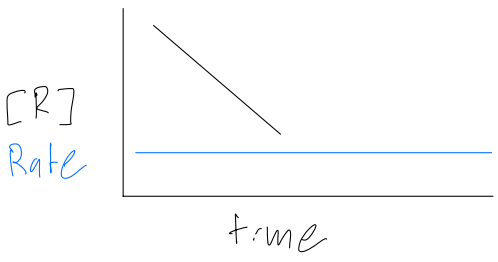
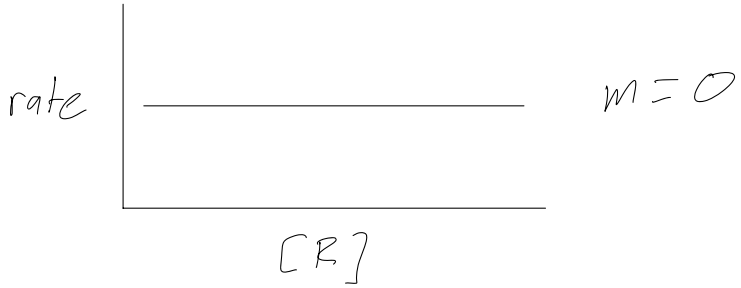
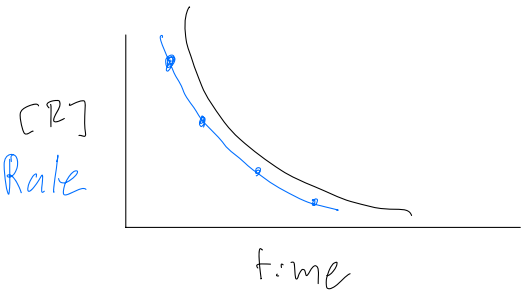
Kinetics → rates of reactions

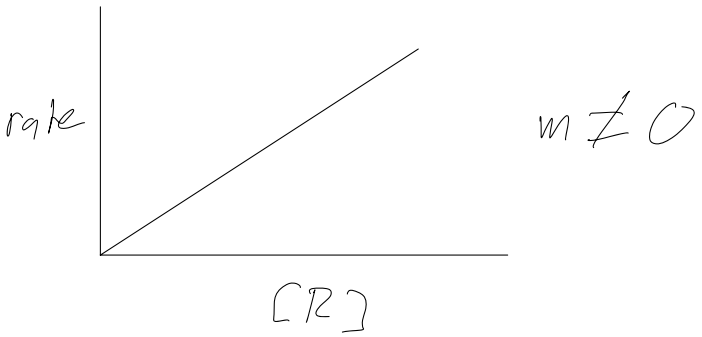
$R \rightarrow P$ $\quad Rate = \frac{\Delta[P]}{\Delta t} = -\frac{\Delta[R]}{\Delta t}$





$m = 0$





$m \neq 0$

Rate Law → $Rate = k[R]^m$ ← "order of reaction" usually 0, 1, 2

↳ rate constant

Goal - figure out rate law

Ideal - measure [R] as function of time
- not feasible for our situation

since we can't directly measure [R] vs time
we will measure initial rate - right at beginning of reaction
- for only a small fraction of R reacting

$[R]_{at\ beginning} \approx [R]_{at\ end}$ because we are only looking at a very small fraction of total time of reaction

Repeat for different $[R]_{init}$

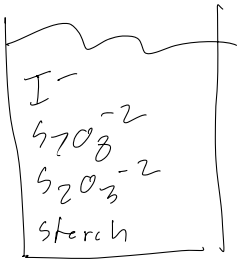
$\Rightarrow$ Rate law

$2 I^- + S_2O_8^{-2} \longrightarrow I_2 + 2 SO_4^{-2}$ $\quad$ Rate $= k[I^-]^m [S_2O_8^{-2}]^n$

Use secondary "clock" reaction to measure initial rate

$I_2 + 2 S_2O_3^{-2} \rightarrow 2 I^- + S_4O_6^{-2}$

starch is an indicator — turns blue in presence of I_2



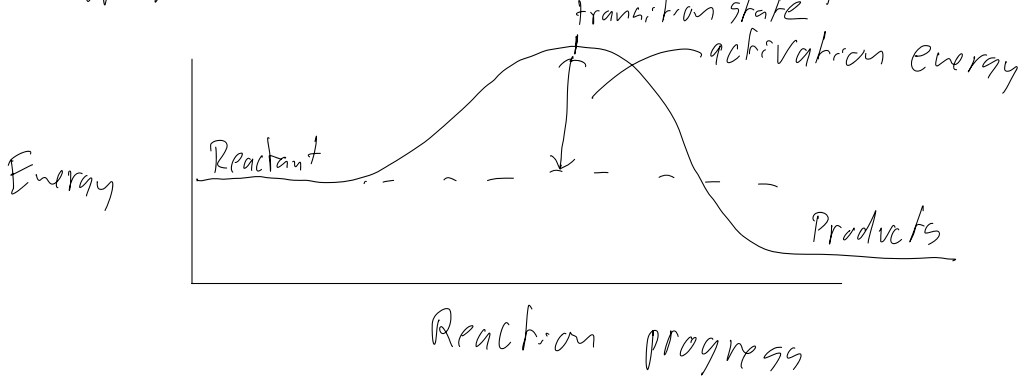
when we run out of $S_2O_3^{-2}$ the I_2 formed doesn't revert to I^- instead it reacts w/starch and color change occurs

Ex: $[I^-]_{init} = 1\ M$

$[S_2O_3^{-2}]_{init} = 0.001\ M$

Use timer to see how long it takes for $S_2O_3^{-2}$ to be used up
use starch to find how much I^- reacted } gives rate

We also will find activation energy



higher temp → molecules have more energy → higher probability of getting over energy barrier

Rate law → k depends on temperature

$$\ln k = \frac{-E_a}{RT} + \text{constant}$$