$= \sqrt{4\times10^{-10}\times2.5\times10^{-1}}$

$= \sqrt{10^{-10}}$

$[H^+] = 10^{-5}$

pH = 5

20. Ans. (4)

pH = 2

$[H^+] = 10^{-2}M$

$[H^+] = \sqrt{K_a \times C}$

$10^{-2} = \sqrt{4.5\times10^{-4}\times C}$

C = 0.2222 M

21. Ans. (4)

$K_w = 10^{-14}$ at (25°C)

On increasing temperature, K_w will increase for example at 90°C, $K_w = 10^{-12}$

22. Ans. (1)

at 25°C

$[H^+] = 10^{-7}$ M

$[OH^+] = 10^{-7}$ M

So, $[H^+] + [OH^-]$

$= 2 \times 10^{-7}$

23. Ans. (1)

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

$$= \frac{10^{-14}}{55.5} = (55.5 \times 10^{14})^{-1}$$

24. Ans. (1)

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

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

$K_d \times [H_2O] = [H^+] \times [OH^-]$

$= K_w$

So, ionic product of water

= dissociation constant of water × H_2O

25. Ans. (3)

At, 90°C, $[H^+] = 10^{-6}M$

$[OH^-] = 10^{-6}M$

$[H^+] + [OH^-] = 10^{-6} + 10^{-6}$

$= 2 \times 10^{-6}$

26. Ans. (4)

For pure water $[H^+] = [OH^-]$

$= 10^{-6.7}$

So, $K_w = [H^+] \times [OH^-]$

$= 10^{-6.7} \times 10^{-6.7}$

$= 10^{-13.4}$

27. Ans. (1)

At 25°C, $[H^+] = 10^{-7}M$,

= pH = 7

On in Creasing temp, $[H^+]\uparrow$

So pH will $\downarrow$

At 373 K,

pH will be less than 7.

28. Ans. (4)

$$[H^+] = 10^{-7}M = \frac{N}{N_A \times V_{(L)}}$$

$N = 10^{-7} \times 6 \times 10^{23} \times 1$

$= 6 \times 10^{16}$

39. Ans. (1)

pH of 0.1 M NaCl will be less than 7 at 90°C

30. Ans. (1)

$\because K_a = 10^{-4}$

$pK_a = 4$

$K_b = 10^{-5}$

$pK_b = 5$

pH of 1M ammonium formate will be

$$= 7 + \frac{1}{2}pK_a - \frac{1}{2}pK_b$$

$$= 7 + \frac{1}{2}[4 - 5]$$

$= 7 – 0.5 = 6.5$

31. Ans. (1)

KCl is a salt of SA + SB, so it will not undergo hydrolysis.

32. Ans. (2)

$HCOO^- + H_2O \rightleftharpoons HCOOH + OH^-$

C	–	–	–
C – Ch	–	Ch	Ch

$K_h = \frac{[HCOOH][OH^-]}{[HCOO^-]} = \frac{Ch.\,Ch}{C - Ch} = Ch^2$

$h = \sqrt{\frac{K_h}{C}}$

33. Ans. (3)

pH of CH_3COONa will be more than 7.

34. Ans. (3)

pK_b of CN^- is 4.7

pK_a of HCN will be $14 - pK_b$

$= 14 - 4.7 = 9.3$

pH of 0.5 M NaCN is

$= 7 + \frac{1}{2} PK_a + \frac{1}{2} \log C$

$= 7 + \frac{1}{2}(9.3) + \frac{1}{2}\log(0.5)$

$= 7 + \frac{1}{2}(9.3) + (-\frac{1}{2}\log 2)$

$= 7 + \frac{1}{2}[(9.3 - \log 2)]$

$= 7 + \left(\frac{1}{2} \times 9\right) = 11.5$

35. Ans. (3)

Salt	pH
0.1 M NaCl	7
0.1 M NH_4Cl 0.1 M CH_3COONa	< 7 > 7
0.1 M CH_3COONH_4	May be less than or slightly more than 7

So, highest pH will be shown by 0.1 M CH_3COONa

36. Ans. (3)

K_2S is a salt of SB + WA,

So, pH will be more than 7.

37. Ans. (3)

For anionic hydrolysis, (WA + SB) type salt,

$pH = 7 + \frac{1}{2} pK_a + \frac{1}{2} \log C$

or

$\frac{1}{2} pK_w + \frac{1}{2} pK_a + \frac{1}{2} \log C$

38. Ans. (2)

Equilibrium constant for the reaction of WA + SB will be

$\frac{1}{K_H} = \frac{1}{(K_w / K_a)} = \frac{K_a}{K_w} = \frac{10^{-4}}{10^{-14}} = 10^{10}$.

39. Ans. (3)

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

C	–	–	–
C – Ch	–	Ch	Ch

$[OH^-] = Ch$

$K_h = \frac{K_w}{K_a} = Ch^2$

$h = \sqrt{\frac{K_w}{K_a \times C}}$

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

$= \sqrt{\frac{K_w \times C}{K_a}}$

40. Ans. (3)

$FeCl_3$ will be acidic & Ammonium acetate in water may be acidic or basic

41. Ans. (4)

Na_3PO_4 & CH_3COONa salt undergoes hydrolysis in water.

42. Ans. (4)

If resulting solution is Alkaline in nature. The salt is made up of WA + SB.

43. Ans. (4)

The solubility product of sparingly soluble AB type salt is defined as the product of ionic concentration a saturated solution.

44. Ans. (4)

K_{sp}	Salt
$4s^3$	M_2X
$4s^3$	QY_2
$4s^3$	PZ_2

If s is equal for M_2X, QY_2 & PZ_2 then K_{sp} will also be equal.

45. Ans. (3)

$Hg_2I_2 \rightarrow Hg_2^{+2} + 2I^-$

$K_{sp} = \left[Hg_2^{+2}\right]\left[I^-\right]^2$

46. Ans. (2)

K_{sp} is only affected by temperature.

47. Ans. (4)

$K_{sp} = 10^{-10} = s^2$

$s = 10^{-5}$ M

$\because BaSO_4 \rightleftharpoons Ba^{+2} + SO_4^{-2}$

– s s

$M = \frac{n}{V_{(L)}} = 10^{-5}$

$\Rightarrow \frac{w}{M_w \times V_{(L)}} = 10^{-5}$

$\Rightarrow \frac{1}{233 \times V_{(L)}} = 10^{-5}$

$V_{(L)} = 430$ L

48. Ans. (3)

$Ag_2CrO_4 \rightleftharpoons 2Ag^+ + CrO_4^{-2}$

2s s

$K_{sp} = [Ag^+]^2 [CrO_4^{-2}]$

$= (2s)^2 (s)$

$= 4s^3$

$[Ag^+] = 1.5 \times 10^{-4} = 2s$

$s = \frac{1.5 \times 10^{-4}}{2}$

$K_{sp} = 4\left(\frac{1.5}{2} \times 10^{-4}\right)^3$

$= \frac{27}{16} \times 10^{-12}$

$= 1.68 \times 10^{-12}$

49. Ans. (3)

$Ag_2CrO_4 \rightleftharpoons 2Ag^+ + CrO_4^{-2}$

2s s

$[CrO_4^{-2}] = s = 2 \times 10^{-4}$

$K_{sp} = 4s^3 = 4(2 \times 10^{-4})^3$

$= 32 \times 10^{-12}$

50. Ans. (3)

$s = \frac{n}{V_{(ml)}} \times 1000$

$= \frac{1.43 \times 10^{-4} \times 1000}{143 \times 100}$

$= 10^{-5}$M

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

– s s

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

$= s^2$

$K_{sp} = (10^{-5})^2 = 10^{-10}$

51. Ans. (4)

$As_2S_3 \rightleftharpoons 2As^{+3} + 3S^{-2}$

– 2s 3s

$K_{sp} = [As^{+3}]^2 [S^{-2}]^3$

52. Ans. (3)

$PbBr_2 \rightleftharpoons Pb^{+2} + 2Br^-$

s 2s

$K_{sp} = [Pb^{+2}] [Br^-]^2$

$= s(2s)^2$

$= 4s^3$

53. Ans. (4)

$K_{sp} = (3)^3(3)^3(2^2)s^{3+3+2}$

$= 2916s^8$

54. Ans. (1)

$CaCO_3 \rightleftharpoons CaO + CO_2$

– s s

$[CaCO_3] = \frac{7}{100 \times 1} = 7 \times 10^{-2}$ M = s

$K_{sp} = s^2 = (7 \times 10^{-2})^2$

$= 4.9 \times 10^{-3}$

55. Ans. (2)

AgBr in 0.1 M $CaBr_2$

$AgBr \rightleftharpoons Ag^+ + Br^-$

– s' s'

$CaBr_2 \rightleftharpoons Ca^{+2} + 2Br^-$

– 0.1 0.2

$[Br^-] = s' + 0.2 \approx 0.2$

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

$K_{sp} = (s') (0.2)$

$s' = \frac{K_{sp}}{0.2}$

56. Ans. (2)

$AgSCN \rightleftharpoons Ag^+ + SCN^-$

If solution is unsaturated

$Q < K_{sp}$

$[Ag^+]\,[SCN^-] < K_{sp}$

57. Ans. (3)

$AB \rightleftharpoons A^+ + B^-$

– s s

$K_{sp} = s = [A^+]\,[B^-]$

$S = s \cdot s$

$S = s^2$

$s = S^{\frac{1}{2}}$

58. Ans. (4)

	K_{sp}	
CuS	$K_{sp} = s_1^2$	$s_1^2 = 10^{-37}$
Ag_2S	$K_{sp} = 4s_2^3$	$4s_2^3 = 10^{-44}$
HgS	$K_{sp} = s_3^2$	$s_3^2 = 10^{-54}$

Order :- $s_2 > s_1 > s_3$

$Ag_2S > CuS > HgS$

59. Ans. (4)

$PbCl_2 \rightleftharpoons Pb^{+2} + 2Cl^-$

– s 2s

(s) = 0.01 M

$K_{sp} = 4s^3 = 4(0.01)^3$

$= 4 \times 10^{-6}$

Solubility of $PbCl_2$ in 0.1 M NaCl will be

$s' = \frac{K_{sp}}{(0.1)^2} = \frac{4\times10^{-6}}{(0.1)^2}$

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

60. Ans. (2)

$M_2SO_4 \rightleftharpoons 2M^+ + SO_4^{-2}$

– 2s s

$K_{sp} = [M^+]^2\,[SO_4^{-2}]$

$= (2s)^2\,(s)$

$= 4s^3 = 1.2 \times 10^{-5}$

$= s^3 = 3 \times 10^{-6}$

$s = (3)^{\frac{1}{3}} \times 10^{-2}$

$[M^+] = 2s = 2(3)^{\frac{1}{3}} \times (10^{-2})$

$= 2.89 \times 10^{-2}M$

61. Ans. (2)

	K_{sp}	Solubility
HgS	$s^2 = 1.6 \times 10^{-54}$	$s \approx 10^{-27}$
$PbSO_4$	$s^2 = 1.3 \times 10^{-8}$	$s \approx 10^{-4}$
ZnS	$s^2 = 7 \times 10^{-26}$	$s \approx 10^{-13}$
AgCl	$s^2 = 1.7 \times 10^{-10}$	$s \approx 10^{-5}$

Hence $PbSO_4$ has maximum solubility.

62. Ans. (4)

	$K_{sp} = s^2$	Solubility
CuS	8.5×10^{-36}	$s \approx 10^{-18}$
CdS	3.6×10^{-28}	$s \approx 10^{-14}$
ZnS	1.2×10^{-28}	$s \approx 10^{-14}$
MnS	1.4×10^{-10}	$s \approx 10^{-5}$

Solubility of MnS in maximum.

63. Ans. (2)

Solubility of AgCl will be maximum in water. In $AgNO_3$, NaCl & KCl, due to common ion solubility of AgCl will decrease.

64. Ans. (4)

	K_{sp}	
PQ	$s_1^2 = 4 \times 10^{-20}$	$s_1 = 2 \times 10^{-10}$
PQ_2	$4s_2^3 = 3.2 \times 10^{-14}$	$s_2 = 2 \times 10^{-5}$
PQ_3	$27s_3^4 = 2.7 \times 10^{-35}$	$s_3 = 10^{-9}$

$s_2 > s_3 > s_1$

65. Ans. (3)

$Gd(OH)_3 \rightleftharpoons Gd^{+3} + 3OH^-$

s 3s

$K_{sp} = 27s^4 = 2.8 \times 10^{-23}$

$s \approx 10^{-6}$

$[OH^-] = 3s = 3 \times 10^{-6}$

$pOH = 6 - \log 3$

$pH = 14 - pOH$

$= 14 - (6 - \log 3)$

$= 8 + \log 3$

$= 8.47$

66. Ans. (3)

K_{sp} of $AgBrO_3 = s_1^2 = 5.5 \times 10^{-5}$

$= s_1 = 7.5 \times 10^{-3}$

K_{sp} of $Ag_2SO_4 = 4s_2^3 = 2 \times 10^{-5}$

$= s_2 \simeq 10^{-2}$

$s_2 > s_1$

Hence $s_{Ag_2SO_4} > s_{AgBrO_3}$

67. Ans. (1)

For precipitation to begin

$Q > K_{sp}$

$Mg(OH)_2 \rightleftharpoons Mg^{+2} + 2OH^-$

0.1 M

$[Mg^{+2}]\,[OH^-]^2 \geq K_{sp}$

$(0.1)\quad (OH^-)^2 \geq 10^{-11}$

$[OH^-]^2 \geq \dfrac{10^{-11}}{0.1}$

$[OH^-]^2 \geq 10^{-10}$

$[OH^-] > 10^{-5}$

$pOH \leq 5$

$pH \geq 9$

68. Ans. (2)

For ZnS

$ZnS \rightleftharpoons Zn^{+2} + S^{-2}$

$Q = [Zn^{+2}]\,[S^{-2}]$

$= [0.01]\,[8.1 \times 10^{-21}]$

K_{sp} of $ZnS = 3 \times 10^{-22}$

$Q < K_{sp}$ No precipitation similarly for CuS

$CuS \rightleftharpoons Cu^{+2} + S^{-2}$

$Q = [Cu^{+2}]\,[S^{-2}]$

$Q = (0.01)\,(8.1 \times 10^{-21})$

$= 8.1 \times 10^{-23}$

K_{sp} of $CuS = 8 \times 10^{-36}$

$Q > K_{sp}$

Hence CuS will precipitate.

69. Ans. (2)

$Mg(OH)_2 \rightleftharpoons Mg^{+2} + 2OH^-$

$pH = 9$

$pOH = 14 - 9 = 5$

$[OH^-] = 10^{-5}$

$[Mg^{+2}] = 0.001$ M (from $Mg(NO_3)_2$)

$Q = [Mg^{+2}]\,[OH^-]^2$

$= [0.001]\,(10^{-5})^2$

$= 10^{-13}$

K_{sp} of $Mg(OH)_2 = 8.9 \times 10^{-12}$

$Q < K_{sp}$

$\therefore$ Precipitation will not take place.

70. Ans. (4)

To prevent precipitation

$Q \leq K_{sp}$

$Mg(OH)_2 \rightleftharpoons Mg^{+2} + 2OH^-$

$Q = [Mg^{+2}]\,[OH^-]^2 \leq 9 \times 10^{-12}$

$(0.01)\,(OH^-)^2 \leq 9 \times 10^{-12}$

$[OH^-]^2 \leq 9 \times 10^{-10}$

$[OH^-] \leq 3 \times 10^{-5}$

So, maximum concentration of OH^- is 3×10^{-5}

71. Ans. (3)

In the presence of HCl, pure NaCl is precipitated because due to common ion $[Cl^-]$, concentration of $[Cl^-]$ increases. & hence ionic product Q exceeds the solubility product of NaCl.

72. Ans. (2)

$[H^+]_1 = [H^+]_2$

$\sqrt{K_{a_1} C_1} = \sqrt{K_{a_2} C_2}$

$K_{a_1} C_1 = K_{a_2} C_2$

$\boxed{C \propto \dfrac{1}{V}}$

So; $K_{a_1} V_2 = K_{a_2} V_1$

73. Ans. (2)

Solutions having same concentration of H^+ ions are called isohydric solutions.

74. Ans. (2)

$[H^+] = C\alpha$

$= (0.1)\left(\dfrac{1}{100}\right)$

$= 10^{-3}$

$pH = 3$

75. Ans. (4)

If pH of Y is more than 7 when it is dissolved in water then Y is a salt of weak acid & strong base.

76. Ans. (3)

$BeCl_2$:- Salt of WB + SA

$pH < 7$

$BaCl_2$:- Salt of SB + SA

$pH = 7$

$Ba(NO_3)_2$:- Salt of SB + SA

$pH = 7$

$Ba(OH)_2$:- strong base

$pH > 7$

So minimum pH will be shown by 0.1 M $BeCl_2$

77. Ans. (3)

0.005 M $H_2SO_4 \Rightarrow [H^+] = 0.005 \times 2$
$= 0.01$ N

0.01 M HCl $\Rightarrow [H^+] = 0.01 \times 1$
$= 0.01$ N

$[H^+]$ for both is same & hence pH will also be same.

78. Ans. (1)

$H_2CO_3 \rightleftharpoons 2H^+ + CO_3^{-2}$ ($\because \alpha = 10\% = 0.1$)

C – Cα 2Cα Cα

$[H^+] = 2C\alpha$
$= 2 \times 5 \times 10^{-3} \times (0.1) = 10^{-3}$

79. Ans. (4)

$M(OH)_4 \rightleftharpoons M^{+4} + 4OH^-$ ($\because \alpha = 50\% = 0.5$)

C – Cα Cα 4Cα

$[OH^-] = 4C\alpha = 4 \times 0.0025 \times 0.5$
$= 5 \times 10^{-3}$

$pOH = 3 - \log 5$
$pH = 14 - pOH$
$= 14 - (3 - \log 5)$
$= 11 + \log 5 = 11.7$

80. Ans. (3)

$\frac{M}{20}$ H_2SO_4, $\frac{M}{10}$ HNO_3, $\frac{M}{40}$ $HClO_4$ will have highest pH.

$$[H^+] = \frac{\sum NV}{\sum V} = \frac{\left[2 \times \frac{1}{20} \times 1\right] + \left[\frac{1}{10} \times 1\right] + \left[\frac{1}{40} \times 1\right]}{1+1+1}$$

$$= \frac{\frac{1}{10} + \frac{1}{10} + \frac{1}{40}}{3} = \frac{9}{120}$$

$$[H^+] = \frac{3}{40}$$

$$pH = \log[H^+] = -\log\left[\frac{3}{40}\right]$$

$= \log 40 - \log 3$
$= 1.6 - 0.48$
$= 1.12$

81. Ans. (2)

$pH = 3, \Rightarrow [H^+]_1 = N_1 = 10^{-3}$
$pH = 3 \Rightarrow [H^+]_2 = N_2 = 10^{-3}$

$$N_{mixture} = \frac{N_1V_1 + N_2V_2}{V_1 + V_2} = [H^+]_{mixture}$$

$$= \frac{10^{-3} \times 100 + 10^{-3} \times 400}{500}$$

$$\frac{5 \times 10^{-3}}{5} = 10^{-3} = [H^+]$$

$pH = -\log[H^+] = 3$

82. Ans. (3)

10^{-6} M = N,

$N_2 = \frac{N_1}{100}$ [on dilution, concentration decreases]

$N_2 = 10^{-8}$N

pH of 10^{-8} M HCl {since its concentration is less than 10^{-7} M, $[H^+]$ ions from water must be added}

$[H^+] = 10^{-8} + 10^{-7}$
$= 10^{-8}[1 + 10]$
$= 11 \times 10^{-8}$

$pH = -\log [H^+] = 8 - \log 11 = 6.95$

83. Ans. (4)

For 0.001M KOH $\Rightarrow [OH^-] = 10^{-3}$M,

$pOH = -\log[OH^-] = 3$

at 90°C pH + pOH = 12

$pH = 12 - pOH = 12 - 3 = 9$

84. Ans. (2)

pH of 0.001 M acetic acid would be more than 3; since it is a weak acid, 100% dissociation will not take place.

85. Ans. (3)

$pH = 3, [H^+] = 10^{-3}M = N_1$

$pH = 6, [H^+] = 10^{-6}M = N_2 \Rightarrow N_2 = \frac{1}{1000} \times N_1$

if pH is changed from 3 to 6, $[H^+]$ ion concentration reduced by 1000 times.

86. Ans. (1)

pOH = 13

$pH = 14 - pOH$ at 298 K

$pH = 14 - 13 = 1$

so, solution will be highly acidic.

87. Ans. (4)

Mixture of strong acid & strong base :-

$\Sigma N_AV_A = 10 \times 0.05 \times 2 = 1$

$\Sigma N_BV_B = 10 \times 0.01 = 1$

If $\Sigma N_AV_A = \Sigma N_BV_B$: solution will be neutral, pH will be 7.

88. Ans. (2)

$pH = 5, [H^+] = 10^{-5}N = N_1$

$pH = 2 \Rightarrow [H^+] = 10^{-2}N = N_2$

$\frac{N_2}{N_1} = \frac{10^{-2}}{10^{-5}} = 10^3$

$N_2 = 1000N_1$

Increase in Hydrogen ion concentration is 1000 times.

89. Ans. (2)

$[H^+] = 6 \times 10^{-4}$

$pH = -\log[H^+] = 4 - \log 6 = 3.22$

90. Ans. (1)

$pOH = 10 \Rightarrow [OH^-] = 10^{-10}$

$pH = 14 - pOH = 4 \Rightarrow [H^+] = 10^{-4}M$

$K_w = 10^{-14} \Rightarrow [H^+][OH^-] = K_w \Rightarrow [H^+] = \frac{K_w}{[OH^-]}$

So, $[H^+] = \frac{K_w}{10^{-10}} = 10^{-4}$

91. Ans. (2)

$pH = 0, [H^+] = 10^{-pH} = 1$ M.

So solution is acidic

92. Ans. (3)

$N_{resultant} = \frac{\text{Total g. eq of KOH}}{\text{Total volume}}$ n_f of KOH = 1

$= \frac{0.1+0.2+0.3+0.4+0.5}{1+2+3+4+5} = 0.1$

$[OH^-] = 0.1$ N

$pOH = -\log(0.1) = 1$

$pH = 14 - pOH = 13$

93. Ans. (2)

pH of 0.02M NH_4OH, $\alpha = 5\%$

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

C 0 0

$C - C\alpha$ $C\alpha$ $C\alpha$

$[OH^-] = C\alpha$

$= 0.02 \times \frac{5}{100} = 10^{-3}M$

$pOH = -\log[OH^-] = 3$

$pH = 14 - 3 = 11$

94. Ans. (1)

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

$= -\log\left(\frac{1}{10}\right) = 1$

95. Ans. (2)

10^{-9}M HCl is very dilute solution, so $[H^+]$ ion form water must be considered.

$[H^+] = [H^+]_{HCl} + [H^+]_{water}$

$= 10^{-9} + 10^{-7}$

$= 101 \times 10^{-9}$

$pH = 9 - \log(101) = 9 - 2.004 = 6.99$

so pH is between 6 and 7.

96. Ans. (3)

$pH = 2$, $[H^+] = 10^{-2}M$

$pH = 3$; $[H^+] = 10^{-3}M$

Number of moles of HCl, which must be removed from 1L solution will be

$= M_1V_1 - M_2V_2$

$= 10^{-2} \times 1 - 10^{-3} \times 1$

$= (10^{-2})\left[\left(1 - \frac{1}{10}\right)\right] = 0.009$

97. Ans. (2)

$N_{NaOH} = \frac{8}{40 \times 1} = 0.2$ N

$N_{H_2SO_4} = \frac{4.9}{49 \times 1} = 0.1$ N

$N_AV_A = 0.1 \times 1 = 0.1$

$N_BV_B = 0.2 \times 1 = 0.2$ $V_A + V_B = 1L = V_T$

$N_BV_B > N_AV_A$

$[OH^-] = \frac{N_BV_B - N_AV_A}{V_A + V_B} = \frac{0.2 - 0.1}{1} = \frac{0.1}{1}$

$[OH^-] = 0.1$ N

$pOH = -\log[OH^-] = 1$

$pH = 14 - pOH = 14 - 1 = 13$

98. Ans. (3)

$[NaOH] = [OH^-] = \frac{w \times 1000}{E_w \times V_{ml}} = \frac{0.2 \times 1000}{40 \times 100}$

$= 0.05M = 5 \times 10^{-2}m$

$pOH = -\log[OH^-] = 2 - \log 5$

$pH = 14 - pOH = 14 - (2 - \log 5)$

$= 12.699$

99. Ans. (3)

$pH = 13$

$pOH = 14 - pH = 14 - 13 = 1$

$[OH^-] = 10^{-1}M$

$M = \frac{w \times 1000}{M_w \times V_{ml}}$

$10^{-1} = \frac{w \times 1000}{40 \times 250}$

$w = 1g$

100. Ans. (1)

$M(OH)_2 \rightleftharpoons M^{+2} + 2OH^-$

0.001 – –

– 0.001 $2 \times 0.001 = 0.002$ mole

$[OH^-] = \frac{n}{V_{ml}} \times 1000$

$= \frac{0.002}{20} \times 1000$

$[OH^-] = 0.1$

$pOH = -\log[OH^-] = 1$

$pH = 14 - pOH$

$= 14 - 1 = 13$

101. Ans. (4)

$[H^+] = [OH^-] = 10^{-7}M$ (at all temperature) is wrong statement

102. Ans. (4)

0.1M HCl; pH = 1

0.1M NaCl pH = 7

0.1M NH_4Cl pH < 7

0.1M NaCN pH > 7

So, pH order

$HCl < NH_4Cl < NaCl < NaCN$

103. Ans. (2)

$pOH = pK_b + \log \frac{[Salt]}{[Base]}$

when $\frac{[Salt]}{[Base]} = 1 : 1$

$pOH = pK_b$ ...(1)

when $\frac{[Salt]}{[Base]} = 2 : 1$

$pOH = pK_b + \log 2$ ...(2)

pOH increases by log2 ; so pH of buffer will decrease.

104. Ans. (3)

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

let V_{ml} of sodium formate is added to 50 ml of 0.05M formic acid.

$[Salt] = \frac{V \times 0.1}{[50+V]}$

$[Acid] = \frac{50 \times 0.05}{(50+V)}$

$pH = pK_a + \log\left[\frac{0.1 \times V}{50 \times 0.05}\right]$

$4 = 3.8 + \log\left[\frac{V}{25}\right]$

$\log\left(\frac{V}{25}\right) = 0.2$

$\frac{V}{25} = 10^{0.2}$

$V_{(ml)} = 39.6ml$

105. Ans. (4)

Phenolphthalein as an indicator is used when base is strong. So it is unsuitable for NH_4OH.

106. Ans. (1)

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

let initially $\frac{[Salt]}{[Acid]} = 1$

$pH = pK_a$

if $\frac{[Salt]}{[Acid]} = \frac{1}{10}$

$pH = pK_a - \log 10$

$pH = pK_a - 1$

So, pH of the solution decrease by one.

107. Ans. (4)

$pOH = pK_b + \log \frac{[Salt]}{[Base]}$

if [salt] = [base]

then

$pOH = pK_b$

$pOH = 4.74 = pK_b$

$pH = 14 - pOH = 14 - 4.74$

$= 9.26$

108. Ans. (3)

For (SB & WA); Phenolphthalein is used as an indicator.

109. Ans. (3)

Phenolphthalein does not act as an indicator for the titration between oxalic acid & $KMnO_4$.

110. Ans. (3)

	NaCN	+ HCl	⟶ HCN	+ NaCl
Mili eq.	40×0.1	20×0.1		
	$= 4$	$= 2$		
	$4 - 2 = 2$	–	2	2

(HCN + NaCN) is an Acidic buffer

111. Ans. (3)

Buffer solution resists the change in pH.

112. Ans. (3)

let V_{mL} of KCN is required.

Then $[KCN] = \frac{5\times V}{(10+V)}$

$[HCN] = \frac{10\times 2}{(10+V)}$

$pH = pK_a + \log\left[\frac{[KCN]}{[HCN]}\right]$

$9 = (10 - \log 5) + \log\left(\frac{5V}{20}\right)$

$\log\left(\frac{4}{V}\right) = 0.3 = \log(2)$

$\frac{4}{V} = 2$

$V = 2mL$

113. Ans. (3)

$pH = pK_a + \log\frac{[Salt]}{[Acid]}$ {∵ max pH = pK_a + 1}

pH of buffer is maximum, when $\frac{[Salt]}{[Acid]} = \frac{10}{1}$

114. Ans. (1)

Pink colour of Phenolphthalein in alkaline medium is due to Negative ion, since in Alkaline medium $[H^+]$ is neutrilised by $[OH^-]$ produced by base.

115. Ans. (1)

Phenolphthalein works in pH range 8 – 9.8.

116. Ans. (3)

$pOH - pK_b = 1$

$pOH = 1 + pK_b$

when $\frac{[Salt]}{[Base]} = 10$

or

$\frac{[Conjugate\ Acid]}{[Base]} = \frac{10}{1}$

117. Ans. (4)

For weak acid – strong base titration, phenolphthalein is used.

118. Ans. (3)

Methyl orange is used for titration of strong acid & weak base, also for strong acid & strong base titration.

119. Ans. (1)

H_3PO_4 + NaOH, total 3 buffer are obtained

(a) $H_3PO_4 + NaH_2PO_4$;

(b) $H_3PO_4 + Na_2HPO_4$;

(c) $H_3PO_4 + Na_3PO_4$;

120. Ans. (1)

If K_b of X^- is 10^{-10}

Then pK_a of HX will be $\frac{10^{-14}}{10^{-10}} = 10^{-4}$

$pH = pK_a + \log\frac{[X^-]}{[HX]}$

if $\frac{[X^-]}{[HX]} = 1$

then $pH = pK_a = 4$

121. Ans. (2)

pH of buffer solution does not change if small amount of acid or base is added.

122. Ans. (2)

Buffer action is to maintain pH if small amount of acid or base is added.

123. Ans. (4)

Phenolphthalein is not a good indicator for titration of SA & WB.

124. Ans. (3)

NH_4Cl + HCl does not Act as buffer because HCl is strong acid.

125. Ans. (1)

$NaOH + CH_3COOH \rightleftharpoons CH_3COONa + H_2O$

initial mole – 100

60%

neutralisation – 100 – 60 = 40 60 60

$[CH_3COONa] = 60$

$[CH_3COOH] = 40$

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

$pH = 4.7 + \log\left(\frac{60}{40}\right)$

$pH = 4.7 + 0.18$

$= 4.88$

126. Ans. (1)

$$[CH_3COOH] = \frac{500\times0.2}{[500+500]} = 0.1M$$

$$[CH_3COONa] = \frac{500\times0.3}{[500+500]} = 0.15M$$

$$pH = pK_a + \log\frac{[CH_3COONa]}{[CH_3COOH]}$$

$$pH = 5 - \log(1.5) + \log\left(\frac{0.15}{0.10}\right)$$

$$pH = 5$$

127. Ans. (1)

At half neutralisation [salt] = [acid]

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

$$pH = 4 - \log2 + \log(1)$$

$$pH = 4 - \log2 = 3.6990$$

128. Ans. (1)

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

[salt] = 0.2 M, [base] = 0.2 M

$$pOH = pK_b$$

$$[OH^-] = K_b = 2 \times 10^{-5}$$

129. Ans. (2)

$$pH - pK_a = 1$$

$pH = 1 + pK_a$ is applicable when

[salt] = 10[acid]

Or

[Conjugate base] = 10 × [acid]

130. Ans. (2)

$$[OH^-] = \frac{K_b\times[base]}{[salt]}$$

$$= \frac{1.8\times10^{-5}\times0.05}{0.001} = 9 \times 10^{-4}$$

131. Ans. (1)

Buffer capacity

$$= \frac{\text{number of moles of } OH^- \text{ or } H^+}{\text{Change in pH}}$$

$$= \frac{0.02}{0.05} = 0.4$$

132. Ans. (3)

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

$$pOH = [5 - \log(1.8)] + \log\frac{[N_SV_S]}{[N_BV_B]}$$

$$= 5 - \log(1.8) + \log\left(\frac{500\times0.5}{300\times0.3}\right)$$

$$= 5 - \log(1.8) + \log\left(\frac{5}{3}\right)^2$$

$$= 5 - \log(1.8) + 2[\log5 - \log3]$$

$$= 5 - \log(1.8) + 2(0.7 - 0.48)$$

$$pH = 14 - pOH$$

$$= 14 - 5 + \log(1.8) - 2\,(0.22)$$

$$pH = 8.81$$

133. Ans. (3)

Case- (I)

$CH_3COONa + HCl \rightleftharpoons CH_3COOH + NaCl$

1	1	–	–
–	–	1	1

pH of 1 mole CH_3COOH in 1 L

$$pH = \frac{1}{2}pK_a - \frac{1}{2}\log C$$

$$= \frac{1}{2}pK_a - \frac{1}{2}\log 1$$

$$(pH)_1 = \frac{1}{2}pK_a \qquad \ldots(1)$$

Case-(II)

1 mole CH_3COONa + 1 mole CH_3COOH is a buffer solution.

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

$$= pK_a + \log 1$$

$$(pH)_2 = pK_a$$

$$\frac{(pH)_1}{(pH)_2} = \frac{\frac{1}{2}pK_a}{pK_a} = 1:2$$

134. Ans. (4)

20 mL $\frac{M}{20}$ NaOH + 10 mL $\frac{M}{10}$ HCl

will give NaCl solution.

which is neutral (pH = 7)

135. Ans. (1)

On adding 1 mL water, total volume will change but concentration of salt & base will remain same. Hence pH would not change.

136. Ans. (1)

$CH_3COOH + NaOH \rightleftharpoons CH_3COONa + H_2O$

	100	–	–	–
(25% titration)	100–25=75	–	25	25
(50% titration)	100–50=50	–	50	50
(75% titration)	100–75=25	–	75	75

At 25% titration

$[CH_3COOH] = 75\ M \qquad [CH_3COONa] = 25\ M$

$$pH = pK_a + \log\frac{[CH_3COONa]}{[CH_3COOH]} = pK_a + \log\left(\frac{25}{75}\right)$$

$$= pK_a + \log\left(\frac{1}{3}\right)$$

Similarly at 50% titration

$pH = pK_a + \log(1) = pk_a$

at 75% titration,

$$pH = pK_a + \log\left(\frac{75}{25}\right) = pK_a + \log 3.$$

137. Ans. (4)

$O^{-2} + H^+ \rightarrow OH^-$

Conjugate base of O^{-2} is OH^-

138. Ans. (2)

If $pK_{b_1} < pK_{b_2}$,

Then BOH is weak base so conjugate acid of BOH will be strong & AOH will be weak, so conjugate of BOH does not show maximum pH

139. Ans. (1)

H_2O can act as Lewis base, Bronsted acid & Bronsted base.

140. Ans. (4)

ClO_3^- does not acidic behaviour.

141. Ans. (3)

SO_3 is Lewis acid.

142. Ans. (1)

H_2O act as Bronsted acid.

143. Ans. (1)

PH_3 will act as Lewis base due to presence of lone pair.

144. Ans. (2)

$SnCl_2$ accepts e^- pair so it act as Lewis acid.

145. Ans. (4)

$$(CH_3)_2\overset{\oplus}{N}H_2 \rightarrow (CH_3)_2NH + H^+$$

Acid (Conjugate base)

146. Ans. (4)

$Cu^{+2} + 4NH_3 \rightleftharpoons [Cu(NH_3)_4]^{+2}$

This is Lewis acid base reaction but not Bronsted acid base reaction, because there is no H^+ loss or gain.

147. Ans. (2)

$N_3H \rightleftharpoons N_3^- + H^+$

(Hydrazoic acid) (Conjugate base)

148. Ans. (1)

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

In this reaction H_2O act as an acid.

149. Ans. (2)

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

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

Due to common ion effect, equilibrium shift in backward direction, concentration of H^+ will decrease.

150. Ans. (1)

$ClO_3(OH)$ or $HClO_4$ is strongest acid

151. Ans. (2)

CH_3COO^- is not a Bronsted acid.

152. Ans. (4)

BF_3, $SnCl_2$, $SnCl_4$ All three are Lewis acid due to availability of vacant orbitals.

153. Ans. (3)

NH_3^- is conjugate base of HNO_3.

154. Ans. (2)

$HBr + H_2O \rightleftharpoons H_3O^+ + Br^-$

S. Acid S. Base W. Acid W. Base

So, conjugate of weak acid is H_2O.

155. Ans. (4)
$AlCl_3$ act as Lewis acid.

156. Ans. (3)
Mg^{+2} weaker Lewis acid than Al^{+3} because charge on Mg is less than charge on Al.

157. Ans. (3)
$C_2O_4^{-2}$ & PO_4^{-3} are two Bronsted conjugate base of $HC_2O_4^{-2}$ & HPO_4^{-2} respectively.

158. Ans. (4)
$HCl + HF \rightleftharpoons H_2^+Cl + F^-$
W.B W. A S. A W. B
Here HCl act as weak base.

159. Ans. (4)
I^+ being a cation is not a Lewis base.

160. Ans. (2)
H_3O^+ is a Bronsted Lowry acid.

161. Ans. (1)
Conjugate base for Bicarbonate ion HCO_3^- is CO_3^{-2}.

162. Ans. (4)
HCl does not behave as acid in C_6H_6 (non-polar solvent)

163. Ans. (1)
I^- is a base according to Bronsted – Lowry concept.
$I^- + H^+ \rightarrow HI$
Base C. acid

164. Ans. (3)
$NH_3 + Na \rightarrow NaNH_2 + \frac{1}{2}H_2$
in this reaction NH_3 act as acid because it releases H atom to combine with Na.

165. Ans. (3)
$NH_3 + H_2O \rightleftharpoons NH_4^+ + OH^-$
↓ ↓
Acid Acid
H_2O & NH_4^+ are two acid.

166. Ans. (2)
CH_3COO^- ion is a strong conjugate base of CH_3COOH, which is a weak acid itself.

167. Ans. (2)
HCO_3^- is strongest conjugate base of H_2CO_3 which is a weak acid itself.

168. Ans. (3)
NH_3 is Lewis base because it has electron pair to donate.

169. Ans. (3)
NH_2CH_2COOH may behave both as an acid & base because N atom has e^- pair to donate & –COOH group can release H^+.

170. Ans. (4)
NH_3 can act both as Bronsted acid & Bronsted base.
$NH_3 \longrightarrow NH_2^- + H^+$
Acid
$NH_3 + H^+ \longrightarrow NH_4^+$
Base

171. Ans. (4)
CH_3COOH is a weak acid so CH_3COO^- is strongest conjugate base.

172. Ans. (3)
$AlCl_3$ is Lewis acid because it has vacant p-orbital.

173. Ans. (3)
Water is Amphiprotic solvent because it can accept H^+ or donate H^+.

174. Ans. (1)
NH_4^+ is a conjugate acid of NH_3

175. Ans. (2)
Na_2CO_3 can not act as Bronsted acid & base both.

176. Ans. (1)
NH_3 is strong Lewis base, Bronsted acid & Bronsted base.

177. Ans. (1)
Arrhenius theory explain the acidic or basic nature of the substance in water.

178. Ans. (2)
NH_3 & S^{-2} are conjugated base of NH_4^+ & HS^- respectively.

179. Ans. (3)
As per theory

180. Ans. (2)
As per theory

Exercise-II (Previous Year Questions)

1. Ans. (3)

KCl	SASB	pH = 7
NaCl	SASB	pH = 7
Na_2CO_3	WASB	pH > 7
$CuSO_4$	SAWB	pH < 7

2. Ans. (3)

Higher the solubility, ppt will form last.

3. Ans. (3)

CH_3COOH + CH_3COONa

WA + WASB salt

4. Ans. (3)

NaOH $N_BV_B = 0.1 \times 1 \times 1 = 0.1$

HCl $N_AV_A = 0.01 \times 1 \times 1 = 0.01$

$[OH^-] = \frac{0.1-0.01}{1+1} = \frac{.09}{2} = .045$

pOH = 1.35

pH = 12.65

5. Ans. (2)

MY (AB) $K_{sp} = s^2 = 6.2 \times 10^{-13}$

$s = \sqrt{6.2\times10^{-13}}$

$s \simeq 10^{-7}$

NY_3 (AB_3) $K_{sp} = 27S^4 = 6.2 \times 10^{-13}$

$s \simeq 10^{-3}$

6. Ans. (4)

$K_b = C\alpha^2$

7. Ans. (4)

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

s s + 0.1

$K_{sp} = (s)(s + 0.1)$

$s << 0.1$

$K_{sp} = s \times 0.1$

$s = \frac{1.6\times10^{-10}}{10^{-1}} = 1.7 \times 10^{-9}$

8. Ans. (3)

$Ag_2CrO_4 \rightleftharpoons 2Ag^+ + CrO_4^{-2}$

2s s

$[Ag^+] = 2s = 2.2 \times 10^{-4}$

$s = 1.1 \times 10^{-4}$

$K_{sp} = 4s^3 = 4(1.1 \times 10^{-4})^3$

$K_{sp} = 5.3 \times 10^{-12}$

9. Ans. (4)

(c) HCl $N_AV_A = \frac{1}{5}\times1\times75 = 15$

NaOH $N_BV_B = \frac{1}{5}\times1\times25 = 5$

$[H^+] = \frac{15-5}{75+25} = \frac{10}{100} = 10^{-1}$

pH = 1

10. Ans. (1)

$BaSO_4$ (AB)

$K_{sp} = s^2 = \left(\frac{2.42\times10^{-5}}{233}\right)^2 \cong 10^{-10}$

11. Ans. (1)

pH = 9 $Ca(OH)_2 \rightleftharpoons Ca^{+2} + 2OH^-$

s 2s

pOH = 5 $[OH^-] = 2s = 10^{-5}$

$s = 5 \times 10^{-6}$

$[OH^-] = 10^{-5}$ $K_{sp} = 4s^3 = 4(5 \times 10^{-6})^3$

$s = 500 \times 10^{-18}$

12. Ans. (3)

	HCl	+ NH_4OH →	NH_4Cl	+ H_2O
Initial	0.1 × 100	200 × 0.3	–	–
	10	60		
Left.	–	50	10	
	Basic	Buffer		

13. Ans. (3)

$[OH^-] = 10^{-2}$

pOH = 2

pH = 12

14. Ans. (3)

HCl → Can donate H^+ ion but not accept.

15. Ans. (3)

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

s 2s + 0.1

$K_{sp} = (s)(2s + 0.1)^2$

$2s << 0.1$

$5.3 \times 10^{-11} = s \times 10^{-2}$

$s = 5.3 \times 10^{-9}$

16. Ans. (2)

$Ni(OH)_2 \rightleftharpoons Ni^{+2} + 2OH^-$

s 2s + 0.1

$K_{sp} = s(2s + 0.1)^2$

$2s << 0.1$

$K_{sp} = s \times 10^{-2} \Rightarrow s = 2 \times 10^{-13}$

17. Ans. (4)

$NH_4Cl \rightarrow$ SAWB Acidic

$(NH_4)_2SO_4 \rightarrow$ SAWB Acidic

$NH_4NO_3 \rightarrow$ SAWB Acidic

$CH_3COONa \rightarrow$ SBWA Basic

18. Ans. (1)

$K_{sp} = s^2$

$s = \sqrt{K_{sp}} = \sqrt{4\times10^{-8}} = 2 \times 10^{-4}$

19. Ans. (3)

$$pH = 7 + \frac{1}{2} pK_a - \frac{1}{2} pK_b$$

20. Ans. (4)

$$pH = pK_a + \log \frac{(M\times V)_{salt}}{(M\times V)_{acid}}$$

$$= 4.57 + \log \left(\frac{\cancel{50}\times 0.1}{0.01\times \cancel{50}} \right)$$

$pH = 5.57$

21. Ans. (3)

$[H^+] = C\alpha$

$$= 0.01 \times \frac{1}{100} = 10^{-4}$$

$pH = 4$

22. Ans. (3)

Sodium hydroxide against oxalic acid titration; phenolphthalein is used as indicator and it's colour changes from colourless to pink.

23. Ans. (1)

$(k_{sp})AgCl = 10^{-10}$

$AgCl_{(s)} \rightleftharpoons Ag^+_{aq} + Cl^-_{aq}$

$k_{sp} = s^2$

$s = (10^{-10})^{1/2} = 10^{-5}$ mol/L

ln 0.1m KCl solution

$AgCl_{(s)} \rightleftharpoons Ag^+_{aq} + Cl^-_{aq}$, Let new solution is s'

$k_{sp} \rightarrow s'(s' + 0.1) = 10^{-10} = s' (s' + 0.1) = s'^2 + 0.1s'$

$$s' = \frac{10+0}{0.1} = 10^{-9} mol/L$$

$$\text{Ratio } \frac{s'}{s} = \frac{10^{-9}}{10^{-5}} = 10^{-4}$$

Exercise-III (Analytical Questions)

1. Ans. (1)

$K_h = h^2$ for hydrolysis of WA + WB type salt.

$h = \sqrt{K_h}$, degree of hydrolysis is independent of concentration of solution.

2. Ans. (2)

pH = 2 for 0.01 M HCl

pH = 12 for 0.01 M NaOH

pH > 7 for 0.01 M CH_3COONa

pH = 7 for 0.01 M NaCl

order : B > C > D > A

3. Ans. (4)

Due to common ion effect, solubility of AgBr decreases in the presence of NaBr. NaBr does not undergoes hydrolysis in water.

4. Ans. (1)

Acid turns blue litmus paper red & acids are sour in taste.

5. Ans. (2)

$NH_3 + H^+ \rightarrow NH_4^+$

Base Conjugate acid

$NH_3 \rightarrow NH_2^- + H^+$

Acid Conjugate base

6. Ans. (3)

Cations of strong base & Anions of strong acid do not get hydrolysed & salts of strong acids & strong bases are neutral.

7. Ans. (1)

Arrhenius acids releases H^+ in water while Bronsted acid releases H^+ in any solvent

So all Arrhenius acids are Bronsted acids but all Bronsted acids are not Arrhenius acid.

8. Ans. (1)

9. Ans. (3)

$Al(OH)_3$	$K_{sp} = (3^3)s^4$	$s = \left(\frac{K_{sp}}{27}\right)^{\frac{1}{4}}$
$BaCrO_4$	$K_{sp} = s^2$	$s = \sqrt{K_{sp}}$
$Zr_3(PO_4)_4$	$K_{sp} = (3)^3 (4)^2 s^7 = 6912\, s^7$	$s = \left(\frac{K_{sp}}{6912}\right)^{\frac{1}{7}}$
Hg_2I_2	$K_{sp} = 4s^3$	$s = \left(\frac{K_{sp}}{4}\right)^{\frac{1}{3}}$

10. Ans. (2)

Relative strength of acid can be compared by $\frac{[H^+]_1}{[H^+]_2}$ if concentration are same.

11. Ans. (1)

(a) $CH_3COOH + NaOH \rightleftharpoons CH_3COONa + H_2O$ if $[CH_3COOH] > [NaOH]$, after reaction $[CH_3COOH] + [CH_3COONa]$ will remain in the solution, so it can act as buffer.

(b) Similarly $HCl + CH_3COONa$ can act as buffer if $[CH_3COONa] > [HCl]$

(d) NH_4CN is a salt of WA + WB, so it is also a buffer.

12. Ans. (3)

pH of 10^{-8} M solution of HCl is 6.96.

13. Ans. (2)

In this, H_2O accepts H^+, hence it is a base according to Bronsted.

14. Ans. (1)

Ammonium hydroxide is a weak base & pH of NH_4Cl is less than 7 (since NH_4Cl is a salt of WB + SA).