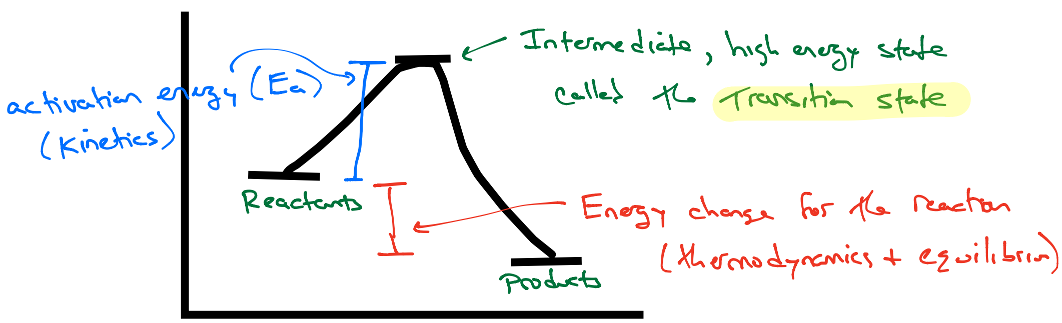
For the next several weeks, our focus will be on understanding some of the physical characteristics of chemical/physical reactions.

3 Pieces:

- Kinetics = the study of reaction rates
- Equilibrium = the study of reaction progress (think % yield)
- Thermodynamics = study of the energy of reactions

All 3 can be related to this image below



Arrhenius Equation

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

↑ rate constant → discussed below

Ways to increase the reaction rate:

① Increase Temp → ↑T = ↑KE = more likely molecules will collide
- In eq. above, ↑T = ↑k = ↑rate

② Increase concentration → ↑ # of collisions

③ ↓ E_a → this can happen only if a catalyst is added to the reaction

Kinetics is the study of reaction RATES. Just like any rate (eg MPH), it is a measure of change vs. time

Speed: Pt A - - - - - - - Pt. B (10 miles)

If it takes 0.1 hours to get from Pt. A to Pt. B, then the rate is

$$\text{rate} = \frac{10 \text{ miles}}{0.1 \text{ hours}} = 100 \text{ mph} = \frac{\Delta \text{distance}}{\Delta \text{time}}$$

For a chemical reaction:

$$2NO_2 \longrightarrow N_2O_4$$

$$\text{The reaction rate} = \frac{\Delta \text{concentration}}{\Delta \text{time}} = \frac{\Delta[N_2O_4]}{\Delta t} = -\frac{1}{2}\frac{\Delta[NO_2]}{\Delta t}$$

Note that 2 NO_2 are consumed for every 1 N_2O_4 produced. So $[NO_2]$ is changing twice as fast as $[N_2O_4]$ → this stoichiometry is incorporated into the expression as ½

- note also that NO_2 is being consumed, so the change (final − initial) is actually negative. For the rate of change to be equal to ΔN_2O_4 (which is (+)), a (−) must be included.

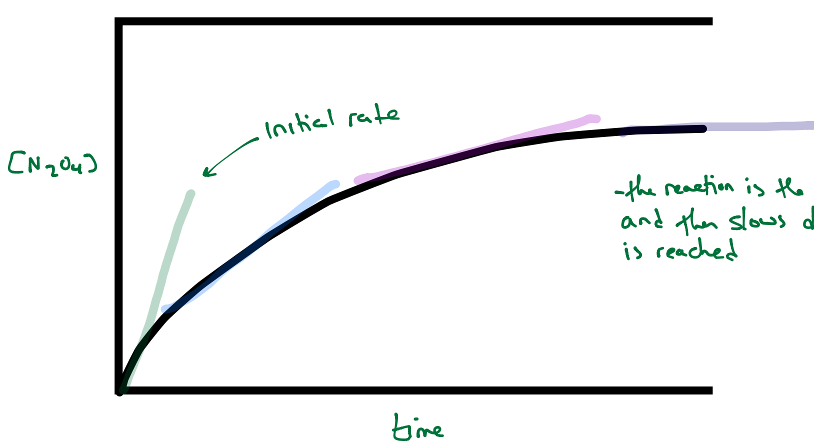
Based on our observations above:

Slope of NO_2 trace changes twice as fast as slope of N_2O_4

$[N_2O_4]$↑ as $[NO_2]$↓

The slope of the line @ any point is the rate of the reaction AT THAT SPECIFIC time.

It's important to note that the slopes change as time progresses. The smaller Δt we look at, the more accurate we can estimate the reaction rate.



- the reaction is the fastest at the beginning and then slows down as equilibrium is reached

This is all great, but it doesn't give us a very convenient way of describing a reaction.

- Recall that one of our observations about rates is that they are concentration dependent.

Another way of saying this is that reaction rate is proportional to concentration

rate ∝ concentration → $rate = k[N_2O_5]^x$ ← this is called a rate law

- we define a proportionality constant → RATE CONSTANT
  - Unique for a given reaction
  - Temperature dependent

note that rate law typically depends on [reactants]

- we don't know how much the reaction rate depends on [ ], so we include a factor (exponent) that allows reactant concentration to carry different weights; again, this is reaction dependent
- this coefficient is NOT directly related to stoichiometry UNLESS it is an elementary reaction

For example:

$4NO_2 + 1O_2 \rightarrow 2N_2O_5$ $rate = k[NO_2]^4[O_2]^1$

[reactant] included in rate law

coefficients are the same as exponents in the rate law. This is an elementary reaction

$N_2 + 3H_2 \rightarrow 2NH_3$ $rate = k[N_2][H_2]$

coefficients ≠ exponents → NOT elementary!

Reaction order: add up the exponents!

$rate = k[N_2][H_2]$ = 2nd order

$rate = k[NO_2]^4[O_2]$ = 5th order

$rate = k$ = 0th order

$rate = k[X]^3$ = 3rd order

Pretty easy, right?

The reaction order absolutely predicts the units of the rate constant!

the key is recognizing that units in a rate law MUST make sense

1st order: $rate = k[X]$, $\frac{M}{s} = \boxed{\frac{1}{s}} M$, $k = s^{-1}$

2nd order: $rate = k[X]^2$, $\frac{M}{s} = \boxed{\frac{1}{Ms}} M^2$, $k = s^{-1}M^{-1}$

3rd order: $rate = k[X]^3$, $\frac{M}{s} = \boxed{\frac{1}{M^2 s}} M^3$, $k = M^{-2}s^{-1}$

Example: the rate constant for a reaction is determined to be $1.86 \times 10^5\ M^{-6} s^{-1}$. What is the order of the reaction?

rate $= k\,[X]^{y}$ need to find y

$\frac{M}{s} = \frac{1}{M^6 s} (M)^y$ $y = 7$ ... this will leave $\frac{M}{s}$ on both sides of the equation.

7th order