

Class: 12th

Subject: Chemistry

**Unit 8: CHEMICAL
EQUILIBRIUM**

❖ Important MCQs:

1. Which statement best describes an irreversible reaction?

- (a) Reaction proceeds in both directions
- (b) Reactants are partially consumed
- (c) Reactants are completely converted into products ✓
- (d) Reaction never stops

2. A reaction stops when:

- (a) Products are formed
- (b) Temperature increases
- (c) Limiting reactant is completely consumed ✓
- (d) Catalyst is added

3. In a reversible reaction at equilibrium, the concentration of reactants and products:

- (a) Becomes zero
- (b) Continuously changes
- (c) Remains constant ✓
- (d) Becomes equal

4. Chemical equilibrium is achieved when:

-
- (a) Reaction stops completely
 - (b) Only forward reaction occurs
 - (c) Rate of forward and reverse reactions become equal
 - (d) All reactants are consumed

5. Which condition is necessary to achieve chemical equilibrium?

- (a) Open container
- (b) Closed system
- (c) High temperature only
- (d) Presence of catalyst

6. Which of the following is a macroscopic event?

- (a) Breaking of bonds
- (b) Collision between molecules
- (c) Change in colour of solution
- (d) Rearrangement of atoms

7. Which of the following represents a microscopic event?

- (a) Formation of precipitate
- (b) Evolution of gas
- (c) Change in pressure

(d) Rearrangement of atoms in molecules ✓

8. At equilibrium, which statement is correct?

(a) Forward reaction stops

(b) Reverse reaction stops

(c) Both reactions continue at equal rates ✓

(d) No reaction occurs

9. Which of the following is NOT a characteristic of macroscopic events?

(a) Visible to naked eye

(b) Includes change in colour

(c) Involves molecular collisions ✓

(d) Includes formation of precipitate

10. Microscopic events mainly involve:

(a) Observable changes

(b) Bulk properties

(c) Particle-level interactions ✓

(d) Colour changes

11. Which statement best explains why equilibrium appears static at macroscopic level?

- (a) No reaction occurs
- (b) Reactants are fully consumed
- (c) Equal rates of forward and reverse reactions cancel observable changes
- (d) Only reverse reaction occurs

12. In a reversible reaction, dynamic equilibrium means:

- (a) Reaction has stopped
- (b) Only forward reaction occurs
- (c) Both reactions occur with equal rates continuously
- (d) Concentrations become zero

13. Why does the forward reaction rate decrease with time?

- (a) Increase in products
- (b) Decrease in temperature
- (c) Decrease in reactant concentration
- (d) Increase in pressure

14. Why does the reverse reaction rate increase initially?

-
- (a) Increase in reactants
 - (b) Formation and accumulation of products ✓
 - (c) Decrease in collisions
 - (d) Decrease in temperature

15. At equilibrium, which condition is always true?

- (a) $[\text{Reactants}] = [\text{Products}]$
- (b) Forward rate = Reverse rate ✓
- (c) Reaction stops
- (d) Only products exist

16. If the system is disturbed after equilibrium, what happens?

- (a) Reaction stops permanently
- (b) No change occurs
- (c) Equilibrium shifts to re-establish balance ✓
- (d) Only forward reaction increases

17. Which of the following is TRUE for microscopic events at equilibrium?

- (a) No bond breaking occurs
- (b) Molecules remain static

(c) Continuous bond breaking and formation occurs ✓

(d) Only products react

18. At the start of reaction $\text{H}_2\text{O} + \text{CO} \rightleftharpoons \text{H}_2 + \text{CO}_2$, why is forward rate maximum?

(a) Product concentration is high

(b) Reactant concentration is high leading to maximum collisions ✓

(c) Reverse reaction is faster

(d) No collisions occur

19. Which graph correctly represents equilibrium?

(a) Concentrations become zero

(b) Concentrations keep increasing

(c) Concentrations become constant after some time ✓

(d) Only reactants remain

20. Which statement is incorrect about equilibrium?

(a) It is dynamic

(b) Molecular activity continues

(c) Reaction stops completely ✓

(d) Rates are equal

21. At equilibrium, if product concentration is greater than reactants, it means:

- (a) System is not at equilibrium
- (b) Only forward reaction occurs
- (c) Equilibrium favors products but rates are still equal
- (d) Reverse reaction has stopped

22. Which factor mainly controls reaction rates at microscopic level?

- (a) Colour change
- (b) Pressure change
- (c) Molecular collisions
- (d) Volume change



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23. In dynamic equilibrium, why do concentrations remain constant?

- (a) Reaction stops
- (b) No molecules react
- (c) Equal rates of formation and consumption of substances
- (d) All reactants are used up

24. Which situation proves equilibrium is dynamic rather than static?

- (a) No visible change
- (b) Continuous interconversion of reactants and products
- (c) Equal concentrations
- (d) Constant temperature

25. Which statement best differentiates macroscopic and microscopic views of equilibrium?

- (a) Both show no reaction
- (b) Macroscopic shows no change, microscopic shows continuous activity
- (c) Both show continuous change
- (d) Microscopic shows no activity

26. Macroscopic events are the result of:

- (a) Single microscopic event
- (b) Multiple simultaneous microscopic events
- (c) Only chemical reactions
- (d) Only physical changes

27. Understanding microscopic events helps to:

-
- (a) Stop reactions
 - (b) Change temperature
 - (c) Explain and predict macroscopic changes ✓
 - (d) Increase pressure

28. Which microscopic factor directly affects reaction rates?

- (a) Colour change
- (b) Molecular collisions ✓
- (c) Volume change
- (d) Shape of container

29. Activation energy mainly influences:

- (a) Colour of reaction
- (b) Rate of microscopic events ✓
- (c) Volume of system
- (d) Pressure only

30. Macroscopic changes in equilibrium are due to:

- (a) Temperature only
- (b) Pressure only
- (c) Combined microscopic events ✓

(d) Catalyst only

31. Dynamic equilibrium between physical states occurs when:

(a) Only one phase exists

(b) Phase change stops

(c) Rates of opposite processes are equal

(d) Temperature becomes zero

32. At 0°C, ice and water exist together because:

(a) Reaction stops

(b) Only melting occurs

(c) Only freezing occurs

(d) Melting and freezing occur at equal rates

33. Which is an example of dynamic equilibrium in physical states?

(a) Boiling water

(b) Freezing water only

(c) $\text{Ice} \rightleftharpoons \text{Water}$ at 0°C

(d) Evaporation only

34. Evaporation is the process in which:

(a) Gas changes into liquid

(b) Liquid changes into gas ✓

(c) Solid changes into liquid

(d) Gas changes into solid

35. Condensation is the process in which:

(a) Liquid changes into gas

(b) Solid changes into gas

(c) Gas changes into liquid ✓

(d) Liquid changes into solid

36. Vapor pressure is caused by:

(a) Solid particles

(b) Liquid molecules only

(c) Gas molecules striking container walls ✓

(d) External pressure only

37. Dynamic equilibrium in vapor pressure is achieved when:

(a) Evaporation stops

(b) Condensation stops

(c) Rates of evaporation and condensation become equal ✓

(d) Temperature becomes zero

38. At dynamic equilibrium, vapor pressure:

- (a) Increases continuously
- (b) Decreases continuously
- (c) Remains constant
- (d) Becomes zero

39. What happens to vapor pressure when temperature increases?

- (a) Decreases
- (b) Remains constant
- (c) Increases
- (d) Becomes zero



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40. Increase in temperature increases evaporation because:

- (a) Molecules lose energy
- (b) Molecules gain kinetic energy
- (c) Pressure decreases
- (d) Volume decreases

41. At higher temperature, equilibrium vapor pressure is:

- (a) Lower
- (b) Same

(c) Higher

(d) Zero

42. Which process occurs simultaneously during vapor equilibrium?

(a) Evaporation only

(b) Condensation only

(c) Both evaporation and condensation

(d) Freezing only

43. Dynamic equilibrium is called "dynamic" because:

(a) System is moving

(b) Continuous molecular activity occurs

(c) Temperature changes

(d) Pressure changes

44. If a system at vapor equilibrium is disturbed, then:

(a) No change occurs

(b) Equilibrium shifts to a new state

(c) Evaporation stops

(d) Condensation stops

45. The best definition of vapor pressure equilibrium is:

- (a) No molecules change phase
- (b) Only evaporation occurs
- (c) Rate of evaporation equals rate of condensation ✓
- (d) Liquid completely evaporates

46. Chemical equilibrium can be established only in:

- (a) Irreversible reactions
- (b) Open systems
- (c) Reversible reactions ✓
- (d) Exothermic reactions



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47. Which condition is necessary for equilibrium?

- (a) Open container
- (b) Closed system ✓
- (c) High temperature
- (d) Presence of catalyst

48. Why must the reaction vessel be closed for equilibrium?

- (a) To increase pressure
- (b) To prevent escape of substances ✓

(c) To increase temperature

(d) To stop reaction

49. At equilibrium, concentrations of reactants and products:

(a) Become zero

(b) Keep changing

(c) Remain constant

(d) Become equal always

50. Equilibrium can be attained starting from:

(a) Reactants only

(b) Products only

(c) Either reactants or products

(d) Catalyst only

51. Effect of catalyst on equilibrium position is:

(a) Shifts equilibrium forward

(b) Shifts equilibrium backward

(c) Does not change equilibrium position

(d) Stops reaction

52. Role of catalyst in equilibrium is to:

-
- (a) Change equilibrium constant
 - (b) Change concentrations
 - (c) Reach equilibrium faster ✓
 - (d) Stop reverse reaction

53. Equilibrium constant depends on:

- (a) Initial concentrations
- (b) Pressure only
- (c) Temperature only ✓
- (d) Catalyst

54. Homogeneous equilibrium is when:

- (a) Reactants only in same phase
- (b) Products only in same phase
- (c) All reactants and products in same phase ✓
- (d) Different phases exist

55. Heterogeneous equilibrium is when:

- (a) All substances are gases
- (b) All substances are liquids
- (c) Substances are in different phases ✓

(d) Only solids are present

56. Which of the following is a heterogeneous equilibrium?

(a) $\text{H}_2 + \text{Cl}_2 \rightleftharpoons 2\text{HCl}$

(b) $\text{N}_2 + 3\text{H}_2 \rightleftharpoons 2\text{NH}_3$

(c) $\text{CaCO}_3(\text{s}) \rightleftharpoons \text{CaO}(\text{s}) + \text{CO}_2(\text{g})$ ✓

(d) $\text{H}_2 + \text{I}_2 \rightleftharpoons 2\text{HI}$

57. The position of equilibrium indicates:

(a) Reaction speed

(b) Temperature change

(c) Relative amounts of reactants and products ✓

(d) Pressure change

58. According to Law of Mass Action, rate of reaction depends on:

(a) Temperature only

(b) Pressure only

(c) Concentration of reactants ✓

(d) Volume only

59. The equilibrium constant expression for $a\text{A} + b\text{B} \rightleftharpoons c\text{C} + d\text{D}$ is:

(a) $([\text{A}][\text{B}])/([\text{C}][\text{D}])$

(b) $\frac{[C]^c [D]^d}{[A]^a [B]^b}$ ✓

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

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

60. When number of moles of reactants and products are equal, equilibrium constant:

(a) Has units

(b) Has no units ✓

(c) Becomes zero

(d) Becomes infinite

61. Which statement is correct about equilibrium constants?

(a) Only one form exists

(b) Depends only on pressure

(c) Can be expressed in different forms (K_c , K_p , etc.) ✓

(d) Depends on catalyst

62. The position of equilibrium depends on:

(a) Reaction speed

(b) Relative concentrations of reactants and products ✓

(c) Activation energy only

(d) Catalyst only

63. If more products are added to an equilibrium system, what will happen?

(a) Forward reaction increases

(b) Reverse reaction increases to consume products

(c) Reaction stops

(d) No effect occurs

64. If products are continuously removed, the system will:

(a) Shift to left

(b) Shift to right to produce more products

(c) Stop reaction

(d) Decrease K

65. According to Le-Chatelier's Principle, a system under stress will:

(a) Ignore the change

(b) Enhance the disturbance

(c) Oppose the disturbance

(d) Stop reaction

66. Which change will NOT disturb equilibrium?

-
- (a) Change in concentration
 - (b) Change in temperature
 - (c) Addition of catalyst (position remains same)
 - (d) Change in pressure

67. When reactant concentration is increased, equilibrium shifts because:

- (a) Reverse reaction becomes faster
- (b) System tries to reduce added reactants
- (c) Products are removed
- (d) K value increases

68. Which situation causes equilibrium to shift left?

- (a) Adding reactants
- (b) Removing products
- (c) Adding products
- (d) Increasing temperature

69. In $\text{BiCl}_3 + \text{H}_2\text{O} \rightleftharpoons \text{BiOCl} + 2\text{HCl}$, adding HCl causes:

- (a) Forward reaction increases
- (b) Reverse reaction increases

-
- (c) No change
 - (d) Reaction stops

70. Why does equilibrium shift when concentration changes?

- (a) To change temperature
- (b) To maintain constant K
- (c) To counteract the disturbance
- (d) To stop reaction

71. Which statement is true about equilibrium constant (K)?

- (a) Changes with concentration
- (b) Changes with catalyst
- (c) Remains constant at constant temperature
- (d) Becomes zero at equilibrium

72. If a reactant is removed at equilibrium, the system will:

- (a) Shift forward
- (b) Shift backward to produce more reactant
- (c) Stop reaction
- (d) Increase K

73. Which factor changes the value of equilibrium constant?

-
- (a) Concentration
 - (b) Pressure
 - (c) Temperature
 - (d) Catalyst

74. Which option best explains Le-Chatelier's Principle practically?

- (a) System speeds up reactions
- (b) System always forms products
- (c) System adjusts itself to minimize external changes
- (d) System stops reacting

75. Which situation causes equilibrium to shift left?

- (a) Adding reactants
- (b) Removing products
- (c) Adding products
- (d) Increasing temperature

76. Effect of pressure change is only significant when:

- (a) No gases are involved
- (b) Gases are involved and moles are unequal
- (c) Liquids are involved

(d) Solids are involved

77. Pressure change has no effect on equilibrium when:

(a) Reaction is exothermic

(b) Gases are present

(c) Number of gaseous moles are equal in reactants and products

(d) Catalyst is present

78. In Haber's process ($\text{N}_2 + 3\text{H}_2 \rightleftharpoons 2\text{NH}_3$), increasing pressure shifts equilibrium:

(a) To left

(b) To right

(c) No change

(d) Stops reaction

79. Increase in pressure favors the side with:

(a) Greater number of moles

(b) Fewer number of moles

(c) Equal moles

(d) No gases

80. Decrease in pressure causes equilibrium to shift towards:

-
- (a) Side with fewer moles
 - (b) Side with greater moles
 - (c) No shift
 - (d) Catalyst side

81. Which condition favors forward reaction in Haber's process?

- (a) Decrease in pressure
- (b) Increase in volume
- (c) Increase in pressure
- (d) Addition of catalyst only

82. Temperature is unique because it affects:

- (a) Only reaction rate
- (b) Only pressure
- (c) Both equilibrium position and equilibrium constant (K)
- (d) Only volume

83. For an exothermic reaction, increasing temperature will:

- (a) Shift equilibrium to forward direction
- (b) Shift equilibrium to reverse direction
- (c) Not affect equilibrium

(d) Increase K

84. For an endothermic reaction, increasing temperature will:

(a) Shift equilibrium to reactants

(b) Shift equilibrium to products

(c) Stop reaction

(d) Decrease K always

85. In exothermic reactions, decrease in temperature:

(a) Favors reverse reaction

(b) Favors forward reaction

(c) Has no effect

(d) Decreases rate only

86. In the reaction $2\text{SO}_2 + \text{O}_2 \rightleftharpoons 2\text{SO}_3$, increase in temperature shifts equilibrium:

(a) To right

(b) To left

(c) No change

(d) Stops reaction

87. For the reaction $\text{N}_2\text{O}_4 \rightleftharpoons 2\text{NO}_2$, increase in temperature:

(a) Favors reactants

(b) Favors products

(c) Has no effect

(d) Decreases K

88. The value of equilibrium constant changes only with:

(a) Pressure

(b) Concentration

(c) Temperature

(d) Catalyst

89. A catalyst affects equilibrium by:

(a) Changing equilibrium position

(b) Changing equilibrium constant

(c) Lowering activation energy and speeding up attainment of equilibrium

(d) Stopping reverse reaction

90. In presence of a catalyst, reaction pathway is characterized by:

(a) Higher activation energy

(b) No activation energy

(c) Lower activation energy

(d) Increased equilibrium constant

91. The synthesis of ammonia in Haber's process is a:

(a) Irreversible reaction

(b) Reversible reaction

(c) Combustion reaction

(d) Displacement reaction

92. Which condition increases the yield of ammonia in Haber's process?

(a) Low pressure

(b) High pressure

(c) High temperature

(d) No catalyst

93. The Haber process involves the reaction of:

(a) Nitrogen and oxygen

(b) Nitrogen and hydrogen

(c) Hydrogen and carbon dioxide

(d) Oxygen and sulfur

94. Increasing pressure shifts the equilibrium of ammonia synthesis because:

- (a) More moles of gas are formed
- (b) Fewer moles of gas are formed
- (c) Reaction becomes irreversible
- (d) Temperature increases

95. The Haber process is exothermic, so decreasing temperature:

- (a) Decreases yield
- (b) Has no effect
- (c) Increases yield
- (d) Stops reaction



96. Which of the following is used as a catalyst in Haber's process?

- (a) V_2O_5
- (b) Fe (iron)
- (c) Pt
- (d) Cu

97. The optimum temperature for Haber's process is approximately:

-
- (a) 200°C
 - (b) 300°C
 - (c) 400°C
 - (d) 600°C

98. In the preparation of sulfur trioxide, the reaction is:

- (a) Endothermic
- (b) Exothermic
- (c) Neutral
- (d) Combustion

99. Which catalyst is used in the manufacture of sulfur trioxide?

- (a) Fe
- (b) V_2O_5
- (c) Ni
- (d) Al_2O_3

100. In the Contact Process, higher yield of SO_3 is favored by:

- (a) High temperature
- (b) Low temperature
- (c) Low pressure

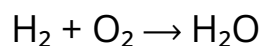
(d) No catalyst

❖ Important Short Questions:

1. What is an irreversible reaction?

Ans: An irreversible reaction is a reaction in which reactants are completely converted into products and the reaction does not proceed in the reverse direction.

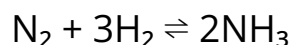
Example:



2. Define a reversible reaction.

Ans: A reversible reaction is a reaction that can proceed in both forward and reverse directions under the same conditions.

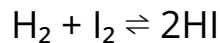
Example:



3. What is chemical equilibrium?

Ans: Chemical equilibrium is a state in a reversible reaction where the concentrations of reactants and products remain constant and the rates of forward and reverse reactions are equal.

Example:



4. What happens to the rates of forward and reverse reactions at equilibrium?

Ans: At equilibrium, the rates of forward and reverse reactions become equal.

5. Differentiate between macroscopic and microscopic events.

Answer:

- **Macroscopic events:** Observable changes (seen with naked eye).
- **Microscopic events:** Changes at molecular level (cannot be seen).

Example:

Colour change (macroscopic), bond breaking (microscopic)

6. Give two examples of macroscopic events.

Answer:

1. Change in colour
2. Formation of precipitate

7. What are microscopic events?

Ans: Microscopic events are changes that occur at molecular level such as collisions, bond breaking and formation.

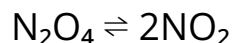
Example:

Collision between H_2 and I_2 molecules

8. What is dynamic equilibrium?

Ans: Dynamic equilibrium is a state where forward and reverse reactions continue at equal rates and concentrations remain constant.

Example:



9. Why is equilibrium called dynamic?

Ans: Because reactions do not stop; both forward and reverse reactions continue at equal rates.

10. Why is a closed system necessary for equilibrium?

Ans: A closed system is necessary to prevent the escape or entry of reactants and products so equilibrium can be maintained.

Example:

Gas reaction in a sealed container

11. What is vapor pressure equilibrium?

Ans: Vapor pressure equilibrium is a state where the rate of evaporation equals the rate of condensation in a closed system.

Example:

Water \rightleftharpoons Water vapor (in a closed container)

12. Differentiate between evaporation and condensation.

Answer:

- **Evaporation:** Liquid → Gas
- **Condensation:** Gas → Liquid

Example:

- Water → vapor (evaporation), vapor → water (condensation)

13. What is equilibrium constant (Kc)?

Ans: Kc is the ratio of the product of concentrations of products to reactants, each raised to their coefficients at equilibrium.

Example:

For: $A + B \rightleftharpoons C + D$

$$K_c = \frac{[C][D]}{[A][B]}$$

14. State the Law of Mass Action.

Ans: The rate of a reaction is directly proportional to the product of the concentrations of reactants, each raised to their coefficients.

Example:

$$\text{Rate} \propto [A][B]$$

15. Why does equilibrium constant depend on temperature?

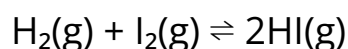
Ans: Because temperature changes the energy of the system, which affects reaction rates and shifts equilibrium position.

Example:

Exothermic reaction shifts backward on increasing temperature

16. What is homogeneous equilibrium?

Ans: It is an equilibrium in which all reactants and products are in the same phase.

Example:**17. Define heterogeneous equilibrium.**

Ans: It is an equilibrium where reactants and products are in different phases.

Example:**18. What is the effect of a catalyst on equilibrium?**

Ans: A catalyst does not change the equilibrium position or K_c ; it only helps to reach equilibrium faster.

Example:

Iron catalyst in Haber process

19. State Le-Chatelier's Principle.

Ans: If a system at equilibrium is disturbed, it shifts in a direction to oppose the disturbance.

Example:

Adding reactant shifts equilibrium forward

20. What is meant by the position of equilibrium?

Ans: It refers to the relative amounts of reactants and products present at equilibrium.

Example:

More products = equilibrium lies to the right

21. How does increase in pressure affect equilibrium?

Ans: Increase in pressure shifts equilibrium towards the side with fewer moles of gas.

Example:

$\text{N}_2 + 3\text{H}_2 \rightleftharpoons 2\text{NH}_3 \rightarrow$ shifts right (fewer moles)

22. What is the effect of temperature on equilibrium constant?

Ans: Temperature changes the value of equilibrium constant (K).

Example:

For exothermic reactions, increasing temperature decreases K

23. What happens when temperature increases in an exothermic reaction?

Ans: Equilibrium shifts to the reverse direction (towards reactants).

Example:

$2\text{SO}_2 + \text{O}_2 \rightleftharpoons 2\text{SO}_3$ (heat released) \rightarrow shifts left on heating

24. What is the effect of temperature on an endothermic reaction?

Ans: Increase in temperature shifts equilibrium to the forward direction (towards products).

Example:

$\text{N}_2\text{O}_4 \rightleftharpoons 2\text{NO}_2$ \rightarrow shifts right on heating

25. What are the optimum conditions for maximum yield of ammonia in Haber's process?

Answer:

- High pressure (200–300 atm)
- Moderate temperature ($\sim 400^\circ\text{C}$)
- Iron catalyst

Example:

$\text{N}_2 + 3\text{H}_2 \rightleftharpoons 2\text{NH}_3$

❖ Important Long Questions:

☀ **Q1: Define reversible reaction and chemical equilibrium.**

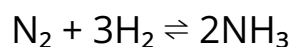
Explain dynamic equilibrium with suitable examples.

1. Reversible Reaction

A reversible reaction is a chemical reaction that can proceed in both forward and reverse directions under the same conditions.

In such reactions, reactants are converted into products in the forward direction, and products can again convert back into reactants in the reverse direction.

Example:



- Forward reaction: $\text{N}_2 + 3\text{H}_2 \rightarrow 2\text{NH}_3$
- Reverse reaction: $2\text{NH}_3 \rightarrow \text{N}_2 + 3\text{H}_2$

At the beginning, the forward reaction is faster because reactants are present in higher concentration. As products form, the reverse reaction also starts.

2. Chemical Equilibrium

Chemical equilibrium is the state of a reversible reaction where the rate of forward reaction becomes equal to the rate of reverse reaction, and the concentrations of reactants and products remain constant with time.

At equilibrium:

-
- No observable change occurs
 - Concentrations remain constant (not necessarily equal)
 - The system appears stable

3. Dynamic Equilibrium

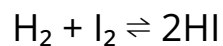
Dynamic equilibrium refers to the condition in which both forward and reverse reactions continue to occur simultaneously at equal rates.

👉 It is called “dynamic” because:

- Reactions are still occurring at the molecular level
- Reactants are continuously forming products
- Products are continuously forming reactants
- But both processes occur at the same rate, so no net change is observed

4. Examples of Dynamic Equilibrium

(a) Chemical Example:



At equilibrium:

Rate of forward reaction = Rate of reverse reaction

(b) Physical Example:

Water \rightleftharpoons Water Vapour (in a closed container)

- Evaporation: liquid \rightarrow gas

-
- Condensation: gas \rightarrow liquid

At equilibrium:

Rate of evaporation = Rate of condensation

◆ Summary:

- Reversible reactions proceed in both directions.
- Chemical equilibrium is the state where forward and reverse reaction rates are equal.
- At equilibrium, concentrations remain constant but reactions continue.
- This state is called dynamic equilibrium because molecular activity never stops.

★ Q2. Differentiate between macroscopic and microscopic events. Explain their relationship in a chemical equilibrium system.

❖ Answer:

1. Macroscopic Events

Macroscopic events are those changes in a chemical system that can be observed with the naked eye or measured using instruments. These changes occur at the bulk level of the substance.

They include visible and measurable properties of the system such as colour change, formation of precipitate, change in temperature, pressure, or volume.

Example: Formation of a precipitate during a reaction or a visible colour change in a solution.

2. Microscopic Events

Microscopic events are those changes that occur at the molecular or atomic level and cannot be observed directly. These involve the behavior of individual particles such as molecules or atoms.

They include processes like collisions between molecules, breaking of chemical bonds, and formation of new bonds.

Example: In a chemical reaction, reactant molecules collide, rearrange, and form product molecules at the molecular level.

3. Difference Between Macroscopic and Microscopic Events

Macroscopic events are observable and deal with bulk properties of matter, while microscopic events involve individual particles and their interactions, which cannot be directly seen.

Macroscopic changes give us an overall picture of the reaction, whereas microscopic events explain how the reaction actually occurs at the particle level.

4. Relationship in a Chemical Equilibrium System

In a chemical equilibrium system, both macroscopic and microscopic events are closely related.

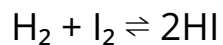
At equilibrium, macroscopic properties such as concentration, colour, pressure, and temperature remain constant with time. There is no visible change in the system.

However, at the microscopic level, reactions are still occurring continuously. Reactant molecules keep converting into products, and product molecules keep converting back into reactants.

The key point is that the rate of the forward reaction is equal to the rate of the reverse reaction. Because of this balance, there is no net change in macroscopic properties, even though microscopic activity continues.

Example to Explain the Relationship

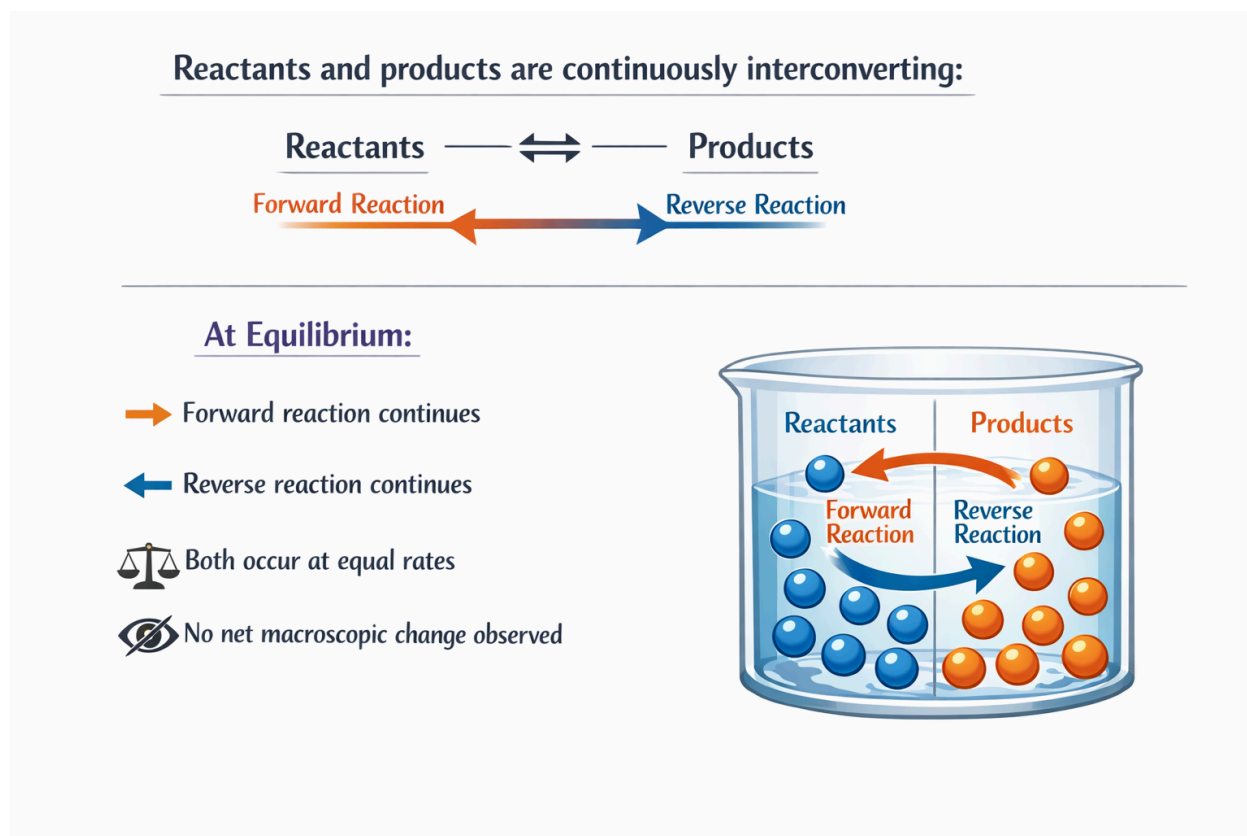
Consider the equilibrium reaction:



At the macroscopic level, the concentrations of H_2 , I_2 , and HI remain constant once equilibrium is reached. No visible change is observed.

At the microscopic level, H_2 and I_2 molecules continue to collide and form HI , while HI molecules decompose back into H_2 and I_2 . Both processes occur continuously at equal rates.

Diagram:



◆ Summary:

Macroscopic events are observable changes at the bulk level, while microscopic events occur at the molecular level and are not visible. In a chemical equilibrium system, macroscopic properties remain constant because microscopic forward and reverse reactions continue at equal rates. This continuous activity at the molecular level makes equilibrium dynamic in nature.

☀ Q3. State and explain Le-Chatelier's Principle.

❖ **Answer:**

1. Statement of Le-Chatelier's Principle

Le-Chatelier's Principle states that:

"If a system at equilibrium is subjected to a change in concentration, pressure, or temperature, the system will adjust itself in such a way as to counteract or oppose the effect of that change."

2. Explanation of the Principle

When a chemical system is at equilibrium, the forward and reverse reaction rates are equal. If any external condition is changed, the equilibrium is disturbed.

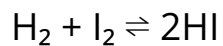
To re-establish equilibrium, the system shifts in a direction that reduces the effect of the disturbance.

- If reactants are added → system shifts forward to consume them
- If products are added → system shifts backward to consume them
- If pressure is increased → system shifts toward fewer moles of gas
- If temperature is changed → system shifts depending on whether the reaction is exothermic or endothermic

This shift continues until a new equilibrium is established.

3. Explanation with Example (Concentration Effect)

Consider the equilibrium reaction:



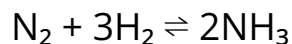
- If the concentration of H_2 or I_2 is increased, the equilibrium shifts to the right to form more HI.
- If the concentration of HI is increased, the equilibrium shifts to the left to produce more H_2 and I_2 .

This shows that the system opposes the change by consuming the added substance or producing more of the removed substance.

4. Pressure Effect (Conceptual Explanation)

In reactions involving gases, increasing pressure shifts equilibrium toward the side with fewer moles of gas.

Example:



Here, increasing pressure shifts the equilibrium to the right because the product side has fewer moles of gas.

5. Temperature Effect

- **For exothermic reactions:** increasing temperature shifts equilibrium to the left (reverse direction).
- **For endothermic reactions:** increasing temperature shifts equilibrium to the right (forward direction).

This happens because the system tries to absorb or release heat to oppose the change.

◆ **Summary:**

Le-Chatelier's Principle explains how a system at equilibrium responds to external changes. The system always shifts in a direction that opposes the disturbance and restores equilibrium. This principle is useful in predicting the direction of equilibrium shifts under changes in concentration, pressure, and temperature.

✨ **Q4. Define equilibrium constant (K_c). Derive the expression for a general reaction and explain its characteristics.**

❖ **Answer:**

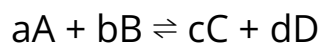
1. Definition of Equilibrium Constant (K_c)

The equilibrium constant (K_c) is defined as the ratio of the product of the concentrations of products to the product of the concentrations of reactants, each raised to the power of their stoichiometric coefficients, at equilibrium.

It is written in terms of molar concentrations (mol/L).

2. General Reaction and Derivation

Consider a general reversible reaction:



According to the law of mass action, the rate of forward reaction is proportional to the concentrations of reactants:

$$\text{Rate (forward)} \propto [A]^a [B]^b$$

$$\text{Rate (reverse)} \propto [C]^c [D]^d$$

At equilibrium:

$$\text{Rate (forward)} = \text{Rate (reverse)}$$

So,

$$[A]^a [B]^b \propto [C]^c [D]^d$$

Rearranging, we get:

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

This is the equilibrium constant expression for the given reaction.

3. Characteristics of K_c

Constant at a given temperature:

- The value of K_c remains constant for a reaction at a fixed temperature.

Independent of initial concentrations:

- K_c does not depend on the starting amounts of reactants or products.

Depends only on temperature:

-
- Changing temperature changes the value of K_c .

Indicates position of equilibrium:

Large $K_c \rightarrow$ products are favored (equilibrium lies to the right)

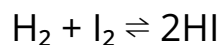
Small $K_c \rightarrow$ reactants are favored (equilibrium lies to the left)

Unit may vary:

The units of K_c depend on the overall change in moles in the reaction.

4. Example

For the reaction:



$$K_c = \frac{[\text{HI}]^2}{[\text{H}_2][\text{I}_2]}$$

◆ Summary:

The equilibrium constant (K_c) is a ratio of product concentrations to reactant concentrations at equilibrium, raised to their stoichiometric powers. It is derived using the law of mass action and is constant at a given temperature. K_c helps predict the extent and position of a reaction at equilibrium.

☀ Q5. Explain the effect of change in pressure, volume, and temperature on chemical equilibrium with suitable examples.

❖ Answer:

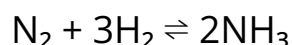
1. Effect of Change in Pressure and Volumes

The effect of pressure and volume is mainly applicable to reactions involving gases.

Pressure Increase

When pressure is increased, the equilibrium shifts toward the side having fewer moles of gas to reduce the pressure.

Example:



Here, 4 moles of reactants form 2 moles of product.

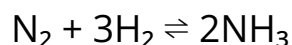
So, increasing pressure shifts equilibrium to the right, favoring ammonia formation.

Pressure Decrease (or Volume Increase)

When pressure is decreased (or volume is increased), the equilibrium shifts toward the side with a greater number of gas moles.

Example:

In the same reaction:



Decreasing pressure shifts equilibrium to the left, favoring the reactants.

Important Note:

If the number of moles of gases is equal on both sides, pressure or volume change has no effect on equilibrium.

2. Effect of Temperature

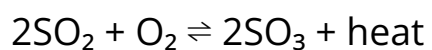
Temperature affects both the position of equilibrium and the value of the equilibrium constant (K).

Exothermic Reactions

Heat is released in exothermic reactions.

- Increasing temperature shifts equilibrium in the reverse direction (towards reactants).
- Decreasing temperature shifts equilibrium in the forward direction (towards products).

Example:



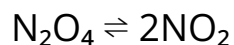
- Increasing temperature shifts equilibrium to the left
- Decreasing temperature shifts equilibrium to the right, favoring SO_3 formation

Endothermic Reactions

- Heat is absorbed in endothermic reactions.

-
- Increasing temperature shifts equilibrium in the forward direction (towards products).
 - Decreasing temperature shifts equilibrium in the reverse direction (towards reactants).

Example:



- Increasing temperature shifts equilibrium to the right
- Decreasing temperature shifts equilibrium to the left

3. Combined Explanation

- Pressure and volume changes affect equilibrium only when gases are involved and the number of moles differs on both sides.
- Temperature changes always affect equilibrium and also change the equilibrium constant (K).
- The system shifts in the direction that opposes the change, according to Le-Chatelier's Principle.

◆ **Summary:**

- Increasing pressure favors the side with fewer gas moles.
- Decreasing pressure favors the side with more gas moles.
- Decreasing volume is equivalent to increasing pressure.
- Temperature changes shift equilibrium depending on whether the reaction is exothermic or endothermic.

-
- These shifts help the system maintain equilibrium by opposing the external disturbance.

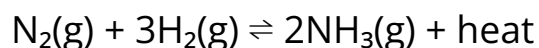
☀ Q6. Describe the industrial application of chemical equilibrium with reference to Haber's process. (10 marks answer)

❖ Introduction:

Chemical equilibrium plays an important role in industry where the aim is to obtain the maximum yield of a desired product. One of the best examples is the Haber's process, which is used for the industrial synthesis of ammonia.

Haber's Process (Reaction)

In this process, nitrogen reacts with hydrogen to form ammonia in a reversible reaction:



This reaction is exothermic and involves a decrease in the number of moles of gas (4 moles \rightarrow 2 moles).

3. Application of Le-Chatelier's Principle

To obtain maximum yield of ammonia, different conditions are adjusted according to Le-Chatelier's Principle:

(a) Effect of Pressure

- Increasing pressure shifts the equilibrium towards the side with fewer moles of gas.

-
- In this reaction, the product side has fewer moles (2 moles of NH_3).
 - Therefore, high pressure favors ammonia formation.

Industrial condition:

Pressure of about 200–300 atmospheres is used.

(b) Effect of Temperature

Since the reaction is exothermic, lowering the temperature favors the forward reaction and increases ammonia yield.

However, too low temperature slows down the reaction rate.

So a moderate temperature is used as a compromise.

Industrial condition:

Temperature is maintained around 673 K (about 400°C).

(c) Effect of Catalyst

A catalyst increases the rate of reaction without affecting the equilibrium position or equilibrium constant.

Industrial catalyst:

Finely divided iron (Fe) with promoters such as Al_2O_3 , K_2O , and MgO is used.

The catalyst helps the system reach equilibrium faster but does not increase the yield directly.

(d) Removal of Ammonia

Ammonia is continuously removed from the reaction mixture.

According to Le-Chatelier's Principle, removal of product shifts the equilibrium to the right, producing more ammonia.

◆ **Summary of Optimum Conditions**

High pressure (200–300 atm)

Moderate temperature ($\sim 400^\circ\text{C}$ / 673 K)

Iron catalyst with promoters

Continuous removal of NH_3

These conditions are selected to balance between high yield and reasonable reaction rate.

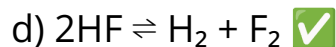
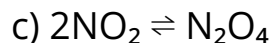
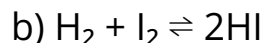
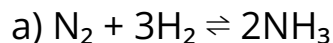
◆ **Summary:**

Haber's process is a key industrial application of chemical equilibrium. By applying Le-Chatelier's Principle, conditions such as high pressure, moderate temperature, use of catalyst, and removal of product are optimized to maximize ammonia production. This ensures both efficient yield and practical reaction rate in industrial settings.

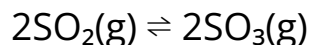
Exercise

Q.1 For each question, select the correct choice.

I. For which system does the equilibrium constant have units of (concentration)⁻¹?



II. Which statement about the following equilibrium is correct?



$$\Delta H = -197.9 \text{ kJ mol}^{-1}$$

a) The value of K_p falls with a rise in temperature. ✓

b) The value of K_p falls with increasing pressure.

c) Adding V_2O_5 catalyst increases the equilibrium yield of SO_3 .

d) The value of K_p is equal to K_c

III. For the reaction: $2\text{SO}_3(\text{g}) \rightleftharpoons 2\text{SO}_2(\text{g}) + \text{O}_2(\text{g})$, the K_c expression is:

a) $\frac{[\text{SO}_2]^2}{[\text{SO}_3]}$

b) $\frac{[\text{SO}_2]^2 [\text{O}_2]}{[\text{SO}_3]^2}$ ✓

c) $[\text{SO}_3]^2 / ([\text{SO}_2]^2 [\text{O}_2])$

d) $[\text{SO}_2][\text{O}_2]$

IV. A saturated solution represents a dynamic equilibrium because:

a) No more solute particles are dissolving

b) The rate of dissolution is zero

c) Solute particles dissolve and precipitate at the same rate

d) All solute particles have dissolved

V. Which statement correctly describes the effect of temperature on equilibrium constant?

a) K_c is directly proportional to temperature

b) K_c is inversely proportional to temperature

c) K_c depends on the enthalpy change of the reaction

d) Temperature has no effect on K_c

VI. For the equilibrium: $2\text{H}_2\text{O} \rightleftharpoons 2\text{H}_2 + \text{O}_2$ (endothermic forward reaction), which change decreases H_2O ?

a) Adding more oxygen

b) Adding a catalyst

c) Decreasing the volume

d) Increasing the temperature

VII. If $K_c = 0.04$ at 450°C for $\text{PCl}_5 \rightleftharpoons \text{PCl}_3 + \text{Cl}_2$, then K_p is:

a) 0.40

b) 0.64

c) 2.4

d) 0.052

VIII. In which gaseous equilibrium does pressure have no effect on equilibrium position?

a) $2\text{NO}_2 \rightleftharpoons \text{N}_2\text{O}_4$

b) $\text{PCl}_5 \rightleftharpoons \text{PCl}_3 + \text{Cl}_2$

c) $\text{CO} + \text{H}_2\text{O} \rightleftharpoons \text{CO}_2 + \text{H}_2$

d) $2\text{SO}_2 + \text{O}_2 \rightleftharpoons 2\text{SO}_3$

IX. For $2\text{H}_2 + \text{O}_2 \rightleftharpoons 2\text{H}_2\text{O}$, if H_2O is increased, H_2 and O_2 will:

a) Increase

b) Decrease

c) Remain the same

d) Change irregularly

X. For a reaction, the equilibrium constant K_c :

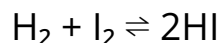
-
- a) Always remains the same under all conditions
 - b) Increases if product concentration increases
 - c) Changes with temperature ✓
 - d) Increases if reactant concentration increases

Q.2 Attempt the following short-answer questions:

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

Chemical equilibrium is the state of a reversible reaction at which the rate of the forward reaction becomes equal to the rate of the reverse reaction, and the concentrations of reactants and products remain constant with time.

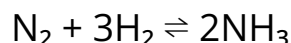
Example:



b. Define reversible reaction. Give an example.

A reversible reaction is a chemical reaction that can proceed in both forward and backward directions under the same conditions, meaning reactants form products and products can again form reactants.

Example:



c. The change of volume disturbs the equilibrium position for some gas phase reactions but not the equilibrium constant. Explain.

A change in volume changes the pressure of the system, which may shift the equilibrium position according to Le-Chatelier's Principle if the number of moles of gases on both sides is unequal. The system shifts toward the side with fewer or more gas moles depending on pressure change.

However, the equilibrium constant (K_c or K_p) depends only on temperature, not on pressure or volume. Therefore, although the position of equilibrium shifts, the value of the equilibrium constant remains unchanged.

d. Mention the characteristics of chemical equilibrium.

- It is a dynamic state where forward and reverse reactions continue simultaneously.
- The rates of forward and reverse reactions are equal.
- The concentrations of reactants and products remain constant (not necessarily equal).
- It can be established only in a closed system.
- It can be achieved from either direction (starting from reactants or products).
- It is affected by changes in temperature, pressure, and concentration, but not by a catalyst (in terms of position).

e. Explain why equilibrium is dynamic and not static.

Equilibrium is called dynamic because both forward and reverse reactions continue to occur at the same rate even after equilibrium is

achieved. Although there is no net change in concentrations, molecules are constantly being converted from reactants to products and vice versa.

It is not static because reactions do not stop; they are continuously happening at the molecular level.

f. Why do the rates of forward reactions slow down as equilibrium is approached?

As the reaction proceeds, the concentration of reactants decreases and the concentration of products increases. Since the rate of the forward reaction depends on the concentration of reactants, the forward reaction slows down. At the same time, the reverse reaction speeds up due to increased product concentration. Eventually, both rates become equal at equilibrium.

g. Why can ice at 0°C be melted by applying pressure without supplying heat?

Ice has a larger volume than liquid water. When pressure is applied, the system shifts toward the side with less volume (liquid water) according to Le-Chatelier's Principle. This causes ice to melt at 0°C without adding external heat because pressure favors the formation of the denser liquid phase.

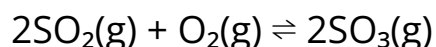
h. Write two conditions of equilibrium constant.

The temperature must remain constant, as the equilibrium constant depends only on temperature.

The system must be closed so that no reactants or products escape.

Q (i)

The reversible reaction:



has come to equilibrium in a vessel of specific volume at a given temperature. Before the reaction began, the concentrations of the reactants were 0.06 mol/dm^3 of SO_2 and 0.05 mol/dm^3 of O_2 . After equilibrium was reached, the concentration of SO_3 was 0.04 mol/dm^3 .

What is the equilibrium concentration of O_2 ?

Solution (Step-by-step explanation)

Step 1: Given data

Initial concentrations:

$$\text{SO}_2 = 0.06 \text{ mol/dm}^3$$

$$\text{O}_2 = 0.05 \text{ mol/dm}^3$$

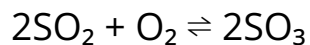
$$\text{SO}_3 = 0 \text{ mol/dm}^3$$

At equilibrium:

$$\text{SO}_3 = 0.04 \text{ mol/dm}^3$$

Step 2: Use stoichiometry of the reaction

From the equation:



This means:

2 moles of SO_3 are formed from 1 mole of O_2

So the ratio is:

O_2 consumed : SO_3 formed = 1 : 2

Step 3: Calculate O_2 consumed

SO_3 formed = 0.04 mol/dm^3

O_2 consumed = $0.04 \div 2 = 0.02 \text{ mol/dm}^3$

Step 4: Calculate equilibrium concentration of O_2

Equilibrium O_2 = Initial O_2 – O_2 consumed

= $0.05 - 0.02$

= 0.03 mol/dm^3

Final Answer:

Equilibrium concentration of O_2 = 0.03 mol/dm^3

❖ DESCRIPTIVE QUESTIONS

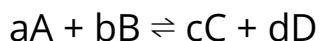
🌟 Q.3 (a) Law of Mass Action & Derivation of Equilibrium Constant

Law of Mass Action (Definition)

The law of mass action states that at a constant temperature, the rate of a chemical reaction is directly proportional to the product of the active masses (concentrations) of the reactants, each raised to the power of their stoichiometric coefficients.

◆ Explanation

Consider a general reversible reaction:



According to the law of mass action:

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

Rate of reverse reaction $\propto [C]^c [D]^d$

At Equilibrium

At equilibrium:

Rate of forward reaction = Rate of reverse reaction

So,

$$k_1 [A]^a [B]^b = k_2 [C]^c [D]^d$$

Rearranging:

$$([C]^c [D]^d) / ([A]^a [B]^b) = k_1 / k_2$$

Let $k_1 / k_2 = K_c$ (equilibrium constant)

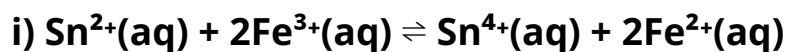
Final Expression of Equilibrium Constant

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

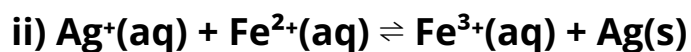
Where:

- K_c = equilibrium constant in terms of concentration
- Square brackets represent molar concentration (mol/dm^3)
- Exponents are stoichiometric coefficients

☀ Q.3 (b) Expressions for K_c

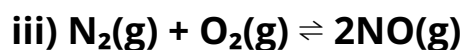


$$K_c = ([\text{Sn}^{4+}] [\text{Fe}^{2+}]^2) / ([\text{Sn}^{2+}] [\text{Fe}^{3+}]^2)$$

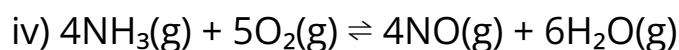


(Pure solid Ag is not included in K_c expression)

$$K_c = ([\text{Fe}^{3+}]) / ([\text{Ag}^+] [\text{Fe}^{2+}])$$



$$K_c = ([\text{NO}]^2) / ([\text{N}_2] [\text{O}_2])$$



$$K_c = ([\text{NO}]^4 [\text{H}_2\text{O}]^6) / ([\text{NH}_3]^4 [\text{O}_2]^5)$$



$$K_c = \frac{[\text{PCl}_3] [\text{Cl}_2]}{[\text{PCl}_5]}$$

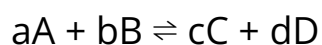
Important Exam Notes

- Pure solids and pure liquids are excluded from equilibrium expressions.
- Only gaseous and aqueous species are included.
- Powers in the expression always match the balanced equation coefficients.

★ **Q.4 Write down the K_c for the following reactions. Suppose that the reaction mixture in all the cases is in dm^{-3} (aqueous or gaseous concentration).**

(General Rule Reminder)

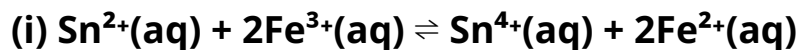
For a reaction:



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

- Only aqueous (aq) and gaseous (g) species are included
- Pure solids and liquids are omitted
- Powers are the stoichiometric coefficients

Expressions of K_c

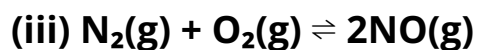


$$K_c = \frac{[\text{Sn}^{4+}] [\text{Fe}^{2+}]^2}{[\text{Sn}^{2+}] [\text{Fe}^{3+}]^2}$$

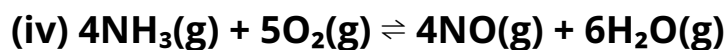


(Pure solid Ag is not included)

$$K_c = \frac{[\text{Fe}^{3+}]}{[\text{Ag}^+] [\text{Fe}^{2+}]}$$



$$K_c = \frac{[\text{NO}]^2}{[\text{N}_2] [\text{O}_2]}$$



$$K_c = \frac{[\text{NO}]^4 [\text{H}_2\text{O}]^6}{[\text{NH}_3]^4 [\text{O}_2]^5}$$



$$K_c = \frac{[\text{PCl}_3] [\text{Cl}_2]}{[\text{PCl}_5]}$$

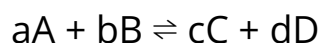
Important Note

Since all concentrations are given in dm^{-3} , the expressions of K_c are written directly using molar concentrations without additional conversion.

✨ **Q.4 Write down the K_c for the following reactions (concentrations in dm^{-3})**

General rule:

For a reaction:

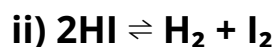


$$K_c = \frac{[C]^c [D]^d}{[A]^a [B]^b}$$

(Products in numerator, reactants in denominator; exponents are stoichiometric coefficients)



$$K_c = \frac{[\text{CH}_3\text{COOC}_2\text{H}_5] [\text{H}_2\text{O}]}{[\text{CH}_3\text{COOH}] [\text{CH}_3\text{CH}_2\text{OH}]}$$



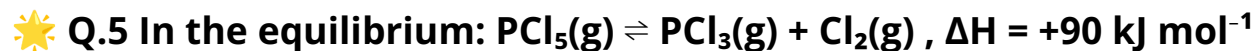
$$K_c = \frac{[\text{H}_2] [\text{I}_2]}{[\text{HI}]^2}$$



$$K_c = \frac{[\text{NH}_3]^2}{[\text{N}_2] [\text{H}_2]^3}$$

Note

- All species are in concentration form (mol dm^{-3}).
- Coefficients in the balanced equation become powers in the K_c expression.
- No solids or pure liquids are present, so all species are included.



(Forward reaction is endothermic)

(i) Effect of increase in temperature

Since the forward reaction is endothermic (absorbs heat), increasing temperature adds heat to the system.

According to Le-Chatelier's Principle, the equilibrium shifts in the direction that absorbs heat (forward direction).

Effect:

- Equilibrium shifts to the right
- More PCl_3 and Cl_2 are formed
- Decomposition of PCl_5 increases
- Equilibrium constant K_c increases (for endothermic reactions)

(ii) Effect of decrease in volume (increase in pressure)

When volume is decreased, pressure increases. The system shifts towards the side with fewer moles of gas.

Left side: 1 mole of gas (PCl_5)

Right side: 2 moles of gas ($\text{PCl}_3 + \text{Cl}_2$)

Effect:

- Equilibrium shifts to the left
- Formation of PCl_5 is favored
- Amount of PCl_3 and Cl_2 decreases

(iii) Effect of adding a catalyst

A catalyst increases the rate of both forward and reverse reactions equally.

Effect:

-
- No change in equilibrium position
 - No change in equilibrium constant (K_c)
 - Only helps the system reach equilibrium faster

(iv) Effect of adding chlorine (Cl_2)

Adding a product increases its concentration.

According to Le-Chatelier's Principle, the system will try to reduce the added chlorine by shifting the equilibrium.

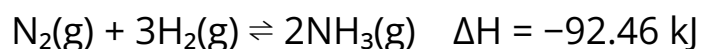
Effect:

- Equilibrium shifts to the left
- More PCl_5 is formed
- Concentration of PCl_3 decreases
- Added Cl_2 is partially consumed

◆ Summary:

- Temperature increase \rightarrow shifts to products (endothermic forward reaction)
- Decrease in volume \rightarrow shifts to side with fewer moles (left side)
- Catalyst \rightarrow no effect on equilibrium position
- Adding Cl_2 \rightarrow shifts equilibrium to the left (reverse direction)

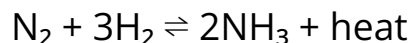
☀ **Q.6 Synthesis of ammonia by Haber's process is an exothermic reaction:**



What is the effect of change in temperature at the equilibrium stage?

❖ **Answer:**

Since the reaction is exothermic (heat is released in the forward direction), heat can be considered as a product:



Effect of Increase in Temperature

- Increasing temperature means adding heat to the system.
- According to Le-Chatelier's Principle, the system will try to consume the added heat.
- Therefore, equilibrium shifts in the reverse direction (left side).

Result:

- Formation of NH_3 decreases
- Decomposition of NH_3 increases
- Yield of ammonia decreases
- Equilibrium constant (K_c) decreases
- Effect of Decrease in Temperature
- Decreasing temperature means removing heat from the system.
- The system shifts to produce more heat.

Therefore, equilibrium shifts in the forward direction (right side).

Result:

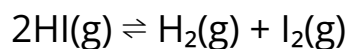
- Formation of NH_3 increases

-
- Yield of ammonia increases
 - Equilibrium constant (K_c) increases

◆ **Summary:**

- Increase in temperature → shifts equilibrium to the left → less NH_3 formed
- Decrease in temperature → shifts equilibrium to the right → more NH_3 formed
- Since the reaction is exothermic, low temperature favors product formation

★ **Q.7 K for the following reaction is 0.016 at 520°C:**



The equilibrium mixture contains:

$\text{HI} = 0.08 \text{ M}$, $\text{H}_2 = 0.01 \text{ M}$ and $\text{I}_2 = 0.01 \text{ M}$.

To this mixture, more HI is added so that its new concentration becomes 0.096 M. What will be the concentrations of HI, H_2 and I_2 when equilibrium is re-established?

◆ **Solution**

Step 1: Given Data

Initial equilibrium concentrations:

$\text{HI} = 0.08 \text{ M}$

$$\text{H}_2 = 0.01 \text{ M}$$

$$\text{I}_2 = 0.01 \text{ M}$$

After adding HI:

$$\text{HI} = 0.096 \text{ M}$$

$$K_c = 0.016$$

Step 2: Direction of Shift

Adding HI (reactant) disturbs equilibrium.

The system shifts to the right to form more products and consume HI.

Step 3: Let the change be x

Since equilibrium shifts to the right:

- HI decreases by $2x$
- H_2 increases by x
- I_2 increases by x

Step 4: New Equilibrium Concentrations

$$\text{HI} = 0.096 - 2x$$

$$\text{H}_2 = 0.01 + x$$

$$\text{I}_2 = 0.01 + x$$

Step 5: Apply Equilibrium Constant Expression

$$K_c = ([H_2][I_2]) / ([HI]^2)$$

$$0.016 = (0.01 + x)(0.01 + x) / (0.096 - 2x)^2$$

Step 6: Approximation Method

Since K_c is small, x is small, so:

$$0.096 - 2x \approx 0.096$$

So:

$$0.016 = (0.01 + x)^2 / (0.096)^2$$

Take square root:

$$\sqrt{0.016} = (0.01 + x) / 0.096$$

$$0.126 = (0.01 + x) / 0.096$$

$$0.01 + x = 0.096 \times 0.126$$

$$0.01 + x \approx 0.0121$$

$$x \approx 0.0021$$

Step 7: Final Concentrations

$$HI = 0.096 - 2x = 0.096 - 0.0042 = 0.0918 \text{ M} \approx 0.092 \text{ M}$$

$$H_2 = 0.01 + x = 0.0121 \text{ M} \approx 0.012 \text{ M}$$

$$I_2 = 0.01 + x = 0.0121 \text{ M} \approx 0.012 \text{ M}$$

Final Answer

HI \approx 0.092 M

H₂ \approx 0.012 M

I₂ \approx 0.012 M

◆ **Summary:**

- Adding HI shifts equilibrium to the right.
- Use ICE (Initial–Change–Equilibrium) method.
- Apply K_c expression to solve for x.
- Substitute x to find final equilibrium concentrations.

★ **Q.8 The equilibrium constant for the reaction between acetic acid and ethyl alcohol is 4. A mixture of 3 moles of acetic acid and 1 mole of ethyl alcohol is allowed to come to equilibrium. Calculate the amount of ethyl acetate present at equilibrium.**

❖ **Answer:**

Solution

Step 1: Write the reaction

Acetic acid + Ethyl alcohol \rightleftharpoons Ethyl acetate + Water

Step 2: Given data

Initial moles:

Acetic acid = 3 mol

Ethyl alcohol = 1 mol

Ethyl acetate = 0 mol

Water = 0 mol

$K_c = 4$

Step 3: Let the change be x

At equilibrium:

Acetic acid = $3 - x$

Ethyl alcohol = $1 - x$

Ethyl acetate = x

Water = x

Step 4: Write equilibrium expression

$K_c = (\text{products}) / (\text{reactants})$

$K_c = (x \times x) / ((3 - x)(1 - x))$

$4 = x^2 / ((3 - x)(1 - x))$

Step 5: Solve the equation

$x^2 = 4(3 - x)(1 - x)$

Expand:

$(3 - x)(1 - x) = 3 - 4x + x^2$

So:

$$x^2 = 4(3 - 4x + x^2)$$

$$x^2 = 12 - 16x + 4x^2$$

Bring all terms to one side:

$$0 = 12 - 16x + 4x^2 - x^2$$

$$0 = 12 - 16x + 3x^2$$

Rearrange:

$$3x^2 - 16x + 12 = 0$$

Step 6: Solve quadratic equation

$$x = [16 \pm \sqrt{(256 - 144)}] / 6$$

$$x = [16 \pm \sqrt{112}] / 6$$

$$\sqrt{112} \approx 10.58$$

$$x = (16 - 10.58) / 6 \approx 5.42 / 6 \approx 0.90 \text{ mol}$$

(Other value is not possible as it exceeds initial moles)

Final Answer

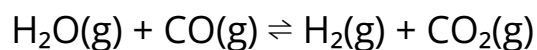
Amount of ethyl acetate at equilibrium = 0.90 moles (approx.)

◆ **Summary:**

- Use ICE (Initial-Change-Equilibrium) method
- Apply equilibrium constant expression

-
- Solve quadratic equation
 - Choose the physically valid root
 - The amount of product formed equals x , which is ≈ 0.90 mol

★ Q.9 Study the equilibrium



(i) Write the expression of K_p .

(ii) When 1.00 mole of steam and 1.00 mole of CO are allowed to reach equilibrium, 33.3% of equilibrium mixture is hydrogen. Calculate the value of K_p . State the units of K_p .

❖ **Answer:**

(i) Expression of K_p

$$K_p = (P_{\text{H}_2} \times P_{\text{CO}_2}) / (P_{\text{H}_2\text{O}} \times P_{\text{CO}})$$

(ii) Calculation of K_p

Step 1: Initial moles

$$\text{H}_2\text{O} = 1.00 \text{ mol}$$

$$\text{CO} = 1.00 \text{ mol}$$

$$\text{H}_2 = 0$$

$$\text{CO}_2 = 0$$

Step 2: Let x be the moles reacted

$$\text{H}_2\text{O} = 1 - x$$

$$\text{CO} = 1 - x$$

$$\text{H}_2 = x$$

$$\text{CO}_2 = x$$

$$\text{Total moles at equilibrium} = (1 - x) + (1 - x) + x + x = 2 \text{ mol}$$

Step 3: Given condition

Hydrogen = 33.3% of mixture

$$\text{Mole fraction of H}_2 = x / 2 = 1/3$$

So,

$$x = 2/3 = 0.667 \text{ mol}$$

Step 4: Equilibrium moles

$$\text{H}_2 = 0.667 \text{ mol}$$

$$\text{CO}_2 = 0.667 \text{ mol}$$

$$\text{H}_2\text{O} = 0.333 \text{ mol}$$

$$\text{CO} = 0.333 \text{ mol}$$

Step 5: Mole fractions (partial pressures proportional)

$$\text{Total moles} = 2$$

$$P_{\text{H}_2} = 0.667 / 2$$

$$P_{\text{CO}_2} = 0.667 / 2$$

$$P_{\text{H}_2\text{O}} = 0.333 / 2$$

$$P_{\text{CO}} = 0.333 / 2$$


Step 6: Substitute in Kp expression

$$K_p = (P_{\text{H}_2} \times P_{\text{CO}_2}) / (P_{\text{H}_2\text{O}} \times P_{\text{CO}})$$

$$K_p = 1$$

Final Answer:

$$K_p = 1$$

Units of Kp: None (dimensionless) 

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