

QB365
Important Questions - Equilibrium
11th Standard CBSE

Chemistry

Reg.No. :

Time : 01:00:00 Hrs

Total Marks : 50

Section-A

- 1) A tank is full of water. Water is coming in as well as going out at same rate. What will happen to level of water in a tank? What is name given to such a state? 1
- 2) Explain , why pure liquids and solids can be ignored while writing the equilibrium constant expression? 1
- 3) What will be the pH of 1M Na_2SO_4 solution? 1
- 4) What is the expression for K_{sp} of Ag_2CrO_4 ? 1
- 5) Is it possible to get a precipitate of $Fe(OH)_3$ at pH=2? given reason. 1
- 6) Will AgCl be more soluble in aqueous solution or NaCl solution and why? 1
- 7) Why pH of our blood remains almost constant at 7.4 though we quite often eat spicy food? 1
- 8) the ionization constant of formic acid is 1.8×10^{-4} . Calculate the ratio of sodium formate and formic acid in a buffer of pH4.25 1
- 9) Identify the substance that get reduced in the following reaction. 1
 $Fe_2O_3(s) + 3CO(g) \rightarrow 2Fe(s) + 3CO_2(g)$
- 10) Write the conjugate acids for the following Bronsted bases NH_2^- , NH_3 and $HCOO^-$. 1
 Base + $H^+ \rightleftharpoons$ Conjugate acid

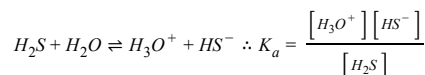
Section-B

- 11) Which of the following reactions involve homogeneous equilibrium and which involve heterogeneous equilibrium? 2
 $Ag_2O(s) + 2HNO_3(aq) \rightleftharpoons 2AgNO_3(aq) + H_2O(l)$
- 12) Which of the following reactions involve homogeneous equilibrium and which involve heterogeneous equilibrium? 2
 $CH_3COOC_2H_5(aq) + H_2O(l) \rightleftharpoons CH_3COOH(aq) + C_2H_5OH(aq)$
- 13) What will be the PH of 0.1 M ammonium acetate solution? $Pk_a = Pk_b = 4.74$ 2
- 14) For the reaction $N_2(g) + 3H_2(g) \rightleftharpoons 2NH_3(g)$, the partial pressure 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? 2
- 15) One mole of H_2O and one mole of CO are taken in 10 L vessel and heated to 725 K. At equilibrium, 40% of water (by mass) reacts with CO according to the equation, 2
 $H_2O(g) + CO(g) \rightleftharpoons H_2(g) + CO_2(g)$
 Calculate the equilibrium constant for the reaction.
- 16) Urine has a pH of 6.0. If a patient eliminates 1300 mL of urine per day, how many gram equivalents of the acid the eliminates per day? 2
- 17) Calculate the percentage hydrolysis of sodium acetate in 0.1 M solution at 25°C assuming the salt to be completely dissociated. $K_a = 1.8 \times 10^{-5}$. 2
- 18) Blood has pH of 7.40. Calculate the ratio of hydrogen carbonate ion HCO_3^- to H_2CO_3 carbonic acid in blood to maintain its pH value. 2
 $pK_a (H_2CO_3) = 6.37$.
- 19) One of the reaction that takes place in producing steel from iron ore is the reduction of iron (II) oxide by carbon monoxide to give iron metal and CO_2 2
 $FeO(s) + CO(g) \rightleftharpoons Fe(s) + CO_2(g); K_p = 0.265 \text{ atm at } 1050 \text{ K}$
 What are the equilibrium partial pressure of CO and CO_2 at 1050 K if the initial partial pressure are $P_{CO} = 1.4 \text{ atm}$ and $PCO_2 = 0.80 \text{ atm}$?
 First Q_p by the given initial partial pressures of [CO] and $[CO_2]$.
- 20) What is the minimum volume of water required to dissolved 1g of calcium sulphate at 298 K? (For calcium sulphate, K_{sp} is 9.1×10^{-6}) 2

Section-C

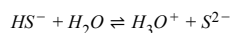
- 21) Which of the following reactions involve homogeneous equilibrium and which ones involve heterogeneous equilibrium? 5
 $2NH_2(g) \rightleftharpoons N_2(g) + 3H_2(g)$
- 22) K_{a_1} , K_{a_2} and K_{a_3} are the respective ionisation constants for the following reactions. 5
 $H_2S \rightleftharpoons H^+ + HS^-$
 $HS^- \rightleftharpoons H^+ + S^{2-}$
 $H_2S \rightleftharpoons H^+ + HS^- \rightleftharpoons H^+ + S^{2-} \rightleftharpoons 2H^+ + S^{2-}$
 Write the correct relationship K_{a_1} , K_{a_2} , K_{a_3} .
 To find out the correct relationship between three ionisation constants (K_{a_1} , K_{a_2} and K_{a_3}) this must be keep in mind that when two reactions are added, their equilibrium constants are multiplied.
- 23) Calculate the pH of a buffer which is 0.1 M in acetic acid and 0.15 M in sodium acetate. Given that the ionisation constants of acetic acid is 1.75×10^{-5} . Also calculate the change in pH of the buffer if to 1L of the buffer 5
 1 cc of 1 M HCl are added. Assume that the change in volume is negligible.

- 24) The first ionisation constant on H_2S is 9.1×10^{-8} . Calculate the concentration of HS^- ion in its 0.1 M solution. How will this concentration be affected if the solution is 0.1 M in HCl also? If the second dissociation constant of H_2S is 1.2×10^{-13} , calculate the concentration of S^{2-} under both conditions.
- H_2S being a weak acid, dissociates as



HCl being a strong acid dissociates completely, so calculate $[HS^-]$ in the presence of 0.1M HCl by taking $[HS^+]$ concentration as 0.1M.

Calculate $[S^{2-}]$ ion concentration by using the equation,



Calculate the $[S^{2-}]$ ion concentration in the presence of 0.1M HCl.

Section-A

- 1) It will remain the same because rate of inflow is equal to rate of outflow. The state is called of 'equilibrium'. 1
- 2) 1
Molar concentration of pure solid or liquid (if in excess) is constant (i.e. independent of the amount present). That's why pure liquids and solids can be ignored while the equilibrium constant expression.
- 3) Na_2SO_4 is a salt of the strong acid and strong base, thus its aqueous solution will be neutral. Therefore, its pH will be 7 1
- 4) $Ag_2CrO_4(s) \rightleftharpoons 2Ag^+(aq) + CrO_4^{2-}(aq); K_{sp} = [Ag^+]^2[CrO_4^{2-}] = (2s)^2(s) = 4s^3$ 1
- 5) No, because $Fe(OH)_3$ will dissolve in the strongly acidic medium. 1
- 6) 1
In NaCl Solution, the Cl^- ions will increase. Since solubility product, $K_{sp} = [Ag^+][Cl^-]$ remains constant, $[Ag^+]$ will decrease. Therefore, the solubility of AgCl will be less in NaCl solution than in water.
- 7) 1
Blood is a buffer containing carbonic acid (H_2CO_3) and bicarbonate ions (H_2CO_3). Small amounts of the acid or base produced from the spicy food do not disturb its pH.
- 8) $pK_a = -\log(1.8 \times 10^{-4}) = 3.74$ $\log \frac{[salt]}{[Acid]} = pH - pK_a = 4.015 - 3.75 = 0.51$ or $\frac{[salt]}{[Acid]} = \text{Antilog } 0.51 = 3.24$ 1
- 9) In the reaction, Fe_2O_3 loses oxygen and is reduced to Fe. 1
- 10) The conjugate acid should have one extra proton in each case and therefore the corresponding conjugate acid are NH_3 , NH_4^+ and HCOOH respectively 1

Section-B

- 11) Heterogeneous equilibrium 2
- 12) Homogeneous equilibrium 2
- 13) $pH = -\frac{1}{2}pK_w + \frac{1}{2}pK_a - \frac{1}{2}pK_b = -\frac{1}{2}(14) + \frac{1}{2}(4.75) - \frac{1}{2}(4.74) = 7$ 2
- 14) $N_2(g) + 3H_2(g) \rightleftharpoons 2NH_3(g)$, 2
Given, at equilibrium, $P_{N_2} = 0.80$ atmosphere
 $P_{H_2} = 0.40$ atmosphere
 $P_{N_2} + P_{H_2} + P_{NH_3} = 2.80$ atmosphere
 $\therefore P_{NH_3} = 2.80 - (0.80 + 0.40) = 1.60$ atmosphere
From, $K_p = \frac{P_{NH_3}^2}{P_{N_2} \times P_{H_2}^3} = \frac{(1.60)^2}{0.80 \times (0.40)^3} = 50.0$
- 15) $H_2O(g) + CO(g) \rightleftharpoons H_2(g) + CO_2(g)$ 2

Initial conc.	1	1	0	0
Equili. conc.	(1-x)	(1-x)	x	x

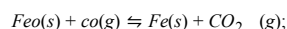
H_2O reacted = 40% of 1 mole of H_2O = 0.4 mol
 $\therefore x = 0.4 \text{ mol}$
 $(1-x) = 1 - 0.4 = 0.6 \text{ mol}$
 $K_c = \frac{[H_2][CO_2]}{[H_2O][CO]} = \frac{0.4 \times 0.4}{0.6 \times 0.6} = 0.444$
- 16) pH=6.0 2
 $[H_3O^+] = 10^{-6} \text{ M}$
 $[Acid] = 10^{-6} \text{ M} = 10^{-6} \text{ N}$
 Thus, 1000 mL of the urine contain acid = 10^{-6} g eq.
 1300 mL of the urine will contain acid = $1.3 \times 10^{-6} \text{ g eq.}$
- 17) Ans. 0.00075% 2

18) **Ans.** 10 mL

2

19)

2



Initial pressure 1.4atm 0.80 atm

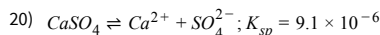
($K_p = 0.265$ at 1050 K)

$$Q_p = \frac{P_{CO_2}}{P_{CO}} = \frac{0.80}{1.4} = 0.571 [\because Fe \text{ and } FeO \text{ are solids}]$$

$\because Q_p > K_p$, the reaction will go in reverse direction. Due to this, pressure of CO_2 will decrease and that of CO will increase to attain equilibrium.

Suppose p is the decrease in pressure of CO_2 and p is the increase in pressure of CO. Hence,

$$p_{CO_2} = (0.80 - p) \text{ and } p_{CO} = (1.4 + p) \text{ Now, from } K_p = \frac{p_{CO_2}}{p_{CO}} \Rightarrow 0.265 = \frac{(0.80 - p)}{(1.4 + p)} \quad p = \frac{0.429}{1.265} = 0.339 \text{ atm Hence, at equilibrium, } p_{CO_2} = 0.80 - 0.3$$



2

S S S

Where s is the solubility of $CaSO_4$

$$K_{sp} = [Ca^{2+}][SO_4^{2-}] = S \cdot S = S^2 S = \sqrt{K_{sp}} = \sqrt{9.1 \times 10^{-6}} \Rightarrow S = 3.017 \times 10^{-3} M$$

Solubility of $CaSO_4 = 3.017 \times 10^{-3} \text{ mol}^{-1}$

$$= 3.017 \times 10^{-3} \times 136 \text{ gL}^{-1}$$

(Molar mass of $CaSO_4 = 136 \text{ g mol}^{-1}$)

$$= 410.3 \times 10^{-3} \text{ gL}^{-1}$$

$410.3 \times 10^{-3} \text{ g } CaSO_4$ is dissolved in = 1L

$$1 \text{ g } CaSO_4 \text{ is dissolved in } = \frac{1 \times 1}{410.3 \times 10^{-3}} = 2.437 \text{ L}$$

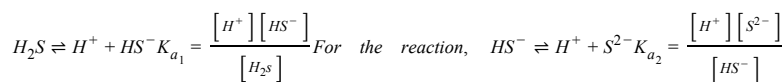
Section-C

21) Homogeneous equilibrium

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22) For the reaction

5



When the above two reactions are added, their equilibrium constants are multiplied, thus

$$K_{a3} = \frac{[H^+]^2[S^{2-}]}{[H_2S]} = K_{a1} \times K_{a2} \text{ Hence, } K_{a3} = K_{a1} \times K_{a2} x = \frac{-b \pm \sqrt{b^2 - 4ac}}{2a}$$

23) 1 cc of 1 M HCl contains HCl = 10^{-3} mol. This will convert 10^{-3} mol CH_3COONa into CH_3COOH .

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Now, [Acid] = $0.10 + 0.001 = 0.101 \text{ M}$

[Salt] = $0.15 - 0.001 = 0.149 \text{ M}$

$$pH = 4.757 + \log \frac{0.149}{0.101}$$

$$= 4.757 + 0.169$$

$$= 4.925$$

Decrease in pH = $4.933 - 0.007$ which is again negligible.

24) In the presence of 0.1M HCl,

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$$K_{a2} = \frac{[H_3O^+][S^{2-}]}{[HS^-]} = \frac{[0.1][S^{2-}]}{[9.1 \times 10^{-8}]} [S^{2-}] = 1.092 \times 10^{-19} M$$