Collegedunia NCERT Formula Sheet

The Ultimate Formula Reference for Class 12 Chemistry

NCERT 2026-27 / Latest Edition

Chapter 3: Chemical Kinetics

Rate Laws | Order & Molecularity | Integrated Rate Equations | Half-Life | Arrhenius | Collision Theory

Key Constants & Conversions for this Chapter

Quantity	Symbol	Value (SI)
Universal gas constant	R	8.314 J K^{-1} mol^{-1}
Log conversion	$\ln x = 2.303 \log x$	$\log 2 = 0.301$
$\ln 2$ value	$\ln 2$	0.693
Standard temperature	T_0	298 K (25°C)
Avogadro number	N_A	6.022×10^{23} mol^{-1}
Typical activation energies	E_a	40–400 kJ mol^{-1}
SI units of k (order n)	—	(mol L^{-1})$^{1-n}$ s^{-1}

1 Rate of a Chemical Reaction

This section covers the definitions of average and instantaneous rate, the stoichiometric form of the rate of reaction, and the units used when concentration or partial pressure is the measured variable.

Average rate of reaction

For $R \rightarrow P$:

$$r_{\text{av}} = -\frac{\Delta[R]}{\Delta t} = +\frac{\Delta[P]}{\Delta t}$$

where $[R]$, $[P]$ = molar concentrations (mol L^{-1}); Δt = time interval (s).

Measures the **mean change** in concentration over a finite interval. The minus sign keeps the rate positive since $[R]$ decreases with time.

Instantaneous rate

$$r_{\text{inst}} = -\frac{d[R]}{dt} = +\frac{d[P]}{dt}$$

where $\frac{d[R]}{dt}$ = slope of the $[R]$ vs t curve at that instant.

Limit of the average rate as $\Delta t \to 0$. Geometrically it is the **slope of the tangent** to the concentration-time curve at the chosen instant.

General stoichiometric rate

For $aA + bB \to cC + dD$:

$$\text{Rate} = -\frac{1}{a}\frac{d[A]}{dt} = -\frac{1}{b}\frac{d[B]}{dt} = +\frac{1}{c}\frac{d[C]}{dt} = +\frac{1}{d}\frac{d[D]}{dt}$$

where a, b, c, d = stoichiometric coefficients.

Dividing by the coefficient gives a **single unique rate** for the reaction, independent of which species is being monitored.

Units of rate

Solution / general: mol L^{-1} s^{-1}

Gas phase (via pressure): atm s^{-1} or bar min^{-1}

Use pressure units when monitoring gas-phase reactions at constant T and V — partial pressure is directly proportional to molar concentration.

Average vs instantaneous rate

Average rate is computed over a measurable interval Δt and is easy to obtain from tabulated data. **Instantaneous rate** is needed when the rate itself changes within Δt; it is obtained graphically as the slope of the tangent to the $[R]$ vs t curve at the instant of interest.

2 Factors Affecting Reaction Rate

The rate of a chemical reaction depends on the nature of the reactants, their concentrations (or partial pressures), the temperature, the presence of a catalyst, and (for heterogeneous systems) the surface area and physical state of the reactants.

Five key factors

(i) Concentration / pressure — higher concentration ⇒ more collisions per second ⇒ faster rate.
(ii) Temperature — 10 K rise roughly **doubles** the rate near room temperature.
(iii) Catalyst — lowers the activation energy by providing an alternative path.
(iv) Nature of reactants — ionic reactions are generally faster than covalent.
(v) Surface area — finely divided solids react faster than lumps.

3 Rate Law, Order & Molecularity

This section gathers the differential rate law, the experimental definition of order, the units of the rate constant by order, and the distinction between order and molecularity.

Rate law (differential form)

For $aA + bB \to cC + dD$:

$$\text{Rate} = k\,[A]^x[B]^y$$

$$-\frac{d[R]}{dt} = k\,[A]^x[B]^y$$

where k = rate constant; x, y = orders w.r.t. A and B (*determined experimentally*).

x, y **need not** equal the stoichiometric coefficients a, b. The rate law cannot be predicted from the balanced equation alone — only experiment fixes it.

Order of reaction

Overall order $n = x + y$

where x = order w.r.t. A; y = order w.r.t. B.

Order can be **0, 1, 2, 3** or even a **fraction**. A zero-order reaction has a rate independent of concentration.

Rate constant — units by order

$$k = \frac{\text{Rate}}{[A]^x[B]^y} = \frac{\text{(conc)}}{\text{time}} \cdot \frac{1}{\text{(conc)}^n}$$

$$\Rightarrow \ \text{units of } k = (\text{mol L}^{-1})^{1-n}\,\text{s}^{-1}$$

The units of k **reveal the order** — a quick way to identify order when only the rate-constant units are given.

Order	Units of k
0	mol L^{-1} s^{-1}
1	s^{-1}
2	L mol^{-1} s^{-1}
3	L^2 mol^{-2} s^{-1}

Order vs Molecularity

Order — experimental; the sum of powers of concentration in the rate law. Can be 0, fractional or integer. Applies to elementary *and* complex reactions.

Molecularity — the number of reacting species (atoms, ions or molecules) colliding in an elementary step. Always a positive integer (1, 2 or rarely 3). Defined only for elementary reactions.

Elementary vs complex reactions

An **elementary reaction** occurs in a single step; its rate law follows directly from its molecularity. A **complex reaction** proceeds through a sequence of elementary steps (a mechanism); its observed rate is governed by the slowest step, called the **rate-determining step**.

Order $\neq$ stoichiometric coefficient

For $2NO_2 + F_2 \rightarrow 2NO_2F$, the experimentally observed rate law is $r = k[NO_2][F_2]$, not $k[NO_2]^2[F_2]$. **Never read order off the balanced equation** — always rely on experimental data.

4 Integrated Rate Equation — Zero Order

For a zero-order reaction the rate is independent of $[R]$. Integrating the differential rate law gives a linear $[R]$ vs t relation that lets us extract k directly from a single straight-line plot.

Zero order: integrated law

$$-\frac{d[R]}{dt} = k[R]^0 = k$$

$$[R] = [R]_0 - kt$$

$$k = \frac{[R]_0 - [R]}{t}$$

where $[R]_0$ = initial concentration; $[R]$ = concentration at time t.

Concentration falls **linearly** with time. A plot of $[R]$ vs t gives a straight line with **slope** $= -k$ and intercept $[R]_0$.

Zero order: half-life

$$t_{1/2} = \frac{[R]_0}{2k}$$

Half-life is **directly proportional** to initial concentration — a larger $[R]_0$ takes longer to halve. Units of k: mol L^{-1} s^{-1}.

Examples of zero-order reactions

Decomposition of NH_3 on hot platinum (high-pressure regime, surface fully saturated); thermal decomposition of HI on gold; many enzyme-catalysed reactions at saturating substrate. In each case the catalyst surface is the bottleneck, so the rate is independent of bulk $[R]$.

5 Integrated Rate Equation — First Order

This section covers the first-order integrated rate law in both natural-log and base-10 forms, the exponential decay form, the gas-phase pressure form, pseudo first-order reactions, and half-life.

First order: integrated law (ln form)

$$-\frac{d[R]}{dt} = k[R]$$

$$\ln[R] = \ln[R]_0 - kt$$

$$k = \frac{1}{t}\ln\frac{[R]_0}{[R]}$$

$$[R] = [R]_0\, e^{-kt}$$

where $[R]_0$ = initial concentration; $[R]$ = concentration at time t.

Concentration falls **exponentially**. A plot of $\ln[R]$ vs t is a straight line of slope $-k$ and intercept $\ln[R]_0$.

First order: $\log_{10}$ form (exam-friendly)

$$k = \frac{2.303}{t}\log\frac{[R]_0}{[R]}$$

$$\log\frac{[R]_0}{[R]} = \frac{kt}{2.303}$$

A plot of $\log([R]_0/[R])$ vs t is a straight line of slope $k/2.303$. Most NCERT numerical problems use this form.

First order: two-point form

$$k = \frac{1}{t_2 - t_1}\ln\frac{[R]_1}{[R]_2} = \frac{2.303}{t_2 - t_1}\log\frac{[R]_1}{[R]_2}$$

where $[R]_1, [R]_2$ = concentrations at times t_1, t_2.

Useful when the initial concentration is unknown — only two concentration-time data points are needed.

First order: half-life

$$t_{1/2} = \frac{0.693}{k} \qquad \text{(since } \ln 2 = 0.693\text{)}$$

Half-life is **independent of initial concentration** — a defining signature of first-order kinetics. The same $t_{1/2}$ holds no matter what $[R]_0$ is.

First order: gas-phase (total-pressure form)

For $A(g) \rightarrow B(g) + C(g)$:

$$k = \frac{2.303}{t}\log\frac{p_i}{p_A} = \frac{2.303}{t}\log\frac{p_i}{(2p_i - p_t)}$$

where p_i = initial pressure of A; p_t = total pressure at time t; $p_A = 2p_i - p_t$.

Used when the reaction is followed by total-pressure measurement instead of concentration.

Derived from $p_t = p_i + x$.

Pseudo first-order reactions

For a bimolecular reaction $A + B \rightarrow$ products with $[B] \gg [A]$ (e.g. $[B]$ = solvent water):

Rate $= k\,[A][B] \;\approx\; k'\,[A]$

$k' = k\,[B] \approx k\,[\mathrm{H_2O}]$ (effectively constant)

where k = true second-order rate constant; k' = observed pseudo-first-order rate constant (s^{-1}).

The reaction is truly second order, but because $[B]$ scarcely changes it *appears* first order. Examples: acid hydrolysis of ester (**$CH_3COOC_2H_5$ + H_2O**) and inversion of cane sugar.

JEE/NEET Extension: time for any fraction

For a first-order reaction, the time to drop from $[R]_0$ to a fraction $f\,[R]_0$ is:

$$t = \frac{2.303}{k}\log(1/f)$$

Time for 99.9% completion: $t = \dfrac{2.303}{k}\log 10^3 = 10\,t_{1/2}$.

6 Integrated Rate Equation — Second Order

NCERT focuses on zero and first order only; second order is summarised here for JEE/NEET completeness and for problems where the rate constant's units (L mol^{-1} s^{-1}) reveal the order.

JEE/NEET Extension: Second-order kinetics

For rate $= k[A]^2$:

Integrated form: $\dfrac{1}{[A]} - \dfrac{1}{[A]_0} = kt$

Half-life: $t_{1/2} = \dfrac{1}{k[A]_0}$ (inversely proportional to $[A]_0$)

Units of k: L mol^{-1} s^{-1}. A plot of $1/[A]$ vs t is linear with slope k.

JEE/NEET Extension: nth-order half-life

For rate $= k[A]^n$ with $n \neq 1$:

$$t_{1/2} = \frac{2^{n-1} - 1}{(n-1)\,k\,[A]_0^{n-1}}$$

So $t_{1/2} \propto [A]_0^{1-n}$ — the dependence of half-life on initial concentration itself tells you the order.

Order	Integrated form	Linear plot	$t_{1/2}$
0	$[R] = [R]_0 - kt$	$[R]$ vs t	$\frac{[R]_0}{2k}$
1	$\ln[R] = \ln[R]_0 - kt$	$\ln[R]$ vs t	$\frac{0.693}{k}$
2	$\frac{1}{[R]} - \frac{1}{[R]_0} = kt$	$\frac{1}{[R]}$ vs t	$\frac{1}{k[R]_0}$

7 Temperature Dependence — Arrhenius Equation

This section presents the Arrhenius equation, its logarithmic form for graphical analysis, and the two-temperature form used to compute E_a from a pair of rate constants. Catalyst action is also covered here.

Arrhenius equation

$$k = A\,e^{-E_a/RT}$$

where A = pre-exponential / frequency factor (same units as k); E_a = activation energy (J mol^{-1}); $R = 8.314$ J K^{-1}mol^{-1}; T = absolute temperature (K).

$e^{-E_a/RT}$ is the **fraction of molecules** with energy $\geq E_a$. A higher T or a lower E_a **both** raise the rate.

Arrhenius — logarithmic form

$$\ln k = \ln A - \frac{E_a}{RT}$$

$$\log k = \log A - \frac{E_a}{2.303\,RT}$$

A plot of $\ln k$ vs $1/T$ is a straight line with **slope** $= -E_a/R$ and intercept $\ln A$. The standard graphical method to extract E_a and A.

Arrhenius plot — slope formulas

$$\ln k \text{ vs } 1/T: \quad \text{slope} = -\frac{E_a}{R}, \quad \text{intercept} = \ln A$$

$$\log k \text{ vs } 1/T: \quad \text{slope} = -\frac{E_a}{2.303\,R}, \quad \text{intercept} = \log A$$

$$E_a = -R \times (\text{slope of } \ln k \text{ vs } 1/T)$$

$$E_a = -2.303\,R \times (\text{slope of } \log k \text{ vs } 1/T)$$

Both slopes are **negative** because k rises as T rises (so $\ln k$ rises while $1/T$ falls). Multiply the magnitude of the slope by R (or $2.303R$) to read off E_a directly.

Two-temperature form (for E_a)

$$\ln\frac{k_2}{k_1} = \frac{E_a}{R}\left[\frac{1}{T_1} - \frac{1}{T_2}\right]$$

$$\log\frac{k_2}{k_1} = \frac{E_a}{2.303\,R}\left[\frac{T_2 - T_1}{T_1 T_2}\right]$$

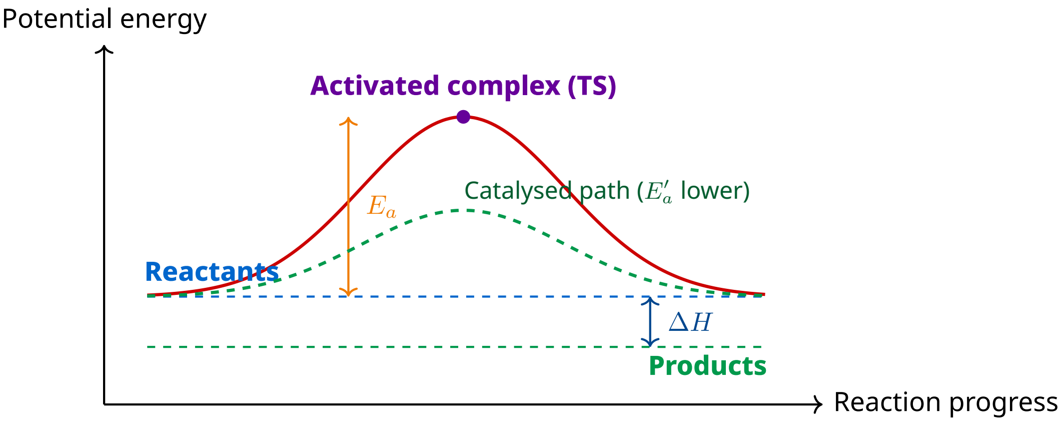
where k_1, k_2 = rate constants at T_1, T_2.

Workhorse formula for NCERT numericals — given two k values at two temperatures, solve for E_a (or vice versa).

Activation energy — definition

$E_a = E_{\text{activated complex}} - E_{\text{reactants}}$

Minimum extra energy reactants need to reach the **activated (transition) complex**. Independent of the path's enthalpy change ΔH.



Energy profile for an exothermic reaction. A catalyst supplies a lower-E_a route (dashed green) without changing ΔH.

Effect of a catalyst

A catalyst provides an **alternative pathway** with a **lower** E_a, so the Boltzmann fraction $e^{-E_a/RT}$ rises and the rate increases. A catalyst does **not** alter ΔG, ΔH or the equilibrium constant K — it speeds up the forward and reverse reactions equally, reaching the same equilibrium faster.

Temperature coefficient

$$\mu = \frac{k_{(T+10)}}{k_T} = \frac{k_{35°\text{C}}}{k_{25°\text{C}}} \approx \text{ 2 to 3}$$

where k_T = rate constant at temperature T (K); $k_{(T+10)}$ = rate constant 10 K higher.

Rule of thumb: near room temperature a **10 K rise roughly doubles to triples** the rate. Consistent with typical E_a values of 50–100 kJ mol^{-1} via the Arrhenius equation.

E_a units — J or kJ?

R is in J K^{-1}mol^{-1}, so E_a must be in **J mol**$^{-1}$ inside E_a/RT. If a problem gives E_a in kJ mol^{-1}, multiply by 10^3 before substituting.

Sign of slope in Arrhenius plot

Plot ln k on the y-axis and $1/T$ on the x-axis. The line slopes **down to the right**: as T rises, $1/T$ falls and ln k rises. Hence the slope $= -E_a/R$ is **negative**.

8 Collision Theory of Bimolecular Reactions

This section covers the collision-theory expression for the rate of a bimolecular reaction, the role of collision frequency, threshold energy, and the steric (orientation) factor.

Collision-theory rate

For bimolecular $A + B \to$ products:

Rate $= Z_{AB}\, e^{-E_a/RT}$

With orientation correction:

Rate $= P\, Z_{AB}\, e^{-E_a/RT}$

$\Rightarrow\ k = P\, Z_{AB}\, e^{-E_a/RT}$

where Z_{AB} = collision frequency of A and B (collisions per unit volume per second); P = steric / probability factor ($0 < P \leq 1$); $e^{-E_a/RT}$ = fraction with energy $\geq E_a$.

Comparing with the Arrhenius equation shows $A \approx P\, Z_{AB}$, so the **pre-exponential factor encodes both** collision frequency and orientation.

Effective collisions

Not every collision yields products. A collision is **effective** only when the colliding molecules have (i) energy $\geq$ threshold energy and (ii) the **proper orientation** for bond reorganisation. The steric factor P quantifies the orientation requirement.

Threshold energy

Threshold energy $= E_a +$ average energy of reactants. Only molecules whose kinetic energy exceeds the threshold can react. The Maxwell-Boltzmann distribution shifts to higher energies as T rises, so the reactive fraction (area beyond E_a) increases sharply.

Quick Reference — Chemical Kinetics at a Glance

Every key formula in this chapter, side-by-side

Topic	Formula	Key fact
Average rate	$r_{\text{av}} = -\frac{\Delta[R]}{\Delta t} = +\frac{\Delta[P]}{\Delta t}$	Over a finite interval
Instantaneous rate	$r_{\text{inst}} = -\frac{d[R]}{dt}$	Slope of tangent
General rate	$\text{Rate} = -\frac{1}{a}\frac{d[A]}{dt} = \frac{1}{c}\frac{d[C]}{dt}$	Divide by coefficient
Rate law	$\text{Rate} = k[A]^x[B]^y$	x, y experimental
Order	$n = x + y$	Can be 0, fraction
Units of k	$(\text{mol L}^{-1})^{1-n}\,\text{s}^{-1}$	Order from units
Zero order	$[R] = [R]_0 - kt$	$[R]$ vs t linear
Zero order $t_{1/2}$	$t_{1/2} = \frac{[R]_0}{2k}$	$\propto [R]_0$
First order	$k = \frac{2.303}{t}\log\frac{[R]_0}{[R]}$	$\ln[R]$ vs t linear
First order exponential	$[R] = [R]_0\, e^{-kt}$	Exponential decay
First order $t_{1/2}$	$t_{1/2} = \frac{0.693}{k}$	Independent of $[R]_0$
First order (gas)	$k = \frac{2.303}{t}\log\frac{p_i}{2p_i - p_t}$	Total-pressure form
Pseudo first order	$k' = k\,[\text{H}_2\text{O}]$; $\text{rate} = k'[A]$	Solvent in excess
Second order (JEE)	$\frac{1}{[A]} - \frac{1}{[A]_0} = kt$	$t_{1/2} = 1/(k[A]_0)$
Arrhenius	$k = A\,e^{-E_a/RT}$	Exponential in T
Arrhenius (log)	$\log k = \log A - \frac{E_a}{2.303\,RT}$	$\log k$ vs $1/T$ linear
Arrhenius slope	$\text{slope}(\ln k \text{ vs } 1/T) = -E_a/R$	E_a from graph
Two-temperature E_a	$\log\frac{k_2}{k_1} = \frac{E_a}{2.303\,R}\frac{T_2 - T_1}{T_1 T_2}$	Solve for E_a
Temperature coefficient	$\mu = k_{(T+10)}/k_T \approx 2\text{–}3$	10 K rule
Catalyst effect	lowers E_a, K unchanged	Speeds both directions
Collision theory	$k = P\,Z_{AB}\,e^{-E_a/RT}$	$A \approx P\,Z_{AB}$

Distinguishing the three orders by signature

Look for	Order
Constant rate; $[R]$ vs t straight line; k in mol L^{-1} s^{-1}	0
$t_{1/2}$ independent of $[R]_0$; k in s^{-1}	1
$t_{1/2}$ doubles when $[R]_0$ halves; k in L mol^{-1} s^{-1}	2