Short Answer Type Questions – I

Q. 1. At 700 K, the equilibrium constant K, for the reaction $2SO_3(g) \rightarrow 2SO_2(g) + O_2(g)$ is ($1 \cdot 80 \times 10^{-3}$ kPa). What is the numerical value in moles per litre of K_C for this reaction at the same temperature?

Ans. $K_p = K_c(RT)^{\Delta n}$

$2SO_3(g) \rightarrow 2SO_2(g) + O_2(g)$

$\Delta n = np - n_R$

$= 3 - 2 = 1$

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

$K_c = \dfrac{1 \cdot 80 \times 10^{-3}}{0 \cdot 0821 \times 700}$

$= 3 \cdot 1 \times 10^{-5}$ mol/lit.

Q. 2. The following concentrations were obtained for the formation of NH_3 from N_2 and H_2 at equilibrium at 500K. $[N_2] = 1.5 \times 10^{-2}$ M, $[H_2] = 3.0 \times 10^{-2}$ M and $[NH_3] = 1.2 \times 10^{-2}$ M. Calculate equilibrium constant.

Ans. For the reaction,

$N_2(g) + 3H_2(g) \rightleftharpoons 2NH_3(g)$
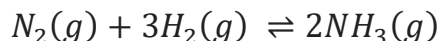
Equilibrium constant

$K_c = \dfrac{[NH_3]^2}{[N_2][H_2]^3}$

$= \dfrac{(1.2 \times 10^{-2})^2}{(1.5 \times 10^{-2})(3 \times 10^{-2})^3}$

$= 0.106 \times 10^4$

$= 1.06 \times 10^3$

Q. 3. At equilibrium, the concentrations of $N_2 = 3.0 \times 10^{-3}\ M, O_2 = 4.2 \times 10^{-3}$ M and NO = 2.8 x 10^{-3} M in a sealed vessel at 800 K. What will be K_c for the reaction-

$N_2(g) + O_2(g) \rightleftharpoons 2NO\ (g)$

Ans. For the reaction, $N_2(g) + O_2(g) \rightleftharpoons 2NO\ (g)$
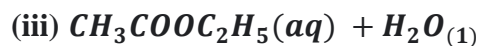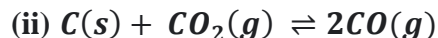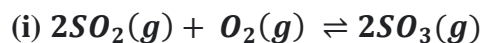
Equilibrium constant $= K_c = \dfrac{[NO]^2}{[N_2][O_2]}$

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

$= 0.622$

Q. 4. PCl₅, PCl₃ and Cl₂ are at equilibrium at 500K and having concentration 1.59 M PCl₃, 1.59 M Cl₂ and 1.41 M PCl₅. Calculate K$_C$ for the reaction, $PCl_5 \rightleftharpoons PCl_3 + Cl_2$

Ans. For the reaction, $PCl_5 \rightleftharpoons PCl_3 + Cl_2$

Equilibrium constant $= K_c = \dfrac{[PCl_3][Cl_2]}{[PCl_5]}$

$$= \dfrac{(1.59) \times (1.59)}{(1.41)}$$

$= 1.79$

Q. 5. Which of the following reactions involve homogeneous equilibrium or heterogeneous equilibrium?

(i) $2SO_2(g) + O_2(g) \rightleftharpoons 2SO_3(g)$

(ii) $C(s) + CO_2(g) \rightleftharpoons 2CO(g)$

(iii) $CH_3COOC_2H_5(aq) + H_2O_{(1)}$

$CH_3COOH(aq) + C_2H_5OH_{(aq)}$

(iv) $BaCO_3(s) \rightleftharpoons BaO(s) + CO_2(g)$
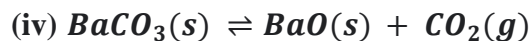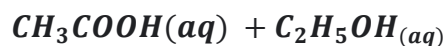
Ans. (i) Homogeneous equilibrium

(ii) Heterogeneous equilibrium

(iii) Homogeneous equilibrium

(iv) Heterogeneous equilibrium

Q. 6. For the reaction, N2(g) + 3H2(g) $\rightleftharpoons$ 2NH3(g), the partial pressures of N_2 and H_2 are 0.80 and 0.40 atmosphere respectively at equilibrium. The total pressure of the system is 2.80 atmosphere. What is K_p for the above reaction?

Ans. For the reaction, $N_2(g) + 3H_2(g) \rightleftharpoons 2NH_3(g)$

Equilibrium constant $= K_p = \dfrac{\left(P_{NH_3}\right)^2}{\left(P_{N_2}\right) \times \left(P_{H_2}\right)^3}$

Given, $P_{N_2} = 0.80$ atm

$P_{N_2} = 0.40$ atm

$P_{total} = 2.80 \, atm$

$P_{N_2} + P_{H_2} + P_{NH_3} = P_{total}$
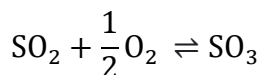
$0.80 + 0.40 + P_{NH_3} = 2.80$

$P_{NH_3} = 2.80 - 1.20$

$= 1.60$ atm

$\therefore K_p = \dfrac{(1.60)^2}{(0.80) \times (0.40)^3}$

$= 50$

Q. 7. K_c for the reaction $SO_2 + \frac{1}{2}O_2 \rightleftharpoons SO_3$ at 600ºC is 61.7. Calculate K_p. What is the unit of K_p for the above equilibrium? ($R = 0.0821$ L atm degree^{-1} mol^{-1})

Ans. For the reaction

$SO_2 + \dfrac{1}{2}O_2 \rightleftharpoons SO_3$

Given, $K_c = 61.7$

$T = 600 + 273 = 873K$

Δn = No. of moles of product - No. of moles of Reactants

$= 1 - \left(1 + \dfrac{1}{2}\right)$

$= -\dfrac{1}{2}$

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

$= 61.7(0.0821 \times 873)^{-1/2}$

$= 7.29$

$K_p = \dfrac{(Pso_3)}{(Pso_2).\left(P_{O_2}\right)^{1/2}}$

$= \dfrac{atm}{atm.(atm)^{1/2}}$

$= \text{atm}^{-1/2}$

$\therefore$ The unit of K_p for the above equilibrium is $\text{atm}^{-1/2}$

Q. 8. The ionic product of water is 0·11 x 10⁻¹⁴ at 273 K, 1 x 10⁻¹⁴ at 298 K and 7·5 x 10⁻¹⁴ at 373K. Deduce from this data whether the ionization of water to hydrogen and hydroxide ion is exothermic or endothermic?

Ans. $K_\omega = (H_3O^+)(OH^-)$

According to the data, the value of K_ω is increasing with temperature. Therefore, according to Le-Chatelier's principle, the ionisation of water is endothermic.

Q. 9. Glycine is an amino acid. It exists in the form of Zwitter ion as $\overset{\oplus}{NH_3}CH_2\overset{\ominus}{COO}$

What are the conjugate acid and conjugate base of this zwitter ion?

Ans. Conjugate acid $\overset{\oplus}{NH_3}CH_2COOH$

Conjugate base $NH_2CH_2\overset{\ominus}{COO}$

Q. 10. Explain ionic product of water. What is the effect of temperature on ionic product of water?   **[DDE, 2017-18]**

Ans. In pure water, one H_2O molecule donates proton and acts as an acid and another H_2O molecule accepts a proton and acts as a base at the same time. So, the following equilibrium exists –

$$H_2O(l) \ + \ H_2O(l) \ \rightleftharpoons \ H_3O^+(aq) \ + \ OH^-(aq)$$
$$\ \ \ Acid \qquad Base \qquad\quad Conjugate\ Acid \qquad Conjugate\ base$$
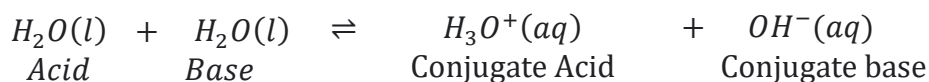
Dissociation constant $K = \dfrac{[H_3O^+][OH^-]}{[H_2O]}$

The concentration of water is omitted from the denominator as water is a pure liquid and its concentration remains constant. $[H_2O]$ is incorporated with in the equilibrium constant to give a new constant, K_w, which is called the ionic product of water.

$$K_w = [H^+] [OH^-]$$

The value of ionic product (K_w) increases with increase in temperature.

Q. 11. What is pH? What is its value for acidic solution?

Ans. The pH of a solution is defined as the negative logarithm to base 10 of the activity (a_{H^+}) of hydrogen ion or concentration of hydrogen ion $[H^+]$.

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

For acidic solution, the value of pH is less than 7.

Q. 12. What is pOH? What is its value for neutral water at 25°C?

Ans. pOH is a measure of hydroxide ion (OH⁻) concentration.

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

For neutral water, the value of pOH is 7 at 25°C.

Q. 13. What will be the conjugate bases for the following Bronsted acids: HE, H₂SO₄, H₃PO₄ and HCO_3^-?

Ans.

Bronsted Acid	Conjugate Base
HF	F^-
H₂SO₄	HSO_4^-
H₃PO₄	$H_2PO_4^-$
HCO_3^-	CO_3^{2-}

Q. 14. Write the conjugate acids for the following Bronsted bases: NH_2^-, NH₃ Cl⁻ and HCOO⁻.

Ans.

Bronsted Base	Conjugate Acid
NH_2^-	NH_3
NH_3	NH_4^+
Cl^-	HCl
$HCOO^-$	$HCOOH$

Q. 15. The species: H_2O, HCO_3^-, HSO_4^-, and NH_3 can act both as Bronsted acids and bases. For each case give the corresponding conjugate acid and conjugate base.

Ans.

Species	Conjugate Acid	Conjugate Base
H_2O	H_3O^+	OH^-
HCO_3^-	H_2CO_3	CO_3^{2-}
HSO_4^-	H_2SO_4	SO_4^{2-}
NH_3	NH_4^+	NH_2^-

Q. 16. Classify the following species into Lewis acids and Lewis bases and show how these act as such:

(a) HO^{--} (b) F^{--} (c) H$^+$ (d) BCl$_3$

Ans. (a) HO^-: It acts as Lewis base as it can donate lone pair of electrons.

(b) F^- : It acts as Lewis base as it can donate lone pair of electrons.

(c) H^+ : It acts as Lewis acid as it can accept lone pair of electrons.

(d) BCl_3 : It acts as Lewis acid as it can accept lone pair of electrons.

Q. 17. The pK_a of acetic acid and pK_b of ammonium hydroxide are 4.76 and 4.75 respectively. Calculate the pH of ammonium acetate solution. [DDE, 2017-18]

Ans. Given,

$pK_a = 4.76$

$pK_b = 4.75$

$\because$ pH $= 7 + \dfrac{1}{2}[pK_a - pK_b]$

$= 7 + \dfrac{1}{2}$ [4.76 - 4.75]

$= 7 + \dfrac{1}{2}$ [0.01]

$= 7 + 0.005$

$= 7.005$

Q.18. Calculate the pH of 0.1 M solution of acetic acid. If the degree of dissociation of acid is 0.0132.   [KVS, Silchar Region, 2016-17]

Ans. $CH_3COOH \rightleftharpoons CH_3COO^- + H^+$

Initial conc.	C	0	0
Equilibrium conc.	$C(1-a)$	Ca	Ca

$Ka = \dfrac{Ca^2}{(1-a)} = Ca^2$

$= 0.1 \times (0.0132)^2$

$= 1.74 \times 10^{-5}$

$pK_a = -log\, pK_a$

$= -\log(1.74 \times 10^{-5})$

$= 5 - 0.2405 = 4.76$

$[H^+] = Ca$

$= 0.1 \times 0.0132$

$= 1.32 \times 10^{-3} M$

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

$= -\log(1.32 \times 10^{-3})$

$= 3 - 0.1206 = 2.88$

Q. 19. Calculate the pH of a 0.01 M solution of acetic acid. K_a for CH_3COOH is 1.8 x 10^{-5} at 25°C. [DDE, 2017-18]

Ans. For the reaction,

$$CH_3COOH \rightleftharpoons CH_3COO^- + H^+$$

Initial conc.   C   0   0
Equilibrium conc. $C(1-a)$ Ca Ca

Conc.
where α = Degree of dissociation

Dissociation constant of acid,

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

$= \dfrac{Ca.Ca}{C(1.a)}$

$= \dfrac{C^2a^2}{C(1-a)}$

∵ For weak acid (1-a) is negligible. So,

$K_a = \dfrac{Ca^2}{1}$

$K_a = Ca^2$

$a = \sqrt{\dfrac{K_a}{C}}$

$= \sqrt{\dfrac{1.8 \times 10^{-5}}{0.01}}$

$$= \sqrt{18 \times 10^{-4}}$$

$$= 4.24 \text{ x } 10^{-2}$$

$$[H^+] = Ca$$

$$= 0.01 \text{ x } 4.24 \text{ x } 10^{-2}$$

$$= 4.24 \text{ x } 10^{-4}$$

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

$$= -log_{10}[4.24 \times 10^{-4}]$$

$$= -log_{10}4.24 + 4 \, log10$$

$$= -0.6273 + 4 = 3.37$$

Q. 20. Account for the following:

(a) A solution of Na₂CO₃ is alkaline.

(b) Ba(OH)₂ is soluble in water while Baso₄ is almost insoluble.　　　　[KVS, 2013]

Ans. (a) $Na_2CO_3 + H_2O \rightleftharpoons H_2CO_3 + 2\,NaOH$

or $CO_3{}^{2-} + 2\,Na^+ + 2H_2O \rightleftharpoons 2Na^+$

$$+2\,OH^- + H_2CO_2$$

Or $CO_3{}^{2-} + 2\,H_2O \rightleftharpoons H_2CO_3 + 2OH^-$

Since N_2CO_3 produces OH^- ions in solution, hence N_2CO_3 solution is alkaline is nature.

(b) The lattice enthalpy of BaSO₄ is much more than its hydration enthalpy and hence it is insoluble in water. But hydration enthalpy of Ba(OH)₂ is much more than lattice enthalpy. Therefore, Ba(OH)₂ is soluble in water.

Q. 21. (a) Define buffer solution.

(b) Give one example each of acidic and basic buffer.　　　[DDE, 2017-18]

Ans. (a) The solution which resists change in pH on dilution or with the addition of small amounts of acid or alkali is called buffer solution.
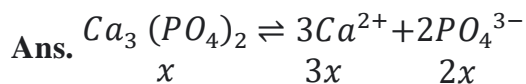
(b) Example of acidic buffer-

$$CH_3COOH + CH_3COONa \, (pH = 4.75)$$

Example of basic buffer-

$$NH_4OH + NH_4CI \, (pH = 9.25)$$

Q. 22. The solubility of $Ca_3(PO_4)_2$ in water is x moles/liter. Calculate its solubility product. [DDE, 2017-18]

Ans. $Ca_3 (PO_4)_2 \rightleftharpoons 3Ca^{2+}+2PO_4{}^{3-}$
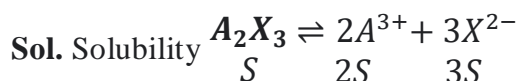$\qquad\quad x \qquad\quad 3x \qquad 2x$

Solubility product $K_{sp} = [Ca^{2+}]^3 [PO_4{}^{3-}]$

$= (3x)^3(2x)^2$

$= 27x^3 \times 4x^2$

$= 128x^5$

Q. 23. Calculate the solubility of A_2X_3 in pure water, assuming that neither kind of ion reacts with water. The solubility product of $A_2X_3, K_{sp} = 1.1 \times 10^{-23}$.

Sol. Solubility $A_2X_3 \rightleftharpoons 2A^{3+}+ 3X^{2-}$
$\qquad\qquad\quad S \qquad 2S \qquad 3S$

∴ Solubility product

$K_{sp} = [A^{3+}]^2 [X^{2-}]^3$
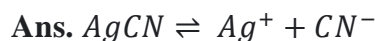
$K_{sp} = (2s)^2(3s)^3$

$= 1.1 \times 10^{-23} = 108 \, S^5$

$or \quad S^5 = \dfrac{1.1 \times 10^{-23}}{108}$

$= 1.0 \times 10^{-25}$

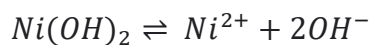$or \quad S = 1.0 \times 10^{-5} mol/L$

Q. 24. The values of K_{sp} of two sparingly soluble salts Ni(OH)$_2$ and AgCN are 2.0×10^{-15} and 6.0×10^{-17} respectively. Which salt is more soluble?
Explain. [NCT, 2008, KVS, 2008, 2010]

Ans. $AgCN \rightleftharpoons Ag^+ + CN^-$

$K_{sp} = [Ag^+][CN^-] = 6 \times 10^{-17}$

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

$K_{sp} = [Ni^{2+}][OH^-]^2 = 2 \times 10^{-15}$

$Let \, [Ag^+] = S, then \, [CN^-] = S_1$

$Let \, [Ni^{2+}] = S_2, then \, [OH^-] = 2S_2$

∴ From eq. (i)

$S_1 \times S_1 = 6 \times 10^{-17}$

$S_1^2 = 6 \times 10^{-17}$

$S_1 = 7.8 \times 10^{-9}$

∴ From eq. (ii)

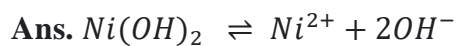$(S_1) \times (2S_2)^2 = 2 \times 10^{-15}$

$4S_2^3 = 2 \times 10^{-15}$

$S_2^3 = \dfrac{2 \times 10^{-15}}{4}$

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

$or \quad S_2 = 0.58 \times 10^{-4}$

Since solubility of Ni(OH)₂ is more than AgCN, so, Ni(OH)₂ is more soluble than AgCN.

Q. 25. Calculate the molar solubility of Ni(OH)₂ in 0.10 M NaOH. The ionic product of Ni(OH)₂ is 20×10^{-15}.

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

Let the solubility of Ni(OH)₂ be equal to S. Dissolution of S mol/L of Ni(OH)₂ provides S mol/L of Ni^{2+} and 2S moVL of OH^-, but the total concentration of $OH^- = (0.10 + 2S)$ mol/L because the solution already contains 0.10 mol/L of OH^- from NaOH.

$K_{sp} = 2.0 \times 10^{-15}$

$K_{sp} = [Ni^{2+}][OH^-]^2$

$2.0 \times 10^{-15} = (S)(0.10 + 2s)^2$

As K_{sp} is small, 2S ◁◁ 0.10

∴ $0.10 + 2S \simeq 0.10$

Hence, $2 \times 10^{-15} = (S)(0.10)^2$

$S = \dfrac{2 \times 10^{-15}}{0.01}$

$S = 2 \times 10^{-13} M = [Ni^{2+}]$

Q. 26. The ionization constant of formic acid is 1.8×10^{-4}. Calculate the ratio of sodium formats and formic acid in a buffer of pH 4.25.

Ans.

$K_a = 1.8 \times 10^{-14}$

pH = 4.25

For acidic buffer,

$$pH = PK_a + \log \frac{[Salt]}{[acid]}$$

$or \quad \log \dfrac{[Salt]}{[acid]} = pH - PK_a$

$PK_a = -\log[K_a]$

$= -\log[1.8 \times 10^{-4}]$

$= -\log 1.8 \times 4 \; log \; 10$

$= -0.2552 + 4$

$= 3.7448 = 3.74$

From eq. (i)
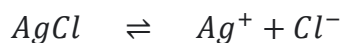
$\log \dfrac{[Salt]}{[acid]} = 4.25 - 3.74$

$= 0.51$

$or \quad \dfrac{[Salt]}{[acid]} = antilog \; 0.51$

$= 3.24$

Q. 27. In one liter saturated solution of AgCl [K_{sp} = 1.6 × 10⁻¹⁰], 0.1 mol of CuCl [K_{sp} = 1.0 × 10⁻⁶] is added. Find out the resultant concentration of Ag^+ in the solution.
[KVS, Agra Region, 2016-17, 2015-16]

Ans. Let the concentration of AgCl is x mol/L and concentration of CuCl is y mol/L

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

$CuCl \quad \rightleftharpoons \quad Cu^+ \; + \; Cl^-$

$\qquad\qquad y = 0.1 \quad y = 0.1$

$K_{sp} \; of \; AgCl = [Ag^+][Cl^-]$

$= x\,(x + y) \qquad\qquad \ldots(i)$

$K_{sp} \; of \; CuCl = [Cu^+][Cl^-]$

$= y\,(x + y) \qquad\qquad \ldots(ii)$

On solving equation (i) and (ii), we get

$$\frac{K_{sp} \; of \; AgCl}{K_{sp} \; of \; CuCl} = \frac{x}{y}$$

$$\frac{1.6 \times 10^{-10}}{1.0 \times 10^{-6}} = \frac{x}{0.1}$$

$$or \quad x = \frac{1.6 \times 10^{-10} \times 0.1}{1.0 \times 10^{-6}}$$

$= 1.6 \times 10^{-5}$ mol/L