Chemical Kinetics Class 12 Formula Sheet

CBSE Class 12 Chemistry · NCERT · for boards, NEET and JEE revision

ChapterNotes

Rate of Reaction

$$\text{Rate} = -\frac{1}{a}\frac{d[A]}{dt} = +\frac{1}{c}\frac{d[C]}{dt} \text{ for } aA \rightarrow cC$$

Units: mol L^{-1} s^{-1}. Reactants get a minus sign, products a plus sign.

- Average rate = Δ[X]/Δt; instantaneous rate = slope of conc-time curve
- Divide each species by its coefficient to get one common rate
- For $N_2 + 3H_2 \rightarrow 2NH_3$, H_2 vanishes 3× and NH_3 forms 2× as fast as N_2

Rate Law

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

k = rate constant (depends on temperature, not on concentration).

- Exponents x, y are experimental, NOT the balanced-equation coefficients
- For multistep reactions the slowest (rate-determining) step fixes the rate law
- k changes with temperature; doubling [A] that doubles rate ⇒ order 1 in A

Order vs Molecularity

$$\text{Order} = x + y \text{ ; Molecularity} = 1, 2, 3 \text{ (whole number only)}$$

Order can be 0, fractional or negative; molecularity is always a positive integer.

- Order comes from experiment; molecularity comes from the mechanism
- Molecularity is never zero or fractional
- Higher order means more sensitive to concentration, not necessarily faster

Units of the Rate Constant

$$k = (\text{mol } L^{-1})^{1-n}\, s^{-1} \text{ for overall order } n$$

Because rate is always mol L^{-1} s^{-1}, the units of k must balance it.

- Zero order: k in mol L^{-1} s^{-1}
- First order: k in s^{-1}
- Second order: k in L mol^{-1} s^{-1}

Zero-Order Integrated Rate

$$[A] = [A]_0 - k\,t$$

Plot of [A] vs t is a straight line with slope −k.

- Rate is constant no matter how much reactant remains
- Examples: decomposition of NH_3 on hot Pt, many photochemical reactions
- Slope of [A]-vs-t line gives −k directly

First-Order Integrated Rate

$$k = \frac{2.303}{t}\log\frac{[A]_0}{[A]} \Leftrightarrow \ln\frac{[A]_0}{[A]} = k\,t$$

Plot of log[A] vs t is linear with slope −k/2.303. The 2.303 converts ln to $\log_{10}$.

- All radioactive decay is first order
- k is independent of initial concentration and of concentration units
- Use fraction REMAINING (not % reacted) in $[A]_0/[A]$

Half-Life (t½) and Order

$$\text{First order: } t_{1/2} = \frac{0.693}{k} \quad \text{Zero order: } t_{1/2} = \frac{[A]_0}{2k}$$

0.693 = ln 2. General rule: $t_{1/2} \propto [A]_0^{1-n}$.

- First-order t½ is independent of initial concentration (its hallmark)
- Zero-order t½ is directly proportional to $[A]_0$
- After n half-lives, fraction remaining = $(1/2)^n$

Arrhenius Equation

$$k = A\,e^{(-E_a/RT)} \Leftrightarrow \log k = \log A - \frac{E_a}{2.303\,R\,T}$$

T in Kelvin; rule of thumb: rate roughly doubles per 10°C rise.

- Plot of log k vs 1/T is a straight line of slope $-E_a/(2.303\,R)$
- Two-temperature form: $\log\frac{k_2}{k_1} = \frac{E_a}{2.303R}\cdot\frac{T_2 - T_1}{T_1 T_2}$
- Lower E_a ⇒ faster reaction; heating shifts Maxwell-Boltzmann to higher energy

Collision Theory

$$\text{Rate} = Z_{AB} \cdot e^{(-E_a/RT)} \cdot P$$

Z_{AB} = collision frequency; P = steric (orientation) factor, usually < 1.

- Two conditions: energy ≥ E_a AND proper molecular orientation
- $e^{(-E_a/RT)}$ is the fraction of energetic (effective) collisions
- P corrects collision theory, which otherwise overpredicts the rate

Catalysis

$$\text{Catalyst lowers } E_a \text{ (forward \& reverse equally)} \cdot \Delta H \text{ and } K_c \text{ unchanged}$$

It is not consumed; it is regenerated at the end.

- Does not change ΔH or the position of equilibrium
- Lowers E_a for both directions, so equilibrium is reached faster
- A small drop in E_a gives a huge rise in rate (E_a sits in an exponent)

Quick revision

- Rate = change in concentration per unit time; always positive
- Order: experimental; Molecularity: theoretical (elementary step only)
- Zero order: $[A] = [A]_0 - kt$, $t_{1/2} = [A]_0/2k$
- First order: $k = (2.303/t)\log([A]_0/[A])$, $t_{1/2} = 0.693/k$ (independent of $[A]_0$)
- Pseudo first order: 2nd order reaction with one reactant in large excess
- Arrhenius: $k = Ae^{(-Ea/RT)}$; plot log k vs 1/T → straight line
- For two temperatures: $\log(k_2/k_1) = (Ea/2.303R)[(T_2 - T_1)/(T_1T_2)]$
- Catalyst lowers Ea but doesn't change ΔH or equilibrium
- 10°C rise ≈ doubles the rate (temperature coefficient ≈ 2)
- All radioactive decays follow first order kinetics

Full chapter notes with solved examples and practice questions: https://chapternotes.in/cbse/class-12-chemistry/chemical-kinetics-class-12-notes/