1. State Law of mass action ?

Ans: It states that the rate of a chemical reaction is directly proportional to the product of molar concentration of the reactants each raised to a power equal to the corresponding stoichiometric coefficient which appear in the balanced chemical equation.

For a reaction : $aA + bB \rightleftharpoons cC + dD$

According to law of mass action,

Rate of reaction $\propto [A]^a [B]^b$

or Rate of reaction $= K[A]^a [B]^b$

where K is a constant of proportionality called rate constant.

2. State Lewis concept of acids and Bases ?

Ans: According to this concept, acids are the substances which can accept a pair of electrons and bases are substances which can donate a pair of electrons.

Acids → electron pair acceptors

Bases → electron pair donors.

3. State Arrhenius concept of acids and bases ?

Ans: According to this concept, acids are the substances which gives hydrogen ion (H^+) in aqueous solution and bases are substances which gives hydroxyl ions (OH^-) in aqueous solution.

3. State Bronsted Lowry concept of acids and bases ?

Ans: According to this concept, acids are the substance which can donate a proton (H^+) and bases are substances which can accept a proton (H^+).

In other words, acids : Proton donors

Bases : Proton acceptors.

4. What are strong electrolytes and weak electrolytes ? Give a suitable example of each.

Ans: Strong electrolytes are the substances that dissociate largely or almost completely (100%) into ions in the solution. The solution consequently becomes very good conductors of electricity. eg: HCl, H_2SO_4, $NaOH$ etc.

Weak electrolytes are the substances that never get completely dissociated. In ordinary conditions, they dissociate into ions in solution to a small extent ($\approx$ 5%) and the solution consequently becomes poor conductor of electricity. eg: H_3PO_4, H_3BO_3, NH_4OH etc.

5. Define pH. How are pH of pure water vary with temperature.

Ans: The pH of a solution is defined as the negative logarithm of the hydrogen ion concentration of the solution.

i.e $pH = -\log [H^+]$

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

The reason temperature affects pure water's pH is that water molecules have a slight tendency to break down into their constituents, hydrogen and oxygen as temperature increases. As the temperature increases, a larger proportion of water molecules break up releasing a few more hydrogen ions, which then decreases the pH of pure water.

6. State the Law of chemical equilibrium ?

Ans: The Law of mass action may be applied to a reversible reaction to derive a mathematical expression for equilibrium constant is known as Law of chemical equilibrium.

Let us consider a simple reversible reaction

$$A + B \rightleftharpoons \overset{C}{X} + \overset{D}{Y}$$

in which an equilibrium exists between the reactant (A and B) and products (X & Y).

At equilibrium,

Rate of forward reaction = Rate of backward reaction

$$K_f [A][B] = K_b [\overset{C}{X}][Y]$$

$$\frac{K_f}{K_b} = \frac{[\overset{C}{X}][Y]}{[A][B]}$$

$$\Rightarrow \quad K = \frac{[X][Y]}{[A][B]} \quad \text{——} (1)$$

∴ The combined constant K is known as equilibrium constant. The above eqn is known as Law of chemical

7. Why BF_3 acts as Lewis acid ?

Ans: A Lewis acid can accept a pair of electrons from a lewis base. The boron in BF_3 is electron deficient and has an empty orbital, so it can accept a pair of electrons, making it a Lewis acid.

8. What is a buffer solution? Give any two application of a buffer?

Ans: A solution which resists any change of pH when strong acid or a strong base is added to it is called a buffer solution or simply a buffer.

eg: CH_3COOH, NH_4OH etc.

Applications:

(i) In Biochemical systems: The pH plays a major role in biochemical systems. The blood in our body is buffered at a pH value of 7.36 - 7.4 due to bicarbonate- carbonic acid buffer. A mere change of 0.2 pH can even cause death.

(ii) Food preservation: The control of pH is very important in the field of food preservation.

(iii) Agriculture: The pH of the soil is very important for having proper crop yield. The choice of fertilizers also depend upon pH of the soil.

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9: Derive the relationship between K_p & K_c for an equilibrium reaction.

Or

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

Ans: Let us consider a reaction,

$$aA + bB \rightleftharpoons cC + dD$$

The equilibrium constant for the reaction in terms of concentration (mole/litre) may be expressed as:

$$K_c = \frac{[C]^c [D]^d}{[A]^a [B]^b} \quad \text{———} \quad (1)$$

The equilibrium constant in terms of partial pressure may be expressed as:

$$K_p = \frac{[P_C]^c . [P_D]^d}{[P_A]^a . [P_B]^b} \quad \text{———} \quad (2)$$

If the gases are assumed to be ideal, according ideal gas equn,

$$P_i V_i = n_i RT$$

$$P_i = \frac{n_i}{V_i} RT$$

$$= C_i RT \quad (3)$$

$$= [i] RT$$

where $[i]$ is the moles i. of the species concentration

$$\text{equn } (2) \Rightarrow K_p = \frac{([C] RT)^c . ([D] RT)^d}{([A] RT)^a . ([B] RT)^b}$$

$$= \frac{[C]^c [D]^d . (RT)^{c+d}}{[A]^a [B]^b . (RT)^{a+b}}$$

or $K_p = K_c \cdot \dfrac{(RT)^{c+d}}{(RT)^{a+b}}$ ..... from ①

or $K_p = K_c \cdot (RT)^{(c+d)-(a+b)}$

or $K_p = K_c \cdot (RT)^{\Delta n}$ ..... $\because (c+d)-(a+b) = \Delta n$

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

10. At 700K, the equilibrium constant K_p for the reaction

2017 $2SO_3(g) \rightleftharpoons 2SO_2(g) + O_2(g)$ is 1.80×10^{-3} kPa. What is the numerical value in moles per litre of K_c for this reaction at the same temperature? [Given, $R = 8.31$ L kPa K^{-1} mol^{-1}]

Ans: Given reaction: $2SO_3(g) \rightleftharpoons 2SO_2(g) + O_2(g)$

Again, $R = 8.31$ L kPa $K^{-1} mol^{-1}$

$K_p = 1.80 \times 10^{-3}$ kPa ; $K_c = ?$

W.K.T: $K_p = K_c \cdot (RT)^{\Delta n}$

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

$= \dfrac{1.80 \times 10^{-3} \text{ kPa}}{(RT)^{\{(2+1)-2\}}}$ ... $[\because \Sigma n_p - \Sigma n_r = \Delta n$

$= \dfrac{1.80 \times 10^{-3} \text{ kPa}}{(RT)^1}$

$= \dfrac{1.80 \times 10^{-3} \text{ kPa}}{8.31 \text{ L kPa } K^{-1} mol^{-1} \times 700 \text{ K}}$

$= 3.09 \times 10^{-7}$ mol L^{-1}

11. Apply Le chatelier's principle to the reaction,

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

Very Imp

$\Delta_r f = -93.6$ KJ, and predict the most favourable conditions for the reaction

Ans: On application of Le-chatelier's principle to the given reaction, the most favourable conditions to maximise the preparation of Ammonia are given below :-

(i) Increase in concentration of N_2 or/and H_2 or decrease in concentration of NH_3

Increase in the conc. of N_2 or H_2 (any one of reactants) $\xrightarrow[\text{equilibrium towards}]{\text{shifts the}}$ Forward direction

Thus, in general, increase in concentration of any of the substances on one side of the equilibrium shifts the equilibrium to produce more of the substances on the other side of it.

(ii) Decrease in temperature :

Decrease in temperature $\xrightarrow[\text{equilibrium towards}]{\text{shifts the}}$ Exothermic reaction

A decrease in temperature of the system will favour the exothermic reaction. Thus when the temperature of the system is lowered, the equilibrium will shift in the forward direction ie more ammonia will be produced.

(iii) Increase in Pressure :

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

1 mol	3 mol	2 mol	($\Delta n = -2$)
1 vol	3 vol	2 vol	

Increase of pressure on a gaseous system will s the equilibrium in a direction of decrease in Volume. Here (Δn = -ve), so equilibrium will shift in the forward direction by increasing the pressure.

Increase in pressure $\xrightarrow[\text{equilibrium towards}]{\text{Shifts the}}$ Lesser no. of gase molecules (ie decrease in

12. Explain solubility product with a reaction.

Ans: The product of the molar concentrations of the ions in a saturated solution of the sparing soluble salt each raised to the power equal to the stoichiometric coefficient of t species in a balanced chemical equation i called solubility product (K_{sp})

eg: $AgCl \rightleftharpoons Ag^+ (aq) + Cl^- (aq)$

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

13. Explain Common ion effect with the help of an example

Ans: The solubility of a sparingly soluble salt decreases when a highly soluble salt having one ion common to the sparingly soluble salt is added to the solution, due to the common – ion effect.

eg: The solubility of AgCl (sparingly soluble salt) in water decreases when a small amount of KCl is added to the solution. This is because KCl dissociate completely in the solution to produce more Cl^- (common ion to AgCl) and thus increase in concentration of Cl^- ions shifts the equilibrium towards left.

$$KCl \longrightarrow K^+ + Cl^-$$

$$AgCl(s) \rightleftharpoons Ag^+(aq) + Cl^-(aq)$$

→ It causes precipitation of more AgCl (s)

→ Thus solubility of AgCl in solution decrease on addition of small quantity of KCl.

ALTERNATE ANSWER FOR Q 5 (2nd part)
→ Choose easy one

Ans: With a rise in temperature, the extent of ionisation of water also increases. Thus, at higher temperature the equilibrium concentration of H^+ would be higher and therefore pH would be lower.

14 Show that $pH^+ + pOH^-$ is equal to K_w. What is the value of $pH^+ + pOH^-$ at 25°C or 298 K:

Ans: W.K.T:

Ionic product of water,

$$K_w = [H^+][OH^-] \quad \text{———} \ (1)$$

Taking log on both sides, we get

$$\text{or} \quad \log K_w = \log [H^+] + \log [OH^-]$$

$$\Rightarrow -\log K_w = -\log [H^+] + \{-\log [OH^-]\} \quad \ldots \text{Multiplying } -1 \text{ on both sides}$$

$$\Rightarrow pK_w = pH^+ + pOH^- \quad \text{———} \ (2)$$

At 25°C / 298 K,

$$K_w = 1.0 \times 10^{-14}$$

Taking negative log on both sides, we get,

$$\text{or} \quad -\log K_w = -\log(10^{-14})$$

$$pK_w = -(-14)$$

$$\text{or} \quad pK_w = 14$$

$$\Rightarrow pH^+ + pOH^- = 14 \quad \text{———} \ \text{from } (2)$$