CLASS 12 CHEMISTRY NCERT SOLUTIONS

NCERT Solutions for Class 12 Chemistry Chapter 3: Chemical Kinetics – Free PDF Download

Chapter 3 covers chemical kinetics — rate of reaction, rate laws and order of reaction, integrated rate equations for zero- and first-order reactions, half-life, the effect of temperature via the Arrhenius equation, and factors affecting reaction rate. All 29 exercises (3.1–3.29) remain current under the 2026-27...

Contents

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NCERT Exercise Solutions

3.1 From the rate expression for the following reactions, determine their order of reaction and the dimensions of the rate constants: (i) $3NO(g) \rightarrow N_2O(g)$, Rate=$k[NO]^2$ (ii) $H_2O_2(aq)+3I^-(aq)+2H^+ \rightarrow 2H_2O(l)+I_3^-$, Rate=$k[H_2O_2][I^-]$ (iii) $CH_3CHO(g) \rightarrow CH_4(g)+CO(g)$, Rate=$k[CH_3CHO]^{3/2}$ (iv) $C_2H_5Cl(g) \rightarrow C_2H_4(g)+HCl(g)$, Rate=$k[C_2H_5Cl]$.
Ans: (i) Order=**2**; k units=$\mathbf{mol^{-1}Ls^{-1}}$. (ii) Order=1+1=**2**; k units=$\mathbf{mol^{-1}Ls^{-1}}$. (iii) Order=**3/2**; k units=$\mathbf{mol^{-1/2}L^{1/2}s^{-1}}$. (iv) Order=**1**; k units=$\mathbf{s^{-1}}$.

3.2 For the reaction $2A+B \rightarrow A_2B$, the rate=$k[A][B]^2$ with $k=2.0\times10^{-6}mol^{-2}L^2s^{-1}$. Calculate the initial rate when [A]=0.1mol L^{-1}, [B]=0.2mol L^{-1}. Calculate the rate of reaction after [A] is reduced to 0.06mol L^{-1}.
Ans: Initial rate=$k[A][B]^2=2.0\times10^{-6}\times0.1\times(0.2)^2=\mathbf{8.0\times10^{-9}mol\ L^{-1}s^{-1}}$. Since A and B react in a 2:1 ratio, when [A] falls by 0.04 (to 0.06), [B] falls by half that, 0.02, to 0.18mol L^{-1}. New rate=$2.0\times10^{-6}\times0.06\times(0.18)^2\approx\mathbf{3.89\times10^{-9}mol\ L^{-1}s^{-1}}$.

3.3 The decomposition of NH_3 on a platinum surface is zero order. If $k=2.5\times10^{-4}mol\ L^{-1}s^{-1}$, what are the rates of production of N_2 and H_2?
Ans: For $2NH_3 \rightarrow N_2+3H_2$ (zero order), rate of reaction=k=$\mathbf{2.5\times10^{-4}mol\ L^{-1}s^{-1}}$. Since 1 mol N_2 forms per reaction event, $d[N_2]/dt=\mathbf{2.5\times10^{-4}mol\ L^{-1}s^{-1}}$; since 3 mol H_2 forms per event, $d[H_2]/dt=3\times2.5\times10^{-4}=\mathbf{7.5\times10^{-4}mol\ L^{-1}s^{-1}}$.

3.4 The decomposition of dimethyl ether leads to CH_4, H_2 and CO, with rate=$k[CH_3OCH_3]^{3/2}$. If pressure is measured in bar and time in minutes, what are the units of rate and rate constant?
Ans: Order=3/2. Rate units=$\mathbf{bar\ min^{-1}}$. Rate constant units=$(bar\ min^{-1})/(bar^{3/2})=\mathbf{bar^{-1/2}min^{-1}}$.

3.5 Mention the factors that affect the rate of a chemical reaction.
Ans: Reaction rate is affected by: the **concentration** of reactants (higher concentration generally increases rate), **temperature** (rate increases with temperature since more molecules gain the activation energy needed to react), the presence of a **catalyst** (which provides an alternate, lower-activation-energy pathway), the **nature of the reactants** (e.g. ionic reactions are typically faster than covalent-bond-breaking reactions), and the **surface area** of reactants in heterogeneous reactions (a finely powdered solid reacts faster than a lump).

3.6 A reaction is second order with respect to a reactant. How is the rate of reaction affected if the concentration of the reactant is (i) doubled (ii) reduced to half?
Ans: Since rate=$k[X]^2$: (i) doubling [X] makes the rate **4 times** the original ($2^2=4$). (ii) halving [X] makes the rate **one-quarter** the original ($(1/2)^2=1/4$).

3.7 What is the effect of temperature on the rate constant of a reaction? How can this temperature effect on the rate constant be represented quantitatively?
Ans: The rate constant of a reaction **increases with temperature** — as a rough rule, it roughly doubles for every 10°C rise for many reactions. This is represented quantitatively by the **Arrhenius equation**, $k=Ae^{-E_a/RT}$, where A is the pre-exponential (frequency) factor, E_a is the activation energy, R is the gas constant, and T is the absolute temperature — higher T makes the exponential term (and hence k) larger.

3.8 A pseudo first order reaction (ester hydrolysis) was followed by measuring ester concentration at intervals: t(s)=0, 30, 60, 90; [Ester](mol L^{-1})=0.55, 0.31, 0.17, 0.085. Calculate the average rate of reaction between t=30s and t=60s, and the pseudo first order rate constant for the hydrolysis.
Ans: Average rate (30–60s)=(0.31–0.17)/30≈**4.67×10^{-3}mol $L^{-1}s^{-1}$**. Using k=(1/t)ln($[Ester]_0$/$[Ester]_t$) for each interval: $k_{(0-30)}\approx1.911\times10^{-2}$, $k_{(30-60)}\approx2.003\times10^{-2}$, $k_{(60-90)}\approx2.310\times10^{-2}s^{-1}$. Average pseudo first order **$k\approx2.07\times10^{-2}s^{-1}$**.

3.9 A reaction is first order in A and second order in B. (i) Write the differential rate equation. (ii) How is the rate affected on increasing the concentration of B three times? (iii) How is the rate affected when the concentrations of both A and B are doubled?
Ans: (i) Rate=−d[A]/dt=**$k[A][B]^2$**. (ii) Tripling [B] makes the rate **9 times** the original (3^2=9). (iii) Doubling both [A] and [B] makes the rate **8 times** the original (2×2^2=8).

3.10 In a reaction between A and B, the initial rate was measured for different concentrations: Expt 1, [A]=0.20M, [B]=0.30M, rate=5.07×10^{-5}; Expt 2, [A]=0.20M, [B]=0.10M, rate=5.07×10^{-5}; Expt 3, [A]=0.40M, [B]=0.05M, rate=1.43×10^{-4} (mol $L^{-1}s^{-1}$). What is the order of the reaction with respect to A and B?
Ans: Comparing Expt 1 and 2, changing [B] (0.30→0.10) with [A] constant does not change the rate, so order with respect to **B=0**. Comparing Expt 2 and 3, doubling [A] (0.20→0.40) increases the rate by a factor of $1.43\times10^{-4}/5.07\times10^{-5}\approx2.82\approx2^{1.5}$, so order with respect to **A=1.5**.

3.11 The following results were obtained for the reaction 2A+B→C+D: Expt I, [A]=0.1,[B]=0.1, rate=6.0×10^{-3}; Expt II, [A]=0.3,[B]=0.2, rate=7.2×10^{-2}; Expt III, [A]=0.3,[B]=0.4, rate=2.88×10^{-1}; Expt IV, [A]=0.4,[B]=0.1, rate=2.4×10^{-2} (mol $L^{-1}min^{-1}$). Determine the rate law and the rate constant.
Ans: Comparing II and III ([A] constant, [B] doubled): rate increases 4×, so order in **B=2**. Comparing I and IV ([B] constant, [A] quadrupled): rate increases 4×, so order in **A=1**. Rate law: **$Rate=k[A][B]^2$**. From Expt I: $k=6.0\times10^{-3}/(0.1\times0.1^2)$=**$6.0L^2mol^{-2}min^{-1}$** (confirmed consistent across all four experiments).

3.12 The following data were obtained for a reaction that is first order in A and zero order in B: Expt 1, [A]=0.1M,[B]=0.1M,rate=2.0×10^{-2}; Expt 2, [B]=0.2M, rate=4.0×10^{-2}; Expt 3, [A]=0.4M,[B]=0.4M, rate=?; Expt 4, [B]=0.2M, rate=2.0×10^{-2} (mol $L^{-1}min^{-1}$). Complete the table.
Ans: Rate=k[A] (B has no effect). From Expt 1: $k=2.0\times10^{-2}/0.1$=**$0.2min^{-1}$**. Expt 2: $4.0\times10^{-2}=0.2\times[A]$ ⇒ **[A]=0.2M**. Expt 3: rate=0.2×0.4=**0.08mol $L^{-1}min^{-1}$**. Expt 4: $2.0\times10^{-2}=0.2\times[A]$ ⇒ **[A]=0.1M**.

3.13 Calculate the half-lives of first-order reactions with rate constants: (i) $200s^{-1}$ (ii) $2min^{-1}$ (iii) $4year^{-1}$.
Ans: Using $t_{1/2}$=0.693/k: (i) 0.693/200≈**3.47×10^{-3}s**. (ii) 0.693/2≈**0.35min**. (iii) 0.693/4≈**0.173year**.

3.14 The half-life for radioactive decay of $^{14}_{6}C$ is 5730 years. An archaeological artifact containing wood had only 80% of the ^{14}C found in a living tree. Estimate the age of the sample.
Ans: k=0.693/5730≈$1.209\times10^{-4}year^{-1}$. Using t=(1/k)ln(100/80): t≈**1845 years**.

3.15 The decomposition of N_2O_5 at 318K was followed by monitoring [N_2O_5] (×10^2mol L^{-1}) at intervals up to 3200s (data: 1.63, 1.36, 1.14, 0.93, 0.78, 0.64, 0.53, 0.43, 0.35 at t=0,400,...,3200). Confirm the reaction is first order and calculate the rate constant and half-life.
Ans: Applying k=(1/t)ln($[N_2O_5]_0$/$[N_2O_5]_t$) at each interval gives a near-constant k (ranging ≈4.5–4.8×$10^{-4}s^{-1}$), confirming **first order**. Average **$k\approx4.7\times10^{-4}s^{-1}$**, giving **$t_{1/2}=0.693/k\approx1470s$**.

3.16 The rate constant for a first-order reaction is $60s^{-1}$. How much time will it take to reduce the initial concentration to its 1/16th value?
Ans: Using t=(1/k)ln(16): t=(1/60)×ln(16)≈**4.62×10^{-2}s**.

3.17 ^{90}Sr has a half-life of 28.1 years. Calculate the amount remaining after (i) 10 years (ii) 60 years, if the initial amount was 1µg.
Ans: $k=0.693/28.1\approx2.466\times10^{-2}\text{year}^{-1}$. (i) $N=1\times e^{-k\times10}\approx$**0.781µg**. (ii) $N=1\times e^{-k\times60}\approx$**0.228µg**.

3.18 Show that for a first-order reaction, the time required for 99% completion is twice the time required for 90% completion.
Ans: $t=(1/k)\ln(100/(100-x))$. For 90% completion (10% remains): $t_{90}=(1/k)\ln(10)=2.303/k$. For 99% completion (1% remains): $t_{99}=(1/k)\ln(100)=4.606/k$. So $t_{99}/t_{90}=4.606/2.303=$**2**, confirming $t_{99}=2\times t_{90}$.

3.19 A first-order reaction takes 40 minutes for 30% decomposition. Calculate its rate constant and half-life.
Ans: $k=(1/40)\ln(100/70)\approx$**$8.92\times10^{-3}\text{min}^{-1}$**. $t_{1/2}=0.693/k\approx$**77.7min**.

3.20 The decomposition of azoisopropane, $(CH_3)_2CHN=NCH(CH_3)_2(g)\rightarrow N_2(g)+C_6H_{14}(g)$, was monitored via total pressure at 543K: t(s)=0,360,720, P(mmHg)=35.0,54.0,63.0. Calculate the rate constant.
Ans: Since 1 mol reactant gives 2 mol product, $P_{reactant}=2P_0-P_{total}$. At t=360s: $P_A=70-54=16.0$mmHg, $k\approx2.17\times10^{-3}\text{s}^{-1}$. At t=720s: $P_A=70-63=7.0$mmHg, $k\approx2.24\times10^{-3}\text{s}^{-1}$. Average **$k\approx2.21\times10^{-3}\text{s}^{-1}$**.

3.21 The first-order thermal decomposition of $SO_2Cl_2(g)\rightarrow SO_2(g)+Cl_2(g)$ at constant volume gave: Expt 1, t=0s, total pressure=0.5atm; Expt 2, t=100s, total pressure=0.6atm. Calculate the rate of the reaction when total pressure is 0.65atm.
Ans: Since $P_{SO2Cl2}=P_0-x$ and $P_{total}=P_0+x$, at t=100s: x=0.1, $P_{SO2Cl2}=0.4$atm, giving $k=(1/100)\ln(0.5/0.4)\approx$**$2.23\times10^{-3}\text{s}^{-1}$**. At total pressure=0.65atm: x=0.15, $P_{SO2Cl2}=0.35$atm, so rate=$k\times P_{SO2Cl2}=2.23\times10^{-3}\times0.35\approx$**$7.81\times10^{-4}$atm s^{-1}**.

3.22 The rate constant for the decomposition of a hydrocarbon is $2.418\times10^{-5}\text{s}^{-1}$ at 546K. If the energy of activation is 179.9kJ/mol, what is the value of the pre-exponential factor?
Ans: Using $k=Ae^{-E_a/RT}$: $A=k/e^{-E_a/RT}=2.418\times10^{-5}/e^{-179900/(8.314\times546)}\approx$**$3.93\times10^{12}\text{s}^{-1}$**.

3.23 The rate constant for a first-order reaction A→Products is $2.0\times10^{-2}\text{s}^{-1}$. Calculate the concentration of A remaining after 100s if the initial concentration was 1.0mol L^{-1}.
Ans: $[A]_t=[A]_0e^{-kt}=1.0\times e^{-2.0\times10^{-2}\times100}=1.0\times e^{-2}\approx$**0.135mol L^{-1}**.

3.24 Sucrose decomposes in acid solution into glucose and fructose following first-order kinetics, with $t_{1/2}$=3.00 hours. What fraction of the sample of sucrose remains after 8 hours?
Ans: $k=0.693/3.00\approx0.231\text{hour}^{-1}$. Fraction remaining=$e^{-kt}=e^{-0.231\times8}\approx$**0.158** (about 15.8%).

3.25 The decomposition of a hydrocarbon follows $k=4.5\times10^{11}e^{-28000K/T}\text{s}^{-1}$. Calculate its activation energy.
Ans: Comparing with the Arrhenius equation, $E_a/R=28000$K, so $E_a=28000\times8.314\approx$**232.8kJ/mol**.

3.26 The rate constant for the first-order decomposition of H_2O_2 is given by log $k=14.34-1.25\times10^4\text{K}/T$. Calculate the activation energy and the temperature at which its half-life is 256 minutes.
Ans: Since $\log k=\log A-E_a/(2.303RT)$, $E_a=1.25\times10^4\times2.303\times8.314\approx$**239.3kJ/mol**. For $t_{1/2}=256\text{min}=15360$s: $k=0.693/15360\approx4.51\times10^{-5}\text{s}^{-1}$, $\log k\approx-4.346$. Solving $14.34-1.25\times10^4/T=-4.346$ gives **T≈669K**.

3.27 The decomposition of A into products has a rate constant of $4.5\times10^3\text{s}^{-1}$ at 10°C, and the energy of activation is 60kJ/mol. At what temperature would k be $1.5\times10^4\text{s}^{-1}$?
Ans: Using $\ln(k_2/k_1)=-(E_a/R)(1/T_2-1/T_1)$ with $T_1=283$K: solving gives **$T_2\approx$297K (24°C)**.

3.28 The time required for 10% completion of a first-order reaction at 298K is equal to that required for its 25% completion at 308K. If $A=4\times10^{10}s^{-1}$, calculate k at 318K and the energy of activation.
Ans: Since the times are equal, $k_{308}/k_{298}=\ln(100/75)/\ln(100/90)\approx2.73$. Using the Arrhenius two-point equation: $\mathbf{E_a\approx76.6kJ/mol}$. Then $k_{318}=Ae^{-E_a/(R\times318)}\approx\mathbf{1.03\times10^{-2}s^{-1}}$.

3.29 The rate of a reaction quadruples when the temperature changes from 293K to 313K. Calculate the energy of activation, assuming it does not change with temperature.
Ans: Using $\ln(4)=(E_a/R)(1/293-1/313)$: $\mathbf{E_a\approx52.9kJ/mol}$.

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