CHAPTER
7

Equilibrium

7.1 Equilibrium in Physical Processes

1. In liquid-gas equilibrium, the pressure of vapours above the liquid is constant at
(a) constant temperature
(b) low temperature
(c) high temperature
(d) none of these.
(1995)

7.3 Law of Chemical Equilibrium and Equilibrium Constant

2. The equilibrium constants of the following are
$N_2 + 3H_2 \rightleftharpoons 2NH_3;\ K_1$
$N_2 + O_2 \rightleftharpoons 2NO;\ K_2$
$H_2 + \frac{1}{2}O_2 \rightleftharpoons H_2O;\ K_3$

$2NH_3 + \frac{5}{2}O_2 \xrightarrow{K} 2NO + 3H_2O$ will be
(a) $K_2K_3^3/K_1$ (b) K_2K_3/K_1
(c) $K_2^3K_3/K_1$ *(d)* $K_1K_3^3/K_2$
(NEET 2017, 2007, 2003)

3. If the equilibrium constant for
$N_{2(g)} + O_{2(g)} \rightleftharpoons 2NO_{(g)}$ is K, the equilibrium constant for $\frac{1}{2}N_{2(g)} + \frac{1}{2}O_{2(g)} \rightleftharpoons NO_{(g)}$ will be
(a) $\frac{1}{2}K$ (b) K
(c) K^2 (d) $K^{1/2}$ *(2015)*

4. Given that the equilibrium constant for the reaction,
$2SO_{2(g)} + O_{2(g)} \rightleftharpoons 2SO_{3(g)}$
has a value of 278 at a particular temperature. What is the value of the equilibrium constant for the following reaction at the same temperature?
$SO_{3(g)} \rightleftharpoons SO_{2(g)} + \frac{1}{2}O_{2(g)}$
(b) 1.8×10^{-3} (b) 3.6×10^{-3}
(c) 6.0×10^{-2} (d) 1.3×10^{-5}
(Mains 2012)

5. Given the reaction between 2 gases represented by A_2 and B_2 to give the compound $AB_{(g)}$.
$A_{2(g)} + B_{2(g)} \rightleftharpoons 2AB_{(g)}$
At equilibrium, the concentration of
$A_2 = 3.0 \times 10^{-3}$ M, of $B_2 = 4.2 \times 10^{-3}$ M, of $AB = 2.8 \times 10^{-3}$ M
If the reaction takes place in a sealed vessel at 527°C, then the value of K_c will be
(a) 2.0 (b) 1.9 (c) 0.62 (d) 4.5
(Mains 2012)

6. For the reaction, $N_{2(g)} + O_{2(g)} \rightleftharpoons 2NO_{(g)}$, the equilibrium constant is K_1. The equilibrium constant is K_2 for the reaction,
$2NO_{(g)} + O_{2(g)} \rightleftharpoons 2NO_{2(g)}$
What is K for the reaction,
$NO_{2(g)} \rightleftharpoons \frac{1}{2}N_{2(g)} + O_{2(g)}$?
(a) $\frac{1}{2K_1K_2}$ (b) $\frac{1}{4K_1K_2}$
(c) $\left[\frac{1}{K_1K_2}\right]^{1/2}$ (d) $\frac{1}{K_1K_2}$ *(2011)*

7. The dissociation constants for acetic acid and HCN at 25°C are 1.5×10^{-5} and 4.5×10^{-10} respectively. The equilibrium constant for the equilibrium,
$CN^- + CH_3COOH \rightleftharpoons HCN + CH_3COO^-$ would be
(a) 3.0×10^{-5} (b) 3.0×10^{-4}
(c) 3.0×10^{4} (d) 3.0×10^{5} *(2009)*

8. The value of equilibrium constant of the reaction,
$HI_{(g)} \rightleftharpoons \frac{1}{2}H_{2(g)} + \frac{1}{2}I_{2(g)}$
is 8.0. The equilibrium constant of the reaction
$H_{2(g)} + I_{2(g)} \rightleftharpoons 2HI_{(g)}$ will be

(a) 16 (b) 1/8

(c) 1/16 (d) 1/64 *(2008)*

9. Equilibrium constants K_1 and K_2 for the following equilibria:

$NO_{(g)} + \frac{1}{2}O_{2(g)} \overset{K_1}{\quad} NO_{2(g)}$ and

$2NO_{2(g)} \overset{K_2}{\quad} 2NO_{(g)} + O_{2(g)}$

are related as

(a) $K_2 = 1/K_1^2$ (b) $K_2 = K_1^2$

(c) $K_2 = 1/K_1$ (d) $K_2 = K_1^{1/2}$ *(2005)*

10. If K_1 and K_2 are the respective equilibrium constants for the two reactions,

$XeF_{6(g)} + H_2O_{(g)} \rightarrow XeOF_{4(g)} + 2HF_{(g)}$

$XeO_{4(g)} + XeF_{6(g)} \rightarrow XeOF_{4(g)} + XeO_3F_{2(g)}$,

the equilibrium constant of the reaction,

$XeO_{4(g)} + 2HF_{(g)} \rightarrow XeO_3F_{2(g)} + H_2O_{(g)}$,

will be

(a) K_1/K_2 (b) $K_1 \cdot K_2$

(c) $K_1/(K_2)^2$ (d) K_2/K_1 *(1998)*

11. The equilibrium constant for the reaction $N_2 + 3H_2$ □ $2NH_3$ is K, then the equilibrium constant for the equilibrium $2NH_3$ □ $N_2 + 3H_2$ is

(a) $\sqrt{K}$ (b) $\sqrt{\frac{1}{K}}$ (c) $\frac{1}{K}$ (d) $\frac{1}{K^2}$ *(1996)*

12. K_1 and K_2 are equilibrium constants for reactions (i) and (ii) respectively.

$N_{2(g)} + O_{2(g)} \quad 2NO_{(g)}$ .... (i)

$NO_{(g)} \quad \frac{1}{2}N_{2(g)} + \frac{1}{2}O_{2(g)}$ ...(ii)

(a) $K_1 = \left(\frac{1}{K_2}\right)^2$ (b) $K_1 = K_2^2$

(c) $K_1 = \frac{1}{K_2}$ (d) $K_1 = (K_2)^0$ *(1989)*

7.4 Homogeneous Equilibrium

13. The reaction, $2A_{(g)} + B_{(g)}$ □ $3C_{(g)} + D_{(g)}$ is begun with the concentrations of A and B both at an initial value of 1.00 M. When equilibrium is reached, the concentration of D is measured and found to be 0.25 M. The value for the equilibrium constant for this reaction is given by the expression

(a) $[(0.75)^3(0.25)] \div [(1.00)^2(1.00)]$

(b) $[(0.75)^3(0.25)] \div [(0.50)^2(0.75)]$

(c) $[(0.75)^3(0.25)] \div [(0.50)^2(0.25)]$

(d) $[(0.75)^3(0.25)] \div [(0.75)^2(0.25)]$ *(Mains 2010)*

14. The dissociation equilibrium of a gas AB_2 can be

$2AB_{2(g)}$ □ $2AB_{(g)} + B_{2(g)}$

The degree of dissociation is x and is small compared to 1. The expression relating the degree of dissociation (x) with equilibrium constant K_p and total pressure P is

(a) $(2K_p/P)^{1/2}$ (b) (K_p/P)

(c) $(2K_p/P)$ (d) $(2K_p/P)^{1/3}$ *(2008)*

15. The values of K_{p_1} and K_{p_2} for the reactions,

X □ $Y + Z$ ...(i)

A □ $2B$ ...(ii)

are in the ratio 9 : 1. If degree of dissociation of X and A be equal, then total pressure at equilibrium (i) and (ii) are in the ratio

(a) 36 : 1 (b) 1 : 1

(c) 3 : 1 (d) 1 : 9 *(2008)*

7.5 Heterogeneous Equilibrium

16. A 20 litre container at 400 K contains $CO_{2(g)}$ at pressure 0.4 atm and an excess of SrO (neglect the volume of solid SrO). The volume of the container is now decreased by moving the movable piston fitted in the container. The maximum volume of the container, when pressure of CO_2 attains its maximum value, will be

(Given that : $SrCO_{3(s)}$ $SrO_{(s)} + CO_{2(g)}$,

$K_p = 1.6$ atm)

(a) 10 litre (b) 4 litre

(c) 2 litre (d) 5 litre *(NEET 2017)*

17. In which of the following equilibrium K_c and K_p are not equal?

(a) $2NO_{(g)}$ □ $N_{2(g)} + O_{2(g)}$

(b) $SO_{2(g)} + NO_{2(g)}$ □ $SO_{3(g)} + NO_{(g)}$

(c) $H_{2(g)} + I_{2(g)}$ □ $2HI_{(g)}$

(d) $2C_{(s)} + O_{2(g)}$ □ $2CO_{2(g)}$ *(2010)*

18. If the concentration of OH^- ions in the reaction

$Fe(OH)_{3(s)}$ □ $Fe^{3+}_{(aq)} + 3OH^-_{(aq)}$

is decreased by 1/4 times, then equilibrium concentration of Fe^{3+} will increase by

(a) 64 times (b) 4 times

(c) 8 times (d) 16 times. *(2008)*

19. Equilibrium constant K_p for following reaction

$MgCO_{3(s)}$ $MgO_{(s)} + CO_{2(g)}$

(a) $K_p = p_{CO_2}$

(b) $K_p = p_{CO_2} \times \frac{p_{CO_2} \times p_{MgO}}{p_{MgCO_3}}$

(c) $K_p = \frac{p_{CO_2} + p_{MgO}}{p_{MgCO_3}}$ (d) $K_p = \frac{p_{MgCO_3}}{p_{CO_2} \times p_{MgO}}$

7.6 Applications of Equilibrium Constant

20. If the value of equilibrium constant for a particular reaction is 1.6×10^{12}, then at equilibrium the system will contain

(a) mostly products

(b) similar amounts of reactants and products

(c) all reactants

(d) mostly reactants. *(2015, Cancelled)*

21. In Haber process, 30 litres of dihydrogen and 30 litres of dinitrogen were taken for reaction which yielded only 50% of the expected product. What will be the composition of gaseous mixture under the aforesaid condition in the end?

(a) 20 litres ammonia, 20 litres nitrogen, 20 litres hydrogen

(b) 10 litres ammonia, 25 litres nitrogen, 15 litres hydrogen

(c) 20 litres ammonia, 10 litres nitrogen, 30 litres hydrogen

(d) 20 litres ammonia, 25 litres nitrogen, 15 litres hydrogen *(2003)*

22. The reaction quotient (Q) for the reaction $N_{2(g)} + 3H_{2(g)}$ $2NH_{3(g)}$ is given by $Q = \frac{[NH_3]^2}{[N_2][H_2]^3}$. The reaction will proceed from right to left if

(a) $Q = K_c$ (b) $Q \le K_c$

(c) $Q > K_c$ (d) $Q = 0$

where K_c is the equilibrium constant. *(2003)*

7.7 Relationship Between *K, Q,* and *G*

23. Hydrolysis of sucrose is given by the following reaction : Sucrose + $H_2O \rightleftharpoons$ Glucose + Fructose

If the equilibrium constant (K_C) is 2×10^{13} at 300K, the value of $\Delta_r G°$ at the same temperature will be

(a) $-8.314\ \text{J mol}^{-1}\text{K}^{-1} \times 300\ \text{K} \times \ln(2 \times 10^{13})$

(b) $8.314\ \text{J mol}^{-1}\text{K}^{-1} \times 300\ \text{K} \times \ln(2 \times 10^{13})$

(c) $8.314\ \text{J mol}^{-1}\text{K}^{-1} \times 300\ \text{K} \times \ln(3 \times 10^{13})$

(d) $-8.314\ \text{J mol}^{-1}\text{K}^{-1} \times 300\ \text{K} \times \ln(4 \times 10^{13})$

(NEET 2020)

24. Which of the following statements is correct for a reversible process in a state of equilibrium?

(a) $\Delta G° = -2.30\,RT \log K$ (b) $\Delta G° = 2.30\,RT \log K$

(c) $\Delta G = -2.30\,RT \log K$ (d) $\Delta G = 2.30\,RT \log K$

(2015, Cancelled)

25. Match List I (Equations) with List II (Type of processes) and select the correct option.

List I (Equations)	List II (Type of processes)
A. $K_p > Q$	(i) Non- spontaneous
B. $\Delta G° < RT \ln Q$	(ii) Equilibrium
C. $K_p = Q$	(iii) Spontaneous and endothermic
D. $T > \frac{\Delta H}{\Delta S}$	(iv) Spontaneous

(a) A - (i), B - (ii), C - (iii), D - (iv)

(b) A - (iii), B - (iv), C - (ii), D - (i)

(c) A - (iv), B - (i), C - (ii), D - (iii)

(d) A - (ii), B - (i), C - (iv), D - (iii) *(Mains 2010)*

7.8 Factors Affecting Equilibrium

26. Which one of the following conditions will favour maximum formation of the product in the reaction

$A_{2(g)} + B_{2(g)}$ □ $X_{2(g)}$, $\Delta_r H = -X$ kJ ?

(a) Low temperature and high pressure

(b) Low temperature and low pressure

(c) High temperature and high pressure

(d) High temperature and low pressure

(NEET 2018)

27. For the reversible reaction,

$N_{2(g)} + 3H_{2(g)}\ 2NH_{3(g)}$ + heat

The equilibrium shifts in forward direction

(a) by increasing the concentration of $NH_{3(g)}$

(b) by decreasing the pressure

(c) by decreasing the concentrations of $N_{2(g)}$ and $H_{2(g)}$

(d) by increasing pressure and decreasing temperature. *(2014)*

28. For a given exothermic reaction, K_p and K'_p are the equilibrium constants at temperatures T_1 and T_2, respectively. Assuming that heat of reaction is constant in temperature range between T_1 and T_2, it is readily observed that

(a) $K_p > K'_p$ (b) $K_p < K'_p$

(c) $K_p = K'_p$ *(d)* $K_p = \frac{1}{K'_p}$ *(2014)*

29. $KMnO_4$ can be prepared from K_2MnO_4 as per the reaction,

$3MnO_4^{2-} + 2H_2O\ 2MnO_4^- + MnO_2 + 4OH^-$

The reaction can go to completion by removing OH^- ions by adding

(a) CO_2 (b) SO_2

(c) HCl (d) KOH *(NEET 2013)*

30. The value of ΔH for the reaction

$X_{2(g)} + 4Y_{2(g)}$ □ $2XY_{4(g)}$ is less than zero.

Formation of $XY_{4(g)}$ will be favoured at
(a) high temperature and high pressure
(b) low pressure and low temperature
(c) high temperature and low pressure
(d) high pressure and low temperature. *(2011)*

31. For the reaction :
$CH_{4(g)} + 2O_{2(g)}\ CO_{2(g)} + 2H_2O_{(l)}$,
$\Delta H_r = -170.8$ kJ mol^{-1}. Which of the following statements is not true?
(a) The reaction is exothermic.
(b) At equilibrium, the concentrations of $CO_{2(g)}$ and $H_2O_{(l)}$ are not equal.
(c) The equilibrium constant for the reaction is given by $K_p = \frac{[CO_2]}{[CH_4][O_2]}$.
(d) Addition of $CH_{4(g)}$ or $O_{2(g)}$ at equilibrium will cause a shift to the right. *(2006)*

32. Reaction $BaO_{2(s)}$ $BaO_{(s)} + O_{2(g)}$; ΔH = +ve. In equilibrium condition, pressure of O_2 depends on
(a) increase mass of BaO_2
(b) increase mass of BaO
(c) increase temperature on equilibrium
(d) increase mass of BaO_2 and BaO both. *(2002)*

33. For any reversible reaction, if we increase concentration of the reactants, then effect on equilibrium constant
(a) depends on amount of concentration
(b) unchange
(c) decrease
(d) increase. *(2000)*

34. According to Le Chatelier's principle, adding heat to a solid and liquid in equilibrium will cause the
(a) temperature to increase
(b) temperature to decrease
(c) amount of liquid to decrease
(d) amount of solid to decrease. *(1993)*

35. Which one of the following information can be obtained on the basis of Le Chatelier principle?
(a) Dissociation constant of a weak acid
(b) Entropy change in a reaction
(c) Equilibrium constant of a chemical reaction
(d) Shift in equilibrium position on changing value of a constraint *(1992)*

7.9 Ionic Equilibrium in Solution

36. Aqueous solution of which of the following compounds is the best conductor of electric current?
(a) Hydrochloric acid, HCl
(b) Ammonia, NH_3
(c) Fructose, $C_6H_{12}O_6$
(d) Acetic acid, $C_2H_4O_2$ *(2015)*

37. Aqueous solution of acetic acid contains
(a) CH_3COO^- and H^+
(b) CH_3COO^-, H_3O^+ and CH_3COOH
(c) CH_3COO^-, H_3O^+ and H^+
(d) CH_3COOH, CH_3COO^- and H^+ *(1991)*

7.10 Acids, Bases and Salts

38. Conjugate base for Bronsted acids H_2O and HF are
(a) H_3O^+ and H_2F^+, respectively
(b) OH^- and H_2F^+, respectively
(c) H_3O^+ and F^-, respectively
(d) OH^- and F^-, respectively. *(NEET 2019)*

39. Which of the following cannot act both as Bronsted acid and as Bronsted base?
(a) HCO_3^- (b) NH_3
(c) HCl (d) HSO_4^-
(Odisha NEET 2019)

40. Which of the following fluoro-compounds is most likely to behave as a Lewis base?
(a) BF_3 (b) PF_3 (c) CF_4 (d) SiF_4
(NEET-II 2016)

41. Which of these is least likely to act as a Lewis base?
(a) BF_3 (b) PF_3 (c) CO (d) F^-
(NEET 2013)

42. Which is the strongest acid in the following?
(a) $HClO_4$ (b) H_2SO_3
(c) H_2SO_4 (d) $HClO_3$ *(NEET 2013)*

43. Which one of the following molecular hydrides acts as a Lewis acid?
(a) NH_3 (b) H_2O
(c) B_2H_6 (d) CH_4 *(2010)*

44. Which of the following molecules acts as a Lewis acid?
(a) $(CH_3)_2O$ (b) $(CH_3)_3P$
(c) $(CH_3)_3N$ (d) $(CH_3)_3B$ *(2009)*

45. Which one of the following statements is not true?
(a) Among halide ions, iodide is the most powerful reducing agent.
(b) Fluorine is the only halogen that does not show a variable oxidation state.
(c) HOCl is a stronger acid than HOBr.
(d) HF is a stronger acid than HCl. *(2003)*

46. Which one of the following compounds is not a protonic acid?

(a) $B(OH)_3$ (b) $PO(OH)_3$
(c) $SO(OH)_2$ (d) $SO_2(OH)_2$ *(2003)*

47. In HS^-, I^-, $R-NH_2$, NH_3 order of proton accepting tendency will be
(a) $I^- > NH_3 > R-NH_2 > HS^-$
(b) $NH_3 > R-NH_2 > HS^- > I^-$
(c) $R-NH_2 > NH_3 > HS^- > I^-$
(d) $HS^- > R-NH_2 > NH_3 > I^-$ *(2001)*

48. Conjugate acid of NH_2^- is
(a) NH_4OH (b) NH_4^+
(c) NH_2^- (d) NH_3 *(2000)*

49. Which compound is electron deficient?
(a) $BeCl_2$ (b) BCl_3
(c) CCl_4 (d) PCl_5 *(2000)*

50. The strongest conjugate base is
(a) SO_4^{2-} (b) Cl^-
(c) NO_3^- (d) CH_3COO^- *(1999)*

51. Which of the following is not a Lewis acid?
(a) SiF_4 (b) C_2H_4
(c) BF_3 (d) $FeCl_3$ *(1996)*

52. Repeated use of which one of the following fertilizers would increase the acidity of the soil?
(a) Ammonium sulphate
(b) Superphosphate of lime
(c) Urea
(d) Potassium nitrate *(1998)*

7.11 Ionization of Acids and Bases

53. Find out the solubility of $Ni(OH)_2$ in 0.1 M NaOH. Given that the ionic product of $Ni(OH)_2$ is 2×10^{-15}.
(a) 2×10^{-13} M (b) 2×10^{-8} M
(c) 1×10^{-13} M (d) 1×10^{8} M
(NEET 2020)

54. The pH of 0.01 M $NaOH_{(aq)}$ solution will be
(a) 7.01 (b) 2 (c) 12 (d) 9
(Odisha NEET 2019)

55. Following solutions were prepared by mixing different volumes of NaOH and HCl of different concentrations :
A. $60\text{ mL }\frac{M}{10}\text{ HCl} + 40\text{ mL }\frac{M}{10}\text{ NaOH}$
B. $55\text{ mL }\frac{M}{10}\text{ HCl} + 45\text{ mL }\frac{M}{10}\text{ NaOH}$
C. $75\text{ mL }\frac{M}{5}\text{ HCl} + 25\text{ mL }\frac{M}{5}\text{ NaOH}$
D. $100\text{ mL }\frac{M}{10}\text{ HCl} + 100\text{ mL }\frac{M}{10}\text{ NaOH}$

pH of which one of them will be equal to 1?
(a) *B* (b) *A* (c) *D* (d) *C*
(NEET 2018)

56. The percentage of pyridine (C_5H_5N) that forms pyridinium ion ($C_5H_5\overset{+}{N}H$) in a 0.10 M aqueous pyridine solution (K_b for $C_5H_5N = 1.7 \times 10^{-9}$) is
(a) 0.0060% (b) 0.013%
(c) 0.77% (d) 1.6% *(NEET-II 2016)*

57. What is the pH of the resulting solution when equal volumes of 0.1 M NaOH and 0.01 M HCl are mixed?
(a) 2.0 (b) 7.0
(c) 1.04 (d) 12.65 *(2015)*

58. Which of the following salts will give highest pH in water?
(a) KCl (b) NaCl
(c) Na_2CO_3 (d) $CuSO_4$ *(2014)*

59. Accumulation of lactic acid ($HC_3H_5O_3$), a monobasic acid in tissues leads to pain and a feeling of fatigue. In a 0.10 M aqueous solution, lactic acid is 3.7% dissociated. The value of dissociation constant, K_a, for this acid will be
(a) 1.4×10^{-5} (b) 1.4×10^{-4}
(c) 3.7×10^{-4} (d) 2.8×10^{-4}
(Karnataka NEET 2013)

60. At 100°C the K_w of water is 55 times its value at 25°C. What will be the pH of neutral solution? (log 55 = 1.74)
(a) 7.00 (b) 7.87 (c) 5.13 (d) 6.13
(Karnataka NEET 2013)

61. Equimolar solutions of the following substances were prepared separately. Which one of these will record the highest pH value?
(a) $BaCl_2$ (b) $AlCl_3$
(c) LiCl (d) $BeCl_2$ *(2012)*

62. What is $[H^+]$ in mol/L of a solution that is 0.20 M in CH_3COONa and 0.10 M in CH_3COOH? (K_a for $CH_3COOH = 1.8 \times 10^{-5}$)
(a) 3.5×10^{-4} (b) 1.1×10^{-5}
(c) 1.8×10^{-5} (d) 9.0×10^{-6} *(2010)*

63. The ionization constant of ammonium hydroxide is 1.77×10^{-5} at 298 K. Hydrolysis constant of ammonium chloride is
(a) 6.50×10^{-12} (b) 5.65×10^{-13}
(c) 5.65×10^{-12} (d) 5.65×10^{-10} *(2009)*

64. What is the $[OH^-]$ in the final solution prepared by mixing 20.0 mL of 0.050 M HCl with 30.0 mL of 0.10 M $Ba(OH)_2$?
(a) 0.40 M (b) 0.0050 M
(c) 0.12 M (d) 0.10 M *(2009)*

65. Equal volumes of three acid solutions of pH 3, 4 and 5 are mixed in a vessel. What will be the H^+ ion concentration in the mixture?
(a) 3.7×10^{-3} M (b) 1.11×10^{-3} M
(c) 1.11×10^{-4} M (d) 3.7×10^{-4} M *(2008)*

66. A weak acid, HA, has a K_a of 1.00×10^{-5}. If 0.100 mol of this acid is dissolved in one litre of water, the percentage of acid dissociated at equilibrium is closest to
(a) 1.00% (b) 99.9%
(c) 0.100% (d) 99.0% *(2007)*

67. Calculate the pOH of a solution at 25°C that contains 1×10^{-10} M of hydronium ions, *i.e.* H_3O^+.
(a) 4.000 (b) 9.000
(c) 1.000 (d) 7.000 *(2007)*

68. The hydrogen ion concentration of a 10^{-8} M HCl aqueous solution at 298 K ($K_w = 10^{-14}$) is
(a) 1.0×10^{-8} M (b) 1.0×10^{-6} M
(c) 1.0525×10^{-7} M (d) 9.525×10^{-8} M *(2006)*

69. At 25°C, the dissociation constant of a base, *B*OH, is 1.0×10^{-12}. The concentration of hydroxyl ions in 0.01 M aqueous solution of the base would be
(a) 1.0×10^{-5} mol L^{-1} (b) 1.0×10^{-6} mol L^{-1}
(c) 2.0×10^{-6} mol L^{-1} (d) 1.0×10^{-7} mol L^{-1} *(2005)*

70. Which has highest pH?
(a) CH_3COOK (b) Na_2CO_3
(c) NH_4Cl (d) $NaNO_3$ *(2002)*

71. Ionisation constant of CH_3COOH is 1.7×10^{-5} and concentration of H^+ ions is 3.4×10^{-4}. Then find out initial concentration of CH_3COOH molecules.
(a) 3.4×10^{-4} (b) 3.4×10^{-3}
(c) 6.8×10^{-4} (d) 6.8×10^{-3} *(2001)*

72. Correct relation between dissociation constants of a dibasic acid is
(a) $K_{a_1} = K_{a_2}$ (b) $K_{a_1} > K_{a_2}$
(c) $K_{a_1} < K_{a_2}$ (d) $K_{a_1} = \frac{1}{K_{a_2}}$ *(2000)*

73. Which statement is wrong about pH and H^+?
(a) pH of neutral water is not zero.
(b) Adding 1 N solution of CH_3COOH and 1 N solution of NaOH, pH will be seven.
(c) $[H^+]$ of dilute and hot H_2SO_4 is more than concentrated and cold H_2SO_4.
(d) Mixing solution of CH_3COOH and HCl, pH will be less than 7. *(2000)*

74. The concentration of $[H^+]$ and concentration of $[OH^-]$ of a 0.1 aqueous solution of 2% ionised weak acid is [ionic product of water = 1×10^{-14}]
(a) 2×10^{-3} M and 5×10^{-12} M
(b) 1×10^{-3} M and 3×10^{-11} M
(c) 0.02×10^{-3} M and 5×10^{-11} M
(d) 3×10^{-2} M and 4×10^{-13} M *(1999)*

75. The hydride ion H^- is stronger base than its hydroxide ion OH^-. Which of the following reaction will occur if sodium hydride (NaH) is dissolved in water?
(a) $H^- + H_2O \rightarrow$ no reaction
(b) $H^-_{(aq)} + H_2O \rightarrow H_2O$
(c) $H^-_{(aq)} + H_2O_{(l)} \rightarrow \overset{-}{O}H + H_2$
(d) None of these. *(1997)*

76. The ionic product of water at 25°C is 10^{-14}. Its ionic product at 90°C will be,
(a) 1×10^{-14} (b) 1×10^{-16}
(c) 1×10^{-20} (d) 1×10^{-12} *(1996)*

77. If α is dissociation constant, then the total number of moles for the reaction, $2HI \rightarrow H_2 + I_2$ will be
(a) 1 (b) $1 - \alpha$
(c) 2 (d) $2 - \alpha$ *(1996)*

78. The pH value of N/10 NaOH solution is
(a) 12 (b) 13 (c) 10 (d) 11 *(1996)*

79. The pH value of a 10 M solution of HCl is
(a) equal to 1 (b) equal to 2
(c) less than 0 (d) equal to 0 *(1995)*

80. At 80°C, distilled water has $[H_3O^+]$ concentration equal to 1×10^{-6} mole/litre. The value of K_w at this temperature will be
(a) 1×10^{-12} (b) 1×10^{-15}
(c) 1×10^{-6} (d) 1×10^{-9} *(1994)*

81. 0.1 M solution of which one of these substances will act basic?
(a) Sodium borate (b) Ammonium chloride
(c) Calcium nitrate (d) Sodium sulphate
(1992)

82. The compound whose water solution has the highest pH is
(a) NaCl (b) $NaHCO_3$
(c) Na_2CO_3 (d) NH_4Cl *(1988)*

7.12 Buffer Solutions

83. Which will make basic buffer?
(a) 100 mL of 0.1 M HCl + 100 mL of 0.1 M NaOH
(b) 50 mL of 0.1 M NaOH + 25 mL of 0.1 M CH_3COOH
(c) 100 mL of 0.1 M CH_3COOH + 100 mL of 0.1 M NaOH
(d) 100 mL of 0.1 M HCl + 200 mL of 0.1 M NH_4OH
(NEET 2019)

84. Which one of the following pairs of solutions is not an acidic buffer?

(a) CH_3COOH and CH_3COONa
(b) H_2CO_3 and Na_2CO_3
(c) H_3PO_4 and Na_3PO_4
(d) $HClO_4$ and $NaClO_4$ *(2015)*

85. The dissociation constant of a weak acid is 1×10^{-4}. In order to prepare a buffer solution with a pH = 5, the [Salt]/[Acid] ratio should be

(a) 4 : 5 (b) 10 : 1 (c) 5 : 4 (d) 1 : 10

(Karnataka NEET 2013)

86. Buffer solutions have constant acidity and alkalinity because

(a) these give unionised acid or base on reaction with added acid or alkali
(b) acids and alkalies in these solutions are shielded from attack by other ions
(c) they have large excess of H^+ or OH^- ions
(d) they have fixed value of pH. *(2012)*

87. A buffer solution is prepared in which the concentration of NH_3 is 0.30 M and the concentration of NH_4^+ is 0.20 M. If the equilibrium constant, K_b for NH_3 equals 1.8×10^{-5}, what is the pH of this solution? (log 2.7 = 0.43)

(a) 9.08 (b) 9.43
(c) 11.72 (d) 8.73 *(2011)*

88. In a buffer solution containing equal concentration of B^- and HB, the K_b for B^- is 10^{-10}. The pH of buffer solution is

(a) 10 (b) 7 (c) 6 (d) 4 *(2010)*

89. Which of the following pairs constitutes a buffer?

(a) HCl and KCl (b) HNO_2 and $NaNO_2$
(c) NaOH and NaCl (d) HNO_3 and NH_4NO_3

(2006)

90. The rapid change of pH near the stoichiometric point of an acid-base titration is the basis of indicator detection. pH of the solution is related to ratio of the concentrations of the conjugate acid (HIn) and base (In^-) forms of the indicator by the expression

(a) $\log\frac{[In^-]}{[HIn]} = pK_{In} - pH$
(b) $\log\frac{[HIn]}{[In^-]} = pK_{In} - pH$
(c) $\log\frac{[HIn]}{[In^-]} = pH - pK_{In}$
(d) $\log\frac{[In^-]}{[HIn]} = pH - pK_{In}$ *(2004)*

91. Solution of 0.1 N NH_4OH and 0.1 N NH_4Cl has pH 9.25. Then find out pK_b of NH_4OH.

(a) 9.25 (b) 4.75
(c) 3.75 (d) 8.25 *(2002)*

92. A physician wishes to prepare a buffer solution at pH = 3.85 that efficiently resists changes in pH yet contains only small concentration of the buffering agents. Which of the following weak acids together with its sodium salt would be best to use?

(a) 2, 5-Dihydroxybenzoic acid ($pK_a = 2.97$)
(b) Acetoacetic acid ($pK_a = 3.58$)
(c) *m*-Chlorobenzoic acid ($pK_a = 3.98$)
(d) *p*-Chlorocinnamic acid ($pK_a = 4.41$) *(1997)*

93. The pH value of blood does not appreciably change by a small addition of an acid or a base, because the blood

(a) can be easily coagulated
(b) contains iron as a part of the molecule
(c) is a body fluid
(d) contains serum protein which acts as buffer.

(1995)

7.13 Solubility Equilibrium of Sparingly Soluble Salts

94. pH of a saturated solution of $Ca(OH)_2$ is 9. The solubility product (K_{sp}) of $Ca(OH)_2$ is

(a) 0.5×10^{-10} (b) 0.5×10^{-15}
(c) 0.25×10^{-10} (d) 0.125×10^{-15}

(NEET 2019)

95. The molar solubility of CaF_2 ($K_{sp} = 5.3 \times 10^{-11}$) in 0.1 M solution of NaF will be

(a) 5.3×10^{-11} mol L^{-1} (b) 5.3×10^{-8} mol L^{-1}
(c) 5.3×10^{-9} mol L^{-1} (d) 5.3×10^{-10} mol L^{-1}

(Odisha NEET 2019)

96. The solubility of $BaSO_4$ in water is 2.42×10^{-3} g L^{-1} at 298 K. The value of its solubility product (K_{sp}) will be

(Given molar mass of $BaSO_4$ = 233 g mol^{-1})

(a) 1.08×10^{-10} mol^2 L^{-2}
(b) 1.08×10^{-12} mol^2 L^{-2}
(c) 1.08×10^{-14} mol^2 L^{-2}
(d) 1.08×10^{-8} mol^2 L^{-2} *(NEET 2018)*

97. Concentration of the Ag^+ ions in a saturated solution of $Ag_2C_2O_4$ is 2.2×10^{-4} mol L^{-1}. Solubility product of $Ag_2C_2O_4$ is

(a) 2.66×10^{-12} (b) 4.5×10^{-11}
(c) 5.3×10^{-12} (d) 2.42×10^{-8}

(NEET 2017)

98. The solubility of $AgCl_{(s)}$ with solubility product 1.6×10^{-10} in 0.1 M NaCl solution would be

(a) 1.26×10^{-5} M (b) 1.6×10^{-9} M

(c) 1.6×10^{-11} M (d) zero. *(NEET-II 2016)*

99. MY and NY_3, two nearly insoluble salts, have the same K_{sp} values of 6.2×10^{-13} at room temperature. Which statement would be true in regard to MY and NY_3?

(a) The salts MY and NY_3 are more soluble in 0.5 M KY than in pure water.

(b) The addition of the salt of KY to solution of MY and NY_3 will have no effect on their solubilities.

(c) The molar solubilities of MY and NY_3 in water are identical.

(d) The molar solubility of MY in water is less than that of NY_3. *(NEET-I 2016)*

100. The K_{sp} of Ag_2CrO_4, AgCl, AgBr and AgI are respectively, 1.1×10^{-12}, 1.8×10^{-10}, 5.0×10^{-13}, 8.3×10^{-17}. Which one of the following salts will precipitate last if $AgNO_3$ solution is added to the solution containing equal moles of NaCl, NaBr, NaI and Na_2CrO_4?

(a) AgBr (b) Ag_2CrO_4

(c) AgI (d) AgCl

(2015, Cancelled)

101. Using the Gibbs' energy change, $\Delta G° = +63.3$ kJ, for the following reaction,

$$Ag_2CO_{3(s)} \qquad 2Ag^+_{(aq)} + CO^{2-}_{3(aq)}$$

the K_{sp} of $Ag_2CO_{3(s)}$ in water at 25 °C is

($R = 8.314$ J K^{-1} mol^{-1})

(a) 3.2×10^{-26} (b) 8.0×10^{-12}

(c) 2.9×10^{-3} (d) 7.9×10^{-2} *(2014)*

102. The values of K_{sp} of $CaCO_3$ and CaC_2O_4 are 4.7×10^{-9} and 1.3×10^{-9} respectively at 25°C. If the mixture of these two is washed with water, what is the concentration of Ca^{2+} ions in water?

(a) 5.831×10^{-5} M (b) 6.856×10^{-5} M

(c) 3.606×10^{-5} M (d) 7.746×10^{-5} M

(Karnataka NEET 2013)

103. Identify the correct order of solubility in aqueous medium.

(a) $Na_2S > CuS > ZnS$ (b) $Na_2S > ZnS > CuS$

(c) $CuS > ZnS > Na_2S$ (d) $ZnS > Na_2S > CuS$

(NEET 2013)

104. pH of a saturated solution of $Ba(OH)_2$ is 12. The value of solubility product (K_{sp}) of $Ba(OH)_2$ is

(a) 3.3×10^{-7} (b) 5.0×10^{-7}

(c) 4.0×10^{-6} (d) 5.0×10^{-6} *(2012, 2010)*

105. In qualitative analysis, the metals of group I can be separated from other ions by precipitating them as chloride salts. A solution initially contains Ag^+ and Pb^{2+} at a concentration of 0.10 M. Aqueous HCl is added to this solution until the Cl^- concentration is 0.10 M. What will the concentrations of Ag^+ and Pb^{2+} be at equilibrium?

(K_{sp} for AgCl = 1.8×10^{-10}, K_{sp} for $PbCl_2$ = 1.7×10^{-5})

(a) $[Ag^+] = 1.8 \times 10^{-7}$ M, $[Pb^{2+}] = 1.7 \times 10^{-6}$ M

(b) $[Ag^+] = 1.8 \times 10^{-11}$ M, $[Pb^{2+}] = 8.5 \times 10^{-5}$ M

(c) $[Ag^+] = 1.8 \times 10^{-9}$ M, $[Pb^{2+}] = 1.7 \times 10^{-3}$ M

(d) $[Ag^+] = 1.8 \times 10^{-11}$ M, $[Pb^{2+}] = 1.7 \times 10^{-4}$ M

(Mains 2011)

106. H_2S gas when passed through a solution of cations containing HCl precipitates the cations of second group of qualitative analysis but not those belonging to the fourth group. It is because

(a) presence of HCl decreases the sulphide ion concentration

(b) solubility product of group II sulphides is more than that of group IV sulphides

(c) presence of HCl increases the sulphide ion concentration

(d) sulphides of group IV cations are unstable in HCl. *(2005)*

107. The solubility product of a sparingly soluble salt AX_2 is 3.2×10^{-11}. Its solubility (in moles/L) is

(a) 5.6×10^{-6} (b) 3.1×10^{-4}

(c) 2×10^{-4} (d) 4×10^{-4} *(2004)*

108. The solubility product of AgI at 25°C is 1.0×10^{-16} mol^2 L^{-2}. The solubility of AgI in 10^{-4} N solution of KI at 25°C is approximately (in mol L^{-1})

(a) 1.0×10^{-16} (b) 1.0×10^{-12}

(c) 1.0×10^{-10} (d) 1.0×10^{-8} *(2003)*

109. Solubility of MX_2 type electrolytes is 0.5×10^{-4} mol/lit., then find out K_{sp} of electrolytes.

(a) 5×10^{-12} (b) 25×10^{-10}

(c) 1×10^{-13} (d) 5×10^{-13} *(2002)*

110. Solubility of M_2S salt is 3.5×10^{-6} then find out solubility product.

(a) 1.7×10^{-6} (b) 1.7×10^{-16}

(c) 1.7×10^{-18} (d) 1.7×10^{-12} *(2001)*

111. The solubility of a saturated solution of calcium fluoride is 2×10^{-4} moles per litre. Its solubility product is

(a) 22×10^{-11} (b) 14×10^{-4}

(c) 2×10^{-2} (d) 32×10^{-12} *(1999)*

112. The solubility product of CuS, Ag_2S and HgS are 10^{-31}, 10^{-44} and 10^{-54} respectively. The solubilities of these sulphides are in the order

(a) HgS > Ag_2S > CuS (b) CuS > Ag_2S > HgS
(c) Ag_2S > CuS > HgS (d) Ag_2S > HgS > CuS *(1997)*

113. The solubility of AgCl will be minimum in

(a) 0.01 M $CaCl_2$ (b) pure water
(c) 0.001 M $AgNO_3$ (d) 0.01 M NaCl *(1995)*

114. Which one of the following is most soluble?

(a) Bi_2S_3 ($K_{sp} = 1 \times 10^{-70}$)
(b) Ag_2S ($K_{sp} = 6 \times 10^{-51}$)
(c) CuS ($K_{sp} = 8 \times 10^{-37}$)
(d) MnS ($K_{sp} = 7 \times 10^{-16}$) *(1994)*

ANSWER KEY

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

Hints & Explanations

1. (a) : Vapour pressure is directly related to temperature. Greater is the temperature, greater will be the vapour pressure. So to keep it constant, temperature should be constant.

2. (a) : From the given equations,

$2NH_3 \square\ N_2 + 3H_2; \frac{1}{K_1}$ ...(i)

$N_2 + O_2 \square\ 2NO; K_2$ ...(ii)

$3H_2 + \frac{3}{2}O_2 \square\ 3H_2O; K_3^3$ ...(iii)

By adding equations (i), (ii) and (iii), we get

$2NH_3 + \frac{5}{2}O_2 \overset{K}{\ } 2NO + 3H_2O, K = \frac{K_2K_3^3}{K_1}$

3. (d) : If the reaction is multiplied by $\frac{1}{2}$, then new equilibrium constant, $K' = K^{1/2}$.

4. (c) : $2SO_{2(g)} + O_{2(g)} \square\ 2SO_{3(g)}, K = 278$ .... (i)

By reversing the equation (i), we get

$2SO_{3(g)} \square\ 2SO_{2(g)} + O_{2(g)}$ ...(ii)

Equilibrium constant for this reaction is,

$$K' = \frac{1}{K} = \frac{1}{278}$$

By dividing the equation (ii) by 2, we get desired equation,

$SO_{3(g)}\ SO_{2(g)} + \frac{1}{2}O_{2(g)}$ ...(iii)

Equilibrium constant for this reaction,

$$K'' = \sqrt{K'} = \sqrt{\frac{1}{K}} = \sqrt{\frac{1}{278}} = 0.0599 \approx 0.06 \text{ or } 6 \times 10^{-2}$$

5. (c) : $A_{2(g)} + B_{2(g)} \square\ 2AB_{(g)}$

$$K_c = \frac{[AB]^2}{[A_2][B_2]}$$

$$= \frac{\square(2.8 \times 10^{-3})^2}{(3.0 \times 10^{-3})(4.2 \times 10^{-3})} = \frac{2.8 \times 2.8}{3.0 \times 4.2} = 0.62$$

6. (c) : $N_2 + O_2 \square\ 2NO; K_1$

$2NO + O_2 \square\ 2NO_2; K_2$

$NO_2 \square\ \frac{1}{2}N_2 + O_2; K$

$$K_1 = \frac{[NO]^2}{[N_2][O_2]}; K_2 = \frac{[NO_2]^2}{[NO]^2[O_2]}$$

$$K = \frac{[N_2]^{1/2}[O_2]}{[NO_2]} = \sqrt{\frac{[N_2][O_2]\times[NO]^2[O_2]}{[NO]^2\times[NO_2]^2}}$$

$$K = \sqrt{\frac{1}{K_1K_2}}$$

7. (c) : Given, $CH_3COOH \rightleftharpoons CH_3COO^- + H^+$

$$K_1 = \frac{[CH_3COO^-][H^+]}{[CH_3COOH]} = 1.5\times10^{-5}$$

$HCN \rightleftharpoons H^+ + CN^-$

$$K_2 = \frac{[CN^-][H^+]}{[HCN]} = 4.5\times10^{-10}$$

$CN^- + CH_3COOH \rightleftharpoons HCN + CH_3COO^-$

$$K = \frac{[HCN][CH_3COO^-]}{[CN^-][CH_3COOH]}$$

$$K = \frac{K_1}{K_2} = \frac{1.5\times10^{-5}}{4.5\times10^{-10}} \approx 0.3\times10^5 \quad \text{or } K = 3\times10^4$$

8. (d) : $HI_{(g)} \rightleftharpoons 1/2H_{2(g)} + 1/2I_{2(g)}$

i.e. $K = \frac{[H_2]^{1/2}[I_2]^{1/2}}{[HI]} = 8$

$H_{2(g)} + I_{2(g)} \rightleftharpoons 2HI_{(g)}$

$$K' = \frac{[HI]^2}{[H_2][I_2]} = \left(\frac{1}{8}\right)^2 \Rightarrow K' = \frac{1}{64}$$

9. (a) : $K_1 = \frac{p_{NO_2}}{p_{NO}\cdot(p_{O_2})^{1/2}}$ ...(1)

$K_2 = \frac{(p_{NO})^2\cdot p_{O_2}}{(p_{NO_2})^2}$ ...(2)

Taking square root on both sides in eq. 2,

$$\Rightarrow \sqrt{K_2} = \frac{p_{NO}\cdot(p_{O_2})^{1/2}}{p_{NO_2}} \Rightarrow \sqrt{K_2} = \frac{1}{K_1} \Rightarrow K_2 = \frac{1}{K_1^2}$$

10. (d) : Given,

$XeF_6 + H_2O \rightleftharpoons XeOF_4 + 2HF$, $K_{eq} = K_1$

$XeOF_4 + 2HF \rightleftharpoons XeF_6 + H_2O$, $K_{eq} = 1/K_1$ ...(1)

and $XeO_4 + XeF_6 \rightleftharpoons XeOF_4 + XeO_3F_2$, $K_{eq} = K_2$ ...(2)

The reaction, $XeO_4 + 2HF \rightleftharpoons XeO_3F_2 + H_2O$, can be obtained by adding eq. (1) and eq.(2).

So, the equilibrium constant for the above reaction can be obtained by multiplying the equilibrium constants of eq. (1) and eq. (2).

Hence, the value is $= \frac{K_2}{K_1}$

11. (c) : The equilibrium constant for the reverse reaction will be $1/K$.

12. (a) : Reaction (ii) is the reversible reaction of (i) and is half of the reaction (i). Thus, rate constant can be given as :

$$K_2 = \sqrt{\frac{1}{K_1}} \quad \text{or } K_1 = \left[\frac{1}{K_2}\right]^2$$

13. (b) :

	$2A_{(g)}$ +	$B_{(g)}$ ⇌	$3C_{(g)}$ +	$D_{(g)}$
Initial moles :	1	1	0	0
Moles at eq. :	1 – (2×0.25) = 0.5	1 – 0.25 = 0.75	3×0.25 = 0.75	0.25 = 0.25

Equilibrium constant, $K = \frac{[C]^3[D]}{[A]^2[B]}$

$$\therefore K = \frac{(0.75)^3(0.25)}{(0.5)^2(0.75)}$$

14. (d) : $2AB_{2(g)} \rightleftharpoons 2AB_{(g)} + B_{2(g)}$

$2AB_{2(g)}$	$2AB_{(g)}$	$B_{2(g)}$	
2	0	0	(initially)
2(1 – x)	2x	x	(at equilibrium)

Amount of moles at equilibrium = 2(1 – x) + 2x + x = 2 + x

$$K_p = \frac{[p_{AB}]^2[p_{B_2}]}{[p_{AB_2}]^2}$$

$$K_p = \frac{\left(\frac{2x}{2+x}\times P\right)^2\times\left(\frac{x}{2+x}\times P\right)}{\left(\frac{2(1-x)}{2+x}\times P\right)^2} = \frac{\frac{4x^3}{2+x}\times P}{4(1-x)^2}$$

$$K_p = \frac{4x^3\times P}{2}\times\frac{1}{4} \quad (\because 1-x\approx1 \text{ and } 2+x\approx2)$$

$$x = \left(\frac{8K_p}{4P}\right)^{1/3} \Rightarrow x = \left(\frac{2K_p}{P}\right)^{1/3}$$

15. (a) :

$X \rightleftharpoons Y + Z$ ...(i)

$A \rightleftharpoons 2B$ ...(ii)

	X ⇌	Y +	Z
Initially	1	0	0
At equilibrium	1 – α	α	α

Total no. of moles at equilibrium = 1 – α + 2α = 1+ α

Similarly,

	A ⇌	$2B$
Initially	1	0
At equilibrium	1 – α	2α

Total no. of moles at equilibrium = 1 – α + 2α = 1 + α

$$\therefore K_{p_1} = \frac{p_Y\times p_Z}{p_X} = \frac{\frac{\alpha}{1+\alpha}\times P_1\times\frac{\alpha}{1+\alpha}\times P_1}{\frac{1-\alpha}{1+\alpha}\times P_1} = \frac{\alpha^2P_1}{(1+\alpha)(1-\alpha)}$$

$$K_{p_2} = \frac{(p_B)^2}{p_A} = \frac{\left(\frac{2\alpha}{1+\alpha}\times P_2\right)^2}{\frac{1-\alpha}{1+\alpha}\times P_2} = \frac{(2\alpha)^2P_2}{(1+\alpha)(1-\alpha)}$$

Now $\frac{K_{p_1}}{K_{p_2}} = \frac{P_1}{4P_2} \Rightarrow \frac{K_{P_1}}{K_{P_2}} = \frac{9}{1} = \frac{P_1}{4P_2} \Rightarrow \frac{P_1}{P_2} = \frac{36}{1} = 36:1$

16. (d) : $SrCO_{3(s)}$ □ $SrO_{(s)} + CO_{2(g)}; K_p = 1.6$ atm

$$K_p = \frac{p_{CO_2} \times p_{SrO}}{p_{SrCO_3}} \Rightarrow 1.6 = p_{CO_2} \quad (\because p_{SrO} = p_{SrCO_3} = 1)$$

∴ Maximum pressure of CO_2 = 1.6 atm

Let the maximum volume of the container when pressure of CO_2 is 1.6 atm be V L

During the process, PV = constant

∴ $0.4 \times 20 = 1.6 \times V$

$$\Rightarrow V = \frac{0.4 \times 20}{1.6} = 5\text{ L}$$

17. (d) : K_p and K_c are related by the equation,

$K_p = K_c(RT)^{\Delta n_g}$

where Δn_g = difference in the no. of moles of products and reactants in the gaseous state.

for $2C_{(s)} + O_{2(g)}$ □ $2CO_{2(g)}$

$\Delta n_g = 2 - (1) = 1 \neq 0$

18. (a) : $Fe(OH)_{3(s)}$ □ $Fe^{3+}_{(aq)} + 3OH^-_{(aq)}$

$$K = \frac{[Fe^{3+}][OH^-]^3}{[Fe(OH)_3]}$$

$K = [Fe^{3+}][OH^-]^3$ ($\because$ activity of solid is taken unity)

Concentration of OH^- ion in the reaction is decreased by 1/4 times then equilibrium concentration of Fe^{3+} will be increased by 64 times in order to keep the value of K constant.

19. (a) : $K_p = p_{CO_2}$

Solids do not exert pressure, so their partial pressure is taken as unity.

20. (a) : The value of K is high which means reaction proceeds almost to completion *i.e.*, the system will contain mostly products.

21. (b) : $3H_2 + N_2 \rightarrow 2NH_3$

$$\begin{array}{ccc} 3 & 1 & 2 \\ 3/2 & 1/2 & 1 \\ 10 \times \frac{3}{2} & 10 \times \frac{1}{2} & 10 \times 1 \\ 15 & 5 & 10 \end{array}$$

Composition of gaseous mixture under the aforesaid condition in the end will be

H_2 = 30 – 15 = 15 litres

N_2 = 30 – 5 = 25 litres ; NH_3 = 10 litres

22. (c) : $N_{2(g)} + 3H_{2(g)}$ □ $2NH_{3(g)}$

$$K_c = \frac{[NH_3]^2}{[N_2][H_2]^3}; \Delta n_{(g)} = 2 - 4 = -2$$

Thus, the reaction will go from right to left when $Q > K_c$.

23. (a) : $\Delta G = \Delta G° + RT \ln Q$

At equilibrium, $\Delta G = 0$ and $Q = K_C$

∴ $0 = \Delta G° + RT \ln K_C$

$\Rightarrow \Delta G° = -RT \ln K_C$

$= -8.314\text{ J K}^{-1}\text{ mol}^{-1} \times 300\text{ K} \times \ln(2 \times 10^{13})$

24. (a)

25. (c) : When $K_p > Q$, rate of forward reaction > rate of backward reaction.

∴ Reaction is spontaneous.

When $\Delta G° < RT \ln Q$, $\Delta G°$ is positive, reverse reaction is feasible, thus reaction is non spontaneous.

When $K_p = Q$, rate of forward reaction = rate of backward reaction.

∴ Reaction is in equilibrium.

When $T\Delta S > \Delta H$, ΔG will be negative only when ΔH = +ve.

∴ Reaction is spontaneous and endothermic.

26. (a) : On increasing the pressure and decreasing the temperature, equilibrium will shift in forward direction.

27. (d) : As the forward reaction is exothermic and leads to lowering of pressure (produces lesser number of gaseous moles) hence, according to Le Chatelier's principle, at high pressure and low temperature, the given reversible reaction will shift in forward direction to form more product.

28. (a) : $\log \frac{K'_p}{K_p} = -\frac{\Delta H}{2.303R}\left[\frac{1}{T_2} - \frac{1}{T_1}\right]$

For exothermic reaction, ΔH = –ve *i.e.*, heat is evolved.

The temperature T_2 is higher than T_1.

Thus, $\left(\frac{1}{T_2} - \frac{1}{T_1}\right)$ is negative.

so, $\log K'_p - \log K_p$ = –ve or $\log K_p > \log K'_p$

or $K_p > K'_p$

29. (a) : HCl and SO_2 are reducing agents. So, they can reduce MnO_4^-.

CO_2 is neither oxidising nor reducing agent, it will provide only acidic medium. It can shift the reaction in forward direction and the reaction can go to completion.

30. (d) : $X_{2(g)} + 4Y_{2(g)}\ 2XY_{4(g)}$

Δn_g = –ve and ΔH = –ve

The reaction is favoured in forward direction at low temperature and high pressure.

31. (c) : $CH_{4(g)} + 2O_{2(g)}$ □ $CO_{2(g)} + 2H_2O_{(l)}$

$$K_p = \frac{p_{CO_2}}{p_{CH_4} \cdot p^2_{O_2}}$$

32. (c) : Pressure of O_2 does not depend on concentration terms of other reactants (because both are in solid state). Since this is an endothermic reaction, if the temperature is raised, dissociation of BaO_2 would occur, more O_2 is produced at equilibrium, pressure of O_2 increases.

33. (b) : For a reaction, $A + B$ □ $C + D$,

$$K_{eq} = \frac{[C][D]}{[A][B]}$$

Increase in conc. of reactants will proceed the equilibrium in the forward direction giving more products so that the equilibrium constant value remains constant and independent of concentration.

34. (d) : When solid and liquid are in equilibrium, the increase in temperature results in increase in volume of liquid or decrease in the amount of solid.

Solid □ Liquid

With increase in temperature equilibrium shifts in forward direction.

35. (d) : According to Le Chatelier's principle, if an equilibrium is subjected to a change in concentration, pressure or temperature, etc. equilibrium shift in such a way so as to undo the effect of a change imposed.

36. (a) : HCl is a strong acid and dissociates completely into ions in aqueous solution.

37. (b) : $CH_3COOH + H_2O$ □ $CH_3COO^- + H_3O^+$
As acetic acid is a weak acid so, it also contains some undissociated CH_3COOH along with CH_3COO^- and H_3O^+ ions.

38. (d) :

Bronsted acid	Conjugate base
H_2O	OH^-
HF	F^-

39. (c) : HCl cannot accept H^+ ion, therefore cannot act as Bronsted Base.

40. (b) : $BF_3 \longrightarrow$ Lewis acid (incomplete octet)
$PF_3 \longrightarrow$ Lewis base (presence of lone pair on P atom)
$CF_4 \longrightarrow$ Complete octet
$SiF_4 \longrightarrow$ Lewis acid (empty *d*-orbital in Si-atom)

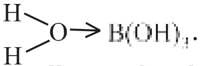
41. (a) : BF_3 is Lewis acid (e^- pair acceptor).

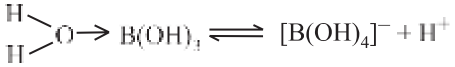
42. (a) : $H\overset{+7}{\underline{Cl}}O_4$ with highest oxidation number and its conjugate base is resonance stabilised, hence it is most acidic. Cl is more electronegative than S.

43. (c) : Compounds that are electron deficient act as Lewis acids. Out of the given hydrides B_2H_6 satisfies this condition and is therefore a Lewis acid.

44. (d) : Lewis acids are electron deficient compounds, since $(CH_3)_3B$ is electron deficient (due to incomplete octet of B), it acts as a Lewis acid.

45. (d) : Due to strong hydrogen-fluorine bond, proton is not given off easily and hence, HF is weakest acid.

46. (a) : $B(OH)_3$ in aqueous medium coordinates a molecule of water to form the hydrated species  $H_2O \rightarrow B(OH)_3$. In this species, B^{3+} ion, because of its small size, has high polarizing power thereby pulling the sigma electron charge of the coordinated O atom towards itself. The coordinated oxygen, in turn, pulls the sigma electron charge of the OH bond of the attached water molecule towards itself. This facilitates the removal of H^+ ion from the O – H bond.



$$H_2O \rightarrow B(OH)_3 \rightleftharpoons [B(OH)_4]^- + H^+$$

Thus, the solution of $B(OH)_3$ in water acts as a weak acid, and it is not a protonic acid.

47. (c) : Proton accepting tendency is known as the strength of basicity.

In $R—\ddot{N}H_2$, N has lone pair of electrons which intensify due to electron releasing *R*-group and increase the tendency to donate lone pair of electrons to H^+.

Secondly as the size of the ion increases there is less attraction for H^+ to form bond with H–atom and are less basic. Thus the order of proton accepting tendency :
$RNH_2 > NH_3 > HS^- > I^-$.

48. (d) : $NH_2^- + H^+ \rightarrow NH_3$ (conjugate acid)

Substance + $H^+ \rightarrow$ conjugate acid
Substance – $H^+ \rightarrow$ conjugate base

49. (b) : In BCl_3 the central atom 'B' is sp^2 hybridised and contains only 'six' electrons in its valence shell. Therefore, it is electron deficient.

50. (d) : CH_3COOH $\quad CH_3COO^- + H^+$
Weak acid $\quad$ Strong conjugate base

As CH_3COOH is the weakest acid, so its conjugate base (CH_3COO^-) is the strongest base. H_2SO_4, HCl, HNO_3 are strong acids, so their conjugate bases are weak.

51. (b) : In BF_3 and $FeCl_3$ molecules, the central atoms have incomplete octet and in SiF_4, the central atom has empty *d*-orbitals. Hence, according to Lewis concept, these are Lewis acids.

52. (a) : Ammonium sulphate is a salt of strong acid (H_2SO_4) and weak base (NH_4OH). Therefore, repeated use of ammonium sulphate would increase the concentration of sulphuric acid, while ammonia from NH_4OH is used up by the plant. Hence, the acidity of soil will increase.

53. (a) : $\underset{s}{Ni(OH)_2}\ \underset{s}{Ni^{2+}} + \underset{2s}{2OH^-}$

where s is the solubility of $Ni(OH)_2$.

$\underset{0.1\text{ M}}{NaOH}\ \underset{0.1\text{ M}}{Na^+} + \underset{0.1\text{ M}}{OH^-}$

$[OH^-] = 2s + 0.1 \approx 0.1$ ($\because 2s <<< 0.1$)

Ionic product of $Ni(OH)_2 = [Ni^{2+}][OH^-]^2$

$2 \times 10^{-15} = s(0.1)^2$

$$s = \frac{2\times10^{-15}}{0.1\times0.1} = 2\times10^{-13}\text{ M}$$

54. (c) : $\underset{0.01\text{ M}}{NaOH} \longrightarrow Na^+ + \underset{0.01\text{ M}}{OH^-}$

$\therefore\ [OH^-] = 0.01$ M

$\therefore\ pOH = -\log[OH^-] = -\log(0.01) = 2$

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

55. (d) : pH = 1, so $[H^+] = 10^{-1}$

For acid base mixture: $N_1V_1 - N_2V_2 = N_3V_3$

(For NaOH and HCl, Normality = Molarity)

$$A.\ M_1(H^+) = \frac{60\times\frac{1}{10} - 40\times\frac{1}{10}}{100} = 2\times10^{-2}\text{ M}$$

i.e. pH = 1.698 ≈ 1.7

$$B.\ M_2(H^+) = \frac{55\times\frac{1}{10} - 45\times\frac{1}{10}}{100} = \frac{1}{100}$$

$= 10^{-2}$ M *i.e.* pH = 2

$$C.\ M_3(H^+) = \frac{75\times\frac{1}{5} - 25\times\frac{1}{5}}{100} = 10^{-1}\text{M } i.e.\ pH = 1$$

$$D.\ M_4(H^+) = \frac{100\times\frac{1}{10} - 100\times\frac{1}{10}}{200} = 0\ i.e.\ pH = 7$$

56. (b): $\underset{0.10\text{ M}}{C_5H_5N} + H_2O\ C_5H_5NH^+ + OH^-$

$$\alpha = \sqrt{\frac{K_b}{C}} = \sqrt{\frac{1.7\times10^{-9}}{0.10}} = 1.30\times10^{-4}$$

$\therefore$ Percentage of pyridine that forms pyridinium ion $= 1.30 \times 10^{-4} \times 100 = 0.013\%$

57. (d) : One mole of NaOH is completely neutralised by one mole of HCl.

Hence, 0.01 mole of NaOH will be completely neutralised by 0.01 mole of HCl.

⇒ NaOH left unneutralised = 0.1 – 0.01 = 0.09 mol

As equal volumes of two solutions are mixed,

$$[OH^-] = \frac{0.09}{2} = 0.045\text{ M}$$

⇒ $pOH = -\log(0.045) = 1.35$

$\therefore\ pH = 14 - 1.35 = 12.65$

58. (c) : Na_2CO_3 which is a salt of NaOH (strong base) and H_2CO_3 (weak acid) will produce a basic solution with pH greater than 7.

59. (b) : Degree of dissociation, $\alpha = \frac{3.7}{100} = 0.037$

According to Ostwald's formula,

$K_a = \alpha^2C = (0.037)^2 \times 0.10 = 1.369 \times 10^{-4} \approx 1.4 \times 10^{-4}$

60. (d) : We know that, at 25°C, $K_w = 1 \times 10^{-14}$

At 100°C, $K_w = 55 \times 10^{-14}$ $[\because K_w = [H^+][OH^-]]$

$\Rightarrow K_w = [H^+]^2$

$H^+ = \sqrt{K_w}$

$H^+ = \sqrt{55\times10^{-14}}$

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

$$pH = -\log\left[\sqrt{55\times10^{-14}}\right]$$

$$= \frac{1}{2}[-\log(55\times10^{-14})] = \frac{1}{2}[-\log 55 + 14\log 10]$$

$$= \frac{1}{2}[-1.74 + 14] = \frac{1}{2}[12.26] = 6.13$$

61. (a) : $BaCl_2$ is made up of $Ba(OH)_2$ and HCl.

$AlCl_3$ is made up of $Al(OH)_3$ and HCl.

LiCl is made up of LiOH and HCl.

$BeCl_2$ is made up of $Be(OH)_2$ and HCl.

$Ba(OH)_2$ is strongest base among the given options thus have maximum pH.

62. (d) : $\underset{C-x}{CH_3COOH} \rightleftharpoons \underset{x}{CH_3COO^-} + \underset{x}{H^+}$

$\underset{0.2\text{ M}}{CH_3COONa} \rightarrow \underset{0.2\text{ M}}{CH_3COO^-} + \underset{0.2\text{ M}}{Na^+}$

$[CH_3COOH] = C - x \approx 0.1$ M

$[CH_3COO^-] = 0.2 + x \approx 0.2$ M

{acetic acid is a weak acid so, dissociation is minimum *i.e.*, x can be neglected.

$$\therefore [H^+] = \frac{K_a[CH_3COOH]}{[CH_3COO^-]}$$

$$= \frac{1.8\times10^{-5}\times10^{-1}}{2\times10^{-1}} = 9\times10^{-6}\text{ M}$$

63. (d) : NH_4Cl is a salt of strong acid and weak base, so hydrolysis constant is

$$K_h = \frac{K_w}{K_b}$$

Given, $K_b(NH_4OH) = 1.77 \times 10^{-5}$

$K_w = 10^{-14}$

$$\therefore K_h = \frac{10^{-14}}{1.77\times10^{-5}} = 0.565\times10^{-9}$$

or $K_h = 5.65 \times 10^{-10}$

64. (d) : Millimoles of H^+ produced = 20 × 0.05 = 1

Millimoles of OH^- produced = 30 × 0.1 × 2 = 6

(∵ Each $Ba(OH)_2$ gives $2OH^-$.)

∴ Millimoles of OH^- remaining in solution = 6 – 1 = 5

Total volume of solution = 20 + 30 = 50 mL

$$\therefore [OH^-] = \frac{5}{50} = 0.1 \text{ M}$$

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

or $[H^+] = 10^{-pH}$; $[H^+]$ of soln. 1 = 10^{-3}

$[H^+]$ of soln. 2 = 10^{-4}; $[H^+]$ of soln. 3 = 10^{-5}

Total concentration of $[H^+]$

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

$$\Rightarrow 10^{-3}\left(\frac{1}{1} + \frac{1}{10} + \frac{1}{100}\right) \Rightarrow 10^{-3}\left(\frac{100+10+1}{100}\right)$$

$$\Rightarrow 10^{-3}\left(\frac{111}{100}\right) = 1.11 \times 10^{-3}$$

So, H^+ ion concentration in mixture of equal volumes of these acid solution = $\frac{1.11 \times 10^{-3}}{3}$ = 3.7×10^{-4} M

66. (a) : For a weak acid, degree of dissociation,

$$\alpha = \sqrt{\frac{K_a}{C}} = \sqrt{\frac{1 \times 10^{-5}}{0.1}} = 10^{-2} \quad i.e. \ 1.00\%$$

67. (a) : Given, $[H_3O^+] = 1 \times 10^{-10}$ or, pH = 10

Now at 25°C, $pH + pOH = pK_w = 14$

or, pOH = 14 – pH = 14 – 10 = 4

68. (c) : 10^{-8} M HCl = 10^{-8} M H^+

Also from water, $[H^+] = 10^{-7}$

Total $[H^+] = 10^{-7} + 0.10 \times 10^{-7} = 1.1 \times 10^{-7}$ M

69. (d) : $C = 0.01$ M

$K_b = 1 \times 10^{-12}$ at 25°C

	BOH	$\rightleftharpoons$	B^+	+ OH^-
Initially	C		0	0
At eq.	$C - C\alpha$		$C\alpha$	$C\alpha$

$[OH^-] = C\alpha$

$$[OH^-] = \sqrt{K_b C} = \sqrt{1 \times 10^{-12} \times 10^{-2}}$$

$[OH^-] = 10^{-7}$ mol L^{-1}

70. (b) : NH_4OH is a weak base but HCl is a strong acid in solution, so pH of NH_4Cl solution is comparatively low.

$NaNO_3$ is a salt of strong base and strong acid, so pH of the solution will be 7.

Hydrolysis of potassium acetate (a salt of a weak acid and a strong alkali) gives a weakly alkaline solution, since the acetate ion acts as a weak base.

$$CH_3COOK + H_2O \rightarrow CH_3COOH + K^+ + OH^-$$

The pH of this solution ≈ 8.8.

Hydrolysis of sodium carbonate (a salt of strong alkali and a weak acid) gives an alkaline solution.

$$Na_2CO_3 + 2H_2O \rightarrow 2(Na^+ + OH^-) + H_2CO_3$$

The pH of this solution is > 10.

71. (d) : $CH_3COOH \rightleftharpoons CH_3COO^- + H^+$

$$K_a = \frac{[CH_3COO^-][H^+]}{[CH_3COOH]}$$

$$[CH_3COOH] = \frac{3.4 \times 10^{-4} \times 3.4 \times 10^{-4}}{1.7 \times 10^{-5}} = 6.8 \times 10^{-3}$$

72. (b) : (i) $H_2A \xrightleftharpoons{K_{a_1}} HA^- + H^+$

(ii) $HA^- \xrightleftharpoons{K_{a_2}} A^{2-} + H^+$

In the 1st step, H^+ ion comes from neutral molecule, while in the 2nd step the H^+ ion comes from negatively charged ions. The presence of –ve charge makes the removal of H^+ ion difficult. Thus, $K_{a_1} > K_{a_2}$.

73. (b) : After mixing 1 N solution of CH_3COOH (weak acid) and 1 N NaOH (strong base), the resulting solution will have free OH^- ions. Thus, pH will be higher than 7.

74. (a) : $[H^+] = C\alpha = 0.1 \times 0.02 = 2 \times 10^{-3}$ M

(As degree of dissociation = 2% = 0.02)

$$\text{Hence, } [OH^-] = \frac{10^{-14}}{2 \times 10^{-3}} = 5 \times 10^{-12} \text{ M}$$

75. (c) : $NaH + H_2O \rightarrow NaOH + H_2$

or, $H^-_{(aq)} + H_2O_{(l)} \rightarrow OH^- + H_2\uparrow$

76. (d) : At high temperature, the value of ionic product increases.

77. (c) :

	2HI →	H_2	+ I_2
Initially :	2	0	0
After dissociation :	2 – 2α	α	α

Total number of moles = 2(1 – α) + 2α = 2

78. (b) : Since NaOH is a strong base, therefore it completely ionises. Thus, the hydroxyl ion concentration is equal to that of the base itself. We know that concentration of OH^- in N/10 NaOH = 0.1 = 10^{-1}.

Therefore value of

$$[H_3O^+] = \frac{K_w}{[OH^-]} = \frac{1 \times 10^{-14}}{10^{-1}} = 1 \times 10^{-13}$$

$pH = -\log[H_3O^+] = -\log[1 \times 10^{-13}] = 13$

79. (c) : Since HCl is a strong acid and it completely ionises, therefore H_3O^+ ions concentration is equal that of the acid itself *i.e.*, $[H_3O^+]$ = [HCl] = 10 M.

Therefore, $pH = -\log[H_3O^+] = -\log[10] = -1$

80. (a) : $[H_3O^+] = [OH^-] = 1 \times 10^{-6}$ mole/litre

$K_w = [H_3O^+][OH^-] = [1 \times 10^{-6}] \times [1 \times 10^{-6}] = 1 \times 10^{-12}$

81. (a) : Sodium borate is a salt formed from strong base (NaOH) and weak acid (H_3BO_3). Hence, sodium borate will act as basic solution.

82. (c) : NH_4Cl and $NaHCO_3$ are acidic in nature and NaCl is neutral. Only Na_2CO_3 is basic and thus, have highest pH.

83. (d) : Acid-base titration :

$$\underset{10\text{ mmol}}{HCl} + \underset{20\text{ mmol}}{NH_4OH} \longrightarrow NH_4Cl$$

∴ HCl is the limiting reagent.

Solution contains NH_4OH (weak base) and NH_4Cl (salt of strong acid and weak base). Therefore, a basic buffer will be formed.

84. (d) : Acidic buffer is a mixture of a weak acid and its salt with a strong base. $HClO_4$ is a strong acid.

85. (b) : $pH = pK_a + \log\frac{[Salt]}{[Acid]}$

$$5 = -\log K_a + \log\frac{[Salt]}{[Acid]} \quad [\because pK_a = -\log K_a]$$

$$5 = -\log[1 \times 10^{-4}] + \log\frac{[Salt]}{[Acid]}$$

$$5 = 4 + \log\frac{[Salt]}{[Acid]}, \; 5 - 4 = \log\frac{[Salt]}{[Acid]}$$

$$1 = \log\frac{[Salt]}{[Acid]}, \frac{[Salt]}{[Acid]} = 10 \; i.e. \; 10:1$$

86. (a)

87. (b) : $[NH_3] = 0.30$ M, $K_b = 1.8 \times 10^{-5}$

$[NH_4^+] = 0.20$ M

$pK_b = -\log(1.8 \times 10^{-5}) = 4.74$

$$pOH = pK_b + \log\frac{[salt]}{[base]} = 4.74 + \log\frac{0.2}{0.3} = 4.56$$

pH = (14 – 4.56) = 9.44

88. (d) : We know, $pOH = pK_b + \log\frac{[B^-]}{[HB]}$

Since, $[B^-] = [HB]$ (given)

∴ $pOH = pK_b \Rightarrow pOH = 10$

∴ pH = 14 – 10 = 4

89. (b) : HNO_2 (weak acid) and $NaNO_2$ (salt of conjugate base) is an example of acidic buffer.

90. (d) : Let us consider the formation of a salt of a weak acid and a strong base.

$In^- + H_2O \; HIn + OH^-$

$$K_h = \frac{[HIn][OH^-]}{[In^-]} \qquad \text{...(i)}$$

Other equations present in the solution are

$HIn \; H^+ + In^-$

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

$$K_{In} = \frac{[H^+][In^-]}{[HIn]} \qquad \text{... (ii)}$$

$$K_w = [H^+][OH^-] \qquad \text{... (iii)}$$

From (ii) and (iii),

$$\frac{K_w}{K_{In}} = \frac{[HIn][OH^-]}{[In^-]} = K_h \qquad \text{... (iv)}$$

$$[OH^-] = \frac{K_w}{K_{In}}\frac{[In^-]}{[HIn]}$$

$$\log[OH^-] = \log K_w - \log K_{In} + \log\frac{[In^-]}{[HIn]}$$

$$-pOH = -pK_w + pK_{In} + \log\frac{[In^-]}{[HIn]}$$

$$pK_w - pOH = pK_{In} + \log\frac{[In^-]}{[HIn]}$$

$$\text{or, } pH = pK_{In} + \log\frac{[In^-]}{[HIn]}$$

$$i.e. \; \log\frac{[In^-]}{[HIn]} = pH - pK_{In}$$

91. (b) : Solution of 0.1 N NH_4OH and 0.1 N NH_4Cl is a buffer solution.

According to Henderson equation, the pH of a basic buffer, $pH = 14 - pK_b - \log\frac{[Salt]}{[Base]}$

$$\Rightarrow pK_b = 14 - pH - \log\frac{[Salt]}{[Base]}$$

$$\Rightarrow pK_b = 14 - 9.25 - \log\frac{0.1}{0.1}$$

$\Rightarrow pK_b = 14 - 9.25 = 4.75$

∴ pK_b of $NH_4OH = 4.75$

92. (b) : pH of an acidic buffer solution is given by Henderson equation :

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

Its buffer capacity = $pK_a \pm 1$

Since a buffer solution is more effective in pH range $pK_a \pm 1$ therefore, the weak acid having $pK_a = 3.58$ together with its sodium salt is chosen. Acetoacetic acid is, therefore, the suitable weak acid.

93. (d) : The pH value of the blood is maintained constant by buffer solution present in the blood itself. Buffer solutions resist the change in pH values.

94. (b) : pH of the saturated solution of $Ca(OH)_2 = 9$

∴ pOH of the saturated solution of $Ca(OH)_2 = 14 - 9 = 5$

$\Rightarrow [OH^-] = 10^{-5}$ $(\because pH + pOH = 14)$

$$Ca(OH)_2 \; \square \; \underset{\substack{s \\ 1/2 \times 10^{-5}}}{Ca^{2+}} + \underset{\substack{2s \\ 10^{-5}}}{2OH^-}$$

$K_{sp} = [Ca^{2+}][OH^-]^2 = [1/2 \times 10^{-5}][10^{-5}]^2$

$= 0.5 \times 10^{-15}$

95. (c) : $\underset{}{CaF_2} \longrightarrow \underset{s}{Ca^{2+}} + \underset{2s}{2F^-}$

$\underset{0.1\text{ M}}{NaF} \longrightarrow \underset{0.1\text{ M}}{Na^+} + F^-$

$[Ca^{2+}] = s$, $[F^-] = (2s + 0.1) \approx 0.1$ M

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

$5.3 \times 10^{-11} = (s)(0.1)^2$

$$s = \frac{5.3 \times 10^{-11}}{(0.1)^2} = 5.3 \times 10^{-9}\text{ mol L}^{-1}$$

$\therefore$ Molar solubility is 5.3×10^{-9} mol L^{-1}

96. (a) : Solubility of $BaSO_4$,

$$s = \frac{2.42 \times 10^{-3}}{233}\text{ mol L}^{-1} = 1.04 \times 10^{-5}\text{ mol L}^{-1}$$

$BaSO_4$ ionizes completely in the solution as :

$BaSO_{4(s)} \rightleftharpoons \underset{s}{Ba^{2+}_{(aq)}} + \underset{s}{SO^{2-}_{4(aq)}}$

$K_{sp} = [Ba^{2+}][SO_4^{2-}] = s^2$

$= (1.04 \times 10^{-5})^2 = 1.08 \times 10^{-10}\text{ mol}^2\text{ L}^{-2}$

97. (c) : Let solubility of $Ag_2C_2O_4$ be s mol L^{-1}

$Ag_2C_2O_{4(s)} \rightleftharpoons \underset{2s}{2Ag^+_{(aq)}} + \underset{s}{C_2O^{2-}_{4(aq)}}$

$K_{sp} = [Ag^+]^2[C_2O_4^{2-}]$

$K_{sp} = (2s)^2(s) = 4s^3$

$K_{sp} = 4 \times (1.1 \times 10^{-4})^3$ $\qquad (\because [Ag^+] = 2s = 2.2 \times 10^{-4})$

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

98. (b) : Let s be the solubility of AgCl in moles per litre.

$\underset{s}{AgCl_{(aq)}} \qquad \underset{s}{Ag^+_{(aq)}} + \underset{(s+0.1)}{Cl^-_{(aq)}}$

($\because$ 0.1 M NaCl solution also provides 0.1 M Cl^- ion).

$K_{sp} = [Ag^+][Cl^-]$; $1.6 \times 10^{-10} = s(s + 0.1)$

$1.6 \times 10^{-10} = s(0.1)$ $\qquad (\because s <<<< 0.1)$

$$s = \frac{1.6 \times 10^{-10}}{0.1} = 1.6 \times 10^{-9}\text{ M}$$

99. (d) : For MY : $K_{sp} = s_1^2$

$\Rightarrow \quad s_1 = \sqrt{K_{sp}} = \sqrt{6.2 \times 10^{-13}} = 7.87 \times 10^{-7}$ mol L^{-1}

For NY_3 : $K_{sp} = (s_2)(3s_2)^3 = 27s_2^4$

$$\Rightarrow \quad s_2 = \sqrt[4]{\frac{6.2 \times 10^{-13}}{27}} = 3.89 \times 10^{-4}\text{ mol L}^{-1}$$

Hence, molar solubility of MY in water is less than that of NY_3.

100. (b) :

Salt	K_{sp}	**Solubility**
Ag_2CrO_4	$1.1 \times 10^{-12} = 4s^3$	$s = \sqrt[3]{\frac{K_{sp}}{4}} = 0.65 \times 10^{-4}$
AgCl	$1.8 \times 10^{-10} = s^2$	$s = \sqrt{K_{sp}} = 1.34 \times 10^{-5}$
AgBr	$5 \times 10^{-13} = s^2$	$s = \sqrt{K_{sp}} = 0.71 \times 10^{-6}$
AgI	$8.3 \times 10^{-17} = s^2$	$s = \sqrt{K_{sp}} = 0.9 \times 10^{-8}$

Solubility of Ag_2CrO_4 is highest thus, it will be precipitated at last.

101. (b) : $\Delta G° = -2.303RT \log K_{sp}$

63.3×10^3 J $= -2.303 \times 8.314 \times 298 \log K_{sp}$

63.3×10^3 J $= -5705.84 \log K_{sp}$

$$\log K_{sp} = -\frac{63.3 \times 10^3}{5705.84} = -11.09$$

K_{sp} = antilog $(-11.09) = 8.128 \times 10^{-12}$

102. (d) : $CaCO_3 \rightarrow \underset{x}{Ca^{2+}} + \underset{x}{CO_3^{2-}}$

$CaC_2O_4 \rightarrow \underset{y}{Ca^{2+}} + \underset{y}{C_2O_4^{2-}}$

Now, $[Ca^{2+}] = x + y$

and $x(x + y) = 4.7 \times 10^{-9}$ ...(i)

$y(x + y) = 1.3 \times 10^{-9}$ ...(ii)

Dividing equation (i) and (ii) we get

$$\frac{x}{y} = 3.6$$

$\therefore \quad x = 3.6y$

Putting this value in equation (ii), we get

$y(3.6y + y) = 1.3 \times 10^{-9}$

On solving, we get $y = 1.68 \times 10^{-5}$

and $x = 3.6 \times 1.68 \times 10^{-5} = 6.05 \times 10^{-5}$

$\therefore \quad [Ca^{2+}] = (x + y) = (1.68 \times 10^{-5}) + (6.05 \times 10^{-5})$

$\therefore \quad [Ca^{2+}] = 7.73 \times 10^{-5}$ M

103. (b) : Sodium sulphide is soluble in water. The solubility product (and hence solubility) of ZnS is larger than that of CuS.

104. (b) : pH of solution = 12

$[H^+] = 10^{-12}$

$$[OH^-] = \frac{10^{-14}}{10^{-12}} = 10^{-2}$$

$Ba(OH)_2 \qquad \underset{s}{Ba^{2+}} + \underset{2s}{2OH^-}$

$$2s = 10^{-2} \Rightarrow s = \frac{10^{-2}}{2}$$

$K_{sp} = (s)(2s)^2 = 4s^3$

$$= 4\times\left(\frac{10^{-2}}{2}\right)^3 = \frac{4}{8}\times 10{-}6 = 5\times 10{-}7$$

105. (c) : $K_{sp}[AgCl] = [Ag^+][Cl^-]$

$$[Ag^+] = \frac{1.8\times10^{-10}}{10^{-1}} = 1.8\times10^{-9}\text{ M}$$

$K_{sp}[PbCl_2] = [Pb^{2+}][Cl^-]^2$

$$[Pb^{2+}] = \frac{1.7\times10^{-5}}{10^{-1}\times10^{-1}} = 1.7\times10^{-3}\text{ M}$$

106. (a) : The cations of group II are precipitated as their sulphides.

Solubility product of sulphide of group II radicals are very low. Therefore, even with low conc. of S^{2-} ions, the ionic product exceeds the value of their solubility product and the radicals of group II gets precipitated. The low conc. of S^{2-} ions is obtained by passing H_2S gas through the solution of the salt in the presence of dil. HCl which suppresses degree of ionisation of H_2S by common ion effect.

$$H_2S \rightleftharpoons \boxed{\begin{array}{c}2H^+\\ H^+\end{array}}\begin{array}{l}+\ S^{2-}\\ +\ Cl^-\end{array}$$

common ion

Note that solubility product of group IV radicals are quite high. It is necessary to suppress the conc. of S^{2-} ions, otherwise radical of group IV will also get precipitated along with group II radicals.

107. (c) : $K_{sp} = 3.2\times10^{-11}$

$$AX_2\ \underset{s}{A^{2+}} + \underset{2s}{2X^-}$$

$K_{sp} = s\times(2s)^2 = 4s^3$; *i.e.*, $3.2\times10^{-11} = 4s^3$

or, $s^3 = 0.8\times10^{-11} = 8\times10^{-12}$

$\therefore\ s = 2\times10^{-4}$

108. (b) : $AgI \ \square\ Ag^+ + I^-$

$$KI\ \square\ \underset{10^{-4}\text{ M}}{K^+} + \underset{10^{-4}\text{ M}}{I^-}$$

[For KI, 1 N = 1 M]

$[I^-] = s + 10^{-4}$

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

$1\times10^{-16} = s(s+10^{-4})$

$1\times10^{-16} = s^2 + 10^{-4}s$

$1\times10^{-16} = 10^{-4}s$ $\quad[\because s^2 <<< 10^{-4}s]$

$$\Rightarrow\ s = \frac{1\times10^{-16}}{10^{-4}} = 1\times10^{-12}\text{ mol L}^{-1}$$

109. (d) : Let s be the solubility of the electrolyte MX_2.

$[M^{2+}] = s$, $[X^-] = 2s$

Solubility product, $K_{sp} = s\times(2s)^2 = 4s^3$;

$s = 0.5\times10^{-4}$ mol/litre

$\therefore\ K_{sp} = 4\times(0.5\times10^{-4})^3$; $K_{sp} = 5\times10^{-13}$

110. (b) : For reaction, $M_2S\ \underset{2s}{2M^+} + \underset{s}{S^{2-}}$

Solubility $= 3.5\times10^{-6}$

Solubility product, $K_{sp} = [M^+]^2[S^{2-}]$

$= (2s)^2 s = 4s^3 = 4\times(3.5\times10^{-6})^3 = 1.7\times10^{-16}$

111. (d) : For CaF_2, decomposition is as follows :

$$CaF_2 \rightarrow \underset{s}{Ca^{2+}} + \underset{2s}{2F^-}$$

$\Rightarrow K_{sp} = [Ca^{2+}][F^-]^2 = s\times(2s)^2$

or $\quad K_{sp} = 4s^3 \Rightarrow K_{sp} = 4\times(2\times10^{-4})^3$

$\Rightarrow K_{sp} = 32\times10^{-12}$

112. (b) : The greater the solubility product, the greater is the solubility.

113. (a) : There are greater number of Cl^- ions in $CaCl_2$ compared to others. Hence, solubility of AgCl will be minimum in 0.01M $CaCl_2$ due to common ion effect.

114. (d) : Higher the value of solubility product, greater is the solubility.