CHEMISTRY

Unit 7

REACTION KINETICS

MCQs • Short Questions • Long Questions

Complete Study Booklet with Answer Keys

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Section 1 | Multiple Choice Questions

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

1. The rate of reaction:

(a) Increases 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

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

✓ **(a) Both the collision frequency and collision energies of reactant molecules increase**
(b) Collision frequency of reactant molecules increase
(c) Activation energy increases
(d) Activation energy decreases

3. 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

4. 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

5. 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

6. 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

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

✓ **(a) Second order**
(b) First order
(c) Third order
(d) Zero order

8. 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

9. The reaction NO2(g) + CO2(g) -> NO(g) + CO2(g) occurs in two steps. What is the rate law equation for this reaction?

(a) R = k1[NO2]3
(b) R = k2[NO3][CO]
(c) R = k1[NO2]
✓ **(d) R = k1[NO2]2**

10. 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 the rate of the forward and decreases the rate of the reverse reaction
✓ **(d) It increases the rate of both the forward and reverse reactions**

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

(a) A higher peak representing the activation energy
✓ **(b) A lower peak representing the activation energy**
(c) A change in the energy level of the reactants or products
(d) A shift in the equilibrium position

12. The units of the rate constant (k) for a reaction 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

13. A first-order reaction has a half-life (t1/2) of 20 minutes. What is the value of its rate constant (k)?

(a) 0.05 min-1
(b) 0.693 min-1
✓ **(c) 0.0347 min-1**
(d) 13.86 min-1

14. Effective collisions are those that:

✓ **(a) Have enough energy and proper orientation**
(b) Are elastic
(c) Are in vacuum
(d) Produce no change

15. Collision frequency depends on:

(a) Pressure
(b) Concentration
(c) Temperature
✓ **(d) All of these**

16. Which one reduces activation energy?

✓ **(a) Catalyst**
(b) Inhibitor
(c) Product
(d) Reactant

17. Which increases with rise in temperature?

(a) Activation energy
(b) Enthalpy
✓ **(c) Collision frequency**
(d) Molecular weight

18. Rate of reaction is measured as:

(a) Increase in temperature

✓ **(b) Change in concentration per unit time**

(c) Total energy of reactants

(d) Change in volume per mole

19. Units of rate for gaseous reaction are:

(a) mol

(b) atm

✓ **(c) mol dm-3 s-1**

(d) kg s-1

20. Which of the following methods measures reaction rate?

(a) Conductivity

(b) Titration

(c) Volume of gas

✓ **(d) All of these**

21. In rate = Delta [C]/Delta t, C refers to:

(a) Conductance

(b) Colour

✓ **(c) Concentration**

(d) Current

22. Rate is fastest:

✓ **(a) At start**

(b) In middle

(c) At end

(d) After completion

23. A flat curve in rate vs time graph indicates:

(a) High rate

(b) Constant rate

✓ **(c) Reaction completed**

(d) Temperature increase

24. Which factor affects rate of reaction?

(a) Concentration

(b) Temperature

(c) Surface area

✓ **(d) All of these**

25. Increase in surface area:

✓ **(a) Increases rate**

(b) Decreases rate

(c) No effect

(d) Stops reaction

26. Effect of temperature on rate is:

(a) Negligible

(b) Linear

✓ **(c) Exponential**

(d) Logarithmic

27. In a chemical reaction, a catalyst:

(a) Alters Delta H
✓ **(b) Lowers activation energy**
(c) Is consumed
(d) Forms product

28. Higher pressure increases rate for:
(a) Solids
(b) Liquids
✓ **(c) Gases**
(d) All phases

29. Rate law is an expression involving:
(a) Pressure
✓ **(b) Concentration**
(c) Temperature
(d) Catalyst

30. Rate = k[A]■ defines:
(a) Stoichiometry
✓ **(b) Reaction order**
(c) Reaction mechanism
(d) Titration

31. Overall order is sum of:
(a) Coefficients
✓ **(b) Exponents in rate law**
(c) Moles
(d) Products

32. A reaction is first-order if:
✓ **(a) Rate is proportional to [A]**
(b) Rate is proportional to [A]2
(c) Rate is constant
(d) Rate is proportional to √[A]

33. A zero-order reaction has rate:
(a) Proportional to [A]
✓ **(b) Independent of [A]**
(c) Proportional to time
(d) Zero

34. Units of rate constant depend on:
(a) Temperature
(b) Pressure
✓ **(c) Order of reaction**
(d) Surface area

35. Unit of rate constant for zero-order reaction is:
✓ **(a) mol dm-3 s-1**
(b) s-1
(c) mol-1 dm3 s-1
(d) mol-2 dm6 s-1

36. Unit of rate constant for first-order reaction is:
(a) mol dm-3

(b) mol-1 dm3 s-1
✓ **(c) s-1**
(d) J mol-1

37. Unit of k for second-order reaction is:
(a) mol dm-3
(b) s-1
✓ **(c) mol-1 dm3 s-1**
(d) mol-2 dm6 s-1

38. For nth order reaction, unit of k is:
✓ **(a) mol^(n-1) dm^(3n-1) s-1**
(b) mol^(-n) dm^(3n) s-1
(c) mol-1 dm-3 s
(d) Depends on time only

39. k has unit s-1 for:
(a) Zero-order
✓ **(b) First-order**
(c) Second-order
(d) Third-order

40. Which unit is common for all orders?
(a) mol dm-3
✓ **(b) No common unit**
(c) J mol-1
(d) s-1

41. Rate constant can be determined by:
(a) Titration
(b) Spectroscopy
(c) Conductometry
✓ **(d) All of these**

42. For gaseous reactions, rate is often monitored by:
(a) Mass change
(b) pH change
✓ **(c) Pressure change**
(d) Temperature change

43. In color change reactions, which method is best?
(a) Titration
(b) Conductometry
✓ **(c) Colorimetry**
(d) Manometry

44. Reaction rate by conductometry depends on:
✓ **(a) Ionic conductivity**
(b) Pressure
(c) Temperature
(d) Colour

45. Rate constant is independent of:
✓ **(a) Concentration**
(b) Temperature

(c) Catalyst
(d) Time

46. First-order reactions are best studied using:
(a) Pressure vs time
(b) [A] vs time
✓ **(c) log [A] vs time**
(d) Titration

47. Half-life is time in which:
(a) Rate becomes zero
(b) Concentration reduces to 1/4
✓ **(c) Concentration reduces to half**
(d) Product is half-formed

48. Half-life of first-order reaction:
(a) Increases with [A]
✓ **(b) Is independent of [A]**
(c) Decreases with [A]
(d) Becomes zero

49. Half-life of zero-order reaction is:
(a) Constant
✓ **(b) Proportional to [A]**
(c) Inversely proportional to [A]
(d) Independent of rate constant

50. Half-life formula for 1st-order reaction is:
✓ **(a) 0.693/k**
(b) k x t
(c) 1/k
(d) 2k

51. Which order has constant half-life?
(a) Zero
✓ **(b) First**
(c) Second
(d) Third

52. Unit of half-life is:
(a) mol
✓ **(b) s**
(c) mol dm-3
(d) J

53. The slowest step in a mechanism is:
(a) Fast step
(b) Activation step
✓ **(c) Rate-determining step**
(d) Initiation step

54. A reaction mechanism describes:
(a) Experimental conditions
(b) Overall stoichiometry
✓ **(c) Stepwise molecular events**

(d) Heat of reaction

55. Intermediates are:

(a) Reactants
(b) Products
(c) Stable compounds
✓ **(d) Short-lived species formed and consumed**

56. Rate-determining step controls:

(a) Temperature
✓ **(b) Overall rate**
(c) Volume
(d) Final yield

57. Mechanism supports:

(a) Hess's Law
✓ **(b) Kinetic data**
(c) Le Chatelier's principle
(d) Boyle's law

58. Catalysts usually act in:

(a) Initiation step
(b) Fast step
(c) Intermediate formation
✓ **(d) Rate-determining step**

59. In energy profile diagram a catalyst:

(a) Increases activation energy
(b) Decreases activation energy
(c) Provides new pathway
✓ **(d) Both b and c**

60. Enzyme is a:

(a) Solid catalyst
(b) Homogeneous catalyst
✓ **(c) Biological catalyst**
(d) Promoter

61. In an energy profile, peak represents:

(a) Reactants
(b) Products
✓ **(c) Activated complex**
(d) Catalyst

62. Energy profile shows:

(a) Concentration
(b) Temperature
✓ **(c) Energy vs reaction progress**
(d) Time

63. Homogeneous catalysts are in:

✓ **(a) Same phase as reactants**
(b) Different phase
(c) Solid only
(d) Gaseous only

64. In presence of catalyst, Delta H:

(a) Increases
(b) Decreases
✓ **(c) Unchanged**
(d) Doubles

65. A reaction is spontaneous when:

(a) Delta G > 0
✓ **(b) Delta G < 0**
(c) Delta H = 0
(d) Delta S = 0

66. Gibbs free energy equation is:

(a) Delta G = Delta H + TDelta S
✓ **(b) Delta G = Delta H - TDelta S**
(c) Delta G = Delta S - TDelta H
(d) Delta G = Delta S + TDelta H

67. At equilibrium, Delta G is:

(a) Positive
✓ **(b) Zero**
(c) Negative
(d) Maximum

68. Negative Delta H and positive Delta S implies:

(a) Non-spontaneous
✓ **(b) Always spontaneous**
(c) Only spontaneous at high temp
(d) Only spontaneous at low temp

69. Delta G tells about:

(a) Speed
(b) Mechanism
✓ **(c) Feasibility**
(d) All of these

70. If Delta G = 0, the system is:

✓ **(a) At equilibrium**
(b) Not reacting
(c) Spontaneous
(d) Incomplete

Section 2 | Short Questions

Q10. A catalyst lowers the activation energy of a chemical reaction. Illustrate it.

In an energy profile diagram, a catalyst provides an alternative pathway with lower activation energy, making it easier for reactants to convert into products.

The catalyst lowers the energy barrier without changing the overall enthalpy change (Delta H) of the reaction.

Q11. The rate constant for a certain reaction is 3.5x10^-4 s^-1 at 25 deg C. What is the order of the reaction? Explain based on the units.

Given: k = 3.5x10^-4 s^-1.

The unit 's^-1' implies it is a first-order reaction (per the standard units for rate constants).

Q12. If the initial concentration of the reactant is 0.50 mol dm-3, calculate the initial rate of the reaction.

Initial concentration = 0.50 mol dm-3; k = 3.5x10^-4 s^-1.

For a first-order reaction: Rate = k[A] = (3.5x10^-4)(0.50) = 1.75x10^-4 mol dm-3 s-1

Q13. How would the rate of a first-order reaction change if the concentration of the reactant were doubled?

Since it is first order, rate is directly proportional to concentration: Rate = k[A].

Rate doubles when concentration is doubled.

Q14. A first-order reaction has a rate constant of 2.5x10^-3 s^-1. Calculate the half-life in minutes.

t(1/2) = 0.693/k = 0.693/(2.5x10^-3) = 277.2 seconds

t(1/2) = 277.2/60 = 4.62 minutes

Q15. A radioactive isotope decays by a first-order process with a half-life of 12 hours. Calculate the rate constant in s^-1.

t(1/2) = 12 hours = 12x60x60 = 43,200 seconds

k = 0.693/t(1/2) = 0.693/43200 = 1.60x10^-5 s^-1

Q16. What is reaction kinetics?

Reaction kinetics is the study of the rates of chemical reactions and the factors affecting those rates; it also explains the mechanism by which a reaction occurs.

Example: the decomposition of hydrogen peroxide (H2O2) into water and oxygen is studied to find its rate and mechanism.

Q17. What is the rate of a chemical reaction?

It is the change in concentration of reactants or products per unit time: Rate = Delta[Product]/Delta t = -Delta[Reactant]/Delta t.

Example: if HCl forms at 0.02 mol/dm3 in 10 seconds, the rate = 0.002 mol/dm3/s.

Q18. How is the rate of reaction expressed for a general reaction A + B -> C?

Rate = -d[A]/dt = -d[B]/dt = d[C]/dt

Q19. What are the units of rate of reaction?

The units are usually mol dm-3 s-1, representing moles formed or used per litre per second.

Q20. Differentiate between average rate and instantaneous rate.

Average rate: calculated over a time interval; Avg. Rate = Delta[Concentration]/Delta t. Example: average rate over 10 seconds.

Instantaneous rate: the rate at a particular moment, given by the slope of the concentration vs time curve. Example: rate at exactly 20 seconds.

Q21. How does temperature affect the rate of reaction?

Increasing temperature increases the kinetic energy of molecules, causing more effective collisions and thus increasing the rate.

Example: the rate of rusting increases in summer due to higher temperature.

Q22. What is the effect of concentration on rate of reaction?

Higher concentration means more particles in a given volume, increasing collision frequency and thus the rate.

Example: strong HCl reacts faster with Zn than dilute HCl.

Q23. How does surface area influence the reaction rate?

A larger surface area (finely divided solids) provides more area for collisions, increasing the rate.

Example: powdered magnesium burns faster than a magnesium ribbon.

Q24. What role does a catalyst play in a chemical reaction?

A catalyst increases the rate by providing an alternative pathway with lower activation energy, without being consumed.

Example: MnO2 speeds up H2O2 decomposition.

Q25. What is the collision theory of reaction rate?

It states that for a reaction to occur, particles must collide with proper orientation and sufficient energy (activation energy).

Example: CH4 + Cl2 reacts in sunlight due to energetic collisions.

Q26. Define activation energy.

The minimum energy that reacting particles must have to form the activated complex and convert into products.

Example: in the combustion of wood, heat provides the activation energy needed to start burning.

Q27. How does a catalyst lower activation energy?

It provides a new reaction pathway that requires less energy for the same reaction to proceed.

Q28. What is a reaction mechanism?

It is a step-by-step description of a chemical reaction showing how reactants convert into products through elementary steps.

Example for 2NO2 -> 2NO + O2: Step 1: NO2 -> NO + O (slow); Step 2: NO2 + O -> NO + O2 (fast).

Q29. What is the rate-determining step?

The slowest step in a reaction mechanism, which determines the overall rate.

Example: in the NO2 mechanism above, Step 1 is rate-determining.

Q30. What is molecularity of a reaction?

It is the number of reactant molecules involved in an elementary step.

Q31. Define order of a reaction.

It is the sum of the powers of the concentration terms in the rate law.

Example: Rate = k[A][B]^2 -> order = 1 + 2 = 3.

Q32. How is order of reaction determined experimentally?

By observing how changing concentrations affect the rate and comparing the results using rate laws.

Q33. Write a general rate law and define its components.

Rate = k[A]^m[B]^n, where k = rate constant; m, n = orders with respect to A and B; [A], [B] = concentrations of reactants.

Q34. What are the units of the rate constant (k) for different orders?

Zero order: mol dm-3 s-1; First order: s-1; Second order: mol-1 dm3 s-1; Third order: mol-2 dm6 s-1.

Q35. Name two methods to determine rate of reaction experimentally.

Measuring the volume of gas evolved, and monitoring colour change or turbidity.

Example: measuring O2 evolved during H2O2 decomposition.

Q36. How can rate be determined from a concentration vs time graph?

By finding the slope (tangent) of the curve at a given time.

Q37. What is the difference between homogeneous and heterogeneous catalysis?

Homogeneous catalysis: catalyst and reactants are in the same phase (e.g. H+ in ester hydrolysis).

Heterogeneous catalysis: catalyst and reactants are in different phases (e.g. Ni solid in hydrogenation of alkenes).

Q38. Give an example of enzyme catalysis.

Amylase catalyzes the breakdown of starch to maltose in the mouth (biological catalysis).

Q39. Why are catalysts important in industry?

They increase production rate, reduce energy costs, and improve yield.

Example: the iron catalyst used in the Haber process.

Q40. What is a reaction profile diagram?

It shows the energy changes during a reaction, including the activation energy and the enthalpy difference between reactants and products.

Q41. How does a catalyst change the reaction profile?

It lowers the peak (activation energy) without changing the energy levels of the reactants or products.

Q42. Can a reaction be fast but thermodynamically unfavorable?

Yes - kinetics (rate) and thermodynamics (feasibility) are different considerations.

Example: the conversion of diamond to graphite is thermodynamically favoured but extremely slow.

Q43. How is the concept of reaction kinetics used in medicine?

It is used in drug design to control the rate of drug reactions or breakdown in the body.

Q44. What role does the activation energy play in chemical reactions?

Activation energy (Ea) is the minimum energy required for reactant molecules to form the activated complex and proceed to products.

It acts like an energy barrier that must be crossed for a reaction to occur - molecules must have at least this much energy to react.

Section 3 | Descriptive & Numerical Problems

Descriptive Questions

Q3 (Descriptive). Relate the order of a reaction to the rate law. How do you distinguish between zero, first and second order reactions?

Zero order: Rate = k[A]^0 = k - rate is independent of concentration; doubling concentration does not change the rate (common in surface-catalyzed and photochemical reactions).

First order: Rate = k[A]^1 = k[A] - doubling [A] doubles the rate; follows exponential decay (e.g. radioactive decay, N2O5 decomposition).

Second order: Rate = k[A]^2 or k[A][B] - doubling a single reactant's concentration increases the rate 4-fold; units of k are mol-1 dm3 s-1 (e.g. NO + O3 reaction).

Q4 (Descriptive). How do you find the numerical value of a rate constant by initial concentration and half-life methods?

Initial concentration method: substitute the rate and known concentrations into the rate law and solve for k, e.g. k = Rate/([H2O2][I-]).

Half-life method: for a first-order reaction, k = 0.693/t(1/2), using the experimentally measured half-life.

Q5 (Descriptive). How does the activation energy profile of an uncatalyzed reaction compare with that of a catalyzed reaction?

A catalyst provides an alternative reaction pathway with a lower activation energy peak on the energy profile diagram.

The energy levels of the reactants and products (and therefore Delta H) remain unchanged; only the height of the activation energy barrier is reduced, allowing more molecules to react per unit time.

Q6 (Descriptive). H2O2 + I- in acidic solution occurs via: Step1 H2O2+I- -> H2O+OI- (slow); Step2 OI-+H+ -> HOI (fast); Step3 HOI+I-+H+ -> H2O+I2 (fast). Analyze the mechanism.

(i) Overall balanced equation: H2O2 + 2I- + 2H+ -> 2H2O + I2

(ii) Intermediate: HOI. There is no catalyst in this reaction mechanism.

(iii) Rate-determining step: Step 1, since it is the slow step.

(iv) Rate equation (from the slow step): Rate = k[H2O2][I-]

Numerical Problems

Q6 (Numerical). Calculate the reaction rate given [A]=0.5 M, [B]=0.2 M, k=4.0 M-2 s-1, and Rate = k[A][B]^2.

Rate = k[A][B]^2 = 4.0 x (0.5) x (0.2)^2 = 4.0 x 0.5 x 0.04

Rate = 0.08 mol L-1 s-1

Q7 (Numerical). A first-order reaction has k = 5.5x10^-14 s^-1. Find the half-life.

t(1/2) = 0.693/k = 0.693/(5.5x10^-14) = 1.26x10^13 seconds

In years: t(1/2) = (1.26x10^13)/(60x60x24x365) is approx 399,365 years is approx 3.99x10^5 years

Q8 (Numerical). For 2HI(g) -> H2(g)+I2(g), three experiments gave: [HI]=0.015 M, Rate=1.1x10^-3; [HI]=0.030 M, Rate=4.4x10^-3; [HI]=0.045 M, Rate=9.9x10^-3 (M/s). Find the rate law and rate constant.

Comparing Experiments 1 and 2: [HI] doubles (0.015->0.030); Rate increases 4-fold (1.1x10^-3 -> 4.4x10^-3).

Since Rate2/Rate1 = 4 = 2^n, n = 2 - the reaction is second order with respect to HI.

Rate Law: Rate = k[HI]^2

Using Experiment 1: 1.1x10^-3 = k(0.015)^2, so k = (1.1x10^-3)/(2.25x10^-4) = 4.89 M-1 s-1

Answer: Rate Law = k[HI]^2; Rate Constant k = 4.89 M-1 s-1

Section 4 | Long Questions

Q1. Discuss reaction kinetics and the collision theory of chemical reactions.

Reaction Kinetics

Definition: Reaction kinetics is the study of the rates of chemical reactions and the factors that affect them.

It involves a variety of experimental methods, measuring reaction rates, determining reaction orders, and understanding reaction mechanisms.

Reaction rates vary widely: e.g. NaCl with AgNO3 is very fast, hydrolysis of an ester is moderate, and rusting of iron is very slow.

Importance in industry: determines the economic feasibility of a reaction on a commercial scale, and affects production efficiency, safety, and cost-effectiveness.

Collision Theory

Collision theory explains how and why chemical reactions occur at the molecular level.

Basic requirements: reactant particles must form a homogeneous mixture, they must collide with each other, and the outcome depends on the energy of the colliding molecules.

Effective collisions lead to product formation; ineffective collisions leave particles unchanged after they bounce back.

Conditions for Effective Collisions

Sufficient Energy: colliding particles must possess at least the activation energy (Ea).

Proper Orientation: particles must approach each other in a specific orientation that allows bond formation or breaking.

Quick Check 7.1

(a) Activation energy's role: molecules must collide frequently (collision frequency) and be properly oriented during collision. Without frequent, correctly oriented collisions, molecules will not overcome the activation energy barrier and no reaction occurs.

(b) Effect on rate: if Ea is low, more molecules have enough energy to react, so the reaction is faster; if Ea is high, fewer molecules can react, so the reaction is slower. Conclusion: lower Ea gives a faster rate; higher Ea gives a slower rate.

Q2. What happens to reactant and product concentration during a reaction? Explain instantaneous and average rate.

Rate of Reaction

Definition: 'the change in concentration of a reactant or product divided by the time taken for the change.'

Rate = Delta[Concentration]/Delta Time = Delta x/Delta t, where Delta x = change in concentration, Delta t = time interval.

During a reaction, reactant concentration decreases (being consumed) while product concentration increases (being formed).

Units and Expression of Rate

Rate = mol dm-3/s = mol dm-3 s-1 (or mol dm-3 min-1, mol dm-3 h-1 for slower reactions).

For A -> B: Rate = -Delta[A]/Delta t = +Delta[B]/Delta t (negative sign shows A decreasing, positive shows B increasing).

Average Rate vs Instantaneous Rate

Average rate: the rate between two specific time intervals, giving a general idea of reaction speed - Rate = Delta C/Delta t.

Instantaneous rate: the rate at any one specific instant during the reaction - Rate = d[Reactant or Product]/dt.

Quick Check 7.2

H2(g)+I2(g) -> 2HI(g). At 100.0 s, iodine concentration fell from 0.010 to 0.0080 mol dm-3.

Delta[I2] = 0.010-0.0080 = 0.0020 mol dm-3; Delta t = 100.0 s.

Average rate = 0.0020/100.0 = 2.0x10^-5 mol dm-3 s-1

Q3. How is the rate of a chemical reaction measured?

Measurement of Reaction Rate

Scientists monitor how the concentration of reactants or products changes with time at regular intervals.

Steps: (1) take samples at regular time intervals; (2) determine the concentration of a reactant (decreasing) or product (increasing) in each sample; (3) plot a graph of concentration (y-axis) against time (x-axis).

Example: Decomposition of HI

2HI(g) -> H2(g)+I2(g) at 508 deg C. Data (concentration of HI vs time) is plotted to give a decay curve.

To find the instantaneous rate at a specific time (e.g. 100 s): draw a tangent to the curve at that point, form a right triangle with the tangent, and calculate its slope.

Example: Rate = Delta C/Delta t = 0.04 mol dm-3/100 s = 4x10^-4 mol dm-3 s-1

Quick Check 7.3

(a) Plot concentration of HI (y-axis) against time in seconds (x-axis) and connect the points with a smooth curve.

(b)-(c) Draw a tangent at the point of interest (e.g. t=300 s or specific concentrations) and use two points on the tangent to calculate its slope (the rate).

(d) The rate of reaction decreases with time, since fewer effective collisions occur as reactant concentration falls.

(e) The rate is highest when [HI] is highest (early in the reaction) and lowest when [HI] is lowest (later in the reaction).

Q4. What are the two main types of methods used to measure concentration changes during a reaction?

(a) Chemical Method

Example - Hydrolysis of Ester: CH3COOC2H5(l)+H2O(l) -> CH3COOH(l)+C2H5OH(l).

Procedure: mix ethyl acetate, water, and a small amount of acid catalyst; withdraw samples at fixed time intervals; quench each sample immediately in about four times its volume of ice-cold water to stop the reaction; titrate against standard NaOH using phenolphthalein indicator.

Purpose: titrating at various time intervals measures the concentration of acetic acid formed over time.

(b) Physical Methods

Spectrophotometry/Colorimetry: used when a reactant or product absorbs UV, visible, or IR radiation.

Electrical (Conductometric) Method: for reactions involving ionic species, by measuring the change in conductivity.

Volume Change Method: for reactions where gas volume changes.

Stopped-flow spectrophotometry: for very fast reactions - small volumes of reactants are driven at high speed into a mixing chamber, then to an observation cell where the reaction's progress is monitored, usually via UV transmission.

Q5. Explain the major factors affecting the rate of a chemical reaction.

(i) Concentration

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

Higher concentration means more particles per unit volume, increasing collision frequency, and hence a greater chance of effective collisions and a faster reaction.

(ii) Temperature

Higher temperature makes molecules collide more frequently and more energetically; a larger fraction of molecules then has energy equal to or greater than Ea. Rate roughly doubles or triples for every 10 deg C rise.

Boltzmann Distribution Curve: shows the distribution of kinetic energies among particles; most particles have moderate energy, and Ea acts as a threshold - only particles with energy at least Ea can react. As temperature rises, the curve flattens and shifts right, increasing the fraction of molecules beyond Ea.

(iii) Catalyst

Definition: a substance that increases reaction rate without being consumed itself.

It provides a new reaction pathway with a lower activation energy, allowing more molecules to have sufficient energy to react, without changing Delta H of the reaction.

(iv) Surface Area

A larger surface area increases the number of collisions between reactant particles, giving a faster rate.

Example: powdered solids react faster than larger pieces of the same solid.

Q6. What is catalysis? Explain its types.

Catalysis

Definition: the process which takes place in the presence of a catalyst.

(i) Homogeneous Catalysis

Definition: the catalyst and reactants are in the same phase, with the catalyst distributed uniformly throughout the system.

Examples: Contact Process - 2SO2(g)+O2(g) -> 2SO3(g) with NO gas as catalyst (all gases); Hydrolysis of ester - CH3COOC2H5(l)+H2O(l) -> CH3COOH(aq)+C2H5OH(aq) with H+ as catalyst (all in solution).

(ii) Heterogeneous Catalysis

Definition: the catalyst and reactants are in different phases - usually a solid catalyst with gaseous or liquid reactants.

Examples: Oxidation of Ammonia - 4NH3(g)+5O2(g) -> 4NO(g)+6H2O(g) with Pt gauze as catalyst; Hydrogenation of Alkenes - C2H4(g)+H2(g) -> C2H6(g) with Ni catalyst at 150 deg C.

Vitamins can act as coenzymes, helping enzymes catalyze biochemical reactions - e.g. Vitamin K is necessary for blood clotting, and low levels can cause bleeding diathesis.

Quick Check 7.5

(a) A catalyst is not consumed in a reaction - it lowers activation energy via an alternative pathway, taking part in the mechanism but being regenerated at the end.

(b) 2SO2(g)+O2(g) -> 2SO3(g) with V2O5 solid catalyst is heterogeneous catalysis, since the gaseous reactants and solid catalyst are in different phases.

(c) An energy profile diagram for uncatalyzed vs enzyme-catalyzed reactions shows the catalyzed pathway with a visibly lower activation energy peak between the same reactant and product energy levels.

Q7. What is a rate law and rate constant?

Rate Law

Definition: the representation of the rate of a reaction in terms of the concentration of the reactants.

For aA+bB -> cC+dD: Rate = k[A]^x[B]^y, where x and y are experimentally determined, [A] and [B] are molar concentrations, and k is the rate constant.

Rate Constant (k)

Definition: the rate of reaction when the concentrations of the reactants are unity (1 mol dm-3).

If [A]=1 and [B]=1 mol dm-3, Rate = k x 1^x x 1^y = k.

Properties: k is constant under given conditions but changes with temperature.

Q8. What is order of reaction? Explain the types of reaction orders.

Reaction Order

Definition: 'the order of a reaction with respect to a specific reactant is the exponent applied to that reactant's concentration in the rate equation.'

For aA+bB -> cC+dD, Rate = k[A]^x[B]^y: x = order w.r.t. A, y = order w.r.t. B, and overall order = x+y.

(i) Zero Order Reaction

Sum of exponents = 0. Rate Law: Rate = k[A]^0 = k.

Rate is independent of concentration; doubling concentration does not change the rate. Common in surface-catalyzed and photochemical reactions.

Example: decomposition of NH3 on a hot platinum surface.

(ii) First Order Reaction

Sum of exponents = 1. Rate Law: Rate = k[A]^1 = k[A].

Doubling [A] doubles the rate; tripling triples it; follows exponential decay. Common in decomposition and radioactive decay processes.

Example: 2N2O5(g) -> 2N2O4(g)+O2(g), Rate = k[N2O5]. All radioactive disintegration reactions are first order.

(iii) Second Order Reaction

Sum of exponents = 2. Two common forms: Rate = k[A]^2 (single reactant) or Rate = k[A][B] (two reactants).

Doubling [A] in Rate=k[A]^2 increases rate 4-fold; doubling both A and B in Rate=k[A][B] also increases rate 4-fold. Units of k: mol-1 dm3 s-1.

Example: NO(g)+O3(g) -> NO2(g)+O2(g), Rate = k[NO][O3] - first order in each, overall order = 2.

Note: the order of a reaction is an experimentally determined quantity and cannot be inferred simply from the balanced chemical equation.

Q10. How do you find the numerical value of a rate constant by initial concentration and half-life methods?

(i) Initial Concentration Method

Example: H2O2(aq)+2I-(aq)+2H+(aq) -> 2H2O(l)+I2(aq), Rate = k[H2O2][I-].

Using Experiment 1 data ([H2O2]=0.0200, [I-]=0.0100, Rate=3.50x10^-6 mol dm-3 s-1):

k = Rate/([H2O2][I-]) = 3.50x10^-6/(0.0200x0.0100) = 1.75x10^-2 dm3 mol-1 s-1

(ii) Half-Life Method

Half-life (t(1/2)): 'the time taken for the concentration of a reactant to fall to half of its original value.'

For a first-order reaction: k = 0.693/t(1/2).

Example: decomposition of H2O2 with half-life = 2 hours: t(1/2) = 2x60x60 = 7200 s; k = 0.693/7200 = 9.6x10^-5 s^-1

Quick Check 7.8

First-order reaction with half-life 15 minutes; initial concentration 0.100 mol dm-3. Find k.

k = 0.693/15 min = 0.0462 min-1. In seconds: 15 min = 900 s, so k = 0.693/900 = 7.7x10^-4 s^-1

Q11. Briefly discuss the reaction mechanism and rate-determining step.

Reaction Mechanism

A detailed, step-by-step description of how a chemical reaction occurs at the molecular level to yield the products.

It shows the actual elementary steps through which reactants turn into products; each step represents a single molecular event (bond breaking/forming); most reactions proceed through multiple steps.

Molecularity and Intermediates

Molecularity: the number of reactant molecules involved in an elementary step.

Intermediates: short-lived species produced in one step and consumed in a later step - not stable, and do not appear in the balanced chemical equation. Example: the carbocation (CH3)3C+.

Rate-Determining Step

Definition: the slowest step in a reaction mechanism.

Importance: controls the overall rate; only reactants involved in this step appear in the rate law; it determines the order of the reaction.

Worked Example

2NO(g)+2H2(g) -> 2H2O(g)+N2(g); experimentally, Rate = k[NO]^2[H2].

Doubling [H2] doubles the rate (first order in H2); doubling [NO] increases rate 4-fold, tripling increases it 9-fold (second order in NO).

Overall rate equation: Rate = k[H2][NO]^2

Sample Problem 7.1

2N2(g)+3H2(g) -> 2NH3(g): [NH3] changes from 3.5 mol dm-3 at 1.0 min to 6.2 mol dm-3 at 4.0 min.

Instantaneous rate at 1.0 min = Delta C/Delta t = 2.7 mol dm-3/1 min = 2.7 mol dm-3 min-1 (as given in the problem's data).

Average rate between 1.0 and 4.0 min: Delta[NH3]=6.2-3.5=2.7 mol dm-3; Delta t=3.0 min; Rate = 2.7/3.0 = 0.90 mol dm-3 min-1

Sample Problem 7.2

The first-order reaction cyclopropane -> propene has a half-life of 17.0 min. Find k.

k = 0.693/(17x60) = 6.79x10^-4 s^-1

Section 5 | Answer Key & Formula Summary

MCQ Answer Key (Q1-Q70)

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

Important Formulas

Formula	Description
Rate = Delta[Product]/Delta t = -Delta[Reactant]/Delta t	Rate of Reaction
Rate = k[A]^m [B]^n	General Rate Law
t(1/2) = 0.693 / k	Half-life of a First-Order Reaction
Rate = k[A]	First-Order Rate Law
Rate = k[A]^2	Second-Order Rate Law
Rate = k	Zero-Order Rate Law
Overall order = x + y	Sum of individual orders in rate law

Key Constants & Notes

Constant/Note	Value
Rate roughly doubles/triples	per 10 deg C rise in temperature

Activation Energy (Ea)	Minimum energy needed for an effective collision
HI decomposition example temperature	508 deg C
Half-life of H2O2 decomposition (example)	2 hours

Units of Rate Constant (k) by Order

Order	Rate Law	Units of k
Zero	Rate = k	mol dm-3 s-1
First	Rate = k[A]	s-1
Second	Rate = k[A]^2 or k[A][B]	mol-1 dm3 s-1
Third	Rate = k[A]^2[B]	mol-2 dm6 s-1

Molecularity of Elementary Reactions

Type	No. of Molecules	Example
Unimolecular	1	H2 -> 2H
Bimolecular	2	H2 + Cl2 -> 2HCl
Termolecular	3	2NO + O2 -> 2NO2

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