

Equilibrium — Assignment 1

Class 11 Chemistry · Topic-wise question bank covering every concept of the chapter · Narayan Gurukul Academy

Instructions: (1) Attempt all questions. This worksheet walks through every concept of the chapter, topic by topic.
(2) The interactive MCQ version of these questions (with instant answer reveal) is available on the notes page.
(3) Universal values: $g = 10 \text{ m/s}^2$ unless stated. Give answers with correct SI units.

Reversible Reactions and Equilibrium State

MCQ Questions (choose the correct option)

- A reversible reaction reaches equilibrium when
 - the rate of the forward reaction equals the rate of the backward reaction
 - the forward reaction stops
 - all reactants are used up
 - the temperature becomes zero
- At equilibrium, the concentrations of reactants and products
 - remain constant
 - become equal
 - become zero
 - keep increasing
- Equilibrium is dynamic because
 - both reactions continue at equal rates
 - reactions stop
 - only the forward reaction occurs
 - only the backward reaction occurs
- For the reaction $\text{N}_2\text{O}_4 \rightleftharpoons 2\text{NO}_2$, an increase in pressure shifts the equilibrium
 - backwards
 - forwards
 - no change
 - to completion

1 Mark Questions

- Define chemical equilibrium.
- Why is equilibrium called dynamic?
- What is a reversible reaction?

2 Mark Questions

- Explain the characteristics of the equilibrium state with an example.
- Distinguish between a reversible and an irreversible reaction with one example each.

3 Mark Questions

- Explain the nature of the equilibrium state, why it is dynamic, and the characteristics of a system at equilibrium.
- Explain the dynamic nature of equilibrium with the help of the reaction $\text{N}_2\text{O}_4 \rightleftharpoons 2\text{NO}_2$.

Equilibrium Constants K_c and K_p

MCQ Questions (choose the correct option)

- For the reaction $aA + bB \rightleftharpoons cC + dD$, the equilibrium constant K_c is
 - $[C]^c[D]^d/[A]^a[B]^b$
 - $[A]^a[B]^b/[C]^c[D]^d$
 - $[A][B]$
 - $[C][D]$
- The relation between K_p and K_c is
 - $K_p = K_c(RT)^{\Delta n}$
 - $K_p = K_c/(RT)$
 - $K_p = K_c$
 - $K_p = K_c(RT)^{-\Delta n}$ only
- The value of the equilibrium constant depends on
 - temperature
 - pressure
 - initial concentration
 - catalyst
- If K_c for a reaction is very large, the reaction
 - proceeds almost to completion
 - hardly proceeds
 - does not occur
 - is at equilibrium with equal amounts

1 Mark Questions

- Write the expression for K_c for a general reaction.
- Write the relation between K_p and K_c .
- On what factor does the value of the equilibrium constant depend?

2 Mark Questions

- Derive the relation between K_p and K_c for a gaseous reaction.
- Calculate K_p for the reaction $N_2(g) + 3H_2(g) \rightleftharpoons 2NH_3(g)$ if $K_c = 2.5 \times 10^{-6}$ at 500 K.

3 Mark Questions

- Derive the relation $K_p = K_c(RT)^{\Delta n}$ for a general gaseous reaction and explain each term.
- For the reaction $H_2(g) + I_2(g) \rightleftharpoons 2HI(g)$, write expressions for K_c and K_p , and show that $K_p = K_c$ for this reaction.

Le Chatelier's Principle**MCQ Questions (choose the correct option)**

- Le Chatelier's principle states that a system at equilibrium shifts to
 - oppose the change imposed on it
 - accept the change
 - remain unchanged
 - go to completion
- For an exothermic reaction, an increase in temperature shifts the equilibrium

- (A) backwards
 - (B) forwards
 - (C) no change
 - (D) to completion
3. For a reaction with an increase in the number of moles of gas, an increase in pressure shifts the equilibrium
- (A) backwards
 - (B) forwards
 - (C) no change
 - (D) to zero
4. Adding an inert gas at constant volume to a system at equilibrium
- (A) does not change the equilibrium
 - (B) shifts it forward
 - (C) shifts it backward
 - (D) stops the reaction

1 Mark Questions

1. State Le Chatelier's principle.
2. State the effect of increasing pressure on an equilibrium involving gases.
3. State the effect of a catalyst on equilibrium.

2 Mark Questions

1. State Le Chatelier's principle and explain the effect of a change in temperature on an exothermic equilibrium.
2. State the effect of pressure and concentration changes on the equilibrium of the ammonia synthesis reaction.

3 Mark Questions

1. State Le Chatelier's principle. Predict the effect of changes in concentration, pressure and temperature on the equilibrium $\text{N}_2(\text{g}) + 3\text{H}_2(\text{g}) \rightleftharpoons 2\text{NH}_3(\text{g})$ ($\Delta H = -92 \text{ kJ}$).
2. Apply Le Chatelier's principle to explain the conditions used in the industrial synthesis of ammonia by the Haber process.

Ionic Equilibrium — Acids and Bases

MCQ Questions (choose the correct option)

1. According to Bronsted–Lowry theory, an acid is a
 - (A) proton donor
 - (B) proton acceptor
 - (C) electron pair donor
 - (D) electron pair acceptor
2. According to Lewis theory, a base is a
 - (A) electron pair donor
 - (B) proton donor
 - (C) electron pair acceptor
 - (D) proton acceptor
3. The pH of a neutral solution at 25 °C is
 - (A) 7

- (B) 0
 - (C) 14
 - (D) 1
4. The conjugate base of H_2O is
- (A) OH^-
 - (B) H_3O^+
 - (C) O^{2-}
 - (D) H_2
5. For a strong acid, the degree of ionisation is
- (A) large
 - (B) small
 - (C) zero
 - (D) infinite

1 Mark Questions

1. Define an acid and a base according to the Bronsted–Lowry theory.
2. Define a Lewis base.
3. What is the pH of a neutral solution at 25°C ?

2 Mark Questions

1. Explain the Bronsted–Lowry concept of acids and bases with an example.
2. Define pH and pOH and derive the relation $\text{pH} + \text{pOH} = 14$.

3 Mark Questions

1. Explain the Arrhenius, Bronsted–Lowry and Lewis concepts of acids and bases with examples, and explain conjugate acid–base pairs.
2. Define pH. Derive the relation between pH and pOH and calculate the pH of 0.001 M HCl.

Ionisation Constants — K_a , K_b and Degree of Dissociation

MCQ Questions (choose the correct option)

1. For a weak acid HA, the ionisation constant K_a is
 - (A) $[\text{H}^+][\text{A}^-]/[\text{HA}]$
 - (B) $[\text{HA}]/[\text{H}^+][\text{A}^-]$
 - (C) $[\text{H}^+]/[\text{HA}]$
 - (D) $[\text{A}^-]/[\text{HA}]$
2. For a weak acid, Ostwald's dilution law states that the degree of dissociation is proportional to
 - (A) the square root of the dilution
 - (B) the dilution
 - (C) the pH
 - (D) the temperature
3. As a weak acid is diluted, its degree of dissociation
 - (A) increases
 - (B) decreases
 - (C) remains constant

- (D) becomes zero
4. The relation between K_a and K_b for a conjugate acid-base pair is
- (A) $K_a \times K_b = K_w$
(B) $K_a + K_b = K_w$
(C) $K_a = K_b$
(D) $K_a/K_b = K_w$

1 Mark Questions

1. Write the expression for the ionisation constant K_a of a weak acid.
2. State Ostwald's dilution law.
3. Write the relation between K_a and K_b .

2 Mark Questions

1. Define the ionisation constant of a weak acid and derive the relation $\alpha = \sqrt{(K_a/c)}$.
2. Explain how the degree of dissociation of a weak acid changes on dilution.

3 Mark Questions

1. State Ostwald's dilution law and derive the expression for the degree of dissociation of a weak electrolyte.
2. Derive the relation between K_a , K_b and K_w for a conjugate acid-base pair.

Buffer Solutions and Salt Hydrolysis

MCQ Questions (choose the correct option)

1. A buffer solution resists a change in
 - (A) pH
 - (B) temperature
 - (C) volume
 - (D) pressure
2. An acidic buffer is a mixture of
 - (A) a weak acid and its salt with a strong base
 - (B) a strong acid and its salt
 - (C) two weak acids
 - (D) two strong acids
3. Salt hydrolysis of a salt of a weak acid and a strong base gives a solution that is
 - (A) basic
 - (B) acidic
 - (C) neutral
 - (D) strongly acidic
4. The buffer capacity is maximum when
 - (A) $\text{pH} = \text{p}K_a$
 - (B) $\text{pH} = 7$
 - (C) $\text{pH} = 0$
 - (D) $\text{pH} = 14$

1 Mark Questions

1. Define a buffer solution.
2. Give one example of an acidic buffer.
3. What is salt hydrolysis?

2 Mark Questions

1. Define a buffer solution and explain how an acidic buffer resists a change in pH.
2. Explain salt hydrolysis and state whether the solution is acidic or basic for the four types of salts.

3 Mark Questions

1. Explain the mechanism of buffer action of an acidic buffer and derive the Henderson–Hasselbalch equation.
2. Explain salt hydrolysis for the four types of salts with one example each and state whether the resulting solution is acidic, basic or neutral.

Solubility Product

MCQ Questions (choose the correct option)

1. The solubility product of AgCl is given by
 - (A) $[\text{Ag}^+][\text{Cl}^-]$
 - (B) $[\text{AgCl}]/[\text{Ag}^+][\text{Cl}^-]$
 - (C) $[\text{Ag}^+]/[\text{Cl}^-]$
 - (D) $[\text{AgCl}]$
2. A sparingly soluble salt precipitates when
 - (A) the ionic product exceeds K_{sp}
 - (B) the ionic product is less than K_{sp}
 - (C) the ionic product equals zero
 - (D) the solution is dilute
3. For a salt AB_2 , the relation between solubility s and K_{sp} is
 - (A) $K_{sp} = 4s^3$
 - (B) $K_{sp} = s^2$
 - (C) $K_{sp} = 2s$
 - (D) $K_{sp} = s^3$
4. The common ion effect causes the solubility of a sparingly soluble salt to
 - (A) decrease
 - (B) increase
 - (C) remain constant
 - (D) become zero

1 Mark Questions

1. Define the solubility product.
2. State when a precipitate is formed from a solution.
3. What is the common ion effect?

2 Mark Questions

1. Define the solubility product and derive the relation between solubility and K_{sp} for AB_2 .
2. Explain the common ion effect on the solubility of a sparingly soluble salt.

3 Mark Questions

1. Define the solubility product. Derive the relation between the solubility and K_{sp} for a salt of the type AB and AB_2 .
2. Explain how the solubility product principle is used to predict precipitation and the common ion effect.