Class 12 Chemistry · Chapter 3: Chemical Kinetics

Answer key and marking scheme · 20 marks · about 40 minutes · based on NCERT

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1. The gas-phase reaction $2NO(g) + 2H_2(g) \rightarrow N_2(g) + 2H_2O(g)$ follows Rate = $k[NO]^2[H_2]$. With concentration in mol L^{-1} and time in seconds, the unit of k is [1]

(a) s^{-1} (b) $mol^{-1}\ L\ s^{-1}$

(c) $mol\ L^{-1}\ s^{-1}$ (d) $mol^{-2}\ L^2\ s^{-1}$

Answer: (d) The order is 2 + 1 = 3, so k = Rate/[conc]^3 has unit (mol L^{-1} s^{-1})/(mol L^{-1})^3 = $mol^{-2}\ L^2\ s^{-1}$.

2. A zero order reaction has k = 0.005 mol L^{-1} min^{-1}. If the initial concentration is 0.40 mol L^{-1}, the half-life is [1]

(a) 80 min (b) 40 min

(c) 20 min (d) 138.6 min

Answer: (b) For zero order $t_{½} = [R]_0/2k = 0.40/(2 \times 0.005) = 40$ min. (138.6 min is 0.693/k, the first order formula wrongly applied.)

3. For $CHCl_3 + Cl_2 \rightarrow CCl_4 + HCl$, the experimental rate law is Rate = $k[CHCl_3][Cl_2]^{1/2}$. The overall order of the reaction is [1]

(a) 1 (b) 2

(c) 3/2 (d) 1/2

Answer: (c) Order = sum of the powers in the rate law = 1 + ½ = 3/2, although the coefficients in the equation add to 2.

4. In the iodide-catalysed decomposition of H_2O_2, the fast step is the elementary reaction $H_2O_2 + IO^- \rightarrow H_2O + I^- + O_2$. The molecularity of this step is [1]

(a) 2 (b) 1

(c) 3 (d) 4

Answer: (a) Two species, H_2O_2 and IO^-, must collide together; the step is bimolecular.

5. Which statement about order and molecularity is correct? [1]

(a) Molecularity is found by measuring the rate at different concentrations.

(b) Molecularity can be zero for a reaction on a metal surface.

(c) Order has meaning only for elementary reactions.

(d) Order can be zero or a fraction, but molecularity cannot.

Answer: (d) Order is an experimental quantity and may be zero or fractional; molecularity counts colliding species in an elementary step, so it is a whole number and has no meaning for a complex reaction.

6. A first order reaction has k = 6.93×10^{-3} min^{-1}. How long will it take for 87.5% of the reactant to be used up? [2]

$t_{½} = 0.693/k = 0.693/(6.93 \times 10^{-3}) = 100$ min. 87.5% used means 12.5% = ⅛ remains: three half-lives. t = 3 × 100 = 300 min.

- $t_{½}$ = 100 min (1)
- ⅛ remaining is three half-lives, 300 min (1)

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7. State the two conditions a collision must meet to give products, and name the factor introduced in collision theory to account for one of them. [2]

An effective collision needs (i) energy equal to or more than the threshold energy, and (ii) proper orientation of the colliding molecules. The probability or steric factor P accounts for orientation: Rate = P Z_{AB} e^(−Ea/RT).

- both conditions: enough energy (threshold) and proper orientation (1)
- P, the probability (steric) factor, for orientation (1)

8. For $4NH_3(g) + 5O_2(g) \rightarrow 4NO(g) + 6H_2O(g)$, NO is formed at 3.6×10^{-3} mol L^{-1} s^{-1} at an instant. Calculate (i) the rate of disappearance of O_2 and (ii) the rate of formation of H_2O at that instant. [2]

Rate = $-\frac{1}{4}\Delta[NH_3]/\Delta t = -\frac{1}{5}\Delta[O_2]/\Delta t = \frac{1}{4}\Delta[NO]/\Delta t = \frac{1}{6}\Delta[H_2O]/\Delta t$. Rate = $\frac{1}{4} \times 3.6 \times 10^{-3} = 9.0 \times 10^{-4}$ mol L^{-1} s^{-1}. (i) $-\Delta[O_2]/\Delta t = 5 \times 9.0 \times 10^{-4} = 4.5 \times 10^{-3}$ mol L^{-1} s^{-1}. (ii) $\Delta[H_2O]/\Delta t = 6 \times 9.0 \times 10^{-4} = 5.4 \times 10^{-3}$ mol L^{-1} s^{-1}.

- O_2 disappears at 4.5×10^{-3} mol L^{-1} s^{-1} (5/4 of the NO rate) (1)
- H_2O forms at 5.4×10^{-3} mol L^{-1} s^{-1} (6/4 of the NO rate) (1)

9. The reaction $H_2(g) + 2ICl(g) \rightarrow I_2(g) + 2HCl(g)$ is proposed to go in two steps: (1) $H_2 + ICl \rightarrow HI + HCl$ (slow); (2) $HI + ICl \rightarrow I_2 + HCl$ (fast). (i) Write the rate law this mechanism predicts and its order. (ii) Name the intermediate and give the molecularity of each step. (iii) Why is the order not 3, though three molecules appear in the overall equation? [3]

(i) The slow step is rate determining: Rate = $k[H_2][ICl]$, second order. (ii) HI is the intermediate, formed in step 1 and used in step 2; both steps are bimolecular. (iii) The overall equation is not an elementary step; the order follows the slow step, which involves one H_2 and one ICl, not the stoichiometric coefficients.

- Rate = $k[H_2][ICl]$, second order (1)
- HI the intermediate; both steps bimolecular (1)
- order is set by the slow step, not by the balanced equation (1)

10. A reaction has Ea = 46.0 kJ mol^{-1} and $k = 1.5 \times 10^{-4}$ s^{-1} at 300 K. Taking A as constant, calculate k at 320 K. (2.303 R = 19.15 J K^{-1} mol^{-1}, antilog 0.500 = 3.16) [3]

$\log(k_2/k_1) = (Ea/2.303R)(T_2 - T_1)/(T_1T_2) = (46000/19.15) \times 20/(300 \times 320) = 2402 \times 2.083 \times 10^{-4} = 0.500$. $k_2/k_1 = 3.16$, so $k_2 = 3.16 \times 1.5 \times 10^{-4} = 4.74 \times 10^{-4}$ s^{-1}.

- the two-temperature Arrhenius relation written (1)
- $\log(k_2/k_1) = 0.500$ (1)
- $k_2 = 4.74 \times 10^{-4}$ s^{-1} (1)

11. A compound X decomposes by first order kinetics. Its concentration falls from 0.80 mol L^{-1} to 0.20 mol L^{-1} in 46.2 min. Calculate (i) the rate constant, (ii) the half-life, and (iii) the time for the concentration to fall to 0.05 mol L^{-1}. (log 2 = 0.3010) [3]

(i) $k = (2.303/t) \log([R]_0/[R]) = (2.303/46.2) \log 4 = (2.303 \times 0.6020)/46.2 = 3.0 \times 10^{-2}$ min^{-1}. (ii) $t_{1/2} = 0.693/k = 0.693/0.030 = 23.1$ min. (iii) 0.80 → 0.05 is a sixteenfold fall: $t = (2.303/k) \log 16 = 2 \times 46.2 = 92.4$ min (four half-lives).

- $k = 3.0 \times 10^{-2}$ min^{-1} with the formula (1)
- $t_{1/2} = 23.1$ min (1)
- 92.4 min for 0.05 mol L^{-1} (1)

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