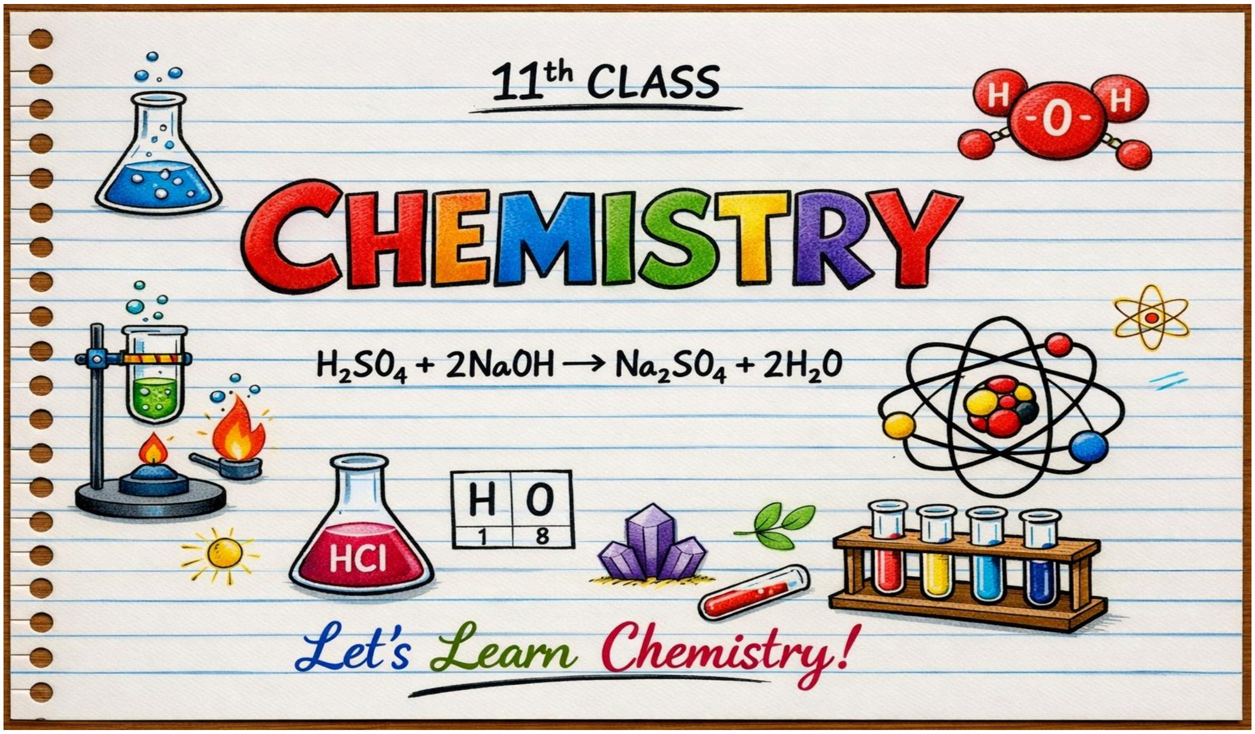

Class: 11th

Subject: Chemistry

Unit 7: REACTION KINETICS

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❖ Important MCQs:

1. According to collision theory, a chemical reaction occurs when:

a) Particles collide randomly

b) Particles collide with sufficient energy and proper orientation ✅

c) Particles remain stationary

d) Temperature remains constant

2. Ineffective collisions are those in which:

a) Energy is greater than activation energy

b) Orientation is correct

c) Energy is insufficient or orientation is improper ✅

d) Pressure is high

3. Activation energy is:

a) Maximum energy of molecules

b) Average kinetic energy

c) Minimum energy required for effective collision ✅

d) Total system energy

4. Most reactions are slow because:

a) All collisions are effective

b) Molecules have equal energy

c) Only a small fraction of collisions are effective ✅

d) Temperature is always low

5. Proper orientation means:

a) Random arrangement

b) Correct alignment of reacting particles during collision ✅

c) High molecular speed

d) Low pressure

6. Rate of reaction is defined as:

a) Change in temperature per unit time

b) Change in concentration per unit time ✅

c) Change in pressure per unit time

d) Change in volume per unit time

7. Standard unit of rate of reaction is:

a) mol dm^{-3} s^{-1} ✅

b) mol s^{-1}

c) dm^{3} mol^{-1}

d) s^{-1}

8. For reaction A ⟶ B, rate can be expressed as:

a) Increase in [A]

b) Decrease in [B]

c) Decrease in [A] or increase in [B] ✅

d) Constant value

9. At the beginning of a reaction, rate is:

a) Minimum

b) Zero

c) Maximum ✅

d) Constant

10. As the reaction proceeds, rate:

a) Increases

b) Remains constant

c) Decreases due to decrease in reactant concentration ✅

d) Becomes infinite

11. The slope of a concentration–time graph represents:

a) Energy

b) Temperature

c) Rate of reaction ✅

d) Pressure

12. A steeper slope at the start indicates:

a) Slow reaction

b) Fast reaction rate ✅

c) No reaction

d) Equilibrium

13. When reactant and product curves become parallel, it indicates:

a) Reaction start

b) Increase in rate

c) Completion of reaction (rate becomes zero) ✅

d) Temperature rise

14. Instantaneous rate is defined as:

a) Rate over a long time

b) Rate at a specific instant of time ✅

c) Average rate

d) Zero rate

15. Average rate is defined as:

a) Rate at a particular instant

b) Rate over a given time interval ✅

c) Constant rate

d) Maximum rate

16. At the beginning of reaction:

a) Instantaneous rate = Average rate

b) Instantaneous rate < Average rate

c) Instantaneous rate > Average rate ✅

d) Both are zero

17. At the end of time interval:

a) Instantaneous rate > Average rate

b) Instantaneous rate < Average rate ✅

c) Both are equal

d) Both are maximum

18. For gaseous reactions, rate is often expressed in terms of:

a) Volume

b) Temperature

c) Pressure ✅

d) Mass

19. Effective collisions lead to:

a) No change

b) Reactants only

c) Formation of products ✅

d) Decrease in energy only

20. Collision theory explains:

a) Chemical equilibrium

b) Reaction rate and mechanism of reactions ✅

c) Atomic structure

d) Thermodynamic stability

21. Measurement of rate of reaction involves:

a) Temperature change

b) Pressure change

c) Determination of concentration at regular time intervals ✅

d) Volume only

22. In a concentration-time graph, time is plotted on:

a) y-axis

b) x-axis ✅

c) Both axes

d) No axis

23. In the decomposition of HI, the concentration curve is:

a) Rising curve

b) Straight line

c) Falling curve ✅

d) Horizontal line

24. The steepness of the curve indicates:

a) Temperature

b) Pressure

c) Rate of reaction ✅

d) Volume

25. Greater slope at the start of reaction shows:

a) Slow reaction

b) Fast reaction rate ✅

c) No reaction

d) Equilibrium

26. The rate at a specific time is determined by:

a) Area under curve

b) Drawing a tangent to the curve ✅

c) Measuring volume

d) Changing temperature

27. The slope of tangent represents:

a) Average rate

b) Instantaneous rate of reaction ✅

c) Total energy

d) Activation energy

28. The formula used to calculate rate is:

a) Δt / Δc

b) Δc / Δt ✅

c) c × t

d) c / t^2

29. If change in concentration is 0.04 mol dm^{-3} in 100 s, rate is:

a) 0.0004 mol dm^{-3} s^{-1} ✅

b) 0.004 mol dm^{-3} s^{-1}

c) 0.04 mol dm^{-3} s^{-1}

d) 4 mol dm^{-3} s^{-1}

30. Chemical method is mainly suitable for:

a) Gas reactions

b) Solid reactions

c) Reactions in solution ✅

d) Nuclear reactions

31. In ester hydrolysis experiment, reaction is stopped by:

a) Heating

b) Cooling and dilution with ice-cold water ✅

c) Adding base

d) Increasing pressure

32. The acid formed in ester hydrolysis is determined by:

a) Filtration

b) Titration with standard NaOH ✅

c) Evaporation

d) Distillation

33. Phenolphthalein is used as:

a) Catalyst

b) Solvent

c) Indicator ✅

d) Reactant

34. Spectrophotometry is used when substance:

a) Is colorless

b) Absorbs radiation (UV/visible/IR) ✅

c) Is solid

d) Is insoluble

35. Electrical conductivity method is used for:

a) Non-ionic reactions

b) Ionic reactions ✅

c) Solid reactions

d) Nuclear reactions

36. The rate of a chemical reaction depends on:

a) Only temperature

b) Only concentration

c) Concentration, temperature, surface area, catalyst ✅

d) Pressure only

37. According to law of mass action, increasing concentration:

a) Decreases rate

b) Does not affect rate

c) Increases rate due to more collisions ✅

d) Stops reaction

38. Increase in concentration leads to:

a) Fewer collisions

b) No collisions

c) More frequent collisions ✅

d) Decrease in energy

39. For every 10°C rise in temperature, rate of reaction:

a) Remains same

b) Decreases

c) Doubles or triples ✅

d) Becomes zero

40. Increase in temperature increases rate because:

a) Molecules slow down

b) Collision frequency and energy increase ✅

c) Activation energy increases

d) Volume decreases

41. Boltzmann distribution curve shows:

a) Pressure vs time

b) Energy distribution of molecules ✅

c) Volume vs temperature

d) Rate vs concentration

42. Area under Boltzmann curve represents:

a) Activation energy

b) Total number of molecules ✅

c) Temperature

d) Rate constant

43. Shaded area in Boltzmann curve represents:

a) Low energy molecules

b) Total molecules

c) Molecules with energy ≥ activation energy ✅

d) Inactive molecules

44. At higher temperature, Boltzmann curve:

a) Becomes steeper

b) Shifts left

c) Flattens and shifts to right ✅

d) Disappears

45. Increase in temperature increases rate mainly due to:

a) Decrease in collisions

b) Increase in successful collisions ✅

c) Decrease in activation energy

d) Decrease in molecules

46. A catalyst is a substance which:

a) Is consumed in reaction

b) Remains unchanged after reaction and changes rate ✅

c) Stops reaction

d) Increases activation energy

47. Catalyst increases rate by:

a) Increasing temperature

b) Increasing concentration

c) Providing lower activation energy pathway ✅

d) Increasing pressure

48. Homogeneous catalysis involves:

a) Different phases

b) Same phase for reactants and catalyst ✅

c) Only solids

d) Only gases

49. Heterogeneous catalysis involves:

a) Same phase

b) Different phases of catalyst and reactants ✅

c) Only liquids

d) Only gases

50. Example of heterogeneous catalysis is:

a) $SO_2 + O_2$ with NO (gas phase)

b) Ester hydrolysis in solution

c) Oxidation of NH_3 using platinum (solid catalyst) ✅

d) Reaction in solution only

51. Rate law is defined as:

a) Relation between temperature and rate

b) Relation between rate and concentration of reactants ✅

c) Relation between pressure and volume

d) Relation between energy and time

52. General form of rate law is:

a) Rate = k[A+B]

b) Rate = $k[A]^x[B]^y$ ✅

c) Rate = A + B

d) Rate = k/AB

53. In rate law, k represents:

a) Concentration

b) Temperature

c) Rate constant ✅

d) Pressure

54. The value of rate constant depends on:

a) Concentration

b) Temperature ✅

c) Volume

d) Pressure only

55. Exponents in rate law (x, y) are:

a) Always equal to coefficients

b) Experimentally determined values ✅

c) Always zero

d) Always one

56. Order of reaction with respect to a reactant is:

a) Coefficient in equation

b) Exponent of its concentration in rate law ✅

c) Temperature value

d) Pressure value

57. Overall order of reaction is:

a) Product of exponents

b) Sum of exponents ✅

c) Difference of exponents

d) Average of exponents

58. For rate = $k[A]^2[B]$, overall order is:

a) 2

b) 1

c) 3 ✅

d) 4

59. In reaction rate = $k[H_2][NO]^2$, order with respect to NO is:

a) 1

b) 2 ✅

c) 3

d) 0

60. A zero-order reaction has rate:

a) Dependent on concentration

b) Independent of concentration ✅

c) Dependent on temperature only

d) Zero always

61. For zero-order reaction, rate law is:

a) Rate = $k[A]$

b) Rate = $k[A]^2$

c) Rate = $k[A]^0$ ✅

d) Rate = k/A

62. A first-order reaction has:

a) Sum of exponents = 0

b) Sum of exponents = 1 ✅

c) Sum of exponents = 2

d) Sum of exponents = 3

63. Example of first-order reaction is:

a) Formation of NH_3

b) Decomposition of N_2O_5 ✅

c) Combustion reaction

d) Neutralization

64. A second-order reaction has rate law:

a) Rate = k[A]

b) Rate = $k[A]^2$ or k[A][B] ✅

c) Rate = k

d) Rate = k/A

65. Fractional order reaction is one in which:

a) Order is zero

b) Order is integer

c) Order is fraction ✅

d) Order is negative

66. Rate constant (k) can be determined using:

a) Only graphical method

b) Initial concentration and half-life methods ✅

c) Only volume method

d) Only pressure method

67. In the initial concentration method, rate constant is determined using:

a) Final concentration

b) Initial concentrations of reactants and rate data ✅

c) Only temperature

d) Only pressure

68. Half-life of a reaction is defined as:

a) Time for completion of reaction

b) Time for concentration to become zero

c) Time for concentration to become half of its initial value ✅

d) Time for doubling of concentration

69. Half-life method is most commonly used for:

a) Zero order reactions

b) First order reactions ✅

c) Third order reactions

d) Fractional order reactions only

70. For a first-order reaction, the relation between half-life and rate constant is:

a) $K = t_1/_2 / 0.693$

b) $K = 0.693 / t_1/_2$ ✅

c) $K = t_1/_2 \times 0.693$

d) $K = 1 / t_1/_2{}^2$

71. If half-life of a first-order reaction is 2 hours, rate constant k is:

a) 0.346 h^{-1}

b) 0.693 h^{-1}

c) 0.693 / 2 = 0.346 h^{-1} ✅

d) 2 / 0.693

72. Reaction mechanism describes:

a) Only final products

b) Step-by-step pathway of reaction ✅

c) Only reactants

d) Only energy changes

73. Each step in a reaction mechanism is called:

a) Overall reaction

b) Elementary step ✅

c) Rate law

d) Intermediate step

74. The number of molecules involved in an elementary step is called:

a) Order of reaction

b) Molecularity ✅

c) Stoichiometry

d) Activation energy

75. A unimolecular reaction involves:

a) Two molecules

b) Three molecules

c) One molecule ✅

d) No molecule

76. Example of unimolecular reaction is:

a) $NO + O_3 \longrightarrow NO_2 + O_2$

b) $N_2O_5 \longrightarrow NO_2 + NO_3$ ✅

c) $H_2 + Cl_2 \longrightarrow 2HCl$

d) $CO + NO_2 \longrightarrow CO_2 + NO$

77. A bimolecular reaction involves:

a) One molecule

b) Two molecules ✅

c) Three molecules

d) Four molecules

78. Termolecular reactions involve:

a) One molecule

b) Two molecules

c) Three molecules ✅

d) Four molecules

79. Species formed in one step and consumed in another are called:

a) Reactants

b) Products

c) Intermediates ✅

d) Catalysts only

80. The slowest step in a reaction mechanism is called:

a) Fast step

b) Intermediate step

c) Rate-determining step ✅

d) Final step

❖ Important Short Questions:

1. Explain collision theory and discuss the conditions required for a reaction to occur.

Answer:

Collision theory states that chemical reactions occur when reactant particles collide with each other. However, not all collisions result in a reaction. For a reaction to occur, collisions must be effective.

Conditions for a successful collision:

- Particles must collide with sufficient energy (equal to or greater than activation energy).
- Particles must have the correct orientation during collision.
- Only effective collisions lead to product formation.

Example:

In the reaction between H_2 and Cl_2 to form HCl, only properly oriented and high-energy collisions result in bond breaking and formation of HCl.

2. Why do all collisions not lead to a chemical reaction?

Answer:

Not all collisions are effective because:

- Some collisions occur with insufficient energy (less than activation energy).
- Some collisions occur with improper orientation of molecules.

Therefore, only a small fraction of collisions are effective enough to form products.

Example:

In a reaction between two gas molecules, many collisions occur randomly, but only those with enough energy and correct alignment produce products.

3. What is the significance of activation energy in collision theory?

Answer:

Activation energy is the minimum energy required for reactants to form products.

Significance:

- Determines whether a collision will be effective or not.
- Higher activation energy means fewer effective collisions ⟶ slower reaction.
- Lower activation energy increases the number of effective collisions ⟶ faster reaction.

Example:

Combustion reactions require a spark because the activation energy must be overcome to start the reaction.

4. How does increasing temperature affect the rate of reaction according to collision theory?

Answer:

Increasing temperature:

- Increases the kinetic energy of particles.
- Increases the frequency of collisions.
- Increases the number of particles having energy ≥ activation energy.

Thus, the number of effective collisions increases, leading to a faster reaction.

Example:

Food spoils faster at higher temperatures because reactions proceed more quickly.

5. Explain the effect of concentration on reaction rate using collision theory.

Answer:

Increasing concentration means:

- More particles are present in a given volume.
- Collision frequency increases.
- Number of effective collisions increases.
- **Hence**, the rate of reaction increases with concentration.

Example:

A reaction between acids and metals occurs faster in a concentrated acid compared to a dilute one.

6. How does the presence of a catalyst affect the rate of reaction in terms of collision theory?

Answer:

A catalyst:

- Provides an alternative pathway with lower activation energy.
- Increases the number of effective collisions at a given temperature.
- Speeds up the reaction without being consumed.

Example:

In the decomposition of hydrogen peroxide, MnO_2 acts as a catalyst and speeds up the reaction.

7. Define rate law and explain why it is determined experimentally.

Answer:

Rate law is an equation that expresses the rate of a reaction in terms of the concentrations of reactants raised to certain powers.

General form:

Rate = k[A]^x[B]^y

It is determined experimentally because:

- The exponents (x, y) cannot be predicted from the balanced equation.
- They depend on the reaction mechanism, not stoichiometry.

Example:

For a reaction, experimentally determined rate law may be Rate = $k[NO]^2[H_2]$.

8. How does rate constant differ from rate of reaction?

Answer:

Rate of reaction:

- Change in concentration per unit time.
- Depends on concentration, temperature, etc.
- Unit: mol dm^{-3} s^{-1}

Rate constant (k):

- Proportionality constant in the rate law.
- Independent of concentration but depends on temperature.
- Unit depends on order of reaction.

Example:

Rate = $k[A]$ ⟶ rate changes with [A], but k remains constant at a given temperature.

9. Explain the relationship between order of reaction and reaction mechanism.

Answer:

- The order of a reaction is determined by the rate-determining step of the mechanism.

- The rate law reflects the molecularity of the slow step.

Overall order provides insight into how the reaction proceeds at the molecular level.

Example:

If the slow step involves two reactant molecules, the reaction may be second order overall.

10. For the rate law Rate = $k[A][B]^2$, determine the order with respect to each reactant and overall order.

Answer:

- Order with respect to A = 1
- Order with respect to B = 2
- Overall order = 1 + 2 = 3

Conclusion:

This is a third-order reaction.

Example interpretation:

Doubling [A] doubles the rate, while doubling [B] increases the rate by a factor of 4.

11. Why can the order of a reaction be different from the stoichiometric coefficients of the balanced equation?

Answer:

The order of a reaction is determined experimentally, while stoichiometric coefficients come from the balanced chemical equation. The balanced equation shows only the overall reaction, not the mechanism. Reaction rate depends on the rate-determining step, not the overall equation. Therefore, the order may differ from coefficients.

Example:

For the reaction:

$2NO + 2H_2 \longrightarrow N_2 + 2H_2O$

Experimental rate law: Rate = $k[H_2][NO]^2$

Here, the order does not directly match stoichiometric coefficients.

12. Explain zero order reactions and give a real-life or experimental example.

Answer:

A zero-order reaction is one in which the rate is independent of the concentration of reactants.

Rate law:

Rate = $k[A]^0 = k$

Characteristics:

- Rate remains constant regardless of concentration.
- The graph of concentration vs time is a straight line.

Example:

Decomposition of ammonia on a hot platinum surface (catalytic reaction) shows zero-order behavior under certain conditions.

13. Derive or state the relation between half-life and rate constant for a first-order reaction and explain its significance.

Answer:

For a first-order reaction, half-life is given by:

$t_1/_2 = 0.693 / k$

Significance:

- Half-life is independent of initial concentration.
- It is constant for a given reaction at a fixed temperature.
- Used to determine the rate constant easily.

Example:

If $t_1/_2$ = 2 hours, then k = 0.693 / 2 = 0.346 h^{-1}.

14. How can half-life be used to identify the order of a reaction?

Answer:

- For first-order reactions, half-life is constant and independent of initial concentration.
- For zero-order reactions, half-life depends on initial concentration.

- For second-order reactions, half-life is inversely proportional to initial concentration.

Conclusion:

By observing how half-life changes with concentration, the order of reaction can be identified.

Example:

If half-life remains the same despite changing concentration, the reaction is first order.

15. Explain the initial concentration method for determining the rate constant.

Answer:

In this method:

- The reaction rate is measured using different initial concentrations of reactants.
- The initial rate is determined for each set of concentrations.
- These values are substituted into the rate law to calculate the rate constant (k) and reaction order.

Example:

In the reaction of H_2O_2 with I^- ions, initial concentrations of reactants are varied and initial rates are measured to determine k.

16. What is a reaction mechanism and why is it important in chemical kinetics?

Answer:

A reaction mechanism is the step-by-step sequence of elementary reactions that describe how reactants are converted into products.

Importance:

- Explains how a reaction occurs at the molecular level.
- Helps identify the rate-determining step.
- Used to derive the rate law.

Example:

The decomposition of N_2O_5 occurs in steps involving intermediates like NO_2 and NO_3.

17. Differentiate between elementary steps and overall reaction.

Answer:

Elementary step:

- A single molecular event.
- Occurs in one step.
- Has its own molecularity.

Overall reaction:

- Sum of all elementary steps.

- Shows only reactants and products.
- Does not show intermediates.

Example:

Elementary steps may involve intermediates like radicals, while the overall equation does not include them.

18. What is molecularity? How is it different from order of reaction?

Answer:

Molecularity is the number of reactant molecules participating in an elementary step.

Difference:

- Molecularity applies only to elementary reactions.
- Order is determined experimentally for overall reaction.
- Molecularity is always a whole number; order can be fractional.

Example:

A bimolecular step involves two molecules colliding simultaneously.

19. Explain unimolecular, bimolecular, and termolecular reactions with examples.

Answer:

1. Unimolecular reaction: Involves one molecule.

Example: $N_2O_5 \longrightarrow NO_2 + NO_3$

2. Bimolecular reaction: Involves two molecules colliding.

Example: $NO + O_3 \longrightarrow NO_2 + O_2$

3. Termolecular reaction: Involves three molecules simultaneously (rare).

Example: $2O_2 + O \longrightarrow O_3 + O_2$

20. What is the rate-determining step and how does it control the overall reaction rate?

Answer:

The rate-determining step is the slowest step in a reaction mechanism.

Importance:

- It controls the overall reaction rate.
- The rate law depends on the reactants involved in this step.
- Faster steps do not affect the overall rate.

Example:

If the slow step involves reactants A and B, then the rate law will depend on [A] and [B].

📘 Exercise

Q.1 Four choices are given for each question. Select the correct choice.

I. The rate of reaction:

a) Increase as the reaction proceeds

b) Decreases as the reaction proceeds ✅

c) Remains the same as the reaction proceeds

d) May decrease or increase as the reaction proceeds

II. Increasing the temperature of a chemical reaction increases the rate because:

a) Both the collision frequency and collision energies of reactant molecules increase ✅

b) Collision frequency of reactant molecules increases

c) Activation energy increases

d) Activation energy decreases

III. Consider two reactions with different activation energies at the same temperature. The reaction with the lower activation energy will have:

a) A smaller rate constant

b) A larger rate constant ✅

c) The same rate constant

d) A rate constant that depends on the enthalpy change

IV. The order of a chemical reaction that is independent of concentration is:

a) Second order reaction

b) First order reaction

c) Zero order reaction ✅

d) Pseudo first order reaction

V. On a Boltzmann distribution curve, the area under the curve represents:

a) Activation energy of the reaction

b) Total number of molecules in the sample ✅

c) Average kinetic energy of the molecules

d) Rate constant of the reaction

VI. On a Boltzmann distribution curve, the activation energy (Ea) is represented by:

a) The height of the peak

b) The area under the entire curve

c) A vertical line drawn at a specific kinetic energy value ✅

d) The difference between the peak and the x-axis

VII. If we double the concentration of a reactant and the rate increases four times, the reaction is:

a) Second order ✅

b) First order

c) Third order

d) Zero order

VIII. The rate determining step in a multi-step reaction is:

a) Always the first step

b) Always the last step

c) The slowest step ✅

d) The fastest step

IX. The reaction: $2NO_2 \longrightarrow NO + NO_3$ (slow), $NO_3 + CO \longrightarrow CO_2 + NO_2$ (fast). The rate law is:

a) $R = k_1 [NO_2]^3$

b) $R = k_2 [NO_3][CO]$

c) $R = k_1 [NO_2]$

d) R = k $[NO_2]^2$ ✅

X. How does the presence of a catalyst affect the rate of a chemical reaction?

a) It always decreases the rate of the reaction

b) It always increases the rate of the reaction

c) It increases forward and decreases reverse reaction rate

d) It increases the rate of both forward and reverse reactions ✅

XI. On an energy profile diagram, the presence of a catalyst is represented by:

a) A higher peak representing activation energy

b) A lower peak representing activation energy ✅

c) A change in energy of reactants or products

d) A shift in equilibrium position

XII. The units of the rate constant (k) depend on the:

a) Activation energy of the reaction

b) Temperature of the reaction

c) Overall order of the reaction ✅

d) Stoichiometry of the balanced chemical equation

XIII. A first-order reaction has a half-life of 20 minutes. The value of its rate constant (k) is:

a) 0.05 min^{-1}

b) 0.693 min^{-1}

c) 0.0347 min^{-1} ✅

d) 13.86 min^{-1}

Q.2 Attempt the following short-answer questions:

a) What do you understand by the rate of a reaction?

The rate of a reaction is the speed at which reactants are converted into products. It is measured as the change in concentration of reactants or products per unit time.

b) Difference between enthalpy change of reaction and energy of activation of reaction:

Enthalpy change (ΔH) is the heat absorbed or released during a reaction, showing the difference between the energy of products and reactants.

Activation energy (Ea) is the minimum energy required for reactants to start a reaction and form products.

c) Difference between order and molecularity of a reaction:

Order of a reaction is the sum of powers of concentration terms in the rate law and is determined experimentally.

Molecularity is the number of reactant molecules that collide in an elementary step and is a theoretical concept.

d) Why does the instantaneous rate change during a reaction?

The instantaneous rate changes because the concentration of reactants decreases as the reaction proceeds, leading to fewer effective collisions between reacting particles.

e) Briefly summarize the effects of temperature and surface area on the rates of reactions:

An increase in temperature increases the rate of reaction because particles gain kinetic energy and collide more frequently with greater energy. Increasing surface area increases the rate because more particles are exposed for collision, leading to more effective reactions.

f) Justify that radioactive decay is always a first order reaction.

Radioactive decay follows first-order kinetics because the rate of decay depends only on the number (concentration) of undecayed nuclei present. It does not depend on external factors like temperature, pressure, or concentration of other substances. Therefore, the rate is directly proportional to the amount of radioactive substance remaining.

g) A reaction is second order with respect to a reactant. How is the rate affected if the concentration is doubled and reduced to half?

For a second-order reaction, rate ∝ $[A]^2$.

- **If concentration is doubled:** rate becomes $2^2 = 4$ times (increases fourfold).
- **If concentration is halved:** rate becomes $(1/2)^2 = 1/4$ (decreases to one-fourth).

h) What is meant by half-life and what is it used for?

Half-life is the time required for the concentration (or amount) of a reactant to reduce to half of its initial value. It is used to determine the stability of a substance, especially in radioactive decay and first-order reactions, and to calculate the rate constant.

i) Why does wood burn more rapidly in pure oxygen than in air?

Wood burns faster in pure oxygen because the concentration of oxygen is higher, leading to more frequent and effective collisions between oxygen and fuel particles, which increases the rate of combustion. In air, oxygen is only about 21%, so the reaction is slower.

j) A catalyst lowers the activation energy of a chemical reaction. Illustrate it.

A catalyst provides an alternative reaction pathway with lower activation energy. This can be illustrated by an energy profile diagram where:

- The uncatalyzed reaction has a higher peak (higher activation energy).
- The catalyzed reaction has a lower peak (reduced activation energy).

Thus, the energy barrier is reduced, allowing more reactant molecules to achieve the required energy and react faster.

k) The rate constant for a certain reaction is 3.5×10^{-4} s^{-1} at 25°C. What is the order of the reaction? Explain based on the units of the rate constant.

Answer:

The unit of the rate constant is s^{-1}, which corresponds to a first-order reaction. Therefore, the reaction is first order.

l) If the initial concentration of the reactant is 0.5 mol dm^{-3}, calculate the initial rate of the reaction.

Answer:

Formula: Rate = k × [A]

Rate = $(3.5 \times 10^{-4}) \times (0.5)$

Rate = 1.75×10^{-4} mol dm^{-3} s^{-1}

m) How would the rate of this reaction change if the concentration of the reactant were doubled?

Answer:

For a first-order reaction, rate is directly proportional to concentration.

If concentration is doubled, the rate will also double.

n) A certain first-order reaction has a rate constant of 2.5×10^{-3} s^{-1}. Calculate the half-life of the reaction in minutes.

Answer:

Formula: t1/2 = 0.693 / k

t1/2 = 0.693 / (2.5×10^{-3})

t1/2 = 277.2 seconds

Convert into minutes:

277.2 / 60 = 4.62 minutes

Final answer: 4.62 minutes

o) A radioactive isotope decays by a first-order process with a half-life of 12 hours. Calculate the rate constant for the decay in s^{-1}.

Answer:

Convert half-life into seconds:

12 hours = 12 × 3600 = 43200 seconds

Formula: $k = 0.693 / t1/2$

$k = 0.693 / 43200$

$k = 1.6 \times 10^{-5}\ s^{-1}$

❖ DESCRIPTIVE QUESTIONS

🌟 Q.3 Relate the order of a reaction to the rate law for the reaction. How do you distinguish between zero order, first order and second order reaction.

❖ **Answer:**

Introduction:

The order of a reaction is directly related to its rate law. The rate law shows how the rate of a reaction depends on the concentration of reactants. It is determined experimentally.

Rate Law and Order of Reaction:

A general rate law is written as:

Rate = k [A]^m [B]^n

In this equation, m and n represent the orders with respect to each reactant, and the overall order of the reaction is:

Overall order = m + n

These exponents indicate how the rate changes when the concentration of reactants changes.

Zero Order Reaction:

In a zero-order reaction, the rate is independent of the concentration of reactants.

Rate = k

This means that changing the concentration does not affect the rate of reaction.

First Order Reaction:

In a first-order reaction, the rate is directly proportional to the concentration of one reactant.

Rate = k[A]

If the concentration is doubled, the rate also doubles.

Second Order Reaction:

In a second-order reaction, the rate depends on the square of the concentration of a reactant or on the product of concentrations of two reactants.

Rate = k[A]^2 or Rate = k[A][B]

If the concentration is doubled, the rate becomes four times.

Conclusion:

The order of a reaction is determined from the rate law and shows the dependence of reaction rate on the concentration of reactants.

✦ **Summary:**

The order of a reaction is obtained from the powers of concentration terms in the rate law. Zero-order reactions have rate independent of concentration (Rate = k), first-order reactions depend linearly on concentration (Rate = k[A]), and second-order reactions depend on the square or product of concentrations (Rate = k[A]^2 or k[A][B]). The order helps predict how the rate changes with concentration.

🌟 Q.4 How do you find the numerical value of a rate constant by initial and half-life methods?

❖ **Answer:**

Introduction:

The rate constant (k) can be determined experimentally using different methods. Two common methods are the initial rate method and the half-life method. These methods depend on the order of the reaction and the data obtained from experiments.

Initial Rate Method:

In the initial rate method, the rate of reaction is measured at the very beginning when the concentration of reactants is known. The rate law is used to calculate the rate constant.

For example, for a reaction:

Rate = k [A]^n

Rearranging the equation:

k = Rate / [A]^n

By substituting the initial rate and initial concentration into this equation, the numerical value of the rate constant can be calculated.

Half-Life Method:

The half-life method uses the time required for the concentration of a reactant to reduce to half of its initial value.

For a first-order reaction, the half-life is given by:

- t1/2 = 0.693 / k
- Rearranging:
- k = 0.693 / t1/2

By measuring the half-life experimentally, the rate constant can be calculated.

Conclusion:

The initial rate method uses the rate and concentration values in the rate law to find k, while the half-life method uses the time taken for half the reactant to react. Both methods are useful depending on the order of the reaction and available data.

✦ **Summary:**

The rate constant can be determined by using the initial rate method (k = Rate / [A]^n) or the half-life method (k = 0.693 / t1/2 for first-order reactions). The initial rate method requires rate and concentration data, while the half-life method uses the time taken for the concentration to reduce to half.

🌟 Q.5 How does the activation energy profile of an uncatalyzed reaction compare with that of the catalyzed reaction?

❖ **Answer:**

Introduction:

The activation energy profile (energy profile diagram) shows the change in energy during a chemical reaction from reactants to products. It also indicates the activation energy required for the reaction to occur.

Uncatalyzed Reaction:

In an uncatalyzed reaction, the activation energy is relatively high. The energy profile diagram shows a single curve with a higher peak. Reactant molecules must gain sufficient energy to overcome this high energy barrier, so the reaction proceeds slowly.

Catalyzed Reaction:

In a catalyzed reaction, a catalyst provides an alternative pathway with lower activation energy. The energy profile diagram shows a curve with a lower peak compared to the uncatalyzed reaction. This lower energy barrier allows more reactant molecules to successfully collide and react, increasing the rate of reaction.

Comparison:

Both reactions have the same initial and final energy levels (reactants and products), meaning the overall enthalpy change (ΔH) remains unchanged. However, the catalyzed reaction has a lower activation energy than the uncatalyzed reaction.

Conclusion:

A catalyst reduces the activation energy by providing an alternative reaction pathway, which results in a lower energy peak in the activation energy profile and a faster reaction rate.

✦ Summary:

In an uncatalyzed reaction, the activation energy is high with a higher energy peak, while in a catalyzed reaction, the activation energy is lower with a smaller peak. The catalyst does not change the energy of reactants or products but only lowers the energy barrier, making the reaction faster.

🌟 Q.6 The reaction between hydrogen peroxide (H_2O_2) and iodide ions (I^-) in acidic solution is believed to occur via the following mechanism:

- Step 1: $H_2O_2 + I^- \longrightarrow H_2O + OI^-$ (slow)

- Step 2: $OI^- + H^+ \longrightarrow HOI$ (fast)
- Step 3: $HOI + I^- + H^+ \longrightarrow H_2O + I_2$ (fast)

i) Overall balanced equation for the reaction:

To obtain the overall reaction, add all steps and cancel intermediates:

$H_2O_2 + 2I^- + 2H^+ \longrightarrow I_2 + 2H_2O$

ii) Identify intermediates and catalysts:

Intermediates:

- Substances that are formed in one step and consumed in another.
- OI^- and HOI are intermediates because they appear in the mechanism but not in the overall equation.

Catalyst:

- No catalyst is present in this mechanism because no substance is used in one step and regenerated in another unchanged.

iii) What is the rate-determining step?

The rate-determining step is the slowest step in the mechanism.

Here, Step 1 is slow.

So, Step 1 is the rate-determining step.

iv) Write the rate equation for the reaction:

Since the slow step determines the rate, the rate law is based on Step 1:

Rate = k $[H_2O_2][I^-]$

This expression is written in terms of the reactants of the overall reaction.

✦ **Summary:**

The overall reaction is:

$H_2O_2 + 2I^- + 2H^+ \longrightarrow I_2 + 2H_2O$

OI^- and HOI are intermediates, no catalyst is present, Step 1 is the rate-determining step, and the rate law is:

Rate = k

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