CHEMISTRY (CODE – 043)

Chapter-wise Test: Chemical Kinetics

(NCERT Class XII Chemistry, Unit 3: Chemical Kinetics)

SOLUTIONS AND MARKING SCHEME

CLASS XII (2026-27)

Time: 1½ hours **Max. Marks: 35**

Step-by-step solutions with the distribution of marks. Marks for each step are shown in square brackets [ ]. Any other correct method or answer should be given full credit.

Quick answer key – Section A

Q.	1	2	3	4	5	6	7	8
Ans.	C	B	D	D	A	C	B	A

SECTION A

Questions 1 to 8 are multiple choice questions. Only one of the choices is correct. Select and write the correct choice as well as the answer to these questions. Each question carries 1 mark.

1	**Answer: C. mol L^{-1} s^{-1}** For a zero order reaction, rate = $k[R]^0 = k$, so k has the same unit as the rate, mol L^{-1} s^{-1}. (First order: s^{-1}; second order: L mol^{-1} s^{-1}.)	1
2	**Answer: B. first order, k = 0.05 min^{-1}** For a first order reaction $\ln [R] = \ln [R]_0 - kt$, so a plot of ln [R] against t is a straight line with slope = $-k$. From Fig. 1: slope = $(-2.0 - (-1.0)) / (20 - 0) = -0.05$ min^{-1}. Hence the reaction is first order and k = 0.05 min^{-1}.	1
3	**Answer: D. rate = $k[A][B]^2$** Experiments 1 and 2: [A] is doubled with [B] constant, and the rate doubles (4.0/2.0 = 2), so the order with respect to A is 1. Experiments 1 and 3: [B] is doubled with [A] constant, and the rate becomes four times (8.0/2.0 = 4 = 2^2), so the order with respect to B is 2. Rate = $k[A][B]^2$ (overall order 3).	1
4	**Answer: D. 10 min** For a first order reaction $t_{1/2} = 0.693 / k = 0.693 / 0.0693$ min^{-1} = 10 min.	1
5	**Answer: A** Order is determined experimentally and can be zero, fractional or an integer. Molecularity is the number of reacting species colliding in an elementary step; it is a theoretical quantity, always a whole number (1, 2 or rarely 3) and can never be zero or fractional. Order and molecularity need not be equal (e.g. pseudo first order reactions).	1
6	**Answer: C** A catalyst provides an alternative reaction mechanism with a lower potential energy barrier (lower E_a), so more molecules can cross it ($k = A\, e^{-Ea/RT}$). It does not change ΔG, ΔH or the equilibrium constant.	1
7	**Answer: B. 40 min** In a first order reaction the half-life is constant. 50% completion takes one half-life (20 min); 75% completion means two half-lives (100 → 50 → 25% left). Time = 2 × 20 = 40 min.	1
8	**Answer: A. Both A and R are true, and R is the correct explanation of A.** $CH_3COOC_2H_5 + H_2O \rightarrow CH_3COOH + C_2H_5OH$ is really second order, but water is present in large excess, so its concentration hardly changes (e.g. 0.01 mol ester with 10 mol water: 10 mol → 9.99 mol). The rate = k'[ester] behaves as first order, i.e. pseudo first order.	1

SECTION B

Question Nos. 9 to 11 are very short answer questions carrying 2 marks each.

9	**(A)** (i) **Curve 2** is at the higher temperature: its peak (most probable kinetic energy) shifts to higher energy and is lower and broader, so a larger fraction of molecules has high energies. [1] (ii) At (T + 10) K the area under the curve to the right of E_a, i.e. the fraction of molecules with energy equal to or greater than E_a, is nearly doubled. More molecules can react on collision, so the rate nearly doubles. [1] **(B)** (i) $k = A\ e^{-Ea/RT}$. A is the Arrhenius (frequency or pre-exponential) factor, a constant for a given reaction related to collision frequency; E_a is the activation energy (J mol^{-1}). [1] (ii) The factor $e^{-Ea/RT}$ corresponds to the fraction of molecules having kinetic energy equal to or greater than the activation energy. [1]	1 + 1
10	**(A)** Any two of the following (1 mark each): • Molecularity is the number of reacting species taking part in an elementary reaction; order is the sum of the powers of the concentration terms in the experimentally determined rate law. [1] • Molecularity is a theoretical concept, whereas order is determined only by experiment. [1] • Molecularity is always a whole number (1, 2, 3) and never zero or fractional; order can be zero, fractional or a whole number. [1] • Molecularity applies only to elementary reactions; order applies to elementary as well as complex reactions. [1] **(B)** (i) Reactions that are not truly first order but obey a first order rate law, because one of the reactants is present in large excess so that its concentration remains practically constant. [1] (ii) Hydrolysis of ethyl ethanoate (acid-catalysed, excess water), or inversion of cane sugar: $C_{12}H_{22}O_{11} + H_2O \rightarrow C_6H_{12}O_6 + C_6H_{12}O_6$; rate = $k[C_{12}H_{22}O_{11}]$. [1]	1 + 1
11	**(A)** $t_{1/2} = 0.693 / k = 0.693 / (2.31 \times 10^{-2}\ min^{-1})$ = **30 min** [1] 8 g → 4 g → 2 g → 1 g takes three half-lives: $t = 3 \times 30$ = **90 min** [1] **(B)** $t_{1/2} = [R]_0 / 2k = 0.8 / (2 \times 2 \times 10^{-2}) = 0.8 / 0.04$ = **20 s** [1] Time for completion: $t = [R]_0 / k = 0.8 / (2 \times 10^{-2})$ = **40 s** [1]	1 + 1

SECTION C

Question Nos. 12 to 15 are short answer questions carrying 3 marks each.

12	**(A)** (i) Experiments 2 and 3: [B] doubled, rate unchanged, so the order in B = 0. Experiments 1 and 2: [A] doubled, rate becomes four times (0.016/0.004 = 4 = 2^2), so the order in A = 2. [1] (ii) Rate = $k[A]^2[B]^0 = k[A]^2$. Overall order = 2 + 0 = **2**. [1] (iii) From experiment 1: k = rate / $[A]^2$ = 0.004 / $(0.10)^2$ = 0.004 / 0.01 = **0.4 L mol^{-1} s^{-1}** (check with experiment 4: 0.036/0.09 = 0.4). [1] **(B)** (i) Overall order = ½ + 2 = **2.5**. [1] (ii) For order n, the unit of k = (mol L^{-1})$^{1-n}$ s^{-1} = (mol L^{-1})$^{-1.5}$ s^{-1} = **mol$^{-3/2}$ L$^{3/2}$ s^{-1}**. [1] (iii) New rate = $k(4[A])^{1/2}(½[B])^2$ = 2 × ¼ × $k[A]^{1/2}[B]^2$ = ½ of the original rate. The rate is **halved**. [1]	1 + 1 + 1
13	**(A)** (i) By the Arrhenius equation ln k = ln $A - E_a/RT$, the slope of the ln k versus 1/T line is $-E_a/R$. [1] (ii) Slope = (−5.0 − (−2.0)) / (0.0035 − 0.0030) = −3.0 / 0.0005 = −6000 K. [½] E_a = −slope × R = 6000 × 8.314 = 49884 J mol^{-1} ≈ **49.9 kJ mol^{-1}** [½] (iii) The intercept is ln A, where A is the Arrhenius (frequency) factor. ln A = −2.0 + 6000 × 0.0030 = **16**. [1] **(B)** log (k_2/k_1) = (E_a/2.303R) × ($T_2 - T_1$)/(T_1T_2) [1] log 2 = 0.301 = E_a × 10 / (2.303 × 8.314 × 300 × 310) [1] E_a = (0.301 × 2.303 × 8.314 × 93000) / 10 ≈ 53600 J mol^{-1} ≈ **53.6 kJ mol^{-1}** [1]	1 + 1 + 1
14	**(A)** (i) **Curve II** is the catalysed reaction. A catalyst provides an alternative path with a lower energy barrier (lower activation energy) [1] (ii) E_a = energy of the peak − energy of the reactants. Curve I: 130 − 50 = **80 kJ mol^{-1}**; Curve II: 90 − 50 = **40 kJ mol^{-1}**. [1] (iii) ΔH = energy of products − energy of reactants = 30 − 50 = **−20 kJ mol^{-1}** (exothermic). A catalyst does **not** change ΔH (or ΔG) or the equilibrium constant; it only helps in reaching equilibrium faster. [1] **(B)** Rate = −d[R]/dt = k[R], so d[R]/[R] = −k dt. [1] Integrating: ln [R] = $-kt + I$. At t = 0, [R] = $[R]_0$, so I = ln $[R]_0$. Hence ln [R] = $-kt$ + ln $[R]_0$, i.e. k = (1/t) ln ($[R]_0$/[R]). [1] Changing to $\log_{10}$ (ln x = 2.303 log x): ***k* = (2.303/*t*) log ($[R]_0$/[R])**. [1]	1 + 1 + 1
15	**(A)** For a first order reaction t = (2.303/k) log ($[R]_0$/[R]). [1] For 90% completion [R] = $[R]_0$/10: t_{90} = (2.303/k) log 10 = 2.303/k. [1] For 99% completion [R] = $[R]_0$/100: t_{99} = (2.303/k) log 100 = 2 × 2.303/k = 2t_{90}. Hence proved. [1] **(B)** t = (2.303/k) log ($[R]_0$/[R]) = (2.303 / 1.15 × 10^{-3}) × log (5/3) [1] log (5/3) = 0.6990 − 0.4771 = 0.2219 [1] t = (2.303 × 0.2219) / (1.15 × 10^{-3}) = 0.5110 / 0.00115 ≈ **444 s** [1]	1 + 1 + 1

SECTION D

Question No. 16 is a case-based/data-based question carrying 4 marks.

16	(i) **First order.** The concentration falls to half (0.80 → 0.40 → 0.20 → 0.10) in every equal interval of 15 min, i.e. the half-life is constant and independent of the initial concentration, which is the characteristic of a first order reaction. [1] (ii) $k = 0.693 / t_{1/2} = 0.693 / 15$ min = **0.0462 min^{-1}** ($\approx 4.62 \times 10^{-2}$ min^{-1}) [1] (iii) **(A)** 60 min = 4 half-lives. $[R] = [R]_0 / 2^4 = 0.80 / 16$ = **0.05 mol L^{-1}** [1 for method + 1 for answer] **(B)** $[R]_0/[R] = 0.80 / 0.025 = 32 = 2^5$, so 5 half-lives are needed. $t = 5 \times 15$ = **75 min** [1 for method + 1 for answer]	1 + 1 + 2

SECTION E

Question No. 17 is a long answer question carrying 5 marks.

17	**(A)** (i) $\log (k_2/k_1) = (E_a / 2.303R)\,(1/T_1 - 1/T_2) = (E_a / 2.303R)\,(T_2 - T_1)/(T_1T_2)$ [1] (ii) $\log 2 = E_a \times (310 - 300) / (2.303 \times 8.314 \times 300 \times 310)$ [1]. $E_a = (0.301 \times 2.303 \times 8.314 \times 93000) / 10 \approx 53600$ J mol^{-1} = **53.6 kJ mol^{-1}** [1] (iii) A catalyst **lowers** the activation energy by providing an alternative path. It has **no effect on ΔG** (nor on ΔH or the equilibrium constant). [1] (iv) Plot ln k (or log k) against $1/T$: a straight line is obtained with slope = $-E_a/R$ (or $-E_a/2.303R$), from which E_a is calculated; the intercept gives ln A. [1] **(B)** (i) Half-life ($t_{1/2}$) is the time in which the concentration of a reactant is reduced to half of its initial value. [1] For a first order reaction, $k = (2.303/t) \log ([R]_0/[R])$. At $t = t_{1/2}$, $[R] = [R]_0/2$, so $t_{1/2} = (2.303/k) \log 2 = 0.693/k$. This does not contain $[R]_0$, so it is independent of the initial concentration. [1] (ii) Take $[R]_0 = 100$, $[R] = 70$ at $t = 40$ min. $k = (2.303/40) \log (100/70)$ [1] $\log (10/7) = 1 - 0.8451 = 0.1549$. $k = (2.303 \times 0.1549) / 40 = 0.3567 / 40 = 8.92 \times 10^{-3}$ min^{-1} [1] $t_{1/2} = 0.693 / k = 0.693 / (8.92 \times 10^{-3}) \approx$ **77.7 min** [1]	5

Topic coverage of the paper

Topic (NCERT section)	Questions
Rate of a reaction; factors affecting rate (3.1, 3.2)	Q1, Q3 (table), Q12 (table)
Order, molecularity and units of rate constant (3.2)	Q1, Q5, Q10, Q12 (table)
Integrated rate laws, half-life, pseudo first order (3.3)	Q2 (graph), Q4, Q7, Q8, Q11, Q15, Q16 (table)
Temperature dependence; Arrhenius equation (3.4)	Q9 (figure), Q13 (graph), Q17 (A)
Effect of catalyst; activation energy profile (3.4.1)	Q6, Q14 (figure), Q17 (A)
First order kinetics: derivation and numericals	Q14 (B), Q15, Q17 (B)

Figure/table-based questions (7): Q2, Q3, Q9, Q12, Q13, Q14 and Q16.