

Chemical Reactions and Their Rates

rate laws

$$\text{rate} = k[A]^x[B]^y[C]^z$$

$$\text{zero-order rate} = k$$

$$\text{first-order rate} = k[A]$$

$$\text{second-order rate} = k[A]^2$$



$$[A]_t = -kt + [A]_0$$



$$\ln[A]_t = -kt + \ln[A]_0$$

$$y = mx + b$$



$$\frac{1}{[A]}_t = kt + \frac{1}{[A]_0}$$

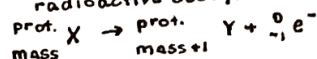
$$y = mx + b$$

half-life the time it takes for half of a substance to react
 ↳ remains constant for a first-order reactant

$$\text{half-life} = \frac{\ln(2)}{k} \quad \text{only applies to first-order (find } k \text{ using rate law)}$$

↳ varies for zero- and second-order reactant
 k remains constant w/ changes in temp

radioactive decay:



collision theory

- ↳ when reactants collide with sufficient energy (activation energy) a reaction occurs
- ↳ for aqueous & gaseous reactants, increased concentration increases rate of reaction (more molecules = more collisions)
- ↳ for solid reactants, reactions proceed faster with higher surface area (more "available" atoms)
- ↳ stirring increases the reaction rate in a heterogeneous mixture (increased likelihood of collisions)
- ↳ reaction rates increase with increased temperature due to higher KE (more molecules have sufficient E_a)
- ↳ reactions only happen with collisions in the correct orientation

Beer's Law

$$A = \epsilon lc$$

A = absorbance

ϵ = molar absorptivity

l = path length

c = concentration of solution

most effective with solutions that visibly change color over time
 spectrophotometer measures the amount of light at a specific wavelength that is absorbed by a solution

reaction mechanisms

reactions happen as sum of elementary steps

intermediates produced but fully consumed in a later step

unimolecular elementary step with one reactant

bimolecular elementary step with two reactants

rate-determining step the slowest step in a reaction

must do substitution w/ previous step if intermediate appears in rate law

catalyst consumed then re-made by overall reaction process

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