UNIVERSAL SELF SCORER

Assertion & Reason

For AIIMS Aspirants

Read the assertion and reason carefully to mark the correct option out of the options given below :

(a)	If both ass explanation	ertion of th	and reason are true and the reason is the correct e assertion							
(b)	<i>If both assertion and reason are true but reason is not the correct explanation of the assertion.</i>									
(c)	If assertion	is tru	e but reason is false.							
(d)	If the asser	tion a	nd reason both are false.							
(<i>e</i>)	If assertion	is fals	se but reason is true.							
1.	Assertion Reason	:	<i>p</i> H of hydrochloric acid solution is less than that of acetic acid solution of the same concentration. In equimolar solutions, the number of titrable protons present in hydrochloric acid is less than that persent in acetic acid.							
2.	Assertion	:	[NDA 1999] A ionic product is used fro any types of electrolytes whereas solubility product is applicable only to sparingly soluble salts.							
	Reason	:	lonic product is defined at any stage of the raction whereas solubility product is only appicable to the saturation stage.							
.	Accortion		[AllMS 2001]							
з.	Assertion	•	A solution of $FeCt_3$ in water produce brown							
	Reason	:	Hydrolysis of $FeCl_3$ takes place in water							
4.	Assertion	:	$BaCO_3$ is more soluble in HNO_3 than in							
	Reason	:	plain water. Carbonate is a weak base and reacts with the							
			H^+ from the strong acid, casuing the barium salt to dissociate.							
5.	Assertion	:	$CHCl_3$ is more acidic than CHF_3 .							
	Reason	:	The conjugate base of ${\it CHCl}_3{\rm is}$ more stable							
6.	Assertion	:	than CHF_3 . Addition of silver ions to a mixture of aqueous sodium chloride and sodium bromide solution will first precipitate $AgBr$ rather than $AgCL$.							
	Reason	:	K_{sp} of $AgCl < K_{sp}$ of $AgBr$. [AllMS 2004]							
7.	Assertion	:	The pK_a of acetic acid is lower than that of							
8.	Reason Assertion	:	phenol. Phenoxide ion is more resonance stabilized.[AIIMS 2004 Sb (111) is not precipitated as sulphide when in its alkaline solution H_2S is passed.							
	Reason	:	The concentration of S^{2-} ion in alkaline medium is inadequate for precipitation. [AIIMS 2004]							
9.	Assertion Reason	:	lonic reactions are not instantaneous. Oppositely charged ions exert strong forces.							

3 1 С 2 b b 4 b 5 b 7 b 8 9 10 6 a d а а 11 С 12 d 13 b 14 b 15 d 16 19 d 17 b 18 b с 20 с 21 22 23 24 25 а с b d с 26 d 27 28 d 29 d 30 с а 31 32 33 34 35 b С С с а 36 С

Acids and Bases

1	а	2	d	3	а	4	С	5	b
6	d	7	b	8	C	9	b	10	а
11	а	12	С	13	a	14	d	15	d
16	b	17	b	18	d	19	а	20	b
21	d	22	b	23	b	24	C	25	d
26	C	27	С	28	d	29	b	30	b
31	b	32	С	33	d	34	а	35	а
36	а	37	C	38	C	39	C	40	а
41	С	42	d	43	C	44	а	45	а
46	С	47	С	48	d	49	a	50	b
51	b	52	а	53	а	54	а	55	d
56	b	57	ab	58	а	59	С	60	d
61	а	62	b	63	b	64	b	65	а
66	а	67	d	68	а	69	а	70	b
71	b	72	d	73	b	74	d	75	b
76	b	77	а	78	а	79	b	80	а
81	d	82	b	83	b	84	d	85	bd
86	а	87	d	88	а	89	d	90	С
91	а	92	а	93	а	94	d	95	b
96	С	97	b	98	C	99	d	100	C
101	а	102	а	103	C	104	C	105	а
106	а	107	d	108	b	109	С	110	d
111	b	112	С	113	C	114	а	115	С
116	d	117	d	118	C	119	C	120	C
121	а	122	d	123	b	124	а	125	b
126	С	127	С	128	d	129	C	130	b
131	b	132	С	133	а	134	C	135	а
136	С	137	С	138	b	139	b	140	b
141	b	142	а	143	b	144	а	145	а
146	а	147	b	148	b	149	С	150	а
151	С	152	С	153	d	154	С	155	а
156	b	157	а	158	d	159	b	160	d
161	b	162	d	163	С	164	d	165	а

Common ion effect, Isohydric solutions, Solubility product, Ionic product of water and Salt hydrolysis

1	b	2	c	3	b	4	b	5	b
6	d	7	С	8	а	9	С	10	а
11	b	12	b	13	а	14	d	15	d
16	b	17	C	18	c	19	d	20	C
21	d	22	d	23	d	24	b	25	a

Electrical conductors, Arrenius theory and Ostwald's dilution law

Answers

UNIVERSAL SELF SCORER

380 Ionic Equilibrium

26	d	27	a	28	d	29	b	30	b
31	а	32	d	33	а	34	С	35	d
36	b	37	b	38	C	39	b	40	а
41	а	42	d	43	b	44	а	45	b
46	d	47	b	48	c	49	С	50	С
51	b	52	d	53	a	54	d	55	С
56	d	57	b	58	C	59	d	60	C
61	b	62	b	63	а	64	bc	65	d
66	a	67	d	68	а	69	d	70	C
71	С	72	d	73	d	74	d	75	а
76	C	77	b	78	d	79	b	80	d
81	C	82	C	83	d	84	С	85	d
86	d	87	b	88	c	89	b	90	а
91	a	92	а	93	c	94	С	95	а
96	d	97	d	98	b	99	b	100	ab
101	d	102	С	103	с	104	а	105	C
106	C	107	С	108	а	109	b	110	b
111	b	112	а	113	а	114	а	115	b
116	а								

Hydrogen ion concentration - *pH* scale and Buffer solution

1	a	2	с	3	d	4	b	5	b
6	d	7	a	8	с	9	b	10	d
11	а	12	d	13	с	14	а	15	b
16	d	17	b	18	с	19	с	20	d
21	b	22	с	23	а	24	d	25	b
26	a	27	a	28	d	29	b	30	b
31	d	32	b	33	а	34	а	35	C
36	C	37	C	38	а	39	а	40	d
41	a	42	a	43	b	44	d	45	b
46	а	47	b	48	а	49	С	50	а
51	а	52	С	53	с	54	с	55	d
56	а	57	С	58	b	59	b	60	С
61	b	62	С	63	b	64	b	65	c
66	а	67	b	68	d	69	а	70	С
71	d	72	а	73	b	74	а	75	b
76	а	77	а	78	С	79	b	80	а
81	b	82	С	83	b	84	b	85	c
86	b	87	a	88	а	89	а	90	d
91	b	92	С	93	а	94	а	95	c
96	b	97	b	98	b	99	b	100	b
101	C	102	С	103	а	104	a	105	С
106	a	107	c	108	С	109	d	110	b
111	d	112	d	113	b	114	a	115	c
116	d	117	b	118	b	119	d	120	С
121	d	122	b	123	С	124	d	125	b
126	b	127	а	128	c	129	b	130	d

131	C	132	d	133	с	134	C	135	а
136	C	137	C	138	c	139	С	140	b
141	b	142	а	143	d	144	b	145	b
146	d	147	С	148	b	149	а	150	d
151	b	152	d	153	b	154	b	155	С
156	d	157	С	158	а	159	С	160	d
161	C	162	d	163	b	164	а	165	d
166	a	167	d	168	b	169	С	170	а
171	C	172	C	173	d	174	b	175	а
176	С	177	а	178	C	179	а	180	b
181	C	182	b	183	b	184	b	185	С
186	b	187	b	188	а	189	а	190	b
191	d	192	b	193	b	194	b	195	а
196	b	197	b	198	b	199	a	200	ad
201	ac	202	a bc	203	с	204	С	205	b
206	b								

Critical Thinking Questions

1	b	2	a	3	b	4	а	5	d
6	а	7	С	8	а	9	а	10	a
11	b	12	d	13	С	14	d	15	C
16	d	17	d	18	b	19	b	20	d
21	а	22	b	23	а	24	b	25	а
26	С	27	d	28	d	29	b	30	d
31	b	32	а	33	С	34	а	35	а
36	b	37	С	38	b	39	а	40	b
41	а	42	а	43	а	44	а	45	a
46	с	47	b	48	b	49	b	50	С
51	d	52	С	53	b	54	C	55	a
56	d	57	d	58	С	59	а	60	а
61	а	62	С	63	d	64	b	65	a
66	а	67	b	68	b	69	d	70	b
71	а	72	С	73	а	74	C	75	C
76	а	77	b	78	а				

Assertion & Reason



Electrical conductors, Arrhenius theory and Ostwald's dilution law

UNIVERS **Ionic Equilibrium 381** SELF SCORE

- $C_{12}H_{22}O_{11}$ is a sugar and non-electrolyte. 1. (c)
- 2. (b) It is a weak electrolyte since it is slightly ionized.
- It is a weak electrolyte since slightly ionized. (h)3.
- Because the degree of dissociation is inversely proportional to 4. (b) the concentration of the electrolyte. (b) Electrolytes are those substances which on dissolving in water 5
- give ions. (a) $K = \frac{\alpha^2 C}{1 - \alpha}; \ \alpha = \frac{0.01}{100} \approx 1 \ \therefore \ K = \alpha^2 C = \left[\frac{0.01}{100}\right]^2 \times 1$ 6.

 $= 1 \times 10^{-8}$

- (b) As NaCl ionises completely to yield free ions. 7.
- (d) $CH_3COONa \Rightarrow CH_3COO^- + Na^+$ 8.

 $H_2 O \Rightarrow H^+ + OH^ CH_3COOH + NaOH$

- (a) NaCl, being a salt, is a strong electrolyte. 9.
- We can determine by measurement of very dilute HF solutions. 10. (a) According to the Ostwald's dilution formula $\alpha^2 = \frac{K(1-\alpha)}{C}$ 11. (c)

But for weak electrolytes α is very small. So that $(1-\alpha)$ can

be neglected. So that $\alpha = \sqrt{\frac{K_a}{C}}$.

- Arrhenius proposed the theory of ionisation. 12 (\mathbf{d})
- higher the dielectric constant of a solvent more of its ionising 13. (b) power.
- (\mathbf{b}) $\alpha \propto \text{dilution of solution.}$ 14.
- (d) Generally ionic compound are conduct electricity in fused state. 15.
- 16. (d) According to Ostwald's dilution law because degree of ionization is directly proportional to the dilution.
- 17. (b) The degree of ionisation of a solute depends upon its nature, concentration, and temperature.
- 18. (b) Mathematical form of Ostwald's dilution law. It is a weak electrolyte because it's ionization is very less. 19. (c)
- When we add NH_4OH in NH_4Cl solution ionization of 20. (c) $NH_{4}OH$ is decreased due to common ion effect.

1

21. (a) Initially After dissociation $a - \alpha$

Total =
$$1 - \alpha + \alpha + 2\alpha = 1 + 2\alpha$$

 $\alpha = \frac{1.98 - 1}{\alpha} = \frac{0.98}{\alpha} = 0.49$
for a mole $\alpha = 0.49$

 $BaCl_2 \rightleftharpoons Ba^{2+} + 2Cl^{-}$

0

 α

0

 2α

For 0.01 mole
$$\alpha = \frac{0.49}{0.01} = 49$$

- 22 (c) It is an ionic salt.
- (b) Due to common ion effect of H^+ . 23.
- Current does not affect the degree of ionization. 24. (d)
- Polar solvent facilitate ionisation of strong electrolytes due to 25. (c) dipole-ion attraction.

26. (d)
$$HA \Rightarrow H^+ + A^-$$

$$K = Ca^2 = 0.1 \times (10^{-4})^2 = 10^{-9}$$

(d) Degree of dissociation $\alpha = ?$ 28.

Normality of solution = 0.1 $N = \frac{1}{10}N$

Volume = 10 *litre*

Dissociation constant $K = 1 \times 10^{-5}$

$$K = \frac{\alpha^2}{V}; \ \alpha = \sqrt{KV} = \sqrt{1 \times 10^{-5} \times 10} ; \ \alpha = 1 \times 10^{-2}$$

- (d) $MgCl_2 \Rightarrow Mg^{++} + 2Cl^-$ in aqueous solution it is dissociated 29. into ions.
- Ostwald's dilution formula is $\alpha^2 = K(1-\alpha)/C$ but for weak 30. (c) electrolyte α is very small. So that $(1 - \alpha)$ is neglected for weak electrolytes. So for weak electrolyte the dilution formula is $\alpha = \sqrt{\frac{K}{C}}$.

31. (b)
$$K_a = C\alpha^2 = 0.2 \times \left(\frac{32}{100}\right)^2 = 2.048 \times 10^{-4}$$

32. (c)
$$H_2SO_4 \rightleftharpoons H^+ + H^-SO_4 \rightleftharpoons H^+ + SO_4^-$$

Because of it is completely ionised.

Colour of electrolyte depends on the nature of both ions. 33. (c)

e.g. $CuSO_4$ is blue because Cu^{2+} ions are blue.

- lonisation depends upon dilution, when dilution increases then 34 (c) ionisation is also increased.
- More is K_a , lesser is $pK_a(pK_a = -\log K_a)$ more is acidic 35. (a) strength.

Acids and Bases

1. (a) CO doesn't have a vacant d-orbital.

2

4

5.

L

(d)
$$HClO_4 + H_2O \Rightarrow H_3O^+ + ClO_4^-$$

Conjugate acid and base pair

(a) $FeCl_3 + 3H_2O \Rightarrow Fe(OH)_3 + 3HCl$. Strong acid and 3. weak base.

(c)
$$Na_2CO_3 + 2H_2O \rightleftharpoons 2NaOH + H_2CO_3$$

(b) Those substance accept the proton are called Bronsted base and which is donate the proton are called Bronsted acid.

$$HCO_3^- + H^+ \rightleftharpoons H_2CO_3$$
 Bronsted base.

$$HCO_3^- \rightleftharpoons H^+ + CO_3^{--}$$
 Bronsted acid.

- (b) The value of pK_a for strong acid is less. 7.
- 8. (c) Because it is a salt of strong base and weak acid.
- Because it is conjugate base of weak acid. 9. (b)

 $CH_3COOH \Rightarrow CH_3COO^- + H^+$.

- (a) Those compound which accept H^+ is called bronstad base 11. NO_3^- accept H^+ and form HNO_3 . So it is a base.
- Larger the size of halogen atom less is the back donation of 13. (a) electrons into empty 2p orbital of B.

$$\textbf{4.} \qquad (\textbf{d}) \quad H_2O + NH_3 \rightleftharpoons NH_4^+ + OH^-$$

16. (b)
$$NH_2^- \Rightarrow NH^{-2} + H^+$$

Conjugate acid, base pair.

- $(b) \quad \text{Those substances which lose proton are called acid.}$ 17. e.g. strong acid have a strong tendency to donate a proton.
- 18. Electron donating species called nucleophile. NH₃ have a lone (d) pair of electron.

ELF SCORER 382 Ionic Equilibrium

 $\textbf{48.} \qquad (d) \quad A \text{ weak acid and strong base.}$

- **49.** (a) In NaH_2PO_2 both hydrogen are bonded with 'P,' so it is not replaceable.
- **50.** (b) Aqueous solution of Na_2CO_3 is alkaline due to hydrolysis of CO_3^{--} .
- **51.** (b) Because they give CO_3^{--} in solution.

(a) M.eq. of 0.2*M* $H_2SO_4 = \frac{2 \times 0.2M}{1000} \times 100 = 0.04 \text{ m/l}$ M.eq. of .2M NaOH = $\frac{0.2}{1000} \times 100 = 0.02 \text{ m/l}$ left $[H^+] = .04 - .02 = .02$. Total volume = $200 = \frac{.02}{200} = .0001 = 10^{-4} M$ pH = 4. (a) H_3BO_3 is a weak monobasic acid it does not act as a H^+ donor but behaves as a Lewis acid. (a) Because $SnCl_2$ is a electron acceptor according to Lewis concept. (d) ROH is a Lewis base because it has an lone pair of electron. (b) $H_2SO_4 + 2H_2O \approx 2H_3O^+ + SO_4^ NaOH \Rightarrow Na^+ + OH^-$ 1 mole of H_2SO_4 acid gives 2 moles of H_3O^+ ions. So 2 moles of OH^- are required for complete neutralization. (ab) Diprotic solvents give $2H^+$ ions or OH^- ions. (a) $N_{NaOH} = 1 \times 1 = 1N$ $N_{H_2SO_4} = 2 \times 10 = 20N$ M.eq. of $NaOH = 1 \times 100 = 100$ M.eq. of $H_2SO4 = 20 \times 10 = 200$ Thus M.eq. of acid are left and therefore pH < 7, so the resulting mixture will be acidic. (d) PH_3 is a Lewis base. Because of CH_3COONa is a salt of weak acid and strong (a) base. (b) Acid $\xrightarrow{-H^+}$ conjugate base.

Base
$$\xrightarrow{+H^+}$$
 conjugate acid.

53. (b)
$$HCl \rightarrow Cl$$

Acid Base

55. (a)
$$HClO_4 > H_2SO_4 > HCl > HNO_3$$

Aciediccharacter decreases

- **56.** (a) Those substances which accept the H^+ are called conjugate base.
- 7. (d) NH_3 is a Lewis base, which donate a lone pair of electron.
- **59.** (a) $FeCl_3 + 3H_2O \Rightarrow Fe(OH)_3 + 3HCl$. Strong acid have less than 7 *pH*.
- **70.** (b) It donates their e^- pair.
- (b) The strength of the acid will depend upon the proton donation.(b) Because it is a salt of strong acid and strong base.

 $H^+_{(aq)} + Cl^-_{(aq)} + Na^+_{(aq)} + OH^-_{(aq)} \rightleftharpoons$

 $H_2O_{(1)} + Na^+_{(aq)} + Cl^-$

74. (d) CCl_4 is not a Lewis or bronsted acid. It does not contain H^+ .

75. (b)
$$NH_4Cl + H_2O \Rightarrow NH_4OH + HCl_{Weak base} + HCl_{strong acid}$$
. So it is acidic in nature.

- **76.** (b) Cu(11) complexes are blue. The four water molecules are attached with secondary valencies of the metal atom e.g. $[Cu(H_2O)_4]SO_4$. H_2O .
- **77.** (a) The species which can accept as well as donate H^+ can act both as an acid and a base.

$$HSO_{4}^{-} + H^{+} \rightleftharpoons H_{2}SO_{4}$$

base
$$HSO_{4}^{-} \rightleftharpoons SO_{4}^{2-} + H^{+}$$

acid

- **78.** (a) NH_4^+ is the weakest acid. So its conjugate base is strongest.
- **79.** (b) Ag^+ is an electron deficient compound and hence is a Lewis acid.

80. (a)
$$H_3 PO_4 \rightleftharpoons H^+ + H_2 PO_4^-$$

Conjugate acid

81. (d)
$$HS_2O_8^- \rightleftharpoons H^+ + S_2O_8^{--}$$

Conjugate acid Conjugate base

- 82. (b) PH₃ donates electron pair to BCl₃.
 83. (b) The conjugate base of weak acid is a strong base.
- **83.** (b) The conjugate base of weak acid is a strong **84.** (d) $OH^- \rightarrow O^{2-} + H^+$

Conjugate base of
$$OH^-$$

- **85.** (bd) Presence of lone pair of electron and they donate two electron pairs.
- **86.** (a) On increasing oxidation number Acidic strength increases.
- **87.** (d) $H_3 PO_4$ is shows +5 maximum oxidation state.
- **88.** (a) $C_2H_5^-$ is a strongest base.
- **89.** (d) NaOCl is a mixture of strong base and weak acid.
- **90.** (c) NH_4OH gives minimum OH^- ion. So it is a weak base.
- **91.** (a) Solution become acidic and methyl orange act on acidic *pH*.
- **92.** (a) Larger is bond length, more is acidic nature (for halogen acids). *HF* bond length is small.
- **93.** (a) *HCl* is accepting proton in *HF* medium and acts as weak base.
- **95.** (b) For oxoacids of the same element, the acidic strength increases with increase in the oxidation number of that element.

$$HClO < HClO_2 < HClO_3 < HClO_3$$

- 96. (c) Because their conjugate base and conjugate acids are strong.
- **97.** (b) *HCl* is a strong acid and their conjugate base is a very weak base.

100. (c)
$$NH_3 \rightleftharpoons NH_2^- + H^+$$

101. (a) $H_2SO_4 \rightleftharpoons H^+ + HSO_4^- \rightleftharpoons H^+ + SO_4^{--}$

 HSO_4^- is conjugate base. But it is also an acid because it lose H^+ .

- **102.** (a) BF_3 is a Lewis acid because '*B* has incomplete octet.
- 104. (c) H_3O^+ , however it exists as $H_9O_4^+$.
- 105. (a) $Al_2(SO_4)_3$ is a salt of weak base $Al_2(OH)_3$ and strong acid H_2SO_4 .
- **106.** (a) Al^{+++} of $AlCl_3$ undergoes hydrolysis.

107. (d)
$$H_2SO_4 \rightleftharpoons H^+ + HSO_4^-$$

- **109.** (c) Conjugate base of H_2SO_4 is HSO_4^- .
- 110. (d) Presence of lone pair.
- **III.** (b) According to Bronsted principle HNO_3 is acid they give H^+ in aqueous solution and form NO_3^- .
- **112.** (c) $H_2O + H_2O \Rightarrow H_3O^+ + OH^-$.
- **113.** (c) NH_4^+ is a conjugate acid;

$$NH_4^+ \Rightarrow NH_3 + H$$

- **114.** (a) $AlCl_3 + 3H_2O \Rightarrow Al(OH)_3 + \frac{3HCl}{Stron acid}$
- **116.** (d) HCl is a strong acid its conjugate base means Cl^- is a weak base.
- **117.** (d) Hydrazoic acid (HN_3) is a Lewis acid.
- **118.** (c) Smaller the pK_a value than. Stronger the acid.
- **120.** (c) BF_3 is acidic because due to Lewis concept it accept a lone pair of electron.
- 122. (d) Because it is a weak electrolyte.
- **126.** (c) $AlCl_3$ and SO_2 both are example of Lewis theory.
- **127.** (c) $Na_2CO_3 + 2H_2O \approx 2NaOH + H_2CO_3$. It is a strong base and weak acid so it is a basic.

129. (c)
$$H_2O + NH_3 \approx NH_4^+ + OH^-$$
.

In this reaction H_2O acts as acid because it donate a proton.

130. (b)
$$H_2SO_4 + H_2O = H_3O^+ + HSO_4^-$$

Conjugate acid and base pair
131. (b) $Al_2(SO_4)_2 = 2Al^{3+} + 3SO_4^{2-}$

31. (b)
$$Al_2(SO_4)_3 \approx 2Al^{3+} + 3SO_4^{2-}$$

 $Al(OH)_3 + H_2SO_4$

Weak base Strong acid
132. (c)
$$NaOH + HCl \xrightarrow{\text{Neutralization}} NaCl + H_2O$$

Reaction

- **133.** (a) Conjugate acid is obtained from the base by gain of H^+ .
- **134.** (c) KCl is a ionic compound
- 135. (a) $H_3BO_3 + H_2O \Rightarrow [B(OH)_4]^- + H^+$

 H_3BO_3 is a weak monobasic acid if does not act as a H^+ donor but behaves as a Lewis acid.

- 136. (c) Because it is not accept the proton.
- **137.** (c) NH_4Cl is a salt of weak base (NH_4OH) and strong acid (*HCl*).

138. (b)Because it accept electron pair from the PH_3

139. (b)
$$NH_4^+ \rightleftharpoons NH_3 + H^+$$

141. (b) Hydrolysis of Fe^{3+}

$$Fe^{3+} + 3H_2O \rightarrow Fe(OH)_3 + 3H^+$$

142. (a) According to the Lewis concept.

- 143. (b) Conjugate base is obtained from the acid by loss of H^+ .
- 144. (a) BF_3 accept electron pair from NH_3 so it is Lewis acid.
- **145.** (a) CH_4 has almost no acidic nature and thus CH_3^- is strongest base.

SCORER 384 Ionic Equilibrium

- **146.** (a) $CuSO_4$ is a salt of weak base, $(Cu(OH)_2)$ and strong acid (H_2SO_4) .
- **147.** (b) Weak acid consists of highest pK_a value and strongest acid consist of less pK_a value.
- 149. (c) Because it gain and also lose the proton $H_2O + H_2O \rightleftharpoons H_3O^+ + OH^-$

150. (a)
$$H_2O + H_2O \Rightarrow H_3O^+ + OH^-$$

- **151.** (c) $HBr + H_2O \Rightarrow H_3O^+ + Br^-$
- 152. (c) Because both are strong acid and strong base.
- **153.** (d) $LiAlH_4$ is a nucleophilic and capable of donating electron pair, thus acts as a Lewis base.
- 154. (c) The solvent which neither accept proton nor donates.
- **155.** (a) Because of F^- is a highly electronegative. So it is easily lose the electron and reaction occur rapidly.
- 156. (b) Strong acid can be used titrate both strong and weak base.
- 159. (b) For a weak acid value of *pKa* will be vary high but in case of strong acid value of *pKa* will be vary low.
- (d) Boron halides behave as Lewis acid because of their electron deficient nature eg., as

 $\begin{array}{c} | \\ B \\ \end{array}$ (Deficiency of two electron for inert configuration)

- **161.** (b) Gaseous *HCl* does not give H^+ but liquid *HCl* gives H^+ in aquous solution there for gaseous *HCl* is not a Arrhenius acid due to covalent bonding in gaseous condition.
- **162.** (d) $H_2O \rightarrow H^+ + OH^-$ [Acid due to donation of proton] Acid $H_1O + H^+ \rightarrow H_2O^+$ [Recip due to gaining of proton]

$$H_2O+H \rightarrow H_3O$$
 [Basic due to gamming of proton]
Base

163. (c) F^- strongest conjugate base due to it smallest size in a group and gain proton due to most electronegative capacity. $H^+ + F^- \rightarrow HF$

164. (d)
$$H_2 PO_4^- \rightarrow H^+ + HPO_4^{2-}$$

Conjugate acid

165. (a)
$$HSO_4^- \rightarrow H^+ + SO_4^{2-}$$

Conjugate acid Conjugate base

Common ion effect, Isohydric solutions, Solubility product, Ionic product of water and Salt hydrolysis

1. (b) Solubility of $Al_2(SO_4)_3$ $Al_2(SO_4)_3 \Rightarrow 2Al^{+++} + 3SO_4^{--}$ $K_{sp} = [Al^{3+}]^2 [SO_4^{2-}]^3$ 2. (c) Due to common ion effect.

3. (b)
$$MX_2 \rightleftharpoons M_S^{2+} + 2X_{2S}^{-}$$

 $K_{sp} = (2S)^2(S) = 4S^3$
 $\Rightarrow S = 2\sqrt[3]{\frac{K_{sp}}{4}} = \sqrt[3]{\frac{4 \times 10^{-12}}{4}} = 1.0 \times 10^{-4} M$

5. (b)
$$MX_2 \rightleftharpoons M_{(S)}^+ + 2X_{(2S)^2}^-$$
; $K_{sp} = 4S^3$

$$S = \sqrt[3]{\frac{K_{sp}}{4}} = \sqrt[3]{\frac{1 \times 10^{-11}}{4}} = 1.35 \times 10^{-4}$$

(a)
$$Mg(OH)_2 \rightleftharpoons Mg^{++} + 2OH_{(X)} + 2OH_{(2X)^2}$$

$$K_{sp} = 4X^{3}$$
(c) $BaSO_{4} \approx Ba^{++} + SO_{4}^{-}$

8.

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14

2

$$K_{sp} = S^{2} = S \times S = 0.01 \times S$$
$$S_{(SO_{4}^{2^{-}})} = \frac{K_{sp}}{S_{(BO_{4}^{++})}} = \frac{1 \times 10^{-9}}{0.01} = 10^{-7} \text{ mole/litre}$$

0. (a)
$$AB_2 \rightleftharpoons A^+_{1 \times 10^{-5}} + \frac{2B^-}{2 \times 10^{-5}}$$

 $K_{sp} = [1 \times 10^{-5}] [2 \times 10^{-5}]^2 = 4 \times 10^{-15}$

(b)
$$CaF_2 \rightleftharpoons Ca^{++}_{(S)} + 2F^-_{(2S)^2}$$
; $K_{sp} = 4S^{-}_{(S)}$

12. (b) Due to common ion effect.

(a)
$$Ag_2CrO_4 \approx [2Ag^+] + [CrO_4^{--}]$$

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

(d)
$$K_{sp}$$
 for $CaF_2 = 4s^3 = 4 \times [2 \times 10^{-4}]^3 = 3.2 \times 10^{-11}$

- 15. (d) The concentration of S^{2-} ions in group II is lowered by maintaining acidic medium in the presence of NH_4Cl . The ionization of H_2S is supressed due to common ion effect. So the ionic product is less than solubility product.
- **16.** (b) NH_4Cl is hydrolysed and give $[H^+]$

 $NH_4Cl + H.OH \Rightarrow NH_4OH + HCl$

$$NH_4^+ + H.OH \Rightarrow NH_4OH + H^+$$

- 17. (c) $FeCl_3$ is a salt of weak base $(Fe(OH)_3)$ and strong acid (HCh).
- 18. (c) For the precipitation of an electrolyte, it is necessary that the ionic product must exceed its solubility product.

19. (d)
$$K_{sp} = [Ag^+]^2 [Cro_4^{--}] = [2S]^2 [0.01]$$

 $=4S^{2}[0.01] = 4[2 \times 10^{-8}]^{2} \times 0.01 = 16 \times 10^{-18}$.

20. (c) Complex salts contain two different metallic elements but give test for only one of them. *e.g.* $K_4 Fe(CN)_6$ does not give test for Fe^{3+} ions.

(d)
$$10^{-3} N \text{ KOH will give } [OH^{-}] = 10^{-2} M$$

 $pOH = 2$
 $\therefore pH + pOH = 14, pH = 14 - 2 = 12$

- **22.** (d) It is $FeSO_4(NH_4)_2SO_4 \cdot 10H_2O$.
- $\label{eq:alpha} \textbf{23.} \qquad (d) \quad \text{Salt of a strong base with a weak acid.}$
- 24. (b) NH_4CN is a salt of weak acid and weak base and thus for it
- $\label{eq:alpha} \textbf{25.} \qquad (a) \quad \text{Because it is a salt of strong base with a weak acid.}$
- **26.** (d) Because CCl_4 is a organic solvent and $AgNO_3$ is insoluble in organic solvent.

27. (a)
$$SnS_2 \Rightarrow Sn^{4+} + 2S^{2-}$$

 $\therefore K_{sp} = [Sn^{4+}] [S^{2-}]^2$ 28. (d) It does not dissociate much or its ionization is very less. NaHCO₃ has one replaceable hydrogen. 29. (b) $CaOCl_2$ has two anions Cl^- and OCl^- along with (b) 30. Ca^{2+} ions. (a) $K_{sp} = 4S^3$, $S^3 = \frac{4 \times 10^{-9}}{4} = 10^{-9}$ 31. $\therefore S = 10^{-3} M \, .$ $Be(OH)_2$ has lowest solubility and hence lowest solubility (d) 32. product. Because it is a salt of strong acid and strong base. 33. (a) (c) $NH_4OH \Rightarrow |NH_4| + OH^-$ 34 $NH_4Cl \Rightarrow \left| NH_4^+ \right| + Cl^-$ Common ion 35. (d) It is a less ionic, so that least soluble in water. (b) pH of 9 means the salt solution should be fairly basic. 36. (b) $CH_3COOH \Rightarrow CH_3COO^- + H^+$ 37. On adding CH_3COONa , $[H^+]$ decreases. (c) 0.01 *M* $CaCl_2$ gives maximum Cl^- ions to keep K_{sp} of AgCl 38. constant, decrease in $[Ag^+]$ will be maximum. Due to the common ion effect. 39. (b) (a) $K_{sp} = 4s^3$ 40.

$$S = \sqrt[3]{\frac{K_{sp}}{4}} = \sqrt[3]{\frac{1.0 \times 10^{-6}}{4}} = 6.3 \times 10^{-3}.$$

- **42.** (d) K_w increases with increase in temperature.
- **43.** (b) It contains two cations and one anion.

4

46.

4. (a)
$$HgSO_4$$
 of $K_{sp} = S^2$
 $S = \sqrt{K_{sp}}$; $S = \sqrt{6.4 \times 10^{-5}}$; $S = 8 \times 10^{-3}$ m/l.

45. (b) The solubility of $BaSO_4$ in *g/litre* is given 2.33×10^{-3}

: in mole/litre.
$$n = \frac{W}{m.wt} = 1 \times 10^{-5} = \frac{2.33 \times 10^{-3}}{233}$$

Because $BaSO_4$ is a compound

$$K_{sp} = S^{2} = [1 \times 10^{-5}]^{2} = 1 \times 10^{-10}$$

(d) $AgCl = Ag^{+} + Cl^{-}_{a}$
 $NaCl = Na^{+} + Cl^{-}_{0.02}$
 $K_{sp} AgCl = 1.20 \times 10^{-10}$
 $K_{sp} AgCl = [Ag^{+}][Cl^{-}] = a \times [a + 0.2] = a^{2} + 0.2a$
 a^{2} is a very small so it is a neglected.

 K_{sp} AgCl = 0.2a

$$1.20 \times 10^{-10} = 0.2a$$
$$a = \frac{1.20 \times 10^{-10}}{0.20} = 6 \times 10^{-10} \text{ mole}$$

$$AgI \Rightarrow Ag^+ + |I^-|$$

 $NaI \Rightarrow Na^+ + |I^-|$
 L_-
Common ion

48. (c)
$$K_{sp}$$
 of $BaSO_4 = 1.5 \times 10^{-9}$; $Ba^{++} = 0.01M$

$$SO_4^{--} = \frac{1.5 \times 10^{-9}}{0.01} = 1.5 \times 10^{-7}$$

49.

(c)
$$AgCrO_4 = 2Ag^+ + CrO_4^-$$

 $(2S)^2 + S^-$
 $K_{sp} = 4S^3$ given $2S = 1.5 \times 10^{-4}$
 $\therefore K_{sp} = (2S)^2 \times S$

$$= (1.5 \times 10^{-4})^2 \times \left(\frac{1.5 \times 10^{-4}}{2}\right) = 1.6875 \times 10^{-12}$$

50. (c)
$$PbCl_2 = Pb_s^{2+} + 2Cl_s^{-}$$

 $K_{sp} \text{ of } PbCl_2 = [Pb^{2+}] \times [Cl^{-}]^2$; $K_{sp} = S \times (2S)^2$

$$K_{sp} = S \times 4S^2 = 4S^3$$
; $S^3 = \frac{K_{sp}}{4}$; $S = \sqrt[3]{\frac{K_{sp}}{4}}$

51. (b)
$$AgCl \Rightarrow [Ag^+][Cl^-]; K_{sp} = S \times S; K_{sp} = S^2$$

$$S = \sqrt{K_{sp}} = \sqrt{1.44 \times 10^{-4}} = 1.20 \times 10^{-2} M.$$

52. (d) By formula
$$BA_2 \rightarrow B^+ + 2A^-$$

$$K_{sp} = 4 x^3$$

53. (a)
$$AgCrO_4 \rightarrow 2Ag^+ + CrO_4^{--}$$

 $K_{sp} = (2S)^2 S = 4S^3$
 $S = \left(\frac{K_{sp}}{4}\right)^{\frac{1}{3}} = \left(\frac{32 \times 10^{-12}}{4}\right)^{\frac{1}{3}} = 2 \times 10^{-4} M.$

- 54. (d) Common ion effect is noticed only for weak electrolyte dissociation. H_2SO_4 is strong electrolyte.
- 55. (c) When we added barium ion in chromate ion solution we obtained yellow ppt of $BaCrO_4$.

$$BaCl_2 + K_2CrO_4 \rightarrow BaCrO_4 \downarrow + 2KCl$$

Yellow ppt.

56. (d) *AB is a* binary electrolyte.

$$S = \sqrt{K_{sp}} = \sqrt{1.21 \times 10^{-6}} = 1.1 \times 10^{-3} M$$

- 57. (b) Precipitation occurs when ionic product > solubility.
- **58.** (c) For a binary electrolyte, so that

$$K_{sp} = S \times S = S^2$$

 $S = \sqrt{K_{sp}}$.

61. (b)
$$CH_{3}COONa$$
 is a salt of weak acid and strong base. Hence its aqueous solution is alkaline.
62. (b) $BaSO_{4} = Ba^{2+} + SO_{4}^{--}$
Solubility constant $= S \times S$
 $1.5 \times 10^{-19} = S^{2}; S = \sqrt{1.5 \times 10^{-19}}; S = 3.9 \times 10^{-5}$
65. (d) $Ca(OH)_{2} = Ca^{++} + 2OH^{-}_{(2S)^{2}}$
 $K_{sp} = 4S^{3} = 4 \times \sqrt{3} \times \sqrt{3} \times \sqrt{3} = 12\sqrt{3}$
66. (a) Due to common ion effect.
67. (d) $PbCl_{2} = Pb^{2+} + 2Cl^{-}_{(2S)^{2}}$
 $K_{sp} = 4S^{3} = 4 \times (2 \times 10^{-2})^{3} = 3.2 \times 10^{-5}$
68. (a) $Ag_{2}S = 2Ag^{+} + S^{--}$
 $K_{sp} = 4S^{3}$
 $S = \sqrt[3]{\frac{K_{sp}}{4}} = \sqrt[3]{\frac{3.2 \times 10^{-11}}{4}} = 2 \times 10^{-6}$
69. (d) $CaCO_{3} = Ca^{++} + CO_{3}^{--}$
Solubility product of $CaCO_{3}$
 $K_{sp} = S^{2}; S = \sqrt{K_{sp}}$
It is a binary electrolyte.
 $S^{2} = K_{sp}; (3.05 \times 10^{-4})^{2} = K_{sp}; K_{sp} = 9.3 \times 10^{-8}$
71. (c) $PbCl_{2} \rightarrow Pb^{++}_{S} + 2Cl^{-}_{2S}$
 $K_{sp} = S \times (2S)^{2} = [6.3 \times 10^{-3}] \times [12.6 \times 10^{-3}]^{2}$.
72. (d) A salt of strong acid and strong base cannot be hydrolysed. In this case the equilibrium cannot shifted towards the backward.
73. (d) If we mixed any substance into the solution. Then the value of pH is increased these substance is a salt of weak acid and strong base.

(d) It is a salt of strong base and weak acid. 74

75. (a)
$$K_{sp} = 4s^3 = 4 \times [2.5 \times 10^{-2}]^3 = 62.5 \times 10^{-6}$$
.

76. (c)
$$NaCl \Rightarrow \frac{Na}{S} + Cl$$

 $S = \frac{S}{S}$
 $K_{sp} = S^2, S = \sqrt{K_{sp}} = \sqrt{36} = 6.$

78. (d)
$$PbI_2 \approx Pb^{++} + 2I^-_{2S}$$

 $K_{sp} = 4S^3 = 4 \times [2 \times 10^{-3}]^3 = 32 \times 10^{-9}.$

(b) When ionic product is greater than K_{sp} then precipitation **79**. occur

$$K_{sp} < 10^{-2} M Ca^{2+} + 10^{-3} M F^{-}$$

(d) In IV^{*} group the S^{2-} concentration increase when added the 80. NH_4OH because

$$NH_4OH \Rightarrow NH_4^+ + OH$$

 $H_2S \Rightarrow 2H^+ + S^{2-}$

 $OH^- + H^+ \Rightarrow H_2O$. So that S^{2-} is increased.

81. (c)
$$BaSO_4 = Ba^{2+} + SO_4^{2-}$$

 $K_{sp} = S^2 \Rightarrow S = \sqrt{K_{sp}}; K_{sp} = [Ba^{2+}] \times [SO_4^{2-}]$
 $4 \times 10^{-10} = [1 \times 10^{-4}] \times [SO_4^{2-}]$
 $[SO_4^{2-}] = \frac{4 \times 10^{-10}}{1 \times 10^{-4}} = 4 \times 10^{-6}$.
82. (c) $AB_2 \Rightarrow A_{(S)}^{2+} + 2B_{(2S)^2}^{-}$
 $K_{sp} = 4S^3$
 $S = \sqrt[3]{\frac{K_{sp}}{4}} = \sqrt[3]{\frac{4 \times 10^{-12}}{4}} = 1 \times 10^{-4} gm.mol/litre$
83. (d) $[B] = \frac{K_{sp}AB}{[A]} = \frac{1 \times 10^{-8}}{10^{-3}} = 1 \times 10^{-5} M$
Where ionic product > K_{sp} , ppt formed
 \therefore 8 should be more then $10^{-5} M$.

(c)
$$NaCl_{(s)} \rightleftharpoons Na_{(aq)}^+ + Cl_{(aq)}^-$$

84.

85.

and

 $HCl \Rightarrow H^+ + Cl^-$. The increase in $[Cl^-]$ brings in an increase in $[Na^+]$ $[Cl^-]$ which will lead for backward reaction because

$$K_{sp}(NaCl) = [Na^+][Cl^-]$$

means lonic product $\geq K_{sp}$

(d)
$$BaSO_4 \approx Ba^{++}_{(S)} + SO_4^{--}_{(S)}$$

 $K_{sp} = S^2; S = \sqrt{K_{sp}} = \sqrt{1.3 \times 10^{-9}}$

$$= 3.6 \times 10^{-5} mol/litre$$

- (b) Alkaline, 87. $CH_3COONa + H_2O \rightleftharpoons CH_3COOH + NaOH_{Weak acid} + NaOH_{Strong base}$ Weak acid
- 88. (c) Because it is a strong base.

89. (b) For pure water
$$[H^+] = [OH^-]$$
, $\therefore K_w = 10^{-12}$ s

90. (a)
$$MX_2 \rightleftharpoons M_{(S)}^{2+} + 2X_{(2S)^2}^{-}$$
; $4S^3 = 4 \times (0.5 \times 10^{-4})^3$
= 5×10^{-13}

- (a) Solubility coefficient $= [Pb^{2+}] [Cl^{-}]^{2}$ 91.
- Solubility of $Al(OH)_3$ is lesser than $Zn(OH)_2$. 92. (a)
- (c) $NaCl_{(s)} \Rightarrow Na^+_{(aq)} + Cl^-_{(aq)}$ 93.

 $HCl \Rightarrow H^+ + Cl^-$ The increase in $\left[Cl^{-}
ight]$ brings in an increase in $[Na^+]$ $[Cl^-]$ which will lead for backward reaction because $K_{sp} NaCl = [Na^+] [Cl^-].$

- (c) Common ion effect. 94.
- (a) $CaF_2 \rightleftharpoons Ca^{++}_S + 2F^-_{(2S)^2}$ 95. $K_{sp} = 4S^{3}$

$$S = \sqrt[3]{\frac{K_{sp}}{4}} = \sqrt[3]{\frac{3.2 \times 10^{-11}}{4}} = 2 \times 10^{-4} m/.$$

96. (d) In aqueous solution following euilibrium is exist. $H_2S \Rightarrow H^+ + HS^-$

While adding the dilute HCl solution

 $(HCl \rightleftharpoons H^+ + Cl^-)$ equilibrium is shift to the left side in $H_2S \rightleftharpoons H^+ + HS^-$

97. (d)
$$M_{2X_{3}} \approx 2M^{+++} + 3X^{--}_{(2y)^{2}} + 3X^{--}_{(3y)^{3}}$$

Solubility product $K_{sp} = 108y^5 mol \frac{d}{m^3}$

98. (b) Solubility is directly proportional to the K_{sp} .

99. (b)
$$PbCl_2 = Pb_S^{++} + 2Cl_{(2S)^2}^{-}$$

 $K_{sp} = S \times (2S)^2 = 4S^3$
 $S = \sqrt[3]{\frac{K_{sp}}{4}} = \sqrt[3]{\frac{1.5 \times 10^{-4}}{4}} = 3.34 \times 10^{-2}$.

- 101. (d) $AlCl_3$ on hydrolysis gives weak base and strong acid among all.
- **102.** (c) Fe^{3+} ions are hydrolysed to develop acidic nature.

$$103. (c) K_h = \frac{K_w}{K_a \times K_b}$$

- **104.** (a) *KCN* is salt of strong base and weak acid.
- 105. (c) Sulphides of Group-II radicals have low solubility product.
- 106. (c) Because NH_3 acts as Lewis acid and they give electron pair to H_3O^+ ion. H_3O^+ is a Lewis base. Which accept the electron pair from NH_3 .
- 107. (c) Due to common ion effect.

108. (a) For
$$Ag_2SO_4 \rightarrow 2Ag^+ + SO_4^-$$

 $_{2x}^{2x} + SO_4^-$
 $_{x}^{x}$
 $K_{sp} = (2x)^2 \cdot x$; $K_{sp} = 4x^3$; $K_{sp} = 4 \times (2.5 \times 10^{-2})^3$
 $K_{sp} = 62.5 \times 10^{-6}$

109. (b) For
$$AgCl \rightarrow Ag^+ + Cl^-$$

 $x \qquad x$
 $K_{sp} = x^2$; $x = \sqrt{K_{sp}}$, $\sqrt{1 \times 10^{-6}} = 1 \times 10^{-3}$ mole / litre.

110. (b)

 $AgCl \to Ag^{=} + Cl^{-}$ $x \qquad x$

After *NaCl* is added $x = x + 1 \times 10^{-4}$

That is why Ag^+ will be less.

III. (b) Because of ionic product of AgI >> solubility product of its.

112. (a)
$$AX_2 \rightarrow A_1 + 2X_2 = X_1 + 2X_2 = 4x^3$$
; $x = \sqrt[3]{\frac{3.2 \times 10^{-11}}{4}}$; $x = 2 \times 10^{-4}$ mole/litre.

113. (c)
$$C_6H_5COONH_4 \rightarrow C_6H_5COO^- + NH_4^+$$

 $0.5 - \frac{.25}{100} \quad \frac{.25}{100} \quad \frac{.25}{100}$
According to Ostwald dilution law.

 $K = \frac{\alpha^2 C}{1 - \alpha} \qquad \qquad \left(\therefore \ \alpha = \frac{.25}{100} \right)$ $K = \alpha^2 C \qquad \qquad (\because \ 1 - \alpha = \text{Very small})$

$$K = \frac{.25}{100} \times \frac{.25}{100} \times .5 \quad ; \quad K = 3.125 \times 10^{-6}$$

(a)
$$Sb_2S_3 \rightarrow 2Sb^{+2} + 3S^{--}; K_{sp} = (2x)^2 . (3x)^3$$

 $2x^{--} 3x^{--}; K_{sp} = (2x)^2 . (3x)^3$

$$K_{sp} = 108x^5$$
; $K_{sp} = 108 \times (1 \times 10^{-5})^5 = 108 \times 10^{-25}$

115. (b) When increasing the temperature the value of ionic product also increases.

16. (a) Hydrolysis constant
$$h = \frac{K_w}{K_a}$$

114.

1

2.

5.

Hydrogen ion concentration - *p*H scale and Buffer solution

1. (a) pH of blood does not change because it is a buffer solution.

(c) 0.001 *M* of *NaOH* means
$$[OH^-] = .001$$

= $10^{-3} M \Rightarrow pOH = 3$

$$pH + pOH = 14 \Longrightarrow pH = 14 - 3 = 11$$

3. (d)
$$[H_3O^+]$$
 means $[H^+] = 6.2 \times 10^{-9} \ mol/l$

$$pH = -\log(6.2 \times 10^{-9}) = 8.21$$

4. (b)
$$CH_{N}H_{1} + HCI \longrightarrow CH_{3}NH_{3}^{+}CI^{-}$$

0.1 0.08 0
0.02 0 0.08
(Basic buffer solution)

$$pOH = pK + \log \frac{0.08}{0.02}$$

 $= pK + 0.602$
 $= 3.30 + 0.602 = 3.902$
 $\therefore pH = 10.09$

$$[H] = 7.99 \times 10^{-1} \approx 8 \times 10^{-1} M$$

(b)
$$pH + pOH = pK_w$$

6. (d)
$$pH = -\log[H^+]$$

$$5.4 = -\log[H^+]; [H^+] = 3.98 \times 10^{-6}$$

7. (a) $KCN + H_2O = KOH + HCN$. KOH is a strong base and HCN is a weak acid.

8. (c)
$$[H^+] = 10^{-3} M$$
, $pH = -\log[10^{-3}]$, $pH = 3$

9. (b)
$$[H^+] = [OH^-]$$

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

$$[H^{+}] = 10^{-7}, pH = -\log[H^{+}] = 7.$$

10. (d)
$$pH = 5$$
 means $[H^+] = 10^{-5}$
 $pOH = 14 - pH = 14 - 5 = 9$
 $[OH^-] = 10^{-pOH} = 10^{-9}$

n. (a) $pH = -\log [H^+]; [H^+] = 0.01 N$ $pH = -\log [10^{-2}]; pH = 2$ UNIVERSAL SELF SCORER

388 Ionic Equilibrium (d) 12. (b) $\frac{N}{100} = 0.01 N HCl; [H^+] = 10^{-2} M; pH = 2$ $BOH \longrightarrow B$ OH 32. Initial *C* – *Cα* Cα At ea Сα $[OH] = 10^{-2} M$ for NaOH $K_b = \frac{C^2 \alpha^2}{C(1-\alpha)} = C \alpha^2$ assuming $\alpha << 1; 1-\alpha \simeq 1$ pH + pOH = 14; pH = 14 - 2; pH = 12It is a buffer solution of strong acid and its weak conjugate (a) 34. $10^{-12} = 10^{-2} \times \alpha^2$; $\alpha^2 = 10^{-10}$; $\alpha = 10^{-5}$ base. $HA \rightleftharpoons H^+ + A^-$ 35. (c) $[OH^{-}] = C\alpha = .01 \times 10^{-5} = 10^{-7}$ $[H^+] = 0.1 M; [H^+]^2 = K_a \times C$ pH = 4 means; $[H^+] = 10^{-4}$ mol 13. (c) $[H^+] = \sqrt{K_- \times C} = \sqrt{1 \times 10^{-5} \times 0.1} = \sqrt{10^{-6}}$ Buffer solution is a mixture of weak acid and its conjugate 14. (a) hase. $[H^+] = 10^{-3} M; pH = 3$ Adding Na_2CO_3 to water makes the solution basic and hence 15. (b) a *pH* increases from 7. 36. (c) As the solution is acidic, pH < 7. This is because $[H^+]$ from NaClO is a salt of strong acid HClO₄. So it is a strong acid 16. (d) H_2O cannot be neglected in comparison to 10^{-8} . salt. Human body contain buffer solution. Its pH= 6.837. (c) NaOH is a base, so that its pH > 717. (b) It is a neutral solution and its pH=738. (a) 18 (c) It is a strong base. pH = 5, means $[H^+] = 10^{-5} M$. (a) 39 1 *M* NaOH has maximum $[OH^-]$ and minimum $[H^+]$ and After dilution $[H^+] = 10^{-5} / 100 = 10^{-7} M$ maximum *pH*. When pH = 7 means neutral, pH < 7 means acidic, pH > 719. (c) $[H^+]$ from H_2O cannot be neglected. means basic. Total $[H^+] = 10^{-7} + 10^{-7} = 2 \times 10^{-7}$ (d) As the solution is acidic, pH < 7. This is because $[H^+]$ from 20 $H_2O[10^{-7}M]$ cannot be neglected in comparison to 10^{-10} pH = 7 - 0.3010 = 6.6990 = 7 (neutral). (d) $[H^+] = \alpha.C = \frac{2}{100} \times .02; [H^+] = 4 \times 10^{-4} M$ 40. (c) It is a strong acid and they lose proton in a solution. 22. (a) $[OH^{-}] = 10^{-2} M; pOH = 2$ 23. $pH = -\log [H^+] = 4 - \log 4$; pH = 3.3979pH + pOH = 14; pH = 14 - pOH(a) $pH = pK_a + \log\left[\frac{\text{salt}}{\text{acid}}\right]$ pH = 14 - 2 = 1241. (d) Order of acidic strength is $H_2Te > H_2Se > H_2S > H_2O$ 24 $= 9.30 + \log \left[\frac{0.2}{0.1} \right] = 9.30 + 0.3010 = 9.6.$ Na_2O is a salt of NaOH + H_2O and H_2O is least acidic among given acids hence pH in this case will be max m. (a) $pH = pK_a + \log \frac{[\text{Salt}]}{[\text{Acid}]}$ *pH* of the solution A = 3(b) 25 42. $[H]_{}=10^{-3} M.$ *pH* of the solution B = 2 $pH = -\log(1.8 \times 10^{-5}) + \log \frac{[10]}{[100]}$ $[H]_{t} = 10^{-1} M$ $[H] = 10^{\circ} + 10^{\circ} = 10^{\circ} + 10 \times 10^{\circ} = 11 \times 10^{\circ}.$ $= -\log 1.8 + 5 + \log 10^{-1}$ $pH = -\log(11 \times 10^{-1}) = 3 - \log 11$ = 3 - 1.04 = 1.95 = -0.2553 + 5 - 1 = 3.7447 or = 4 $CN^- + H_2O \rightarrow HCN + OH^-$ 26. (a) (b) 20 ml of 0.1 NHCl= $\frac{0.1}{1000} \times 20g$ eq. = $2 \times 10^{-3}g$ eq. 43. Because OH^- concentration is increased. (a) On dilution the pH of acid A increases while pH of base B 27. 20*ml.* of 0.001 *KOH* = $\frac{0.001}{1000} \times 20 \text{ gm}$ eq. decreases. CH_3COONa is a salt of weak acid, (CH_3COOH) and 28 (d) strong base (NaOH). $= 2 \times 10^{-5} g eq.$ The equilibrium will shift in the backward direction. 29. (b) \therefore *HCl* left unneutralised = $2(10^{-3} - 10^{-5})$ (b) $K_a = 10^{-5}$; pH = 630. $= 2 \times 10^{-3} (1 - 0.01) = 2 \times 0.99 \times 10^{-3} = 1.98 \times 10^{-3} g \, eq.$ $pH = -\log K_a + \log \frac{[\text{Salt}]}{[\text{Acid}]}$; $6 = -\log 10^{-5} + \log \frac{[\text{Salt}]}{[\text{Acid}]}$ Volume of solution = 40 ml. $\therefore [HCI] = \frac{1.98 \times 10^{-3}}{40} \times 1000M = 4.95 \times 10^{-2}$ $6 = 5\log 10 + \log \frac{[Salt]}{[Acid]}; \ 6 = 5 + \log \frac{[Salt]}{[Acid]}$ $\log \frac{[Salt]}{[Acid]} = 6 - 5 = 1; \frac{[Salt]}{[Acid]} = \frac{10}{1}$ $\therefore pH = 2 - \log 4.95 = 2 - 0.7 = 1.3.$ (b) $10^{-7} M NaOH$ means $[OH^{-}] = 10^{-7}; \quad pOH = 7$ 45. (d) All are true - (a) is true for acid buffer, (b) for basic buffer, 31. pH = 14 - 7 = 7(c) is called buffer solution.

UNIVERSAL SELF SCORE **Ionic Equilibrium 389**

46. (a)
$$[H^+] = c \times \alpha = 0.1 \times \frac{30}{100} = 0.03 M$$

48. (a) The *pH* of buffer solution never changed.
49. (c) $[H^+] = \frac{10^{-14}}{10^{-1}} = 10^{-13} mol/litre pH = 13.$
50. (a) *pH* = -log[*H*⁺]; 7.4 = -log[*H*⁺]; [*H*⁺] = 4 × 10⁻⁸ *M*
51. (a) The *pH* of 0.1*M HCl* = 1 tonization of *H*₂*SO*₄ takes place in
two steps.
*H*₂*SO*₄ = *H*⁺ + *HSO*₄⁻ ; *HSO*₄⁻ = *H*⁺ + *SO*₄⁻⁻
52. (c) 1*N NaOH* solution have highest *pH*
[*OH*⁻] = 1; *pOH* = 0 ; *pH* + *pOH* = 14
pH = 14 - 0 = 14
53. (c) *H*₂*O* = [*H*⁺][*OH*⁻]
HCl = [*H*⁺][*PH*⁻]
HCl = [*H*⁺][*PH*⁻]
HCl = [*H*⁺][*PH*⁻]
HCl = [*H*⁺][*PH*⁻]
[*H*⁺] = 10⁻⁷ × $\frac{11}{10}$
pH = -log[*H*⁺] = -log $\left(10^{-7} + \frac{11}{10}\right)$; *pH* = 6.958
54. (c) *pK* = -log *K*, *pK* = -log *K*
pH = $-\frac{1}{2}$ [log *K*_a + log *K*_w - log *K*_b]
= $-\frac{1}{2}$ [-5 + log() × 10⁻¹⁴) - (-5)]
= $-\frac{1}{2}$ [-5 - 14 + 5] = $-\frac{1}{2}$ (-14) = 7
55. (d) *BaO*, *CaO* and *Na*₂*O* are shows more than 7 *pH* because of
their basic nature.
56. (a) *MgCl*₂ + 2*H*₂*O* = *Mg*(*OH*)₂ + 2*HCl*
57. (c) *H*₂*SO*₄ ionized in two step.
58. (b) *pH* = *pK*_a + log $\frac{[Salt]}{[Acid]}$
5.8 = 4.8 + log $\frac{[Salt]}{[Acid]}$ or log $\frac{[Salt]}{[Acid]}$ = 1.0
 $\frac{[Salt]}{[Acid]}$ = antilog 1.0 = 10
 $\therefore \frac{[Acid]}{[Salt]} = \frac{1}{10} = 0.1$

(c) (i) 20 ml of 0.5 N HCl 60.

 $0.5N \Rightarrow 1000 \, ml \, 0.5 \, mole \, HCl$ is present in $20 \, ml$

$$=\frac{20\times0.5}{1000}=1.0\times10^{-2}$$

(ii) 35 ml of 0.1 N NaOH

$$0.1N \Rightarrow 1000ml$$
 of 0.1 mole $NaOH$ is 35 ml

$$=\frac{35\times0.1}{1000}=0.35\times10^{-2}$$

Total = 20 + 35 = 55 ml.

$$\Rightarrow (1.0-0.35)10^{\circ}=0.65 \times 10^{\circ} \text{mole } HCl$$

HCl = *H*+*Cl*

$$\Rightarrow$$
 [*HCl*] = [*H*]+[*Cl*]

55 ml contains $0.65 \times 10^{\circ}$ mole of H ions

$$1000\,ml - \frac{0.65 \times 10^{-2} \times 10^3}{55} = \frac{6.5}{55}$$

 $pH = -\log[H^+] = -\log(6.5/55)$

 $= \log 55 - \log 6.5 = 0.92$

Due to acidic nature of solutions the colour of phenolphthalein becomes pink.

61. (b)
$$[H^+] = 2 \times 10^{-2} M$$

 of

$$\therefore pH = -\log\left[2 \times 10^{-2}\right];$$

pH = 1.7 *i.e.* in between 1 and 2.

63. (b)
$$pH = 4$$
, $(H^+) = 10^{-pH} = 10^{-4} M$

65. (c)
$$NaOH \Rightarrow Na^{+} + OH^{-}$$

 $[OH^{-}] = 10^{-5} M; [H^{+}] [OH^{-}] = 10^{-14}$
 $[H^{+}] = \frac{10^{-14}}{10^{-5}}; [H^{+}] = 10^{-9} M; pH = 9.$
67. (b) $pH = pK_{a} + \log \frac{[Salt]}{[Acid]}; pH = 4.75 + \log \frac{0.1}{0.1}$
 $pH = 4.75 + \log 1; pH = 4.75$
68. (d) A weak acid and its salt with a strong base maintain $pH = 4.55$
69. (a) $NaOH \Rightarrow Na^{+} + OH^{-} = [OH^{-}] = 10^{-8} MMM.$
70. (c) $[OH^{-}] = 0.0001 N, pOH = 4, pH + pOH = 14$

$$pH = 14 - pOH = 14 - 4 = 10$$

71. (d) 0.001 *M KOH* solution

$$[OH^{-}] = 0.001 M = 1 \times 10^{-3} M$$

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

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

$$[H^{+}] = \frac{1 \times 10^{-14}}{1 \times 10^{-3}} = 1 \times 10^{-14} \times 10^{+3}$$

$$[H^{+}] = 10^{-11} M$$

$$pH = 11$$
72. (a) An acid buffer solution consists of solution of strong base of its salt.

73. (b) An acid buffer solution consists of a weak acid and its salt with strong base. i.e. $CH_3COOH + CH_3COONa$

weak acid with

74. (a)
$$pOH = pK_b + \log \frac{[\text{salt}]}{[\text{base}]}$$

UNIVERSAL 390 Ionic Equilibrium

$$= 5 + \log \frac{0.02}{0.2} = 5 + \log \frac{1}{10} = 5 + (-1) = 4$$

$$pH = 14 - pOH = 14 - 4 = 10$$

75. (b) [Salt]= 0.1 *M*, [Acid]= 0.1 *M*

$$K_a = 1.8 \times 10^{-5} ; pH = -\log K_a + \log \frac{[Salt]}{[Acid]}$$

$$= -\log 1.8 \times 10^{-5} + \log \frac{0.1}{0.1} = -\log 1.8 \times 10^{-5}$$

$$pH = 4.7$$
.

- **76.** (a) NH_4Cl and NH_4OH is a buffer solution (weak base and salt of strong acid).
- **77.** (a) pH + pOH = 14; pH = 14 pOH

- ::[OH⁻] = 10⁻⁷pOH = 7 ∴ pH = 14 - 7 = 7.
- **78.** (c) 0.01 *M* $Ba(OH)_2 = 0.02N Ba(OH)_2$ $N_1V_1 = N_2V_2$

$$[0.02N] \times [50 ml] = N_2 \times 100 ml$$

$$N_2 = \frac{0.02 \times 50}{100} = 10^{-2} N ; [OH^-] = 10^{-2} N$$

pOH = 2 or pH = 12

- **79.** (b) $pH = -\log [H^+]$.
- **80.** (a) Na_2CO_3 is a mixture of weak acid and strong base, so it is a base.

81. (b)
$$10^{-7} N HCl$$
 means $(H^+) = 10^{-7} M$
 $pH = -\log(H^+), \quad pH = 7$

82. (c)
$$pH = 2$$
; $pH = -\log [H^+]$; $2 = -\log [H^+]$
 $[H^+] = 10^{-2} = 0.01 N$

- **83.** (b) pH does not change on addition of some concentration of HCl.
- **84.** (b) Solution of CH_3COONa on addition to acid shows a decrease in dissociation of acid due to common ion effect. To decrease in $[H^+]$ or increase *pH*.
- **85.** (c) pH + pOH = 14; pH = 14 pOH; pH = 14 6 = 8.
- **86.** (b) $[H^+]_{\rm I} = 10^{-5} [H^+]_{\rm II} = 10^{-2}$

Thus increase in $[H^+] = \frac{10^{-2}}{10^{-5}} = 1000$ times

87. (a) The *HCl* is a strong acid and they lose easily H^+ in solution.

88. (a)
$$X^- + H_2 O \rightleftharpoons OH^- + HX$$

$$K_{b} = \frac{[OH^{-}][HX]}{[X^{-}]}$$

$$HX \Rightarrow H^{+} + X^{-}$$

$$K_{a} = \frac{[H^{+}][X^{-}]}{[HX]}$$

$$\therefore K_{a} \times K_{b} = [H^{+}][OH^{-}] = K_{w} = 10^{-1}$$
Hence $K_{a} = 10^{-4}$
Now as $[X^{-}] = [HX], pH = pK_{a} = 4$.

- (d) Buffer solution is formed. So the *pH* will not change.
- **91.** (b) Na_2CO_3 when react with water form strong base and weak acid. So its aqueous solution is basic.

92. (c)
$$K_w = [H_3 O^+] [OH^-]$$

90.

Concentration of $H_3 O^+$ in distilled water $= 1 \times 10^{-6}$ mol/l.

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

$$K_w = [1 \times 10^{-6}] \times [1 \times 10^{-6}] = 1 \times 10^{-12}$$

- **93.** (a) $[OH^-] = 10^{-1}M$; pOH = 1pH + pOH = 14; pH = 14 - 1 = 13.
- **94.** (a) Maximum *pH HClO* is a weak acid all of these. So that the salt of weak acid is also weak.
- 95. (c) As the solution is acidic pH < 7. This is because $[H^+]$ from H_2O $[10^{-7}]$ cannot be neglected in comparison to 10^{-12} M.
- $\textbf{96.} \qquad (b) \quad [\text{Normal salt} + acidic salt] is a buffer solution.$
- 97. (b) 100 ml of $\frac{M}{10}$ NaOH = 50ml of $\frac{M}{5}$ NaOH. They exactly neutralise 50 ml $\frac{N}{5}$ HCl. Hence pH of resulting solution = 7.

98. (b)
$$M_1 = 6.0 \text{ M of } HCl$$
; $V_1 = ?$

 $M_{\rm 2}$ =0.30 M is $\,H^+$ concentration in solution.

$$V_2 = 150 \text{ ml of solution.}$$
$$M_1 V_1 = M_2 V_2; \ 6.0 \times V_1 = .30 \times 150$$
$$V_1 = \frac{.30 \times 150}{6} = 7.5 \text{ ml.}$$

99. (b)
$$pH = 3$$
, $[H^+] = 10^{-3} M$
 $\therefore [H^+] = \sqrt{K \times c}$
 $[10^{-3}]^2 = K \times c$; $\frac{[10^{-6}]}{0.1} =$

100. (b)When ratio of concentration of acid to salt is increased $p{\cal H}$ decrease. 101. (c) For NH_4OH .

 $K = 10^{-5}$

$$[OH^{-}] = C.\alpha$$
; $C = \frac{1}{10}M$, $\alpha = 0.2$

$$\begin{split} & [OH^{-}] = \frac{1}{10} \times 0.2 = 2 \times 10^{-2} \, M \\ & pOH = -\log \left[OH^{-} \right] = \log \left[2 \times 10^{-2} \right]; \ pOH = 1.7 \\ & pH = 14 - pOH = 14 - 1.7 = 12.30 \, . \\ & pH = pK_a + \log \frac{\left[\text{Salt} \right]}{\left[\text{Acid} \right]} \text{. For small concentration of buffering} \\ & \text{agent and for maximum buffer capacity } \frac{\left[\text{Salt} \right]}{\left[\text{Acid} \right]} \approx 1 \, . \\ & [H^{+}] = \text{ increased ten fold means } pH \text{ of solution decreased by} \\ & \text{one.} \\ & pH = \log \frac{1}{\left[H^{+} \right]} \\ & \text{Because the } pH \text{ of buffer are not changed.} \\ & pH = pK_a + \log \frac{\left[\text{Salt} \right]}{\left[\text{Acid} \right]}; \ 5.5 = 4.5 + \log \frac{\left[\text{Salt} \right]}{\left[0.1 \right]} \\ & \text{in6.} \quad (d) \\ & \log \frac{\left[\text{Salt} \right]}{0.1} = 5.5 - 4.5 = 1 \\ & \frac{\left[\text{Salt} \right]}{0.1} = \text{ antilog } 1 = 10; \quad \left[\text{Salt} \right] = 1 \\ & \text{inf.} \quad (d) \\ & \text{Moles of } H_2 SO_4 = \frac{0.49}{98} = 5 \times 10^{-3} \\ & \text{moles of } H_2 SO_4 = \frac{0.49}{98} = 5 \times 10^{-3} \\ & \text{moles of } H_2 SO_4 + 2H_2 O = 2H_3 O^+ + SO_4^{--} \\ & \text{inf.} \end{aligned}$$

one mole of H_2SO_4 give 2 moles of H_3O^+ ions.

$$H_3O^+ = 2 \times (H_2SO_4) = 2 \times 0.005 = 0.01 M$$

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

 $[OH^{-}] =$

102.

103.

104.

105.

106.

(c)

(a)

(a)

(c)

one.

0.1

(a) Moles of

107. (c) CH_3COONH_4 is a simple buffer and called salt of weak acid.

108. (c) N.eq. for
$$HCl = \frac{0.4}{1000} \times 50 = 0.02$$

N.eq. for $NaOH = \frac{0.2}{1000} \times 50 = 0.1$
Now $[OH^-]$ left $= 0.1 - 0.02$
 $[OH^-] = .08 = 8 \times 10^{-2} M$
 $pOH = -\log 8 \times 10^{-2} M$; $pOH = 1.0$
109. (d) Buffer is mixture of weak base and its acid salt.
110. (b) $[NaOH] = 0.4/40 \ mole/l. = 0.1 M$

$$[OH^{-}] = 10^{-1}M$$
, $[H^{+}] = 10^{-13}M$, $pH = 13$

111. (d)
$$pH + pOH = 14$$
, $pH = 4$, $H^+ = 10^{-4}$ mole/litre.

Buffer solution have constant *pH*. When we add the water into 112. (d) this buffer solution. So no effect on it.

113. (b)
$$Ba(OH)_2 \Rightarrow Ba^{2+} + 2OH^-$$

One molecule on dissociation furnishes $2 O H^-$ ions.

So,
$$[OH^-] = 2 \times 10^{-4} N$$

 $N = M \times 2$; $M = \frac{N}{2} = \frac{2 \times 10^{-4}}{2} = 10^{-4}$

 $pOH = -\log[OH^{-}] = -\log(1 \times 10^{-4}) = -4$ pH + pOH = 14; pH = 14 - 4 = 10.

(a) M.eq. of 0.10
$$M$$
 $HCl = \frac{0.10}{1000} \times 40 = 0.004 M$

M.eq. of 0.45 *M* $NaOH = \frac{0.45 \times 10}{1000} = 0.0045M$

Now left $[OH^-] = 0.0045 - 0.004 = 5 \times 10^{-4} M$ Total volume = 50 *ml*.

$$[OH^{-}] = \frac{5 \times 10^{-4}}{50} \times 1000 ; [OH^{-}] = 1 \times 10^{-2}$$

pOH = 2 ; pH = 14 - pOH = 12.

5. (c) 0.001
$$M HCl = 10^{-3} M [H^+], pH = 3.$$

6. (d)
$$[NaOH] = \frac{0.4}{40} = 0.01M; [OH^-] = 10^{-2} M$$

$$[H^+] = 10^{-12}, pH = -\log[H^+] = 12$$

b) Those substance which give a proton is called Bronsted acid while CH_3COO^- doesn't have proton so it is not a Bronsted acid.

20. (c)
$$pH = -\log [H^+]$$

121. (d)
$$pH + pOH = 14, pH = 4.0$$

 $pOH = 14 - pH$; $pOH = 14 - 4.0 = 10.0$

- (b) pH = 0 means $[H^+] = 10^\circ = 1M$. Hence solution is strongly 122. acidic.
- (c) As the solution is acidic, pH < 7. This is because $[H^+]$ from 123. H_2O (10⁻⁷ M) cannot be neglected in comparison to 10^{-10} *M* HCl.

124. (d)
$$H_3O^+ \Rightarrow OH^- + H_2$$

 $pOH + pH = 14$; 7 + 7 = 14; $[H^+] + [OH^-] = 10^{-14}$
 $10^{-7} + 10^{-7} = 10^{-14}$; $[OH^-] = 10^{-7}$ gm ion/l.

(b) When pH = 2, $[H^+] = 10^{-2} M$ 126.

127. (a)
$$[OH^-]$$
 ion conc. $= 0.05 \frac{mol}{l} = 5 \times 10^{-2} \frac{mol}{l}$

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

 $pOH = 1.30; pH + pOH = 14$
 $pH = 14 - pOH = 14 - 1.30 = 12.7$

- (c) When pH = 3, then $[H^+] = 10^{-3} M$ after that we increased the pH from 3 to 6 then $[H^+] = 10^{-6} M$ means reduced 1000 times.
- (b) CO_2 is acidic oxide which on dissolution in water develops 129. acidic nature.
- (d) If pH of any solution is 2. 130.

128.

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

SCORER 392 Ionic Equilibrium

If pH of any solution is just double then pH = 4 and $[H^+]$ will be 10^{-4} .

131. (c) A strong acid is not used to make a buffer.

132. (d)
$$pH = 1$$
 means $[H^+] = 10^{-1} M$

Hence
$$[H_2 SO_4] = \frac{10^{-1}}{2} = \frac{1}{20} = 0.05 M$$

- **133.** (c) The pH of blood is 7.4 due to presence of bicarbonates ions
- 134. (c) As the solution is acidic, pH < 7. This is because $[H^+]$ from H_2O $[10^{-7}]$ cannot be neglected in comparison to 10^{-8}
- **135.** (a) pH will decrease because $[OH^-]$ increased due to this pOH is decreased.

136. (c)
$$[H^+] = 6 \times 10^{-4} M$$

 $pH = -\log [H^+] = -\log [6 \times 10^{-4}] = 3.22$.

- 137. (c) $0.01 M HCl = 10^{-2} M [H^+], pH = 2.$
- **138.** (c) Because buffer solution have a constant *pH*.

139. (c)
$$10^{-6} M HCl = 10^{-8} M [H^+]$$
. Also from H_2O

 $[H^+] = 10^{-7} M$ Total $[H^+] = 10^{-7} + 10^{-8} = 10^{-7} [1 + 0.1] = 10^{-7} [1.1]$ Hence pH = 7 - 0.0414 = 6.96.

- 140. (b) $10^{-10} M HCl = 10^{-10} M [H^+]$. But $pH \neq 10$ because solution is acidic. This is because H^+ from $H_2 O(10^{-7} M)$ cannot be neglected.
 - Total $[H^+] = 10^{-7} + 10^{-10}$

$$=10^{-7} + (1 + 10^{-3}) = 10^{-7} (1.001)$$

That is why pH = 7 (slightly less than 7)

141. (b) $[H^+] = 1.00 \times 10^{-6}$ mole/litre

 $pH = -\log [1.00 \times 10^{-6}]$; pH = 6.

- **142.** (a) $[H^+]$ is in moles per litre.
- 143. (d) As the solution is acidic, pH < 7. This is because $[H^+]$ from H_2O $(10^{-7}M)$ cannot be neglected in comparison to 10^{-8} .
- **145.** (b) pH of 0.001 M $HCl = 10^{-3} M[H^+]$, pH = 3.
- **146.** (d) Because it can furnish H^+ ions in solutions.
- **147.** (c) Because it is a strong acid.

$$H^+ = 10^{-1}$$

$$pH = -\log[H^+] = -\log[10^{-1}]; pH = 1.$$

148. (b) Buffer solution is a combination of weak acid and conjugate base. NaCl is a salt and NaOH is the base.

149. (a)
$$[H^+] = \sqrt{Kc} = \sqrt{10^{-5} \times 0.1} = 10^{-3}, \ pH = 3.$$

150. (d) In
$$\frac{N}{10}$$
 NaOH have $[OH^-] = 10^{-1}$ M means $pOH = 1$ and
then $pH + pOH = 14$
 $pH = 14 - pOH = 13$.

- 151. (b) Borate ions are hydrolyzed to develop alkaline nature in solution.
- **152.** (d) Less the pH, more acidic is the solution.
- **154.** (b) The equal conc. of salt and acid.

157.

166

3.7

155. (c)
$$pH = -\log K_a + \log \frac{[KCN]}{[HCN]}$$

 $pH = -\log[5 \times 10^{-10}] + \log \left(\frac{0.15}{1.5}\right) = 8.302$

(c)
$$pH = pK_a + \log \frac{[Salt]}{[Acid]}$$
 equimolar means

$$\frac{[Salt]}{[Acid]} = 1; \quad pH = 4.74 + 0 = 4.74$$

- **158.** (a) Because of *NaCl* is a salt of strong acid and strong base. So that it is neutral.
- 159. (c) When strong acid and strong base are react neutral salt are formed. So that NaCl is a neutral salt.

162. (d)
$$pH = -\log K_b + \log \frac{[Salt]}{[Acid]}$$

 $pH = -\log[1.8 \times 10^{-5}] + \log \frac{[Salt]}{1.0}$
 $9 = 4.7 + \log \frac{[Salt]}{1.0}$; $\log \frac{[Salt]}{1.0} = 4.7 - 9 = -4.3$
 $\frac{[Salt]}{1.0} = \operatorname{Antilog} \frac{1}{4.3}$; $[Salt] = 1.8$
163. (b) $pH = -\log K_b + \log \frac{[Salt]}{[acid]}$
 $5 = -\log 10^{-4} + \log \frac{[Salt]}{[acid]}$
 $\log \frac{[Salt]}{[acid]} = 1$
 $\frac{[Salt]}{[acid]} = 1$

164. (a) 1 *M KOH* show highest *pH* value because it is a strong base.
165. (d) *NH*₄*OH* is a weak acid and *NH*₄*Cl* is a strong base salt.

(a)
$$pH = 13.6$$

$$pOH = 14 - 13.6 = 0.4$$

 $[OH^{-}] = Antilog(-0.4) = 0.3979$. So the value of

$$[OH^{-}]$$
 between 0.1 M and 1 M

167. (d) Aspirin is a weak acid. Due to common ion effect it is unionised in acid medium but completely ionised in alkaline medium.

168. (b)
$$[H^+][OH^-] = 10^{-14}; (10^{-7})(10^{-7}) = 10^{-14}$$

- **169.** (c) $HCl = 10^{\circ} M$ has pH = 0. The value of pH decreases as concentration further increases.
- **170.** (a) Because pure water has a 7 *pH*.
- **171.** (c) When concentration of $[H^+]$ increased then the value of pH is decreases.

Ionic Equilibrium 393 SELF SCOR

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

(c) The concentration of $[H^+] = 10^{-2}$ mole/litre 172.

$$pH = -\log[H^+] = -\log[10^{-2}]; pH = 2$$

- (d) Due to common ion effect. 173. 174.
 - (b) In water solution.
 - $NH_3 + H_2O \Rightarrow NH_4^+ + OH^-$

concentration of OH^- is increased so that solution become more basic and the *pH* is increased.

. . .

- Na_2CO_3 is basic in nature. So its *pH* is greater than 7. 175. (a)
- (c) It is not a mixture of weak acid or base and their strong salt. 176.

177. (a)
$$[H^+] = \text{Antilog}(-4.58);$$

 $[H^+] = 2.63 \times 10^{-5} \text{ moles}/li$

$$[H^+] = 2.63 \times 10^{-5} \text{ moles/litre}$$

(c) $10^{-2} M NaOH$ will give $[OH^{-}] = 10^{-2}$ 178. $\therefore pOH = 2$, Also pH + pOH = 14 $\therefore pH = 12$. ra 1.1

179. (a)
$$pH = pK_a + \log \frac{|\text{Salt}|}{|\text{Acid}|} = -\log 2 \times 10^{-5} + \log \frac{10 \times 1}{50 \times 2} = 4$$
.

(b) 0.001 *M* NaOH means $[OH^{-}] = 10^{-3}$; pOH = 3180. pH + pOH = 14; pH = 14 - 3pH = 11; $[H^+] = 10^{-11}$ mole-litre

181. (c)
$$[H^+] = C.a$$

$$[H^+] = 0.1 \times \frac{1}{100} = 10^{-3}$$

$$pH = -\log[H^+] = -\log 10^{-3} = 3$$

(b) pH = 4182.

$$pH + pOH = 14$$
; $pOH = 14 - pH$

$$pOH = 14 - 4 = 10; [OH^{-}] = 10^{-10} M$$

183. (b)
$$pH = \log \frac{1}{[H^+]} = \log \frac{1}{[3 \times 10^{-3}]} = 2.523$$

- (c) It is a strong base. 185.
- (b) $[H^+]$ Concentration in 0.01*M HCl* is 10^{-2} *M* because 0.01 *M* 186 *HCl* have only H^+ $HCl \Rightarrow H^+ + Cl^-.$
- (b) The value of H_3O^+ ions will not changed. 187.

$$CH_3COOH + H_2O \Rightarrow CH_3COO^- + H_3O^+$$
.

(a) H_2O_2 (Hydrogen peroxide) is a corrosive volatile liquid. It is 188. slightly acidic in nature. Its pKa value is approximately 10^{-12}

192. (b)
$$pH = \log \frac{1}{H^+}$$
; $pH = \log \frac{1}{10^{-3}}$; $pH = 3$.

(b) NH₃ is Lewis base because of one lone pair of electron. 193.

94. (b)
$$Ba(OH)_2 \rightarrow Ba^{+2} + 2OH^-$$

.05*M* 2×0.5*M*

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

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

- If concentration of acid is increases ten times in a buffer then 195. (a) pH of the solution is increase by one.
- pH > 7 = Basic108 (b) It means contain more hydroxide ions than carbonate ions.

199. (a) At
$$7pH$$
 the concentration of OH^- and H^+ are equal.

200. (a,d) M.eq. of 0.01
$$M HCl = \frac{.01 \times 100}{1000} = 1 \times 10^{-3}$$

 $pH = 3$
M.eq. of .02 $M H_2SO_4 = \frac{.04 \times 50}{1000} = 2 \times 10^{-3}$
M.eq. of .02 $M NaOH = \frac{0.02 \times 50}{1000} = 1 \times 10^{-3}$
Left $[H^+] = 2 \times 10^{-3} - 1 \times 10^{-3} = 1 \times 10^{-3}; pH = 3$

- (a,b,c) Because buffer solution are mixture of weak acid or weak 202. base and their salt.
- (c) Because pH = 8 is basic nature but HCl is a strong acid. 203.

204. (c)
$$H_2 SO_4 = 0.05 \times 2$$

$$\therefore [H^+] = 0.1$$
 and $pH = 1$

205. (b)
$$Mg(OH)_2 \Rightarrow Mg^{2+} + 2OH^-$$

 $K_{sp} = [Mg^{2+}][OH^-]^2$
 $1 \times 10^{-12} = 0.01[OH^-]^2$
 $[OH^-]^2 = 1 \times 10^{-10} \Rightarrow [OH^-] = 10^{-5}$
 $[H^+] = 10^{-14} / 10^{-5} = 10^9$
 $pH = -\log[H^+] = -\log[10^{-9}] = 9$
206. (b) $[OH^-] = 1 \times 10^{-5}$
 $pOH = -\log[OH^-] = 5$

 $pH + pOH = 14 \Longrightarrow pH = 14 - 5 = 9$.

Critical Thinking Questions

- (b) K_{sp} of $AgI = 1.5 \times 10^{-16}$ $10^{-8} M Ag^+$ and $10^{-8} M I^-$ Ionic product = 10^{-16} K_{sp} = Ionic product
- HClO is the weakest acid. Its conjugate base ClO⁻ is the 2. (a) strongest base.

(b)
$$B(OH)_3$$
 not have H^+

1.

з.

4.

(a)
$$h = \frac{k_w}{k_a} = \frac{1 \times 10^{-14}}{4.5 \times 10^{-10}} = 2.22 \times 10^{-5}$$

5. (d)
$$MX_4 \rightarrow M + 4X_{s}$$
; $K_{sp} = (4s)^4 s$; $K_{sp} = 256s^5$

$$s = \left(\frac{K_{sp}}{256}\right)^{1/5}.$$

6. (a) Lewis acid are electrophiles because they accept election pairs.

SELF SCORER 394 Ionic Equilibrium

7. (c) $\underset{\text{Initial}}{\text{Initial}} 2HI = H_2 + I_2 \\ _0 0$

At equilibrium
$$2-\alpha \quad \frac{\alpha}{2} \quad \frac{\alpha}{2} = 2-\alpha + \frac{\alpha}{2} + \frac{\alpha}{2} = 2$$
.

- **8.** (a) Because they have vacant *d*-orbital in central atom.
- $\label{eq:general} \textbf{9.} \qquad (a) \quad \text{Energy has to be spent for the total dissociation of weak acid.}$
- **10.** (a) pK_a of acid A = 4; pK_a of acid B = 5

We know that $pK_a = -\log K_a$

 \therefore Acid A $K_a = 10^{-4}$

Acid *B* $K_a = 10^{-5}$

Hence A is ten times stronger than that of B.

n. (b)
$$\frac{\alpha_1}{\alpha_2} = \sqrt{\frac{K_{a_1}}{K_{a_2}}} = \sqrt{\frac{3.14 \times 10^{-4}}{1.6 \times 10^{-5}}} = 4:1$$

 $\label{eq:linear} \textbf{12.} \qquad (d) \quad \textbf{lt is a salt of weak acid and weak base.}$

13. (c)
$$HA \rightleftharpoons H^+ + A^-$$
; $K_a = \frac{[H^+][A^-]}{[HA]}$ (i)

neutralization of the weak acid with strong base is

$$HA + OH^{-} \rightleftharpoons A^{-} + H_2O$$
$$K = \frac{[A^{-}]}{[HA] [OH^{-}]}$$

dividing (i) by (ii)
$$\frac{K_a}{K} = [H^+][OH^-] = K_w = 10^{-14}$$

$$K = \frac{K_a}{K_w} = \frac{10^{-5}}{10^{-14}} = 10^9.$$

.....(ii)

14. (d) $NH_4 Cl$ undergoes cationic hydrolysis hence pH is >7 because the solution due to cationic hydrolysis in acids. *NaCN* undergoes anionic hydrolysis hence pH is >7. *HCl* is strong acid and *NaCl* is neutral solution. Hence the pH of given solutions will increases. $HCl < NaCl < NaCN < NH_4 Cl$

15. (c) It is a $HClO_4$

- **16.** (d) Smaller the ${}_{p}K_{a}$ value, strong the acid.
- 17. (d) It involves gain and loss of electron pair (Lewis concept).

18. (b)
$$H^{-}_{(aq.)} + H_2 O_{(l)} \rightarrow OH^{-} + H_2$$
.

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

- **20.** (d) Acidity is directly propotional to oxidation number. As the O.No. of *S*, *P* and *Cl* in H_2SO_3 , H_3PO_3 & $HClO_3$ is +4, +3 & +5 respectively so decreasing order of acidity will be III > I > II.
- **21.** (a) A substance which can donate a proton is known as acid so NH_4^+ will be a acid.
- **22.** (b) Acetic acid is a weak acid.
- 23. (a) Because it is a acidic oxide.

24. (b)
$$[Ca^{2+}][F^{-}]^2 = \left[\frac{10^{-2} \times V}{2V}\right] \times \left[\frac{10^{-3} \times V}{2V}\right]^2 = 1.25 \times 10^{-9}$$

25. (a) The degree of hydrolysis of a salt of weak acid and weak base is independent of concentration of salt.

26. (c)
$$C = 0.1M; \ \alpha = 1\%; \ (H^+) = C \times \alpha$$

=
$$0.1 \times \frac{1}{100} = 10^{-3}$$
; (H⁺) = 10^{-3} ; pH = 3.

- **27.** (d) $HClO_4$ is a strong acid, because its oxidation no is + 7.
- **28.** (d) Acceptor of electron pair is known as lewis acid. *S*, : $CH_2, (CH_3)_3 B$ all can accept an electron pair so answer is (d).

29. (b)
$$Mg(OH)^2 \approx Mg^{2+} + 2OH^{-}_{(2s)^2}$$

 $K_{sp} = 4S^3 \Rightarrow S = \sqrt[3]{\frac{K_{sp}}{4}} = \sqrt[3]{\frac{1.2 \times 10^{-11}}{4}}$
 $S = 8.16 \times 10^{-4}$
30. (d) $K_a = \frac{[H^+][CN^-]}{[HCN^-]}$
 $6.2 \times 10^{-10} = \frac{[H^+][0.02]}{[0.01]}$
 $[H^+] = \frac{6.2 \times 10^{-10} \times 0.01}{0.02} = 3.1 \times 10^{-10}$

31. (b) NH_3 presence of lone pair of electrons.

32. (a)
$$CuBr \approx Cu^+ + Br^-_{(S)}$$

 $K_{sp} = S^2 = (2 \times 10^{-4})^2 = 4 \times 10^{-8} \frac{mol^2}{r^2}$

33. (c)
$$Na_2SO_4 = 2Na^+ + SO_4^2$$

(0.004-x) $2x x$

Since both the solution are isotonic 0.004 + 2x = 0.01 $\therefore x = 3 \times 10^{-3}$

: Percent dissociation
$$=\frac{3 \times 10^{-3}}{0.004} \times 100 = 75\%$$

34. (a)
$$Cr(OH)_3 \rightarrow C_x^{+3} + 3OH^-$$

 $K_{sp} = x.(3x)^3 = 27x^4$
 $x = 4\sqrt{\frac{K_{sp}}{27}} ; x = 4\sqrt{\frac{2.7 \times 10^{-31}}{27}}$
 $x = 1 \times 10^{-8} \text{ mole/litre.}$
35. (a) $H^+ = c.\alpha = \sqrt{K_a.c}$

$$pH = -\log(\sqrt{K_a.c})^{1/2} = \frac{1}{2}[-\log K_a - \log c]$$

$$= \frac{1}{2} [4.74 - \log 10^{-2}] = \frac{1}{2} [4.74 + 2] = 3.37.$$
36. (b) $Ag_2SO_4 \approx 2Ag^+ + SO_4^{--}$
 $K_{sp} = 4S^3; K_{sp} = 2 \times 10^{-5}$
 $S = \sqrt[3]{\frac{2 \times 10^{-5}}{4}} = 0.017 \, m/l = 1.7 \times 10^{-2}$
 $Ag BrO_3 \approx Ag^+ + BrO_3^-$
 $K_{sp} = S^2; K_{sp} = 5.5 \times 10^{-5}$
 $S = \sqrt{5.5 \times 10^{-5}} = 7.4 \times 10^{-3} \, m/l.$
37. (c)

37 (c)

> Phenol are more acidic than ethanol ОН because of stabilization of phenoxide as compare to ethoxide ion.

38. (b)
$$X^- + H_2 O \Rightarrow HX + OH^-$$

 $K_h = \frac{10^{-14}}{10^{-5}}$ so $h = \sqrt{\frac{10^{-9}}{10^{-1}}} = 10^{-4}$

 $100 \times 10^{-4} = 10^{-2}$

So, 0.01%.

(a) $K_a = 1.0 \times 10^{-5}$ 39.

 $K_h = hydrolysisconstant$

$$K_h = \frac{K_w}{K_a} = \frac{10^{-14}}{10^{-5}} = 10^{-9}$$

degree of hydrolysis (*h*) = $\sqrt{\frac{K_h}{C}}$

$$=\sqrt{\frac{10^{-9}}{0.001}} = \sqrt{10^{-6}} = 10^{-3}; \quad h = 10^{-3}$$

- (b) Basic radicals of group 11 & 1V are precipitated by $\,H_2S\,$ in the 40. form of their sulphides. 11nd group in acidic medium & 1V group in alkaline medium. They precipitate when ionic product increases than solubility product.
- (a) After mixing $[Ag^+][Cl^-] > K_{sp}$ 41.
- (a) $K_{sp} > \text{ ionic product } 1 \times 10^{-10} > 5 \times 10^{-11}$ 42.
- (a) 1 litre of water contains 1000/18 mole. 43.

So degree of ionization
$$=\frac{10^{-7} \times 18}{1000} = 1.8 \times 10^{-7}$$
 %.

44. (a) AgCl
$$K_{sp} = 1.2 \times 10^{-10}$$

$$S = \sqrt{1.2 \times 10^{-10}} ; S = 1.09 \times 10^{-5}$$

AgBr $K_{sp} = 3.5 \times 10^{-13}$
 $S = \sqrt{3.5 \times 10^{-13}} = 5.91 \times 10^{-6}$

45. (a)
$$K_{sp} = 4S^{3}$$

 $4S^{3} = 3.2 \times 10^{-8}$; $S = 2 \times 10^{-3} M$.
46. (c) $\frac{2.8 \times 10^{-10}}{0.1} = 2.8 \times 10^{-9} ML^{-1}$.
47. (b) $AB \Rightarrow A^{+} + B^{-}$; $K_{sp} = S^{2}$

$$S = \sqrt{K_{sp}} = \sqrt{4 \times 10^{-10}} = 2 \times 10^{-5}$$

48. (b)
$$S = 1.435 \times 10^{-3} g/l$$
, $= \frac{1.435 \times 10^{-3}}{143.5} = 10^{-5} M$

$$K_{sp} = S \times S = 10^{-10}$$

49. (b)
$$pKa = 5$$
, so $Ka = 1 \times 10^{-5}$

$$\alpha = \sqrt{\frac{Ka}{C}} = \sqrt{\frac{1 \times 10^{-5}}{0.1}} = 1 \times 10^{-2}$$

50. (c) Presence of common ion decreases the solubility of salt.

51. (d)
$$Ag_2S > CuS > HgS$$
.

Solubility of
$$CuS = \sqrt{10^{-31}} = 3.16 \times 10^{-16} mol/lit$$

Solubility of Ag_2S

$$=\sqrt[3]{\frac{K_{sp}}{4}} = \sqrt[3]{\frac{10^{-42}}{4}} = 6.3 \times 10^{-5} mole/litre$$

Solubility of $HgS = \sqrt{K_{sp}} = \sqrt{10^{-54}} = 10^{-27} \ mol \,/ \, litre$

(c) $pH = 3.82 = -\log[H^+]$ 52.

$$(H^+) = 1.5 \times 10^{-4}$$
 mole/litre.

53. (b)
$$pH = pK_a + \log \frac{[\text{salt}]}{\text{acid}} = 4.57 + \log \frac{0.10}{0.03} = 5.09$$

55.

$$[H^+] = C\alpha$$
$$= \frac{1}{10} \times 0.001 = 10^{-4} \implies pH = 4$$

(a)
$$K_{sp} = [As^{3+}][S^{2-}], S = 5\sqrt[5]{\frac{sp}{108}}$$

= $5\sqrt{\frac{2.8 \times 10^{-72}}{108}} = 1.09 \times 10^{-15}$

(d) Dissociation constant of $HA = 10^{-9}$ 56.

$$HA \Rightarrow H^{+} + A^{-}$$

$$[H^{+}] = \sqrt{\frac{K_{a}}{C}} = \sqrt{\frac{10^{-9}}{0.1}} ; [H^{+}] = 10^{-4}$$

$$\therefore pH = 4$$

$$\therefore pH + pOH = 14$$

$$pOH = 14 - pH = 14 - 4 ; pOH = 10$$

UNIVERSAL SELF SCORER 396 Ionic Equilibrium

57. (d)
$$\alpha = 1.9 \times 10^{-9}$$
; $C = \frac{1000}{18}$
 $K = \frac{[H^+][OH^-]}{(H_2O)} = C\alpha^2$
 $= 1.9 \times 10^{-9} \times 1.9 \times 10^{-9} \times \frac{1000}{18} = 2.0 \times 10^{-16}$. 66.
58. (c) $K = Ka_1 \times Ka_2 = 4.5 \times 10^{-3} \times 1.7 \times 10^{10}$
 $H^+ = \sqrt{Kc} = \sqrt{4.5 \times 10^{-3} \times 1.7 \times 10^{-10} \times .01}$
 $= .87 \times 10^{-7}$
 $pH = -10g \ 0.87 \times 10^{-7} = 7 - 0.93 = 6.07$.
59. (a) Given that
Concentration of solution =.1
Degree of ionisation $= 2\% = \frac{2}{100} = .02$
Ionic product of water $= 1 \times 10^{-14}$
Concentration of $[H^+] = \text{Concentration of solution X degree of}$
ionisation $= .1 \times .02 = 2 \times 10^{-3} M$
69. Concentration of $[OH^-] = \frac{\text{Ionic product of water}}{[H^+]}$
 $= \frac{1 \times 10^{-14}}{2 \times 10^{-3}} = 0.5 \times 10^{-11} = 5 \times 10^{-12} M$.
60. (a) $HgSO_4 \rightarrow Hg^{++} SO_4^{--}$
 $K_{\psi} = x^2$; $x = \sqrt{K_{\psi}}$; $x = \sqrt{6.4 \times 10^{-5}}$
 $x = 8 \times 10^{-3} \text{ mole/litre.}$
61. (a) $K = c\alpha^2 = 0.1 \times (\frac{1.34}{100})^2 = 1.8 \times 10^{-5}$.
62. (c) $[H^+] = 1 \times 10^{-4} M \Rightarrow [H^+][OH^-] = 1 \times 10^{-14}$
 $\therefore [OH^-] = \frac{1 \times 10^{-14}}{0.5 \times 10^{-4}} = 2 \times 10^{-10} M$.
63. (d) Because $HCIO_4$ is a strong acid. While buffer is a mixture of weak acid and their salt.
64. (b) Initial concentration = 0.006 M $\Rightarrow K_a = 6 \times 10^{-5}$
equilibrium reaction.
 $C_6H_5COOH = C_6H_5COO^- + H^+$
 $C(1^{-0}) \qquad C_a \qquad C_a \qquad C_a \ initial$
 $K_a = C\alpha^2 \qquad (:1 - \alpha \approx 1)$ for weak electrotyte
 $\alpha = \sqrt{\frac{K_a}{c}}$; $\alpha = \sqrt{\frac{6 \times 10^{-5}}{0.006}}$; $\alpha = 10^{-1}$
 $\therefore [H^+] = C\alpha = 0.006 \times 10^{-1} = 6 \times 10^{-4} M$.
65. (a) $pOH = pKb + \log \frac{[Sall]}{[Base]}$

 $14 - 9.35 = -\log(1.78 \times 10^{-5}) + \log\frac{[Salt]}{100}$ $[Salt] = 79.9 \Rightarrow \frac{w}{132} \times 1000 = 79.9 \Rightarrow w = 10.56$ (a) pH = 6 means $[H^+] = 10^{-6} M$ pH = 3 means $[H^{=}] = 10^{-3}M$ After mixing, Total $[H^+] = \frac{(10^{-6} + 10^{-3})}{2} = \frac{1.001 \times 10^{-3}}{2}$ $= 5.005 \times 10^{-4}$ $pH = 4 - \log 5.005$; pH = 3.301. (b) $[H^+] = \sqrt{K \times C}$ $[H^+] = \sqrt{4 \times 10^{-10} \times 1}$: $[H^+] = 2 \times 10^{-5}$ mole/litre (b) Moles of $[OH^-] = M \times V$ Number of $NaOH = 0.3 \times 0.005 \times 2 = 0.0030$ (d) $[H^+][OH^-] = 10^{-13.26}$ $\therefore [H^+] = [OH^-], \therefore [H^+]^2 = 10^{-13.26}$ $[H^+] = 10^{\frac{-13.26}{2}}$ *pH* = 6.63. (b) $pH = pKa + \log \frac{[Salt]}{[Acid]}$ pH = pKa $Ka = 0.1 \times (10^{-3.5})^2 = 0.1 \times 10^{-7} = 10^{-8} \implies pH = 8$ (a) $pH = -\log K_a + \log \frac{[\text{Salt}]}{[\text{Acid}]}$ $[Salt] = \frac{0.2 \times 50}{1000} = 0.01; \ [Acid] = \frac{0.5 \times 40}{1000} = 0.02$ $pH = -\log(1.8 \times 10^{-4}) + \log\frac{0.01}{0.02}$ $pH = 4 - \log(1.8) + \log 0.5$ $pH = 4 - \log(1.8) - 0.301$ pH = 3.4(c) HCO_3^{-} and H_2O (a) $[H^+] = C \times \alpha = 0.1 \times 0.1 = 10^{-2} M$ $pH = 2; pOH = 12; [OH^{-}] = 10^{-12} M.$

- **14.** (c) Decreasing order of acidic character is $H_2SO_4 > CH_3COOH > H_2CO_3$
- 75. (c) The acidic nature increases in the order. $NaNO_2 < NaCl < H_2S < H_2SO_4$

76. (a) $[OH^{-}] = 0.05 = 5 \times 10^{-2} M$

 $pOH = 2 - \log 5 = 1.3$ pH + pOH = 14pH = 14 - 1.3 = 12.7

77. (b) Na_2O form *NaOH*. So that it is basic oxide.

78. (a) $CH_3COO^- + H_2O \Rightarrow CH_3COOH + OH^-$

$$\therefore [OH^{-}] = c \times h; \quad h = \sqrt{\frac{K_w}{K_a} \times c} = \sqrt{\frac{10^{-14}}{1.8 \times 10^{-5}} \times 1}$$
$$= 2.35 \times 10^{-5}$$

: pOH = 4.62; $pH = 9.38 \approx 9.4$

Assertion & Reason

- 1. (c) HCl is a strong electrolyte since it will produce more H^+ , comparison than that of CH_3COOH . Hence assertion is true but reason false.
- (b) For sparingly soluble salts, reason is not a correct explanation. Hence both assertion and reason are true but reason is not a correct explanation of assertion.
- **3.** (a) Aq. solution of $FeCl_3$ on standing produce brown ppt. Due to hydrolysis it produce ppt. of $Fe(OH)_3$ which is of brown colour. Hence both are correct and reason is a correct explanation.
- 4. (a) Barium carbonate is more soluble in HNO_3 than in water become carbonate is a weak base and reacts with the H^+ ion of HNO_3 causing the barium salt to dissociate.

 $BaCO_3 + HNO_3 \rightarrow Ba(NO_3)_2 + CO_2 + H_2O_3$

- 5. (a) The conjugate base of $CHCl_3$ is more stable than conjugate base of $CHF_3(CF_3)$. CCl_3 stabilized by -1 effect of chlorine atoms as well as by the electrons. But conjugate base of $CH_3(CH_3)$ is stabilized only by -1 effect of flourine atoms. Here both assertion and reason are true and reason is correct explanation of assertion.
- 6. (c) lonic product of AgBr is greater than that of AgCl in comparison with there solubility product AgBr will precipitate. First rather than that of AgCl.
- (e) It is fact that ionic reactions are instantaneous due to the fact that oppositely charged ions exert strong forces and combine immediately.