CHEMISTRY

Unit 8

CHEMICAL EQUILIBRIUM

MCQs • Short Questions • Long Questions

Complete Study Booklet with Answer Keys

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Section 1 | Multiple Choice Questions

Tick the correct option. Correct answers are highlighted in green with a check mark.

1. For which system, does the equilibrium constant, Kc has units of (concentration)-1?

(a) N2 + 3H2 <=> 2NH3
(b) H2 + I2 <=> 2HI
✓ **(c) 2NO2 <=> N2O4**
(d) 2HF <=> H2 + F2

2. Which statement about the following equilibrium is correct? 2SO3(g) <=> 2SO2(g) + O2(g) Delta H = -197.9 kJ mol-1

✓ **(a) The value of Kp falls with a rise in temperature**
(b) The value of Kp falls with increasing pressure
(c) Adding V2O5 catalyst increase the equilibrium yield of sulphur trioxide
(d) The value of Kp is equal to Kc

3. For the reaction 2SO3(g) <=> 2SO2(g) + O2(g), the Kc expression is:

(a) [SO2]2/[SO3]
✓ **(b) [SO2]2[O2]/[SO3]2**
(c) [SO3]2/[SO3]2[O2]
(d) [SO2][O2]

4. A saturated solution represents a dynamic equilibrium. Macroscopically, the concentration of dissolved solute is constant. Microscopically, this occurs because:

(a) No more solute particles are dissolving
(b) The rate of dissolution of solute is zero
✓ **(c) Solute particles are dissolving and precipitating at the same rate**
(d) All solute particles have dissolved

5. Which of the following statements correctly describes the effect of temperature on the equilibrium constant?

(a) Kc is directly proportional to temperature
(b) Kc is inversely proportional to temperature
✓ **(c) Kc depends on the enthalpy change of the reaction**
(d) Temperature has no effect on the value of Kc

6. Consider the gas-phase equilibrium system: 2H2O(g) <=> 2H2(g) + O2(g). Given that the forward reaction is endothermic, which change will decrease the equilibrium amount of H2O?

(a) Adding more oxygen
(b) Adding a solid phase catalyst
(c) Decreasing the volume
✓ **(d) Increasing the temperature**

7. Kc = 0.040 at 450 deg C for PCl5(g) <=> PCl3(g) + Cl2(g). Evaluate Kp.

(a) 0.40
(b) 0.64
✓ **(c) 2.4**
(d) 0.052

8. In which gaseous equilibrium, pressure has no effect on the equilibrium position?

(a) 2NO2(g) <=> N2O4(g)
(b) PCl5(g) <=> PCl3(g) + Cl2(g)

✓ (c) CO(g) + H2O(g) <=> CO2(g) + H2(g)
(d) 2SO2(g) + O2(g) <=> 2SO3(g)

9. Consider the equilibrium 2H2(g) + O2(g) <=> 2H2O(g). If the concentration of H2O(g) is increased, the concentrations of H2(g) and O2(g) will:

(a) Increase
✓ (b) Decrease
(c) Remain the same
(d) Change irregularly

10. For a specific reaction, the value of Kc:

(a) Always remains the same at different reaction conditions
(b) Increases if the concentration of one of the product is increased
✓ (c) Changes with changes in the temperature
(d) Increases if the concentration of one of the reactants is increased

11. Which one is not an example of reversible reaction?

(a) Formation of ammonia
✓ (b) Formation of water
(c) Decomposition of PCl5
(d) Decomposition of NO2

12. Which one is a microscopic event?

(a) Formation of precipitate
(b) Evolution of gas
(c) Evolution of heat
✓ (d) Formation of bond

13. The reversible reaction cannot be achieved in:

✓ (a) Open system
(b) Closed system
(c) Both (a) and (b)
(d) None of these

14. A reversible reaction is one which:

(a) Proceeds to completion
(b) Occurs only in one direction
✓ (c) Proceeds in both directions
(d) Has no products

15. Which is an irreversible reaction?

(a) Haber process
✓ (b) Precipitation of AgCl
(c) Synthesis of ammonia
(d) Esterification

16. Which of the following reactions reaches equilibrium?

(a) Irreversible
✓ (b) Reversible
(c) Combustion
(d) Neutralization

17. In a reversible reaction:

(a) Products do not reform reactants
(b) Rate of forward reaction is always higher

✓ (c) Both forward and reverse reactions occur
(d) Products are in excess

18. Reversible reactions are indicated by:
(a) ->
(b) ←
✓ (c) <=>
(d) ≠

19. Combustion of methane is:
(a) Reversible
(b) Endothermic
✓ (c) Irreversible
(d) Equilibrium process

20. Dynamic equilibrium is attained when:
(a) Forward reaction stops
(b) Reverse reaction stops
✓ (c) Forward and reverse reactions continue at equal rates
(d) Concentration of products becomes zero

21. Which statement is true at dynamic equilibrium?
(a) No reaction is occurring
(b) Concentrations are changing
✓ (c) Rates of forward and reverse reactions are equal
(d) Rate of forward reaction > reverse

22. At equilibrium, the observable properties:
(a) Keep changing
(b) Fluctuate randomly
✓ (c) Remain constant
(d) Oscillate

23. Which is NOT a feature of dynamic equilibrium?
(a) Closed system
(b) Constant temperature
✓ (c) Unequal reaction rates
(d) No net change

24. Dynamic equilibrium occurs:
(a) In open systems
(b) Only in gases
✓ (c) In closed systems
(d) Only at low temperature

25. At equilibrium:
(a) Products dominate
(b) Reactants dominate
(c) Forward and reverse rates are zero
✓ (d) Forward rate = reverse rate

26. Law of mass action was proposed by:
(a) Le Chatelier
(b) Arrhenius
✓ (c) Guldberg and Waage

(d) Dalton

27. Law of mass action is applicable to:

✓ **(a) Reversible reactions**

(b) Combustion reactions

(c) Irreversible reactions

(d) Endothermic reactions only

28. According to law of mass action, rate of reaction is proportional to:

(a) Temperature

(b) Pressure

✓ **(c) Product of active masses**

(d) Atomic mass

29. Active mass means:

(a) Moles

(b) Volume

(c) Mass

✓ **(d) Molar concentration**

30. Law of mass action applies to:

✓ **(a) Closed systems only**

(b) Open systems

(c) Irreversible systems

(d) Combustion reactions

31. For reaction: aA + bB <=> cC + dD, the equilibrium constant expression is:

(a) [A]^a[B]^b/[C]^c[D]^d

✓ **(b) [C]^c[D]^d/[A]^a[B]^b**

(c) [C][D]/[A][B]

(d) [A]+[B] = [C]+[D]

32. Kc is expressed in terms of:

(a) Pressure

(b) Mole fraction

✓ **(c) Concentration**

(d) Volume

33. Equilibrium constant depends on:

(a) Pressure

✓ **(b) Temperature**

(c) Volume

(d) Concentration

34. If Kc > 1, the reaction:

✓ **(a) Favors products**

(b) Favors reactants

(c) Is at equilibrium

(d) Does not occur

35. Kp is used for:

(a) Solid-state reactions

(b) Aqueous reactions

✓ **(c) Gaseous reactions**

(d) Liquid reactions

36. Kp = Kc(RT)^Delta n is used to relate:

(a) Temperature and pressure

✓ **(b) Kp and Kc**

(c) Energy and volume

(d) Gibbs free energy

37. For reaction: N2 + 3H2 <=> 2NH3, Delta n = ?

(a) 2

✓ **(b) -2**

(c) -1

(d) +2

38. If Q = K, then:

(a) Reaction is irreversible

✓ **(b) System is at equilibrium**

(c) Forward reaction dominates

(d) Reverse reaction dominates

39. If Q > K, the reaction:

(a) Moves forward

(b) At equilibrium

✓ **(c) Shifts left**

(d) Stops

40. Reaction quotient Q is calculated using:

✓ **(a) Initial concentrations**

(b) Equilibrium concentrations

(c) Temperature

(d) Catalyst

41. A large K value indicates:

✓ **(a) Products are favored**

(b) Reactants are favored

(c) No reaction

(d) Slow reaction

42. Q < K implies:

✓ **(a) Reaction proceeds forward**

(b) Equilibrium is established

(c) System stops

(d) Reaction shifts in reverse

43. Position of equilibrium is affected by:

(a) Catalyst

✓ **(b) Temperature**

(c) Inert gas

(d) Surface area

44. Le Chatelier's Principle applies to:

(a) Irreversible reactions

(b) Static equilibrium

✓ **(c) Dynamic equilibrium**

(d) Precipitation reactions

45. Increasing temperature favors:

(a) Exothermic reaction
✓ **(b) Endothermic reaction**
(c) Formation of solid
(d) Reverse in all cases

46. Decreasing pressure shifts equilibrium towards:
✓ **(a) Side with more gas molecules**
(b) Side with fewer gas molecules
(c) Liquids
(d) Solids

47. Addition of inert gas at constant volume:
(a) Affects equilibrium
(b) Shifts reaction left
(c) Shifts reaction right
✓ **(d) No effect**

48. Catalyst affects:
(a) Value of K
(b) Equilibrium position
✓ **(c) Activation energy**
(d) Final concentrations

49. Removing product from equilibrium:
(a) Shifts equilibrium left
(b) Stops reaction
✓ **(c) Shifts equilibrium right**
(d) Has no effect

50. Increase in concentration of reactants:
(a) Increases K
(b) Decreases K
✓ **(c) Shifts equilibrium forward**
(d) Stops reverse reaction

51. Increase in pressure shifts equilibrium to:
(a) Side with more moles of gas
✓ **(b) Side with fewer moles of gas**
(c) Liquid phase
(d) Solid phase

52. Lowering temperature in an exothermic reaction:
(a) Favors reverse
✓ **(b) Favors forward**
(c) No effect
(d) Stops the reaction

53. Catalyst affects:
✓ **(a) Activation energy**
(b) Equilibrium position
(c) Kp value
(d) Enthalpy

54. Which does not alter equilibrium constant?
(a) Catalyst

(b) Temperature
(c) Pressure
✓ **(d) Both a and c**

55. Removing a reactant shifts equilibrium to:
(a) Right
✓ **(b) Left**
(c) No shift
(d) Depends on temperature

56. Catalyst used in Haber process is:
✓ **(a) Fe**
(b) Pt
(c) Cu
(d) Zn

57. Reaction in Haber process is:
(a) Endothermic
✓ **(b) Exothermic**
(c) Irreversible
(d) Neutral

58. Pressure used in Haber process:
(a) 2 atm
(b) 5 atm
✓ **(c) 200 atm**
(d) 760 mmHg

59. Contact process is used for:
(a) Ammonia synthesis
✓ **(b) Sulfuric acid production**
(c) Nitric acid production
(d) Hydrogenation

60. Catalyst used in contact process:
(a) Fe
✓ **(b) V2O5**
(c) Ni
(d) Al2O3

61. Optimum temperature in Haber process is:
(a) 50 deg C
(b) 200 deg C
✓ **(c) 450 deg C**
(d) 1000 deg C

62. The relationship Kp = Kc(RT)^Delta n is valid for:
(a) Liquid reactions
(b) Aqueous reactions
✓ **(c) Gaseous reactions**
(d) Precipitation

63. For Delta n = 0, Kp =
✓ **(a) Kc**
(b) 1/Kc

(c) Infinite
(d) Zero

64. In the equation Kp = Kc(RT)^Delta n, R is:
(a) 1.38 x 10-23
(b) 8.31 J/mol·K
(c) 0.0821 L·atm/mol·K
✓ **(d) Either b or c depending on units**

65. If Delta n is positive, then Kp is:
✓ **(a) Greater than Kc**
(b) Less than Kc
(c) Equal to Kc
(d) Zero

66. In Kp = Kc(RT)^Delta n, Delta n is:
(a) Change in concentration
✓ **(b) Change in number of moles of gas**
(c) Change in pressure
(d) Change in entropy

67. Units of Kc depend on:
✓ **(a) Reaction stoichiometry**
(b) Catalyst
(c) Activation energy
(d) Delta H

68. NH4Cl in water makes solution:
(a) Neutral
(b) Basic
✓ **(c) Acidic**
(d) Amphoteric

69. Na2CO3 in water gives:
(a) Acidic solution
✓ **(b) Basic solution**
(c) Neutral solution
(d) Buffer

70. Salts of weak acid + strong base are:
(a) Neutral
(b) Acidic
✓ **(c) Basic**
(d) Amphoteric

71. CH3COONa in water forms:
(a) Acidic solution
✓ **(b) Basic solution**
(c) Neutral solution
(d) Salt bridge

72. Strong acid + strong base forms a:
(a) Acidic salt
(b) Basic salt
✓ **(c) Neutral salt**

(d) Buffer

73. Which undergoes hydrolysis?

(a) NaCl

(b) KNO3

✓ **(c) NH4Cl**

(d) NaBr

Section 2 | Short Questions

Q1. What is meant by the state of chemical equilibrium?

It is the state in a reversible reaction when the rates of forward and reverse reactions become equal and the concentrations of reactants and products remain constant.

Q2. Define reversible reaction. Give an example.

A reversible reaction is one that proceeds in both forward and reverse directions.

Example: N2(g) + 3H2(g) <=> 2NH3(g)

Q3. Why does a change of volume disturb the equilibrium position for some gas-phase reactions but not the equilibrium constant?

Change in volume affects the equilibrium position, but not the equilibrium constant, which only changes with temperature.

Q4. Mention the characteristics of chemical equilibrium.

Forward and reverse reaction rates are equal; concentrations remain constant; equilibrium can be reached from either side; the system is dynamic in nature.

Q5. Reversible reactions attain a position of equilibrium that is dynamic, not static. Explain.

Reactants and products continuously convert into each other at equal rates; the system remains active at the microscopic level even though no visible (macroscopic) change occurs.

Q6. Why do the rates of forward reactions slow down as a reversible reaction approaches equilibrium?

Because the concentration of reactants decreases as the reaction proceeds, lowering the rate of the forward reaction until it equals the reverse reaction rate.

Q7. Why can ice at 0 deg C be melted by applying pressure without supplying heat from outside?

Because ice has a lower density than water; applying pressure lowers the melting point of ice, causing it to melt even without added heat.

Q8. Write two conditions of the equilibrium constant.

Temperature must remain constant.

It applies only to a system in dynamic equilibrium.

Q10. What is chemical equilibrium?

Chemical equilibrium is a state in a reversible reaction where the rate of the forward reaction equals the rate of the backward reaction, and the concentrations of reactants and products remain constant.

Example: N2(g) + 3H2(g) <=> 2NH3(g)

Q11. What is a reversible reaction? Give an example.

A reversible reaction is one in which the products can react to re-form the original reactants.

Example: H2(g) + I2(g) <=> 2HI(g)

Q12. Why is chemical equilibrium called dynamic equilibrium?

Because at equilibrium, both the forward and reverse reactions continue to occur, but at equal rates, resulting in no net change in the concentrations of reactants and products.

Q13. Does equilibrium mean equal amounts of reactants and products? Explain.

No. Equilibrium means equal reaction rates, not equal concentrations. The amounts present depend on the reaction conditions and the value of the equilibrium constant.

Q14. State the Law of Mass Action.

At constant temperature, the rate of a chemical reaction is directly proportional to the product of the molar concentrations of the reactants, each raised to the power of its stoichiometric coefficient.

Q15. How is the equilibrium constant expression written for a reaction?

For aA + bB <=> cC + dD: Kc = [C]^c[D]^d / [A]^a[B]^b

Q16. What is Kc?

Kc is the equilibrium constant expressed in terms of concentration - the ratio of product concentrations to reactant concentrations at equilibrium.

Q17. What does the value of Kc indicate about a reaction?

A large Kc (>1) indicates products are favored at equilibrium; a small Kc (<1) indicates reactants are favored; Kc near 1 indicates comparable amounts of both.

Q18. Is Kc affected by pressure, volume, or concentration?

No. Kc depends only on temperature. Changing pressure, volume, or concentration affects the equilibrium position, not the value of Kc.

Q19. Differentiate between homogeneous and heterogeneous equilibrium.

Homogeneous equilibrium: all reactants and products are in the same phase, e.g. H2(g)+I2(g) <=> 2HI(g).

Heterogeneous equilibrium: reactants and products are in different phases, e.g. CaCO3(s) <=> CaO(s)+CO2(g).

Q20. State Le Chatelier's Principle.

If a system at equilibrium is disturbed by a change in temperature, pressure, or concentration, the system shifts its equilibrium position to oppose that change.

Q21. How does an increase in concentration affect equilibrium?

The system shifts to consume the added species.

Example: adding more H2 to H2(g)+I2(g) <=> 2HI(g) shifts the equilibrium to the right.

Q22. What is the effect of pressure on equilibrium in gaseous reactions?

Increasing pressure favors the side with fewer gas molecules; decreasing pressure favors the side with more.

Example: N2(g)+3H2(g) <=> 2NH3(g) - increased pressure shifts the equilibrium to the right.

Q23. How does temperature affect equilibrium?

For exothermic reactions, increasing temperature shifts equilibrium to the left.

For endothermic reactions, increasing temperature shifts equilibrium to the right.

Q24. Why is high pressure used in the Haber process?

Because the forward reaction (formation of NH3) produces fewer gas molecules than it consumes, so high pressure favors ammonia production.

Q25. What is the compromise temperature used in the Haber process?

About 450-500 deg C is used to balance rate and yield, since higher temperature decreases yield (the reaction is exothermic) but increases the reaction rate.

Q26. What is the role of temperature in an exothermic reversible reaction?

Increasing temperature favors the reverse reaction; decreasing temperature favors the forward reaction. In both cases, Kc changes.

Q27. What is the effect of a change of catalyst on equilibrium?

A change of catalyst does not affect or disturb the equilibrium - a catalyst only helps the system reach equilibrium faster by equally increasing the rates of the forward and reverse reactions.

Q28. How does temperature affect an endothermic reversible reaction?

Increasing temperature favors the forward direction; decreasing temperature favors the reverse direction. In both cases, Kc changes.

Q29. What is the unit of Kc?

For aA+bB <=> cC+dD: unit of Kc = (mol dm-3)^[(c+d)-(a+b)]

Q30. What does a small Kc value mean?

The reaction has very low product concentration at equilibrium - the reaction does not proceed much in the forward direction.

Q31. Why are solids and liquids not included in the Kc expression?

Because the concentrations of pure solids and liquids remain constant and do not affect the equilibrium.

Q32. How does a catalyst affect equilibrium?

A catalyst increases the rate of both the forward and reverse reactions equally, but does not change the position of equilibrium or the value of Kc.

Q33. Write the equilibrium expression for: 2SO2(g) + O2(g) <=> 2SO3(g)

Kc = [SO3]^2 / [SO2]^2[O2]

Q34. What will happen if CO2 is removed from: CaCO3(s) <=> CaO(s) + CO2(g)?

The system will shift to the right to replace the lost CO2, favouring further decomposition of CaCO3.

Q35. Why is equilibrium important in industrial processes?

Equilibrium controls the yield of products; by adjusting conditions like pressure and temperature, industries can maximize product formation (e.g. ammonia, sulfuric acid).

Q36. How can you increase the yield of an endothermic reaction?

By increasing the temperature, since heat effectively acts like a reactant and the equilibrium shifts to the right (forward direction).

Q37. Define equilibrium mixture.

A mixture of reactants and products present together at the equilibrium state.

Q38. Define equilibrium position.

The relative concentrations of reactants and products at equilibrium, indicating which side (reactants or products) is favored.

Q39. Define the reaction quotient (Q).

Q is calculated using the same expression as Kc, but for a system not yet at equilibrium; it helps predict the direction in which the reaction will proceed to reach equilibrium.

Section 3 | Descriptive & Numerical Problems

Descriptive Questions

Q5 (Descriptive). For PCl5(g) <=> PCl3(g) + Cl2(g), Delta H = +90 kJ/mol, what is the effect of: (i) increasing temperature, (ii) decreasing volume, (iii) adding a catalyst, (iv) adding chlorine?

(i) Temperature increased: the forward dissociation is endothermic, so increasing temperature shifts equilibrium to the right (more PCl3 and Cl2); Kc increases.

(ii) Volume decreased: moles are 1 (reactant) vs 2 (products); higher pressure shifts equilibrium left (towards fewer moles), forming more PCl5; Kc is unchanged.

(iii) Catalyst added: no change in equilibrium composition; Kc is unchanged.

(iv) Chlorine added: adding a product shifts equilibrium left to consume some Cl2, forming more PCl5; Kc is unchanged.

Q6 (Descriptive). Haber's process, N2(g)+3H2(g) <=> 2NH3(g), Delta H = -92.46 kJ, is exothermic. What is the effect of a change in temperature at the equilibrium stage?

(i) Temperature increased: heat behaves like a product in an exothermic reaction, so increasing temperature shifts equilibrium left (reverse direction) to absorb the excess heat - this decreases NH3 yield, and Kc decreases.

(ii) Temperature decreased: lower temperature favors the exothermic (forward) direction, shifting equilibrium right and producing more NH3; Kc increases.

Optimum conditions: too low a temperature gives high yield but a very slow rate, so an optimum compromise temperature of about 450 deg C is used to balance rate and yield.

Numerical Problems

Q9 (Numerical). 2SO2(g)+O2(g) <=> 2SO3(g) reaches equilibrium. Initial [SO2]=0.060 mol/dm3, [O2]=0.050 mol/dm3. At equilibrium, [SO3]=0.040 mol/dm3. Find equilibrium [O2].

Change in [SO3] = +0.040 mol/dm3; Change in [SO2] = -0.040 mol/dm3; Change in [O2] = -1/2 x 0.040 = -0.020 mol/dm3

[SO2]eq = 0.060-0.040 = 0.020 mol/dm3; [O2]eq = 0.050-0.020 = 0.030 mol/dm3

Answer: [O2] = 0.030 mol/dm3

Q7 (Numerical). Kc = 0.016 at 520 deg C for 2HI(g) <=> H2(g) + I2(g). Equilibrium mixture has HI=0.08 M, H2=0.01 M, I2=0.01 M. More HI is added to reach 0.096 M. Find new equilibrium concentrations.

Kc expression: Kc = [H2][I2]/[HI]^2 = 0.016

(0.010+x)(0.010+x)/(0.096-2x)^2 = 0.016; taking square root: (0.010+x)/(0.096-2x) = sqrt(0.016) = 0.1265

0.010+x = 0.1265(0.096-2x) = 0.012143 - 0.252982x; 1.252982x = 0.002143; x is approx 0.001704 M

[HI]eq = 0.096-2(0.001704) = 0.09258 M; [H2]eq = [I2]eq = 0.010+0.001704 = 0.01171 M

Answer: [HI] = 0.09258 M, [H2] = [I2] = 0.01171 M

Q8 (Numerical). Kc for acetic acid + ethanol <=> ethyl acetate + water is 4. A mixture of 3 mol acetic acid and 1 mol ethanol reaches equilibrium. Find moles of ethyl acetate formed.

Kc expression: x^2 / [(3-x)(1-x)] = 4

x^2 = 4(3-x)(1-x) = 4(3-4x+x^2) = 12-16x+4x^2

Rearranging: 3x^2 - 16x + 12 = 0

Quadratic formula: x = [16 +/- sqrt(256-144)]/6 = [16 +/- sqrt(112)]/6 = [16 +/- 10.58]/6, giving x = 4.43 or x = 0.903

Since x cannot exceed 1 (limited by ethanol available), x = 0.903

Answer: amount of ethyl acetate at equilibrium = 0.903 mol

Section 4 | Long Questions

Q1. Define and explain: Reversible Reactions, Irreversible Reactions, Macroscopic Events, Microscopic Events.

(i) Reversible Reactions

Definition: reactions in which reactants convert into products and products convert back into reactants, proceeding in both directions without a net change in the concentrations of reactants and products.

Examples: N2(g)+3H2(g) <=> 2NH3(g); H2(g)+I2(g) <=> 2HI(g)

Represented by a double arrow (<=>); proceed in both forward and reverse directions; do not go to completion but reach a state of balance; occur in a closed system (no substance enters or leaves); concentrations remain constant; forward rate equals reverse rate.

(ii) Irreversible Reactions

Definition: reactions in which the reactants are completely consumed and converted into products.

Examples: CH4+2O2 -> CO2+2H2O+Heat (combustion of methane); 2H2+O2 -> 2H2O (formation of water)

Represented by a single arrow (->); reactants are completely converted to products; the reaction stops once the limiting reactant is used up.

(iii) Macroscopic Events

Definition: phenomena observable with the naked eye, without considering individual particles or molecules.

Examples: colour change (e.g. rusting of iron), evolution or absorption of heat, formation of a precipitate, gas evolution (bubbles/fizzing), change in volume or pressure, change in composition.

(iv) Microscopic Events

Definition: phenomena that cannot be observed with the naked eye.

Examples: molecular collisions, breaking and forming of chemical bonds, rearrangement of atoms, electron transfer (in redox reactions), changes in molecular structure or geometry.

Q2. Why is equilibrium in a reversible reaction considered dynamic and not static?

Dynamic Equilibrium

Definition: the state in which the rate of the forward reaction and the rate of the reverse reaction proceed continuously in both directions at similar (equal) rates.

Dynamic Nature of Chemical Equilibrium

When a reversible reaction reaches equilibrium, it may appear static macroscopically (concentrations no longer visibly change), but at the microscopic level the system is still highly active:

Reactant molecules continue to collide and form products; product molecules simultaneously revert back to reactants; the forward and reverse reactions occur at equal rates, so there is no net change in concentration.

Example: Reaction of Steam and Carbon Monoxide

Forward: $H_2O(g)+CO(g) \rightarrow H_2(g)+CO_2(g)$. At the start, collisions and forward rate are maximal due to high reactant concentration; as reactants are consumed, forward rate slows.

Reverse: $H_2(g)+CO_2(g) \rightarrow H_2O(g)+CO(g)$. As products accumulate, collisions between H_2 and CO_2 molecules become more frequent, breaking and re-forming bonds to regenerate CO and H_2O.

Eventually both reactions proceed at the same rate - the system is then said to be in chemical equilibrium.

Q3. Write a note on dynamic equilibrium between two physical states.

Dynamic Equilibrium in Phase Changes

A reversible phase change shows dynamic equilibrium between two physical states of a substance.

Example: at 0 deg C, ice and water coexist in equilibrium - water freezes into ice and ice melts into water at the same rate. No net change is observed in the amount of each phase, and the macroscopic properties of the system remain constant over time.

Dynamic Equilibrium and Vapour Pressure

(i) Evaporation and Condensation in a Closed System: Liquid <=> Gas. Some molecules escape into the gas phase (evaporation); gas molecules collide with the liquid surface and return (condensation).

(ii) Establishing Dynamic Equilibrium: a point is reached where the rate of evaporation equals the rate of condensation - this is dynamic equilibrium, and the vapour pressure remains constant.

(iii) Effect of Temperature: increasing temperature raises the average kinetic energy of liquid molecules, giving faster evaporation and higher vapour pressure; a new dynamic equilibrium is established at the higher temperature.

Quick Check 8.2

(a) Dynamic equilibrium exists between water and its vapour at 100 deg C: water boils, so liquid and vapour coexist; molecules evaporate and condense continuously; at equilibrium, evaporation rate equals condensation rate, so the amounts of liquid and vapour stay constant even though the process is dynamic.

(b) Yes, dynamic equilibrium also exists between ice and water at 0 deg C: ice melts to water and water freezes back to ice simultaneously; the rates of melting and freezing become equal at equilibrium, keeping the amounts of ice and water constant.

Q4. What are the conditions for chemical equilibrium?

Conditions of Chemical Equilibrium

(i) Reversibility of Reaction: equilibrium occurs only in reversible reactions - the reaction must proceed in both forward and reverse directions.

(ii) Closed System: the reaction must take place in a closed vessel, with no reactants or products allowed to escape from the system.

Q5. What are the characteristics of chemical equilibrium?

Characteristics of Chemical Equilibrium

Concentration: at equilibrium, the concentrations of reactants and products remain constant.

Starting Point: the equilibrium state can be approached from either side (starting from reactants or products).

Role of Catalyst: a catalyst does not change the equilibrium position or the equilibrium constant - it only helps the system reach equilibrium earlier.

Dependence of Kc on Temperature: the value of the equilibrium constant does not depend on the initial concentrations of reactants; it depends only on temperature.

Effect of Pure Solids/Liquids on Kc

If pure solids or pure liquids are involved in an equilibrium system, their concentrations are not included in the equilibrium constant expression, because their concentration does not change and so has no effect on the equilibrium constant.

Q6. Derive the equilibrium constant expression.

Law of Mass Action

Statement: the rate of a chemical reaction is directly proportional to the product of the concentrations of the reactants, each raised to a power equal to its coefficient in the balanced equation.

Derivation of Kc

For A+B <=> C+D: Forward rate = kf[A][B]; Reverse rate = kr[C][D].

At equilibrium, forward rate = reverse rate: kf[A][B] = kr[C][D], so kf/kr = [C][D]/[A][B].

Since kf/kr is constant (at constant temperature): Kc = [C][D]/[A][B]

General reaction aA+bB <=> cC+dD: Kc = [C]^c[D]^d / [A]^a[B]^b

Example Expressions

Synthesis of Ammonia: N2(g)+3H2(g) <=> 2NH3(g), Kc = [NH3]^2/[N2][H2]^3

Decomposition of N2O5: 2N2O5(g) <=> 4NO2(g)+O2(g), Kc = [NO2]^4[O2]/[N2O5]^2

Quick Check 8.4

(i) Sn2+(aq)+2Fe3+(aq) <=> Sn4+(aq)+2Fe2+(aq): Kc = [Sn4+][Fe2+]^2 / [Sn2+][Fe3+]^2

(ii) 4NH3(g)+5O2(g) <=> 4NO(g)+6H2O(g): Kc = [NO]^4[H2O]^6 / [NH3]^4[O2]^5

Q7. What are the units of the equilibrium constant?

When Delta n = 0 (Equal Moles)

Units cancel out and Kc is dimensionless (no units).

Example: H2+I2 <=> 2HI, Kc = [HI]^2/[H2][I2] = (mol dm-3)^2/(mol dm-3)(mol dm-3) = no units

When Delta n is not 0 (Unequal Moles)

The units of Kc depend on the units of concentration used.

Example: N2(g)+3H2(g) <=> 2NH3(g), Kc = [NH3]^2/[N2][H2]^3 = mol-2 dm6

Q8. Explain the relationships between the various equilibrium constants.

Types of Equilibrium Constants

Kc: based on concentrations (mol dm-3) - Kc = [C]^c[D]^d/[A]^a[B]^b

Kp: based on partial pressures - Kp = P(C)^c x P(D)^d / P(A)^a x P(B)^b

Kn: based on number of moles - Kn = n(C)^c x n(D)^d / n(A)^a x n(B)^b

Kx: based on mole fractions - Kx = X(C)^c x X(D)^d / X(A)^a x X(B)^b

Relationships

Kp = Kc(RT)^Delta n; Kp = Kx(P)^Delta n; Kp = Kn(N)^Delta n

Where R = gas constant, T = absolute temperature, P = system pressure, N = total moles, Delta n = (c+d)-(a+b).

Position of Equilibrium

[Products] > [Reactants]: equilibrium shifted to the right.

[Reactants] > [Products]: equilibrium shifted to the left.

Quick Check 8.5

(i) Ammonia synthesis: N2(g)+3H2(g) <=> 2NH3(g); Delta n = 2-(1+3) = -2; Kp = Kc(RT)^-2, so Kp < Kc.

(ii) Dissociation of PCl5: PCl5(g) <=> PCl3(g)+Cl2(g); Delta n = 2-1 = +1; Kp = Kc(RT)^+1, so Kp > Kc.

Q9. State Le Chatelier's Principle and discuss its applications.

Le Chatelier's Principle

Statement: 'If a system at equilibrium is disturbed by changing temperature, pressure, or concentration, the system shifts its equilibrium position to oppose the change.'

1. Effect of Change in Concentration

Example - Hydrolysis of BiCl3: BiCl3+H2O <=> BiOCl+2HCl. Adding BiCl3 (reactant) shifts equilibrium right; adding BiOCl or HCl (products) shifts it left. Removing BiCl3 shifts left; removing BiOCl or HCl shifts right.

2. Effect of Change in Pressure or Volume

Example - Synthesis of Ammonia: N2+3H2 <=> 2NH3; total moles of reactants = 4, products = 2. Increasing pressure/decreasing volume shifts equilibrium toward fewer moles (forward); decreasing pressure/increasing volume shifts it toward more moles (reverse).

3. Effect of Change in Temperature

Exothermic reactions (Delta H negative), e.g. 2SO2(g)+O2(g) <=> 2SO3(g), Delta H = -198 kJ: increasing temperature shifts equilibrium left (reverse) and decreases Kc; decreasing temperature shifts it right (forward) and increases Kc.

Endothermic reactions (Delta H positive), e.g. N2O4(g) <=> 2NO2(g), Delta H = +57.2 kJ: increasing temperature shifts equilibrium right (forward) and increases Kc; decreasing temperature shifts it left (reverse) and decreases Kc.

4. Effect of a Catalyst on Equilibrium

A catalyst does not change the position of equilibrium and does not affect Kc - it only speeds up the rate at which equilibrium is achieved, so the yield stays the same but is reached faster.

Q10. Discuss the industrial applications of chemical equilibrium.

Haber's Process (Synthesis of Ammonia)

N2(g)+3H2(g) <=> 2NH3(g), Delta H = -46.11 kJ/mol

Optimum conditions: Temperature 450-500 deg C (a compromise - high temperature gives faster rate but lower yield since the reaction is exothermic; low temperature gives higher yield but slower rate); Pressure 200-300 atm (high, since the forward reaction has fewer moles of gas); Catalyst: iron (Fe) with promoters, which speeds up the reaction without affecting the equilibrium position; unreacted gases are recycled to increase overall efficiency and yield.

Contact Process (Synthesis of SO3)

2SO2(g)+O2(g) <=> 2SO3(g), Delta H = -198 kJ

Conditions: Temperature 400-450 deg C (the reaction is exothermic, so lower temperature favours products); Pressure 1-2 atm (no significant effect on this equilibrium, though increasing pressure does shift it forward); Catalyst: V2O5 (vanadium pentoxide), which speeds up the reaction.

Sample Problem 8.1

CH3COOH(l)+C2H5OH(l) <=> CH3COOC2H5(l)+H2O(l). 500 cm3 of equilibrium mixture contains 0.235 mol acetic acid, 0.0350 mol ethanol, 0.182 mol ethyl ethanoate, and 0.182 mol water. Find Kc.

Concentrations: [CH3COOH]=0.235/0.500=0.470, [C2H5OH]=0.0350/0.500=0.070, [Ester]=[H2O]=0.182/0.500=0.364 (mol dm-3).

Kc = (0.364)(0.364)/[(0.470)(0.070)] = 0.132496/0.0329 = 4.03 (no units)

Sample Problem 8.2

CH3COCH3(l)+HCN(g) <=> CH3COH(CN)CH3(l). Starting concentrations 0.0500 mol dm-3 each; at equilibrium, product = 0.0233 mol dm-3. Find Kc.

[CH3COCH3]=[HCN]=0.0500-0.0233=0.0267 mol dm-3.

Kc = 0.0233/[(0.0267)(0.0267)] = 0.0233/0.000712 = 32.7 dm3 mol-1

Sample Problem 8.3

N2(g)+3H2(g) <=> 2NH3(g). Kc = 6.0x10^-2 at 500 deg C. Find Kp.

Delta n = 2-4 = -2; T = 773 K; R = 0.0821 atm dm3/mol/K.

Kp = Kc(RT)^Delta n = 6.0x10^-2 (0.0821x773)^-2 = 6.0x10^-2 (63.5)^-2 = 6.0x10^-2/4032.25 = 1.49x10^-5

Sample Problem 8.4

2SO2(g)+O2(g) <=> 2SO3(g). Equilibrium partial pressures: SO2=1.0x10^6 Pa, O2=7.0x10^6 Pa, SO3=8.0x10^6 Pa. Find Kp.

Kp = P(SO3)^2 / [P(SO2)^2 x P(O2)] = (8.0x10^6)^2 / [(1.0x10^6)^2 x (7.0x10^6)]

Kp = 6.4x10^13 / 7.0x10^18 = 9.1x10^-6 Pa-1

Section 5 | Answer Key & Formula Summary

MCQ Answer Key (Q1-Q73)

Q.No	Ans	Q.No	Ans	Q.No	Ans	Q.No	Ans
1	C	20	C	39	C	58	C
2	A	21	C	40	A	59	B
3	B	22	C	41	A	60	B
4	C	23	C	42	A	61	C
5	C	24	C	43	B	62	C
6	D	25	D	44	C	63	A
7	C	26	C	45	B	64	D
8	C	27	A	46	A	65	A
9	B	28	C	47	D	66	B
10	C	29	D	48	C	67	A
11	B	30	A	49	C	68	C
12	D	31	B	50	C	69	B
13	A	32	C	51	B	70	C
14	C	33	B	52	B	71	B
15	B	34	A	53	A	72	C
16	B	35	C	54	D	73	C
17	C	36	B	55	B		
18	C	37	B	56	A		
19	C	38	B	57	B		

Important Formulas

Formula	Description
Kc = [C]^c[D]^d / [A]^a[B]^b	Equilibrium Constant (concentrations) for aA+bB<=>cC+dD
Kp = P(C)^c x P(D)^d / P(A)^a x P(B)^b	Equilibrium Constant (partial pressures)
Kp = Kc(RT)^Delta n	Relating Kp and Kc
Kp = Kx(P)^Delta n	Relating Kp and Kx (mole fraction constant)
Kp = Kn(N)^Delta n	Relating Kp and Kn (moles constant)
Delta n = (c+d) - (a+b)	Change in moles of gas, products - reactants
Unit of Kc = (mol dm-3)^Delta n	Units of the equilibrium constant

Key Constants & Notes

Constant/Note	Value

Kc greater than 1	Products are favored (reaction goes nearly to completion)
Kc less than 1	Reactants are favored (reaction hardly proceeds)
Kc around 1	Significant amounts of both reactants and products present
Gas Constant (R)	0.0821 atm dm3 mol-1 K-1
Haber Process Temperature (compromise)	450-500 deg C
Haber Process Pressure	200-300 atm
Contact Process Temperature	400-450 deg C
Contact Process Pressure	1-2 atm

What the Value of Kc Indicates

Kc Value	Indication
Kc > 1	Products are favored (reaction goes nearly to completion)
Kc < 1	Reactants are favored (reaction hardly proceeds)
Kc = 1	Significant amounts of both reactants and products present

Homogeneous vs Heterogeneous Equilibrium

Type	Description
Homogeneous Equilibrium	All reactants and products are in the same phase. Example: H2(g)+I2(g) <=> 2HI(g)
Heterogeneous Equilibrium	Reactants and products are in different phases. Example: CaCO3(s) <=> CaO(s)+CO2(g)

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