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Subject: Chemistry

Weightage: High

Title: Chemical & Ionic Equilibrium

Equilibrium Notes

Reversible and Irreversible Reactions

Reversible reactions: These are reactions where the conversion of reactants to products does not go to completion.

$$CH_3COOH + NaOH \rightleftharpoons CH_3COONa + H_2O$$

Irreversible reactions: The reactants are completely converted into products.

$$NaOH + HCl \rightarrow NaCl + H_2O$$

Chemical Equilibrium and Its Dynamic Nature

Equilibrium is a condition where concentrations of reactants and products remain unchanged with time, indicating a dynamic balance.

Law of Mass Action and Equilibrium Constant

For a reaction:

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

$$K_c = \frac{[C]^c[D]^d}{[A]^a[B]^b}$$

Relationship Between K_p, K_c, and K_x

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

$$K_p = K_x(P)^{\Delta n}$$

Where Δn is the change in moles of gaseous species.

Relation Between ΔG° and K

$$\Delta G = \Delta G^\circ + 2.303RT \log Q$$

$$\Delta G^\circ = -2.303RT \log K$$

Le-Chatelier's Principle

If a change is applied to a system at equilibrium, the system adjusts to counteract the imposed change.

Effect of Concentration

- Adding reactants drives reaction forward.
- Adding products drives reaction backward.
- Removing products shifts equilibrium forward.

Effect of Pressure

- Increase favors fewer gas moles.
- Decrease favors more gas moles.
- No Δn: no pressure effect.

Effect of Temperature

- Raising temperature favors endothermic direction.
- Lowering temperature favors exothermic direction.

Inert Gas Addition

Inert Gas Addition

- At constant volume: no effect.
- At constant pressure: increases volume, shifts to side with more moles.

Effect of Catalyst

Catalysts do not affect equilibrium position; they speed up both forward and reverse reactions equally.

Ionic Equilibrium

Electrolytes: Substances conducting electricity in solution.

- **Strong:** Fully dissociate.
- **Weak:** Partially dissociate.

Acid-Base Theories

Arrhenius: Acids produce H^+; bases produce OH^- in water.
Bronsted-Lowry: Acid = H^+ donor, Base = H^+ acceptor.
Lewis: Acid = electron pair acceptor, Base = donor.

Ionic Product of Water

$$H_2O \rightleftharpoons H^+ + OH^-, \quad K_w = [H^+][OH^-] = 1.0 \times 10^{-14} \text{ at } 25°C$$

pH and pOH Scale

$$pH = -\log_{10}[H^+]$$
$$pOH = -\log_{10}[OH^-]$$
$$pK_w = pH + pOH = 14 \text{ at } 25°C$$

Buffer Solutions

- **Acidic buffer:** Weak acid + salt of weak acid and strong base.
- **Basic buffer:** Weak base + salt of weak base and strong acid.

Solubility Product (K_{sp})

For salt A_xB_y:

$$K_{sp} = [A^{y+}]^x[B^{x-}]^y$$

- If K > K_{sp}: precipitation occurs.
- If K < K_{sp}: no precipitation.
- If K = K_{sp}: saturated solution.

Common Ion Effect

Adding a strong electrolyte with a common ion suppresses the ionization of a weak electrolyte.

Acid Strength Comparisons

Binary Acids

- Across a period: Increases with electronegativity. $CH_4 < NH_3 < H_2O < HF$
- Down a group: Increases with atomic size. $HF < HCl < HBr < HI$

Oxyacids

- Different halogens: Electronegativity order $HOI < HOBr < HOCl$
- Same halogen, increasing oxidation state: $HOCl < HClO_2 < HClO_3 < HClO_4$
- Across period: Strength increases.
- Down group: Strength decreases.