Big Idea #4

Chemical Reactions and Their Rates

rate laws

$rate = k[A]^x[B]^y[C]^z$

zero-order $rate = k$

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

first-order $rate = k[A]$

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

$y = mx + b$

second-order $rate = k[A]^2$

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

$y = mx + b$

half-life the time it takes for half of a substance to react

↳ remains constant for a first-order reactant

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

↳ varies for zero- and second-order reactant

k remains constant w/ changes in temp

radioactive decay:

prot. mass $X \rightarrow$ prot. mass +1 $Y + {}^{0}_{-1}e^{-}$

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

$A = \varepsilon \ell c$

A = absorbance

ε = molar absorbtivity

ℓ = path length

c = concentration of solution

most effective with solutions that visibly change color over time

spectrophotometer measures the amount of light of 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