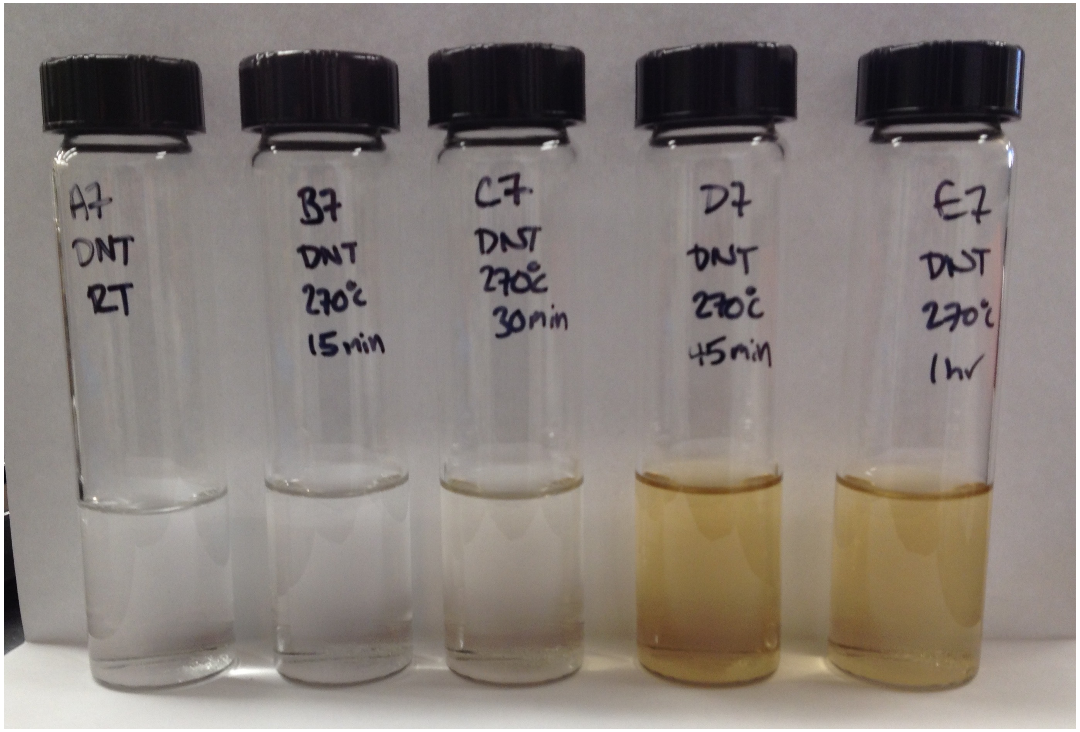
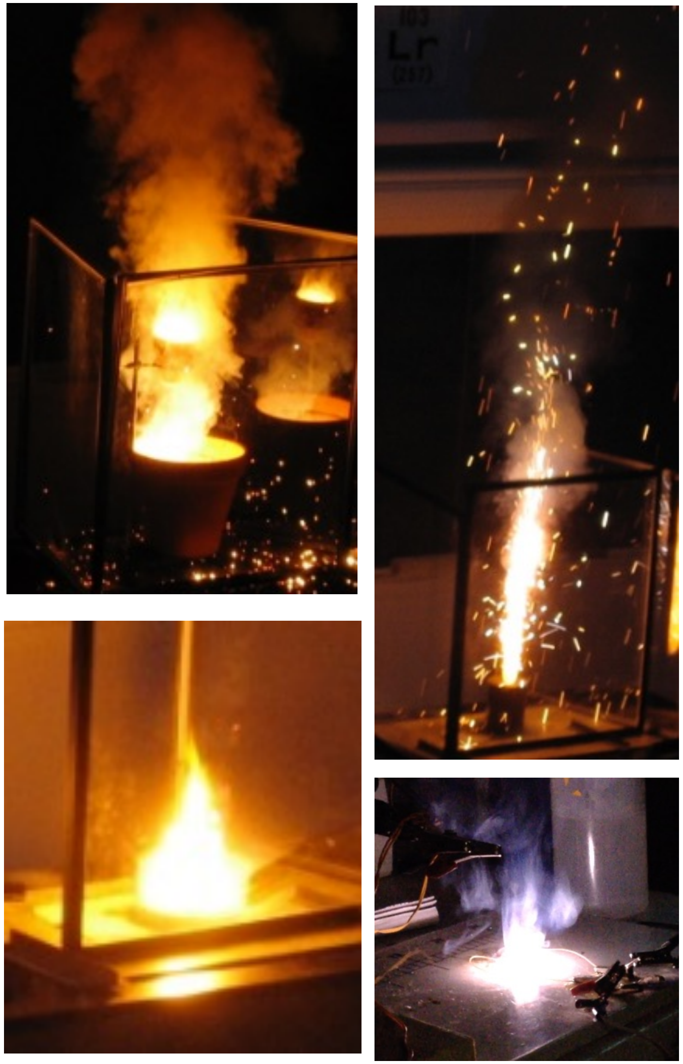
Chapter Fourteen:

Chemical Kinetics



Rate of Reaction = speed



Kinetics = study of how fast reactions take place

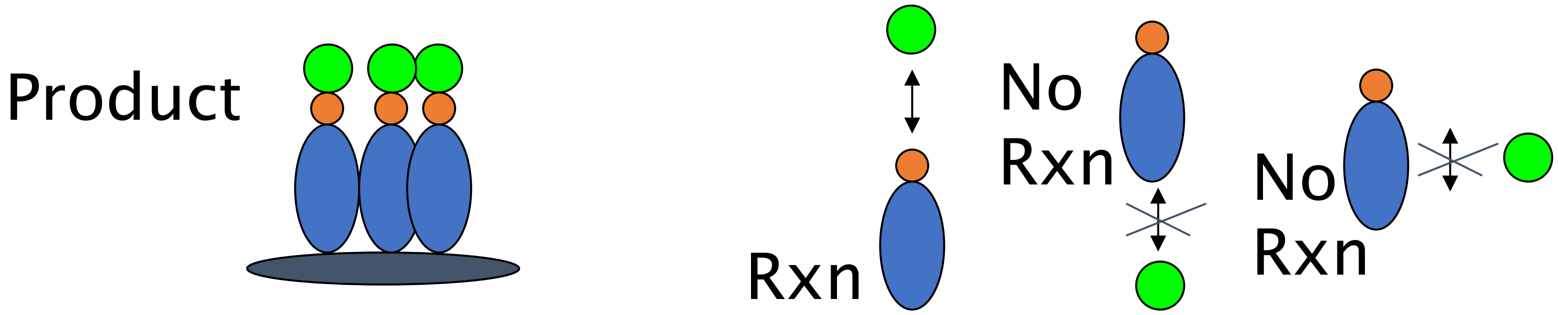
Collision Theory

Three things must happen for a reaction to occur:

1. Reacting molecules must collide

   Can increase the number of collisions with
   - increased reactant concentration
   - higher temps (= faster molecules)

2. Molecules must have the correct orientation



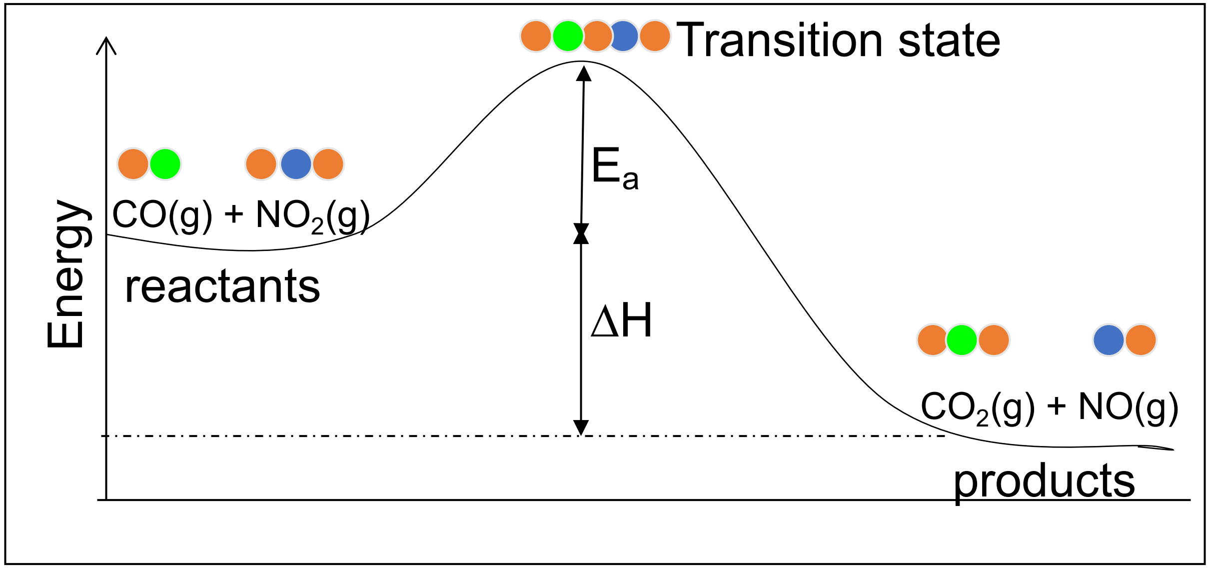
3. Activation Energy (E_a) must be exceeded

E_a = minimum energy required for a reaction to occur

Energy Diagrams: Activation Energy & Transition State

Given the following reaction:

$CO(g) + NO_2(g) \rightarrow CO_2(g) + NO(g)$



Reactants: Original components before a reaction

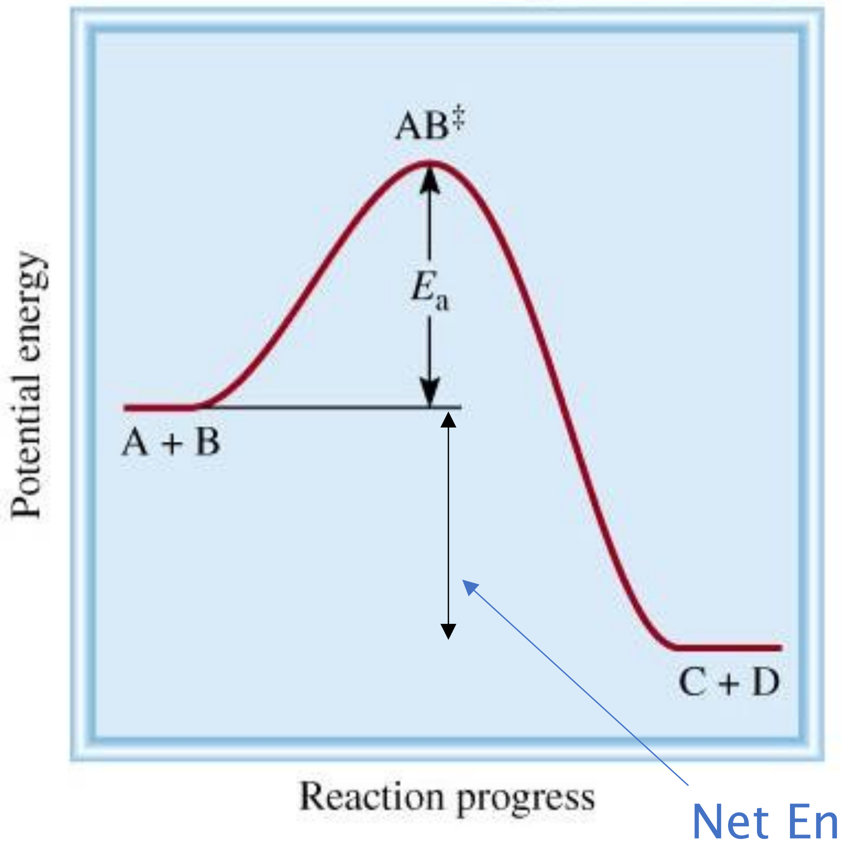
Products: New elements/compounds formed from reaction

Transition State (aka Activated Complex): Temporary & generally unstable combination of atoms formed from molecular collision. Will be transformed into products as the reaction proceeds.

Energy Diagrams: Thermodynamics & Reaction Progress

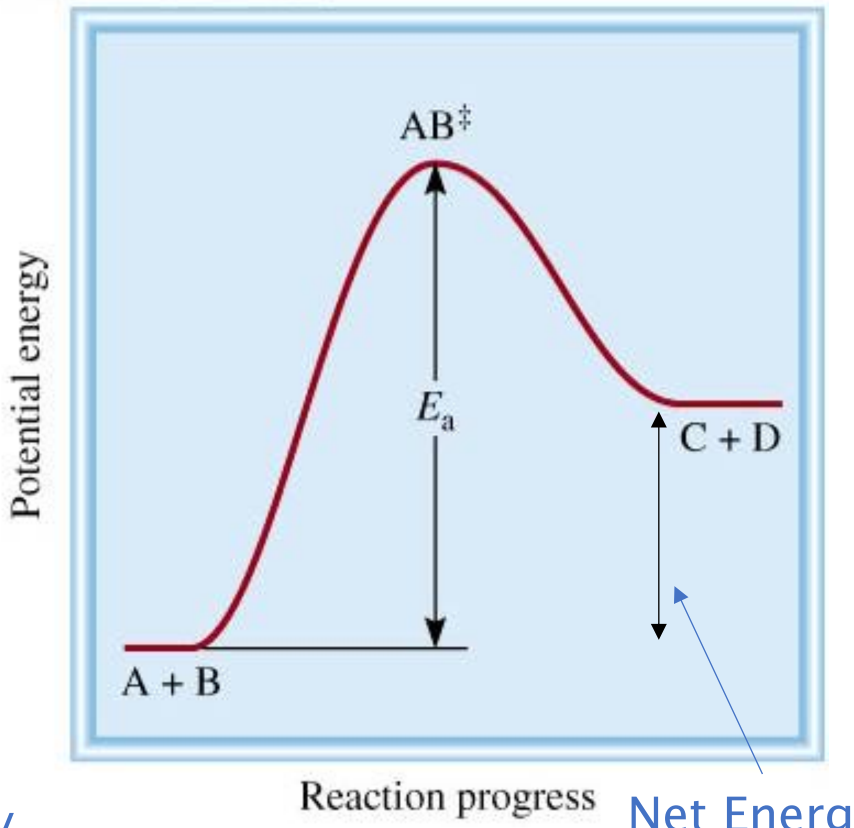
$A + B \rightarrow AB^{\ddagger} \rightarrow C + D$

Exothermic Reaction



Net Energy Released

Endothermic Reaction

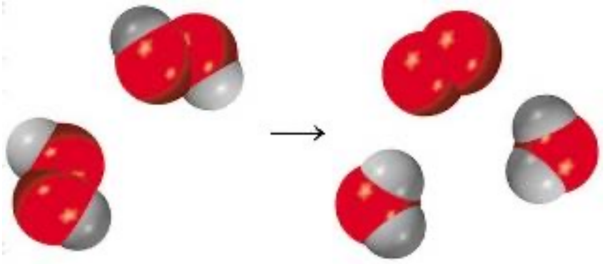


Net Energy Absorbed

Reaction Rates

Kinetics:

- Study of rates of chemical reactions
- How fast does a reaction proceed?



Reaction Rate:

$$\text{Rate} = \frac{\Delta\text{Concentration}}{\Delta\text{Time}} = \frac{[\text{mol/L}]}{\Delta T} = \frac{\text{M}}{\text{sec}}$$

$$2\ H_2O_2\ (l) \rightarrow 2\ H_2O\ (l)\ +\ O_2\ (g)$$

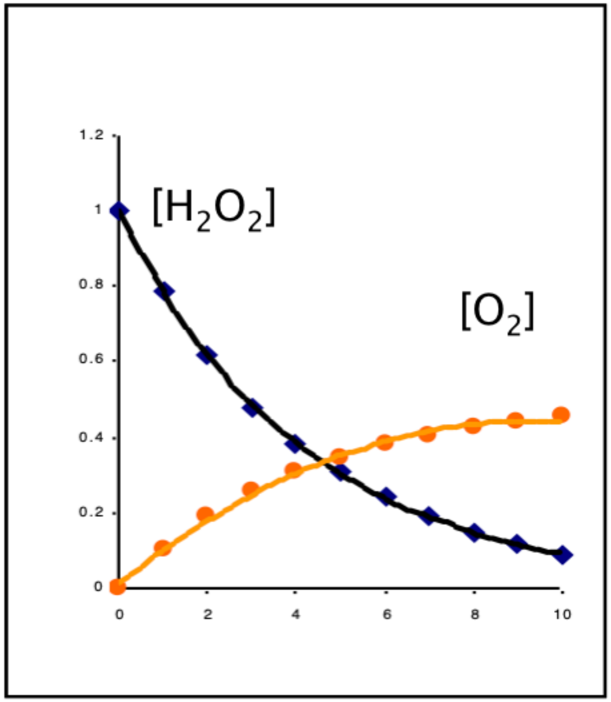
Rate of formation of product:

$\text{Rate} = +\Delta[O_2]/s$

Rate of disappearance of reactant:

$\text{Rate} = -\Delta[H_2O_2]/s$

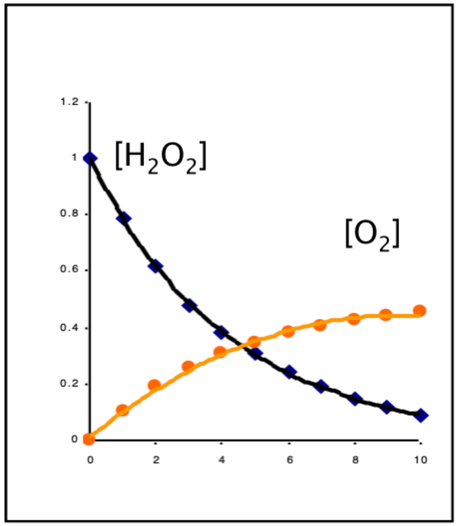
(Use of negative makes rate positive, sometimes rate is written as negative)



Linearity of Reaction Rates

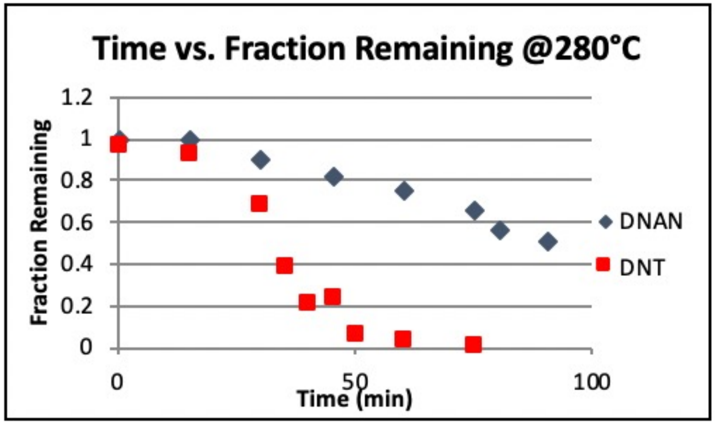
Reaction rates vary greatly

- Very fast; ex: burning
- Very slow; ex: disintegration of plastic in sunlight



Often not linear

- Rate is often faster at the beginning
- Rate is slowest at the end
- Some reactant often left over
- Especially pronounced for slow reactions



Average Rate of Reaction:

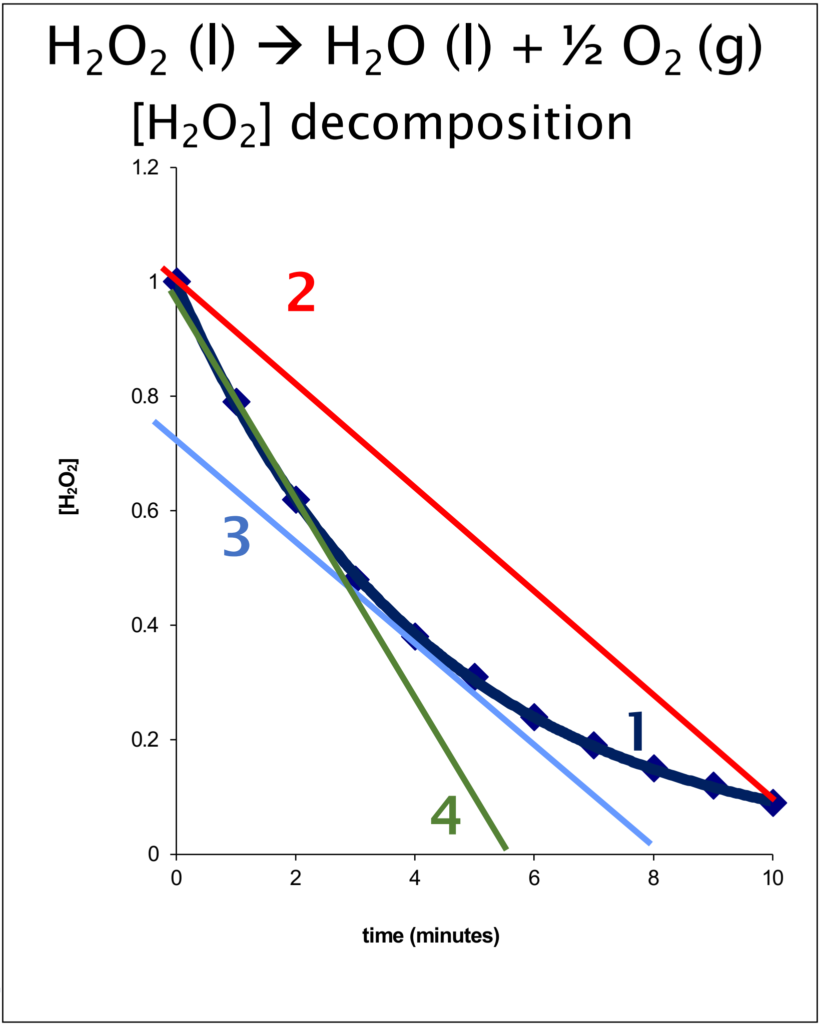
$$\text{Rate} = \frac{\text{Final Concentration} - \text{Initial Concentration}}{\text{Elapsed Time}}$$

Types of Reaction Rates

Reaction rate = ΔConcentration/Δtime = slope

1. Plot of original data
   - [Conc.] vs time
   - If curved = rate changes
     - Will reach equilibrium
   - If linear = rate is constant
     - Will run out of reactant
2. Average rate of reaction
3. Instantaneous rate
   - Tangent to curve 1
   - Can pick any time
4. Initial rate
   - t=0 to t=given time
   - Generally portion of curve 1 that is linear

$H_2O_2\ (l) \rightarrow H_2O\ (l) + \frac{1}{2}\ O_2\ (g)$



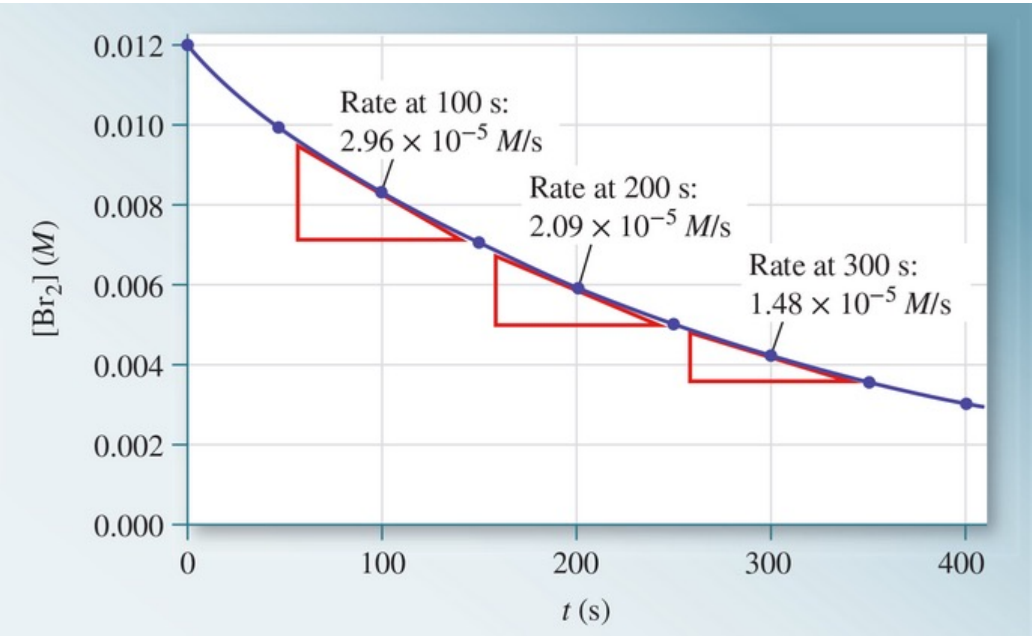
Instantaneous Rate

$Br_2\ (aq) + HCOOH\ (aq) \rightarrow 2\ Br^-\ (aq) + 2H^+\ (aq) + CO_2\ (g)$

Brown color of Br_2 disappears as reaction progresses

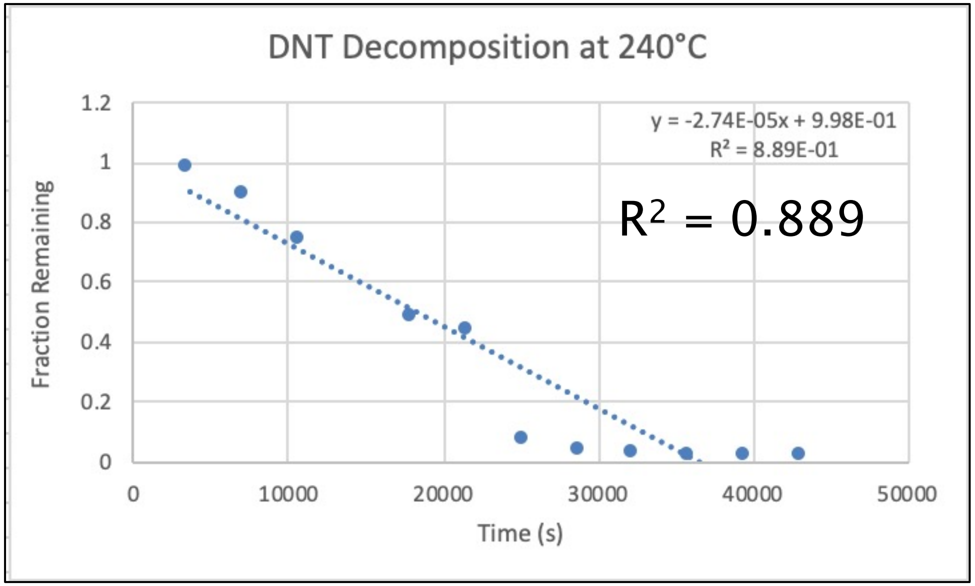
To Find Instantaneous Rate:

- Select time
- Draw line tangent to curve at that time
- Find slope of the tangent line
- Units of [conc.]/time

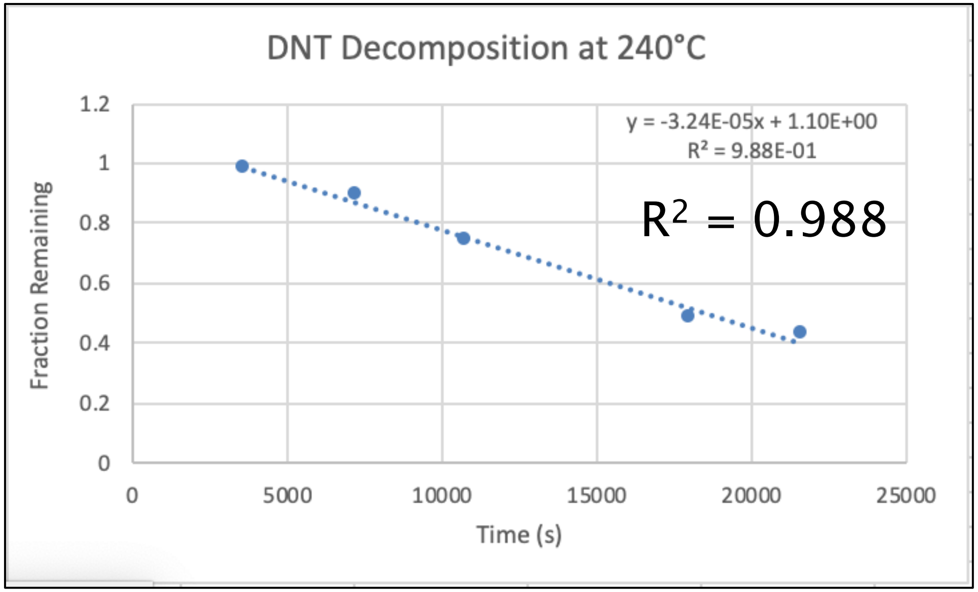


Initial Rate

Portion of original data where curve is linear



$R^2 = 0.889$



$R^2 = 0.988$

Stoichiometry & Reaction Rate

When writing rates based on different reactants or products, think about what is happening

$$\mathbf{Br_2\ (aq) + HCOOH\ (aq) \rightarrow 2\ Br^{-}\ (aq) + 2H^{+}\ (aq) + CO_2\ (g)}$$

Rate at which Br_2 disappears: X M/s

In terms of Br_2:

What is the rate of formation of Br^-? 2X M/s

What is the rate of disappearance of HCOOH? X M/s

Rate at which Br^- forms: Z M/s

In terms of Br^-:

What is the rate of disappearance of Br_2? ½ Z M/s

What is the rate of formation of CO_2? ½ Z M/s

This method works for zero order & instantaneous rates, not for more complex rates.

Stoichiometry & Reaction Rate

1. The rate of decomposition of N_2O_5 at a particular instant in a reaction vessel is 4.2×10^{-7} M/s. What is the rate of appearance of (a) NO_2 and (b) O_2?

$$2\,N_2O_5(g) \rightarrow 4NO_2(g) + O_2(g)$$

(a) 8.4×10^{-7} M/s

(b) 2.1×10^{-7} M/s

2. Consider the reaction:

$$N_2(g) + 3H_2(g) \rightarrow 2NH_3(g)$$

Suppose that at a particular moment during the reaction hydrogen is reacting at the rate of 0.074M/s. (a) At what rate is ammonia being formed? (b) At what rate is nitrogen reacting?

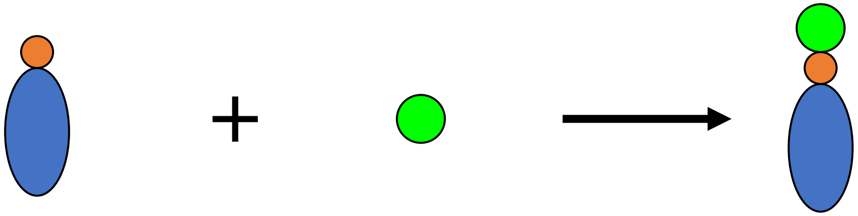
(a) 0.049 M/s

(b) –0.025 M/s

Factors Affecting Reaction Rates

Increasing # collisions increases rate

Reaction:



1. Increase concentration of reactant

More particles in same volume = greater chance of collision

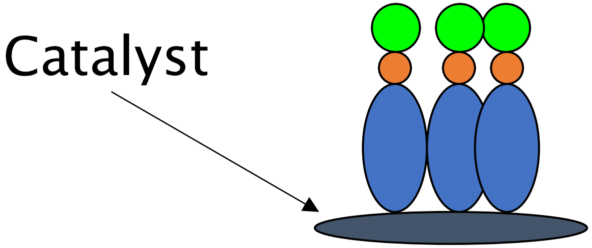
2. Increase temperature

Particles move faster – more likely to collide with enough E_A

3. Add a catalyst

Catalyst: chemical that does not change the amount of overall product or reactants, only the rate of the reaction

- fix molecular orientation
- alter transition state; lower E_a
- increase desired surface area



4. Improve Mixing

More interactions = greater chance of effective collision

Rate Laws:

Show Impact of Reactant Concentration on Rate

Rate Law Format: For the reaction A + B + C … → Products

Rate Law: $\text{Rate} = k[A]^m[B]^n[C]^p\ldots$

Based on <u>initial</u> concentrations and rate

Variables required in Rate Law:
k = rate constant
m = order of reaction in reactant A
n = order of reaction in reactant B
p = order of reaction in reactant C
other letters as needed for additional reactants

Overall Order of Reaction = sum of the orders of reaction for each reactant
Overall Order = $m + n + p$

Rate Laws account for non-linearity of reactions

Rate Laws con't

Properties of Rate Laws:

- Orders do not need to be positive integers
  - Ex: if concentration does not impact rate, order is zero
- Orders must be found experimentally
- Catalyst concentrations are included
- Rate Laws hold only for a specific temperature
- Rate Laws do **NOT** come from the balanced equation!

Not squared!

Example:

$2H_2(g) + 2NO(g) \rightarrow N_2(g) + 2H_2O(g)$ Rate = $k[H_2][NO]^2$

Rate Constant, k

- Depends on reaction, temperature, and catalyst
- Also found experimentally

Determining Rate Laws: Isolation Method

Conduct a series of experiments

- Change the concentration of just one reactant each time
- Measure the impact on the rate from each reactant
- Use the ratios of the rates to determine m, n, etc.
- Rearrange the Rate Law & plug in values to solve for k.

Example: For the reaction A + B → Products Rate $= k[A]^m[B]^n$

Conduct 3 experiments:

1. Use [A] & [B], measure Rate 1
2. Use $[A]_2$ & [B], measure Rate 2
3. Use [A] & $[B]_2$, measure Rate 3

To determine m, divide $$\frac{\text{Rate 1}}{\text{Rate 2}} = \frac{\cancel{k}[A]^m\cancel{[B]^n}}{\cancel{k}[A]_2^m\cancel{[B]^n}} = \frac{[A]^m}{[A]_2^m}$$

To determine n, divide $$\frac{\text{Rate 1}}{\text{Rate 3}} = \frac{\cancel{k}\cancel{[A]^m}[B]^n}{\cancel{k}\cancel{[A]^m}[B]_2^n} = \frac{[B]^n}{[B]_2^n}$$

To determine k, solve $$k = \frac{\text{Rate}}{[A]^m[B]^n}$$

When reporting Rate Laws, use the numbers for k, m, & n

Determining Rate Laws: Isolation Method

A set of experiments was run to study the rate of the reaction:

$BrO_3^-\ (aq) + 5Br^-\ (aq) + 6H^+\ (aq) \rightarrow 3Br_2\ (aq) + 3H_2O\ (l)$

The following data was obtained:

Expt #	$[BrO_3^-]$	$[Br^-]$	$[H^+]$	Rate (M/s)
1	0.10	0.10	0.10	-1.2×10^{-3}
2	0.20	0.10	0.10	-2.4×10^{-3}
3	0.20	0.30	0.10	-7.4×10^{-3}
4	0.20	0.10	0.15	-5.4×10^{-3}

$\text{Rate} = k[BrO_3^-]^m[Br^-]^n[H^+]^p$

Determining Rate Laws: Order of Reactants

$\text{Rate}=k[\text{BrO}_3^-]^m[\text{Br}^-]^n[\text{H}^+]^p$

Expt #	$[\text{BrO}_3^-]$	$[\text{Br}^-]$	$[\text{H}^+]$	Rate (M/s)
1	0.10	0.10	0.10	-1.2×10^{-3}
2	0.20	0.10	0.10	-2.4×10^{-3}
3	0.20	0.30	0.10	-7.4×10^{-3}
4	0.20	0.10	0.15	-5.4×10^{-3}

Order of reaction in BrO_3^- (m):

Order of reaction in Br^- (n):

Order of reaction in H^+ (p):

Determining Rate Laws: k & Rate Law

Expt #	$[BrO_3^-]$	$[Br^-]$	$[H^+]$	Rate (M/s)
1	0.10	0.10	0.10	-1.2×10^{-3}
2	0.20	0.10	0.10	-2.4×10^{-3}
3	0.20	0.30	0.10	-7.4×10^{-3}
4	0.20	0.10	0.15	-5.4×10^{-3}

$\text{Rate}=k[BrO_3^-]^m[Br^-]^n[H^+]^p$

$m = 1$

$n = 1$

$p = 2$

k: Solve using data from each experiment (4X!) & <u>average</u>

$$k = \frac{\text{Rate}}{[BrO_3^-]^1[Br^-]^1[H^+]^2}$$

Rate Law =

Rate Laws can be used to determine the initial speed of a reaction with given concentrations of reactants.

For the reaction:

$BrO_3^- (aq) + 5Br^- (aq) + 6H^+ (aq) \rightarrow 3Br_2 (aq) + 3H_2O (l)$

we determined that the Rate Law was:

$\text{Rate} = -12M^{-3}s^{-1}[BrO_3^-]\,[Br^-]\,[H^+]^2$

If we start with the following concentrations:

$[BrO_3^-] = 0.4M$

$[Br^-] = 0.8M$

$[H^+] = 0.2M$

How fast will the initial reaction proceed?

$$\begin{aligned}\text{Rate} &= -12M^{-3}s^{-1} \text{ x } [0.4M] \text{ x } [0.8M] \text{ x } [0.2M]^2\\ &= 0.15\ M^{-3}s^{-1}\ M^4\\ &= 0.15\ Ms^{-1}\end{aligned}$$

Types of Reactions Based on Rate: Zero Order

Change in concentration over time IS linear

Rate Law for A→ Product: rate = k $[A]^0$ = k

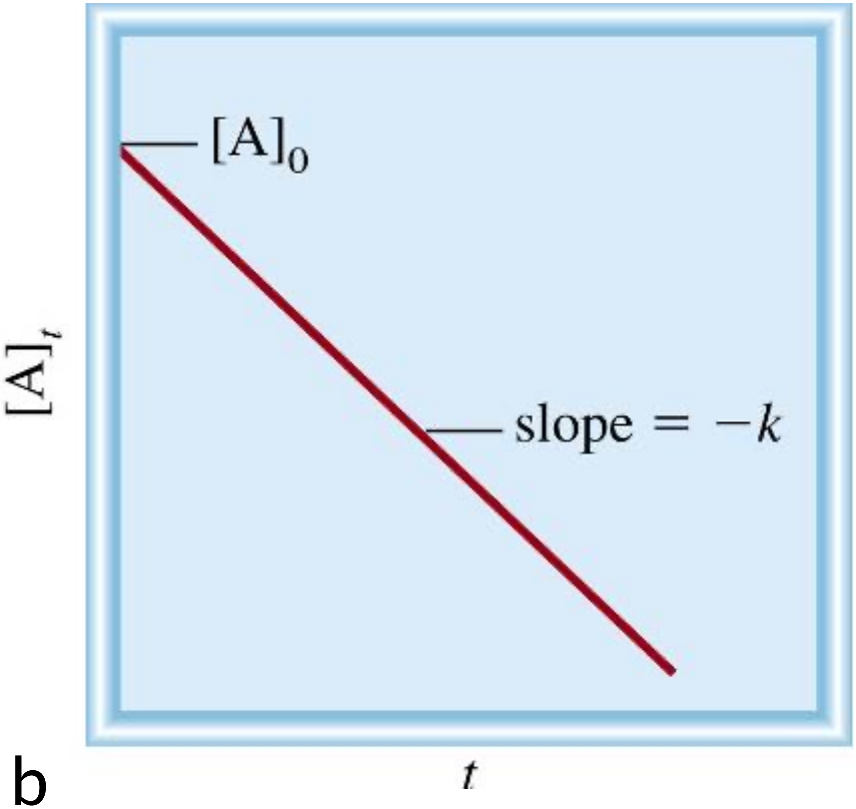
Units of k: k = rate = M/s

Linear Rate Equation:

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

Eventually [A] will go to zero (all reactant will be used up)

Note that this equation is in the form y = mx + b, where b is the y-intercept (i.e. the initial concentration of A).



Types of Reactions Based on Rate: First Order

Change in concentration over time is NOT linear

Change in the natural log (ln) of concentration over time IS linear.

Rate Law for A→ Product: $$\text{rate} = -\frac{\Delta[A]}{\Delta t} = k[A]$$

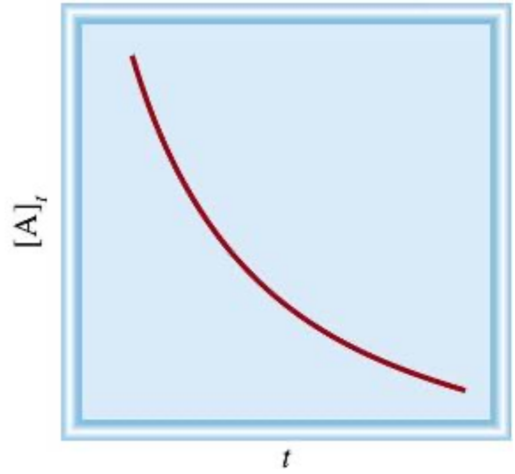
Units of k: k = rate/[A] = (M/s)/M = 1/s = s^{-1}

[A] vs. time is Nonlinear

ln [A] vs time is Linear:

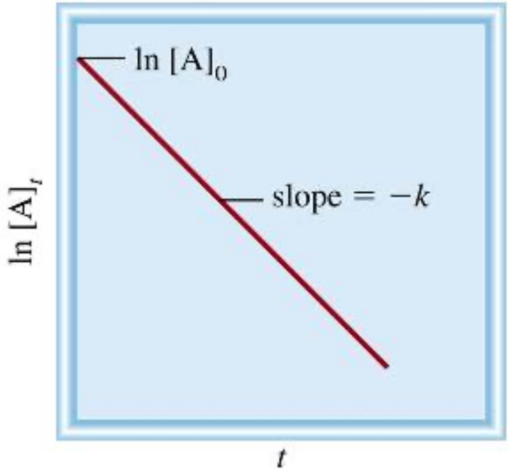
$$[A] = [A]_0 e^{(-kt)}$$

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



$$\ln\left(\frac{[A]}{[A]_0}\right) = -kt$$

$[A]_t$ = [A] at time *t*
$[A]_0$= [A] at *t*=0



Types of Reactions Based on Rate: Second Order

Change in concentration over time is NOT linear

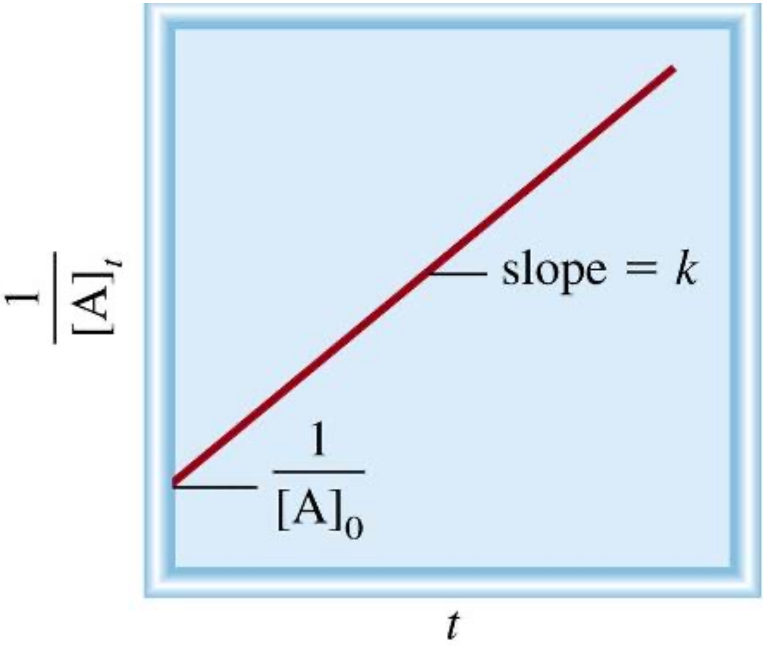
Change in the inverse of concentration (1/[A]) over time IS linear.

Rate Law for A→ Product: $\text{rate} = -\dfrac{\Delta[A]}{\Delta t} = k[A]^2$

Units of k: $k = \text{rate}/[A]^2 = (M/s)/M^2 = 1/Ms = M^{-1}s^{-1}$

The linear equation is:

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

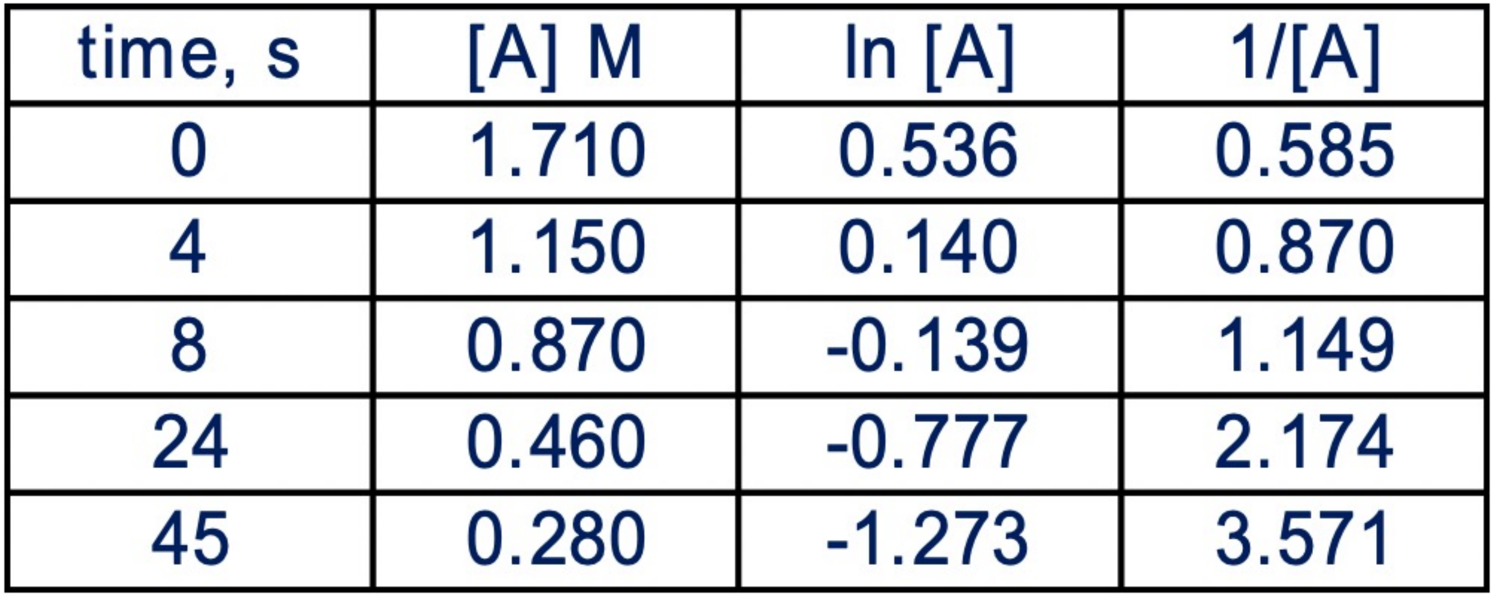


Predicting Reaction Order Graphically

For the reaction A → Products:

1. Calculate ln[A] and 1/[A]
2. Make 3 graphs: [A], ln[A], and 1/[A] vs. time
3. The graph with the best straight line is the reaction order

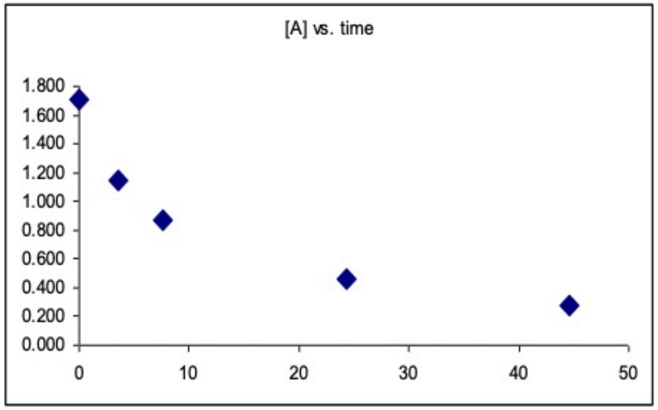
Zero First Second



time, s	[A] M	ln [A]	1/[A]
0	1.710	0.536	0.585
4	1.150	0.140	0.870
8	0.870	-0.139	1.149
24	0.460	-0.777	2.174
45	0.280	-1.273	3.571

Predicting Reaction Order Graphically

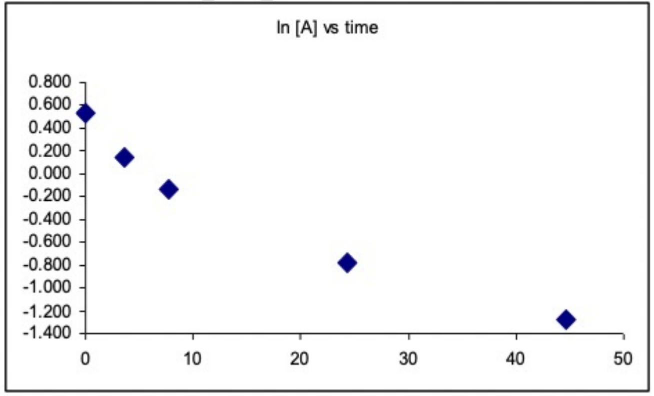
[A] vs. time



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

Not linear, so not Zero Order

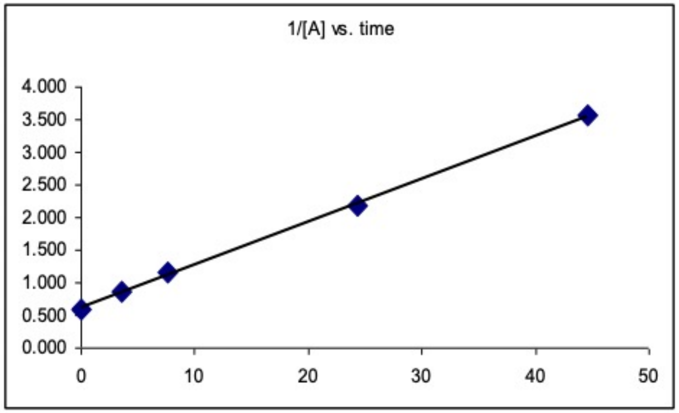
ln[A] vs. time



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

Not linear, so not First Order

1/[A] vs. time



$1/[A] = kt + 1/[A]_0$

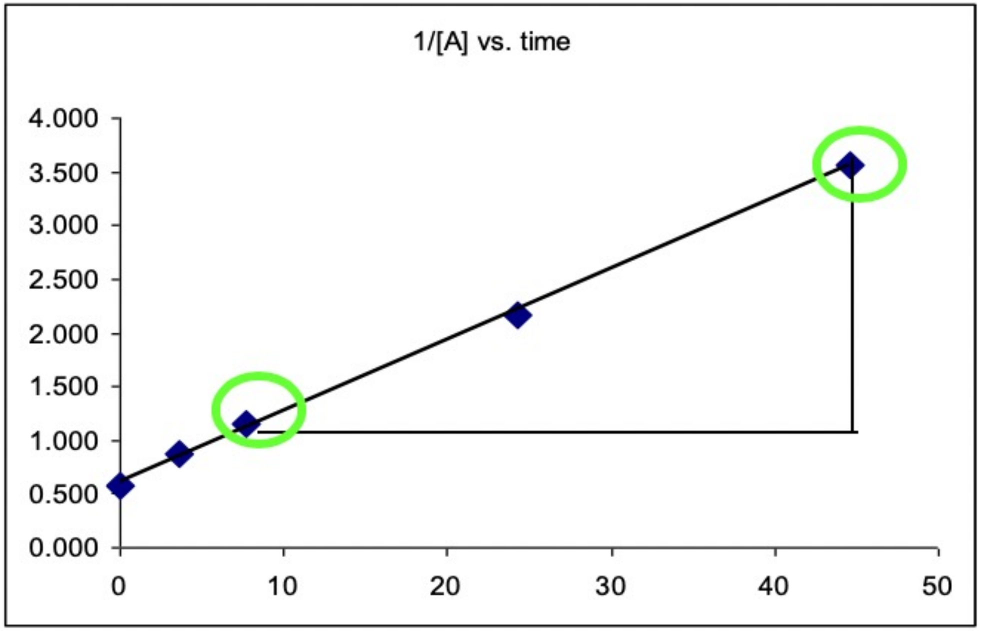
Linear, so reaction is <u>Second Order</u>

Using a Graph to Find Rate Constant, k

Equation of the line is: $y = mx + b$

$1/[A] = kt + 1/[A]_0$

$k = m =$ slope!



$k = m = (y - y_1)/(x - x_1)$

$= (3.57 - 1.15)/(45 - 8)$

$= 0.066\ M^{-1}s^{-1}$

$$\frac{1}{[A]} = 0.066t + \frac{1}{[A]_0}$$

Using Line Equations to Find Information

1. For the first order reaction: $2\ N_2O_5 \rightarrow 2\ N_2O_4 + O_2$ at 45°C, $k = 6.22 \times 10^{-4}\ s^{-1}$. If $[N_2O_5] = 0.100M$:

a.) How long does it take for the concentration to drop to 0.010M?

A: $3.70 \times 10^3 s$

b.) What is the concentration after one hour?

A: 0.0107M

2. The decomposition of sulfuryl chloride (SO_2Cl_2) is first order in SO_2Cl_2. The rate constant for the decomposition at 660K is 4.5×10^{-2} s^{-1}.

a.) If we begin with an initial SO_2Cl_2 pressure of 450 torr, what is the pressure of this substance after 60.s? A: 30. torr

b.) At what time will the pressure of SO_2Cl_2 decline to 1/10 its initial value?

A: 51s

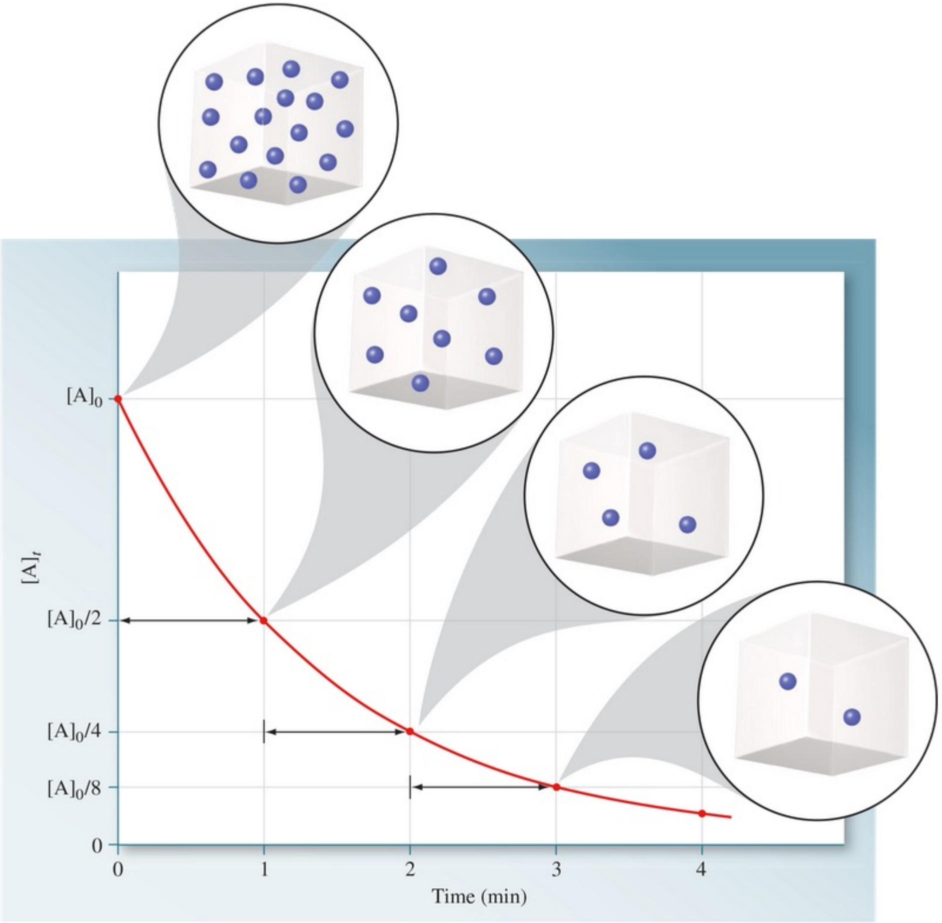
3. Compound A reacts according to second order kinetics with a rate constant of $2.86\times10^{-2}s^{-1}M^{-1}$ at 20°C. If the concentration of A was 1.36M after 15.0s, what was the initial concentration?

A=3.26M

Half-Life ($t_{1/2}$)

The time required for $[A]_0$ to decrease by 50%

- NOT the time for half of the reaction to occur
- Provides another way to get the rate constant



- Symbol is $t_{1/2}$
- $t_{1/2}$ depends on reaction order
- Can solve for $t_{1/2}$ by plugging $[A]_0 = 1A$ and $[A] = ½A$ into rate equation

Half-Life of a First Order Reaction

Linear Equation:

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

Rearranged Equation:

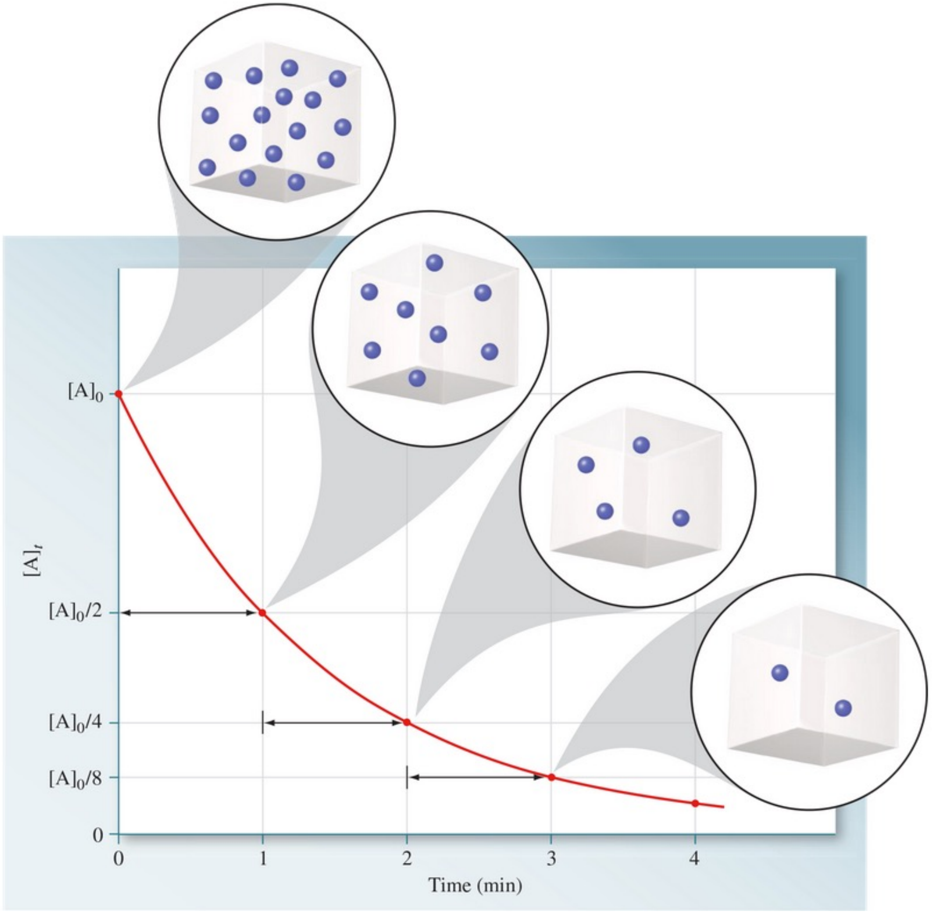
$$\ln \frac{[A]}{[A]_0} = -kt$$

Plug in: $[A]_0 = 1A$ & $[A] = ½A$

$$\ln(½A/A) = \ln(1/2) = -kt_{1/2}$$

Solve for $t_{1/2}$:

$$t_{1/2} = \ln 2/k = \mathbf{0.693/k}$$



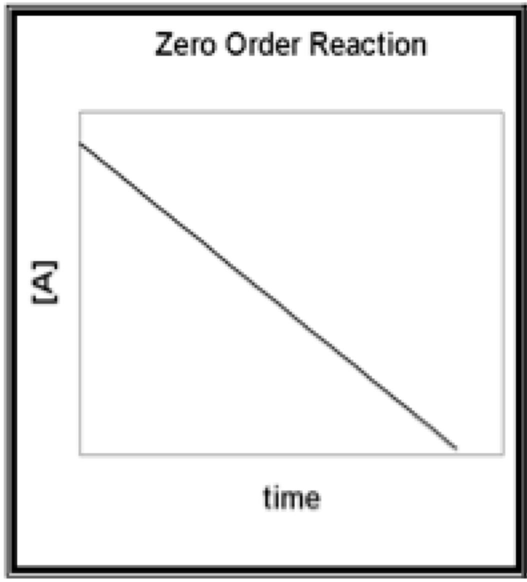
For a <u>First Order</u> Reaction:

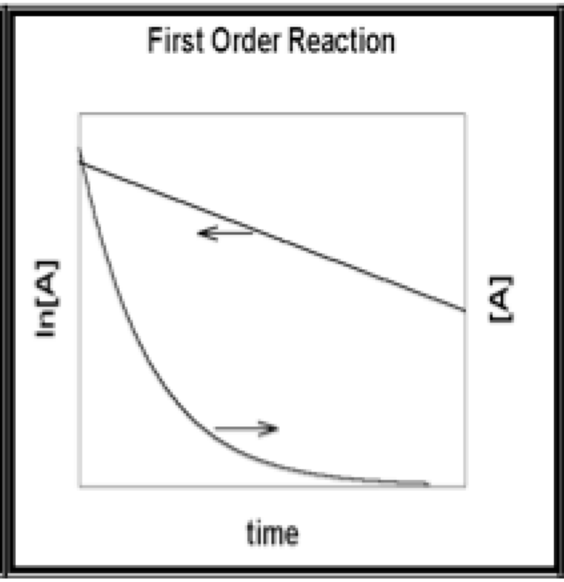
$t_{1/2}$ is independent of initial reactant concentration

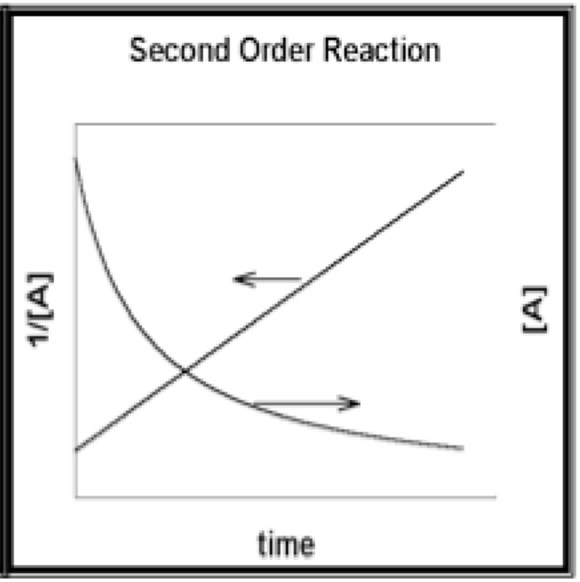
If you know $t_{1/2}$, you know the rate constant

Summary of Kinetic Data

Order	Rate Law	Integrated Rate Law	Linear Plot	k	Half-life
0	Rate = k	$[A]_t = -kt + [A]_0$	[A] vs. t	-slope	$[A]_0/2k$
1	Rate = k[A]	$\ln[A]_t = -kt + \ln[A]_0$	ln[A] vs. t	-slope	0.693/k
2	Rate = $k[A]^2$	$1/[A]_t = kt + 1/[A]_0$	1/[A] vs. t	+slope	$1/k[A]_0$







Half-life Examples

1. For a 1st order reaction, determine the time it will take for only 1/8 of a material to be left if $t_{½} = 3.47 \times 10^2 s$

A: $1.04 \times 10^3 s$

2. The reaction below is first order in $[H_2O_2]$:

$$2H_2O_2\ (l) \rightarrow 2\ H_2O\ (l) + O_2\ (g)$$

A solution originally at 0.600M is found to be 0.075M after 54min. Determine the half-life.

A: 18 min

3. The thermal decomposition of phosphine (PH_3) into phosphorus and molecular hydrogen is a first order reaction:

$$4PH_3\ (g) \rightarrow P_4\ (g) + 6\ H_2\ (g)$$

The half-life of the reaction is 35.0s at 680°C. Calculate (a) the first order rate constant for the reaction and (b) the time required for 95% of the phosphine to decompose.

A: (a) 0.0198s^{-1} (b) 151s

4. The rate constant for the second-order reaction

$$2\ NO_2\ (g) \rightarrow 2\ NO\ (g) + O_2\ (g)$$

is 0.54 $M^{-1}s^{-1}$ at 300°C.

(a) How long, in seconds, would it take for the concentration of NO_2 to decrease from 0.62M to 0.28M?

A: 3.6s

(b) Calculate the half-lives at these two concentrations.

A: 3.0s, 6.6s

Effect of Temperature on Reaction Rate: Arrhenius Equation

Arrhenius Equation – Shows how temp. affects the rate constant (k)

$$k = Ae^{-E_a/RT} \qquad \text{or} \qquad \ln k = (-E_a/RT) + \ln A$$

k = rate constant
A = frequency factor (based on collision frequency & the probability of correct orientation; constant for a given system)
E_a = activation energy
R = gas constant in energy units (8.314 J mol^{-1} K^{-1})
T = temperature in Kelvin

Graphing **ln k vs. 1/T** gives a straight line with
slope = $-E_a/R$

Provides a method to determine E_a for a reaction

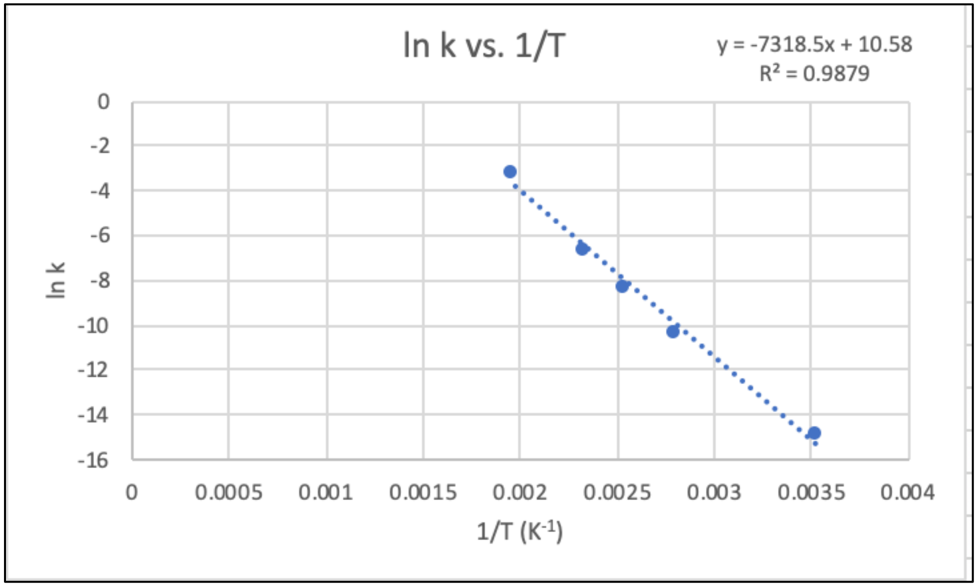
Arrhenius Equation Example

Reaction: $2\ HI(g) \rightarrow H_2(g) + I_2(g)$

Step 1: Find k at several temps.

Temp. (K)	k ($M^{-1}s^{-1}$)
283	3.52×10^{-7}
356	3.02×10^{-5}
393	2.19×10^{-4}
427	1.16×10^{-3}
508	3.95×10^{-2}

Step 2: Graph ln k vs. 1/T



Step 3: Calculate slope ($=-E_a/R$)

Step 4: Multiply slope by $-R$ to get E_a

2 –Point Arrhenius Equation Example

Use rate constants at just 2 temps to solve for E_a

- This method is less accurate but only requires 2 k values
- Best to use **only if there is limited data**
- Equation: $$\ln\frac{k_1}{k_2} = \frac{E_a}{R}\left(\frac{1}{T_2} - \frac{1}{T_1}\right)$$

Note:
$\ln(k_1/k_2) = \ln k_1 - \ln k_2$

Find E_a for the reaction: $2HI(g) \rightarrow H_2(g) + I_2(g)$

Data:
$k_1 = 0.0395\ M^{-1}s^{-1}$
$T_1 = 508\ K$

$k_2 = 0.00116\ M^{-1}s^{-1}$
$T_2 = 427\ K$

$E_a = 78.5$ kJ/mol – note difference from previous slide!

More uses of the Arrhenius Equation

1. The activation energy of a certain reaction is 31.5 kJ/mol. At 30°C, the rate constant is 0.0190 s^{-1}. At what temperature in degrees Celsius would this reaction go twice as fast?

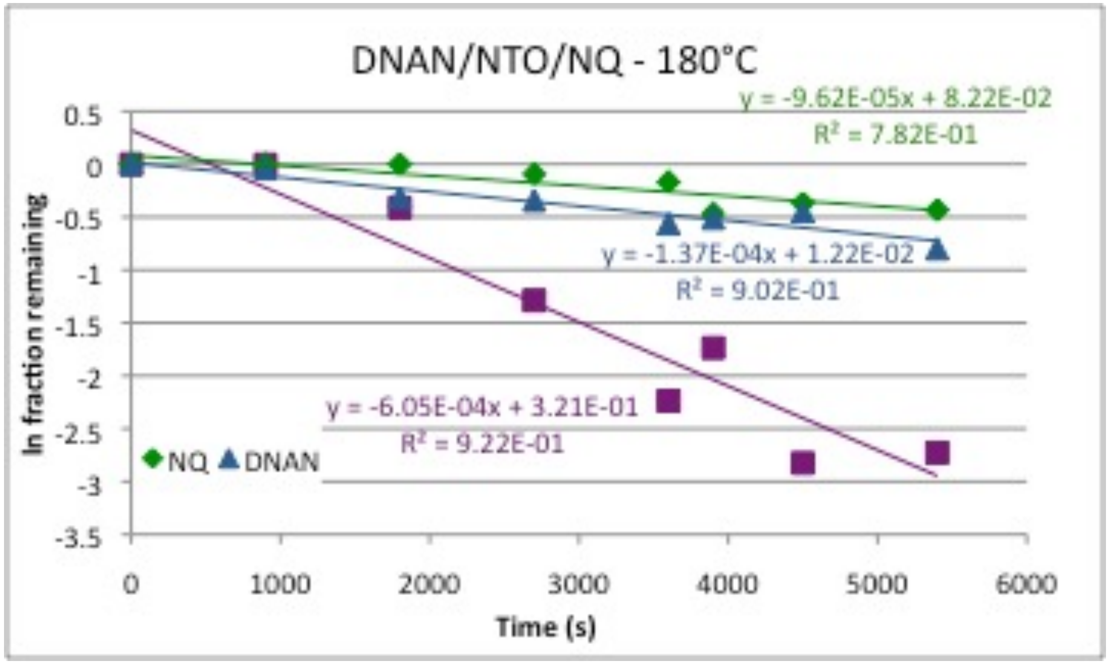
A: 48°C

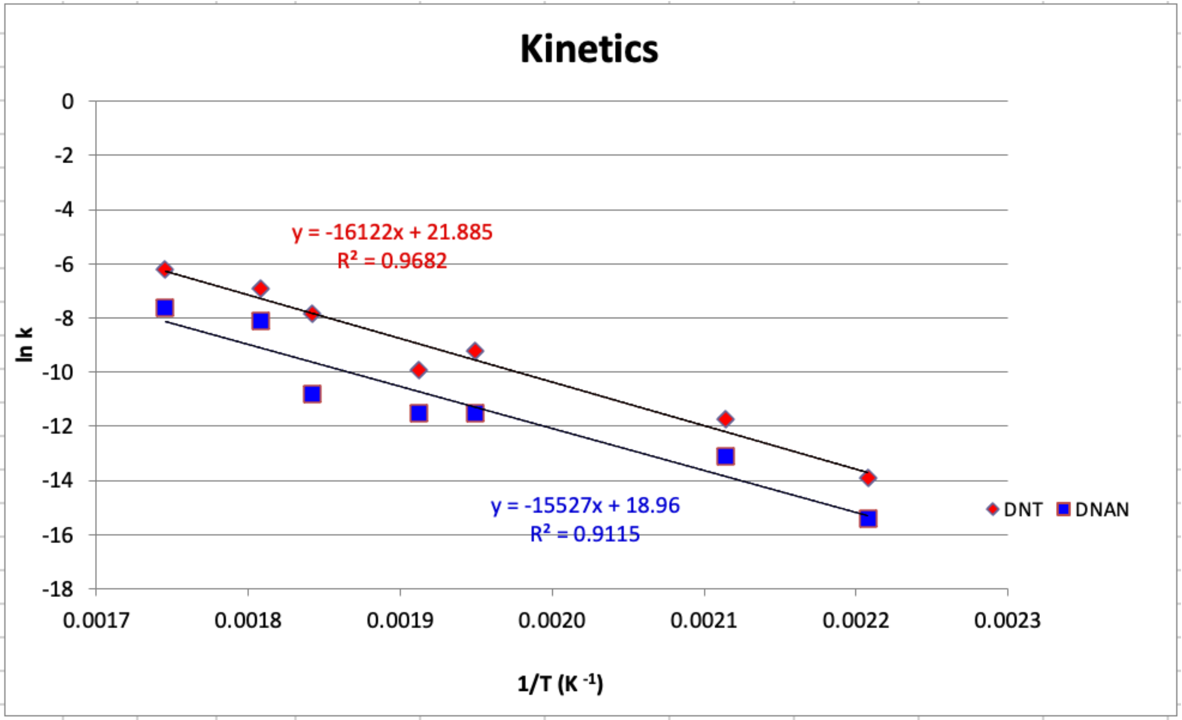
2. The gas-phase reaction Cl(g) + HBr(g) → HCl(g) + Br(g) has an overall enthalpy change of −66kJ. The activation energy for the reaction is 7 kJ.

(a) Sketch the energy profile for the reaction, and label E_a and ΔE.

(b) What is the activation energy for the reverse reaction?

A: 73 kJ





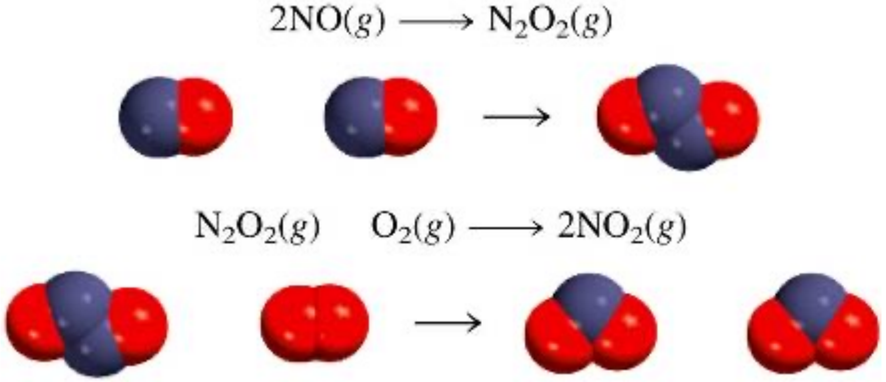
Scientists do use these graphical techniques for real research.

Reaction Mechanism: What (we think) actually happens during a reaction

Series of elementary reactions that lead to an overall rxn

Mechanism info can be used to modify reactions to increase yield, speed, etc.

Ex: 2 $NO(g) + O_2(g) \rightarrow 2NO_2(g)$

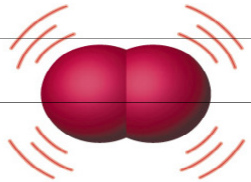


N_2O_2 is detected during the reaction!

$$\begin{array}{rl} \text{Elementary Step:} & NO + NO \rightarrow \cancel{N_2O_2} \\ + \text{ Elementary Step:} & \cancel{N_2O_2} + O_2 \rightarrow 2NO_2 \\ \hline \text{Overall Reaction:} & 2NO + O_2 \rightarrow 2NO_2 \end{array}$$

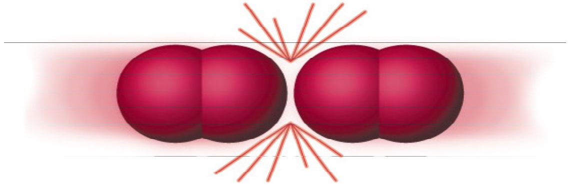
Some Terminology: **Molecularity**

Number of molecules reacting in an elementary step


Unimolecular

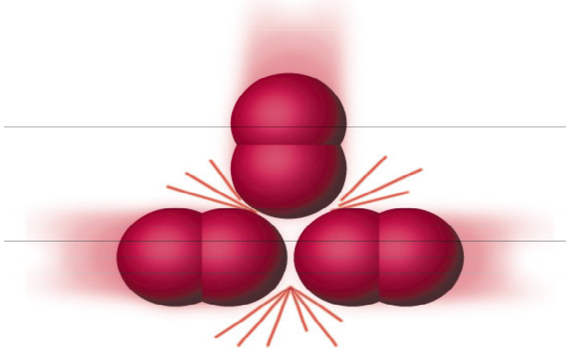
Unimolecular – one

- Fastest
- Molecule simply breaks apart
- No molecular alignment needed


Bimolecular

Bimolecular – two

- Most common
- May require correct alignment


Termolecular

Termolecular – three

- Usually rate limiting (slow!)
- Low probability that all 3 particles will collide simultaneously (esp. with correct alignment and energy)

Some More Terminology

Intermediates: **Species that appear in a reaction mechanism but <u>not in the overall reaction</u>**

- Formed in an early elementary step
- Consumed in a later elementary step

N_2O_2 is an intermediate

Ex:

Elementary Step: $NO + NO \rightarrow N_2O_2$

+ Elementary Step: $N_2O_2 + O_2 \rightarrow 2NO_2$

Overall Reaction: $2NO + O_2 \rightarrow 2NO_2$

Rate Determining Step: **Slowest step in the sequence of steps leading to the overall reaction**

- Often determined by molecularity

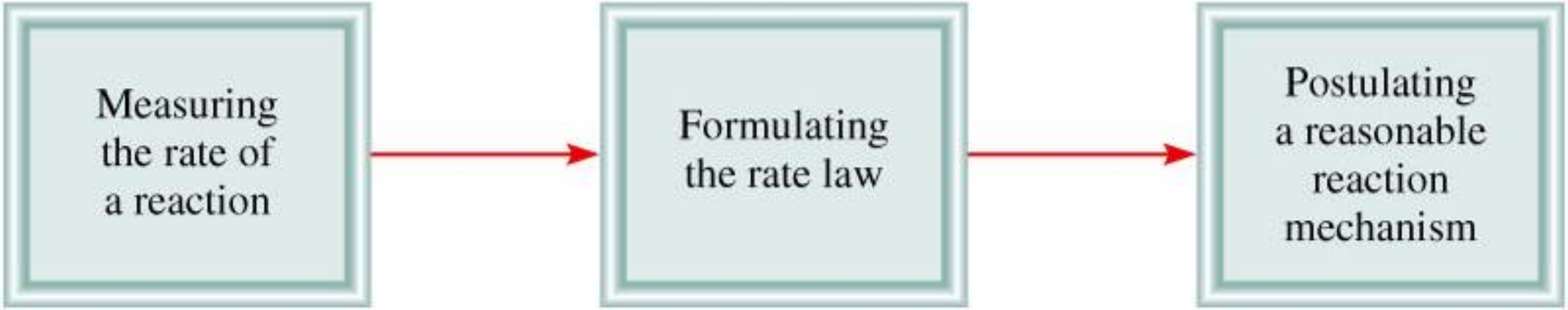
Rate Laws & Intermediate Steps

Unimolecular reaction:	$A \rightarrow \text{Products}$	rate = $k[A]$
Bimolecular reaction:	$A + B \rightarrow \text{Products}$	rate = $k[A][B]$
Bimolecular reaction:	$A + A \rightarrow \text{Products}$	rate = $k[A]^2$

Writing reaction mechanisms:

- The sum of the elementary reactions **<u>must</u>** give the overall balanced equation for the reaction.
- The rate-determining step should predict the same rate law that is determined experimentally
- The rate law for each step is found from the coefficients of the reactants. <u>This is **ONLY** true for **ELEMENTARY** reactions!</u>

Developing a Reaction Mechanism



Measure the rate of the reaction:

- Run a set of isolation method experiments
  - Run 1 experiment for each reactant + a control
  - Measure [A] vs. time

Formulate the rate law

- Determine the order of each reactant based on data
- Use the results to calculate an average value for k
- Determine the overall rate of the reaction

Suggest a reaction mechanism

- Elementary steps must match stoichiometry of overall rxn
- Use molecularity to determine rate determining step
- Order of rate-determining step must match overall rxn order

Ex: Developing a Reaction Mechanism for $NO_2 + CO \rightarrow NO + CO_2$

Based on the results of a set of isolation method experiments, the overall rate was found to be:

$$\text{Rate} = k[NO_2]^2$$

Suggested reaction mechanism:

- Need to account for rate being 2nd order in NO_2
- Need to account for rate being zero order in CO

Step1: $NO_2 + \cancel{NO_2} \rightarrow NO + \cancel{NO_3}$ rate = $k[NO_2]^2$

Step2: $\cancel{NO_3} + CO \rightarrow \cancel{NO_2} + CO_2$ rate = $k[NO_3][CO]$

$NO_2 + CO \rightarrow NO + CO_2$

Step 1 would be the rate-determining step, because the rate matches the overall reaction rate.

You might also be asked to evaluate possible mechanisms.

Evaluating Mechanisms

Ex: $H_2 + 2ICl \rightarrow 2HCl + I_2$ Rate = $k[H_2][ICl]$

Option #1

Step 1: $ICl + ICl \longrightarrow I_2 + \cancel{Cl_2}$ (slow)

Step 2: $\cancel{Cl_2} + H_2 \longrightarrow 2HCl$

Sum: $2\,ICl + H_2 \rightarrow 2HCl + I_2$ ✓

Rate = $k[ICl]^2$ ✗

Option #2

Step 1: $H_2 + ICl \longrightarrow HI + \cancel{HCl}$ (slow)

Step 2: $ICl + \cancel{HCl} \longrightarrow HI + Cl_2$

Sum: $2\,ICl + H_2 \rightarrow 2HI + Cl_2$ ✗

Rate = $k[H_2][ICl]$ ✓

Option #3

Step 1: $H_2 \longrightarrow \cancel{2H}$ (slow)

Step 2: $ICl + \cancel{H} \longrightarrow HCl + \cancel{I}$

Step 3: $\cancel{H} + \cancel{I} \longrightarrow HI$

Sum: $ICl + H_2 \rightarrow HCl + HI$ ✗

Rate = $k[H_2]$ ✗

Option #4

Step 1: $H_2 + ICl \longrightarrow HCl + \cancel{HI}$ (slow)

Step 2: $\cancel{HI} + ICl \longrightarrow HCl + I_2$

Sum: $2\,ICl + H_2 \rightarrow 2HCl + I_2$ ✓

Rate = $k[H_2][ICl]$ ✓

Catalysis

Catalysts alter the rate of a chemical reaction

- Not part of the overall reaction (not a reactant or product)
- Often work by lowering activation energy
  - Since less energy is needed, rate increases
- Materials that slow reaction rates are called **inhibitors**

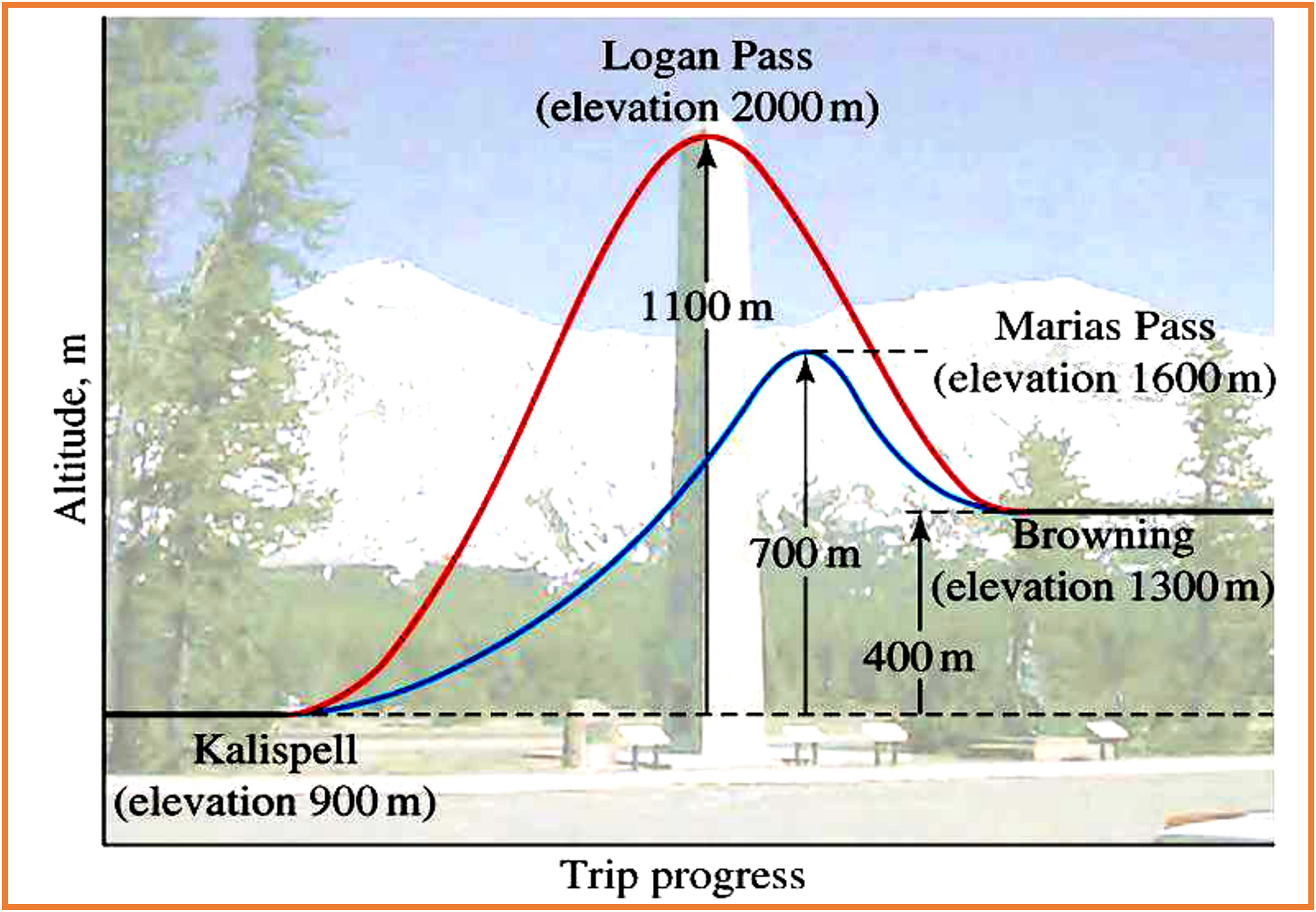
Heterogeneous catalysts

- In a different phase than the reacting medium
- Solids often used to catalyze reactants in solution
- Often used as a surface to align molecules correctly
- Benefit – often easy to remove & re-use

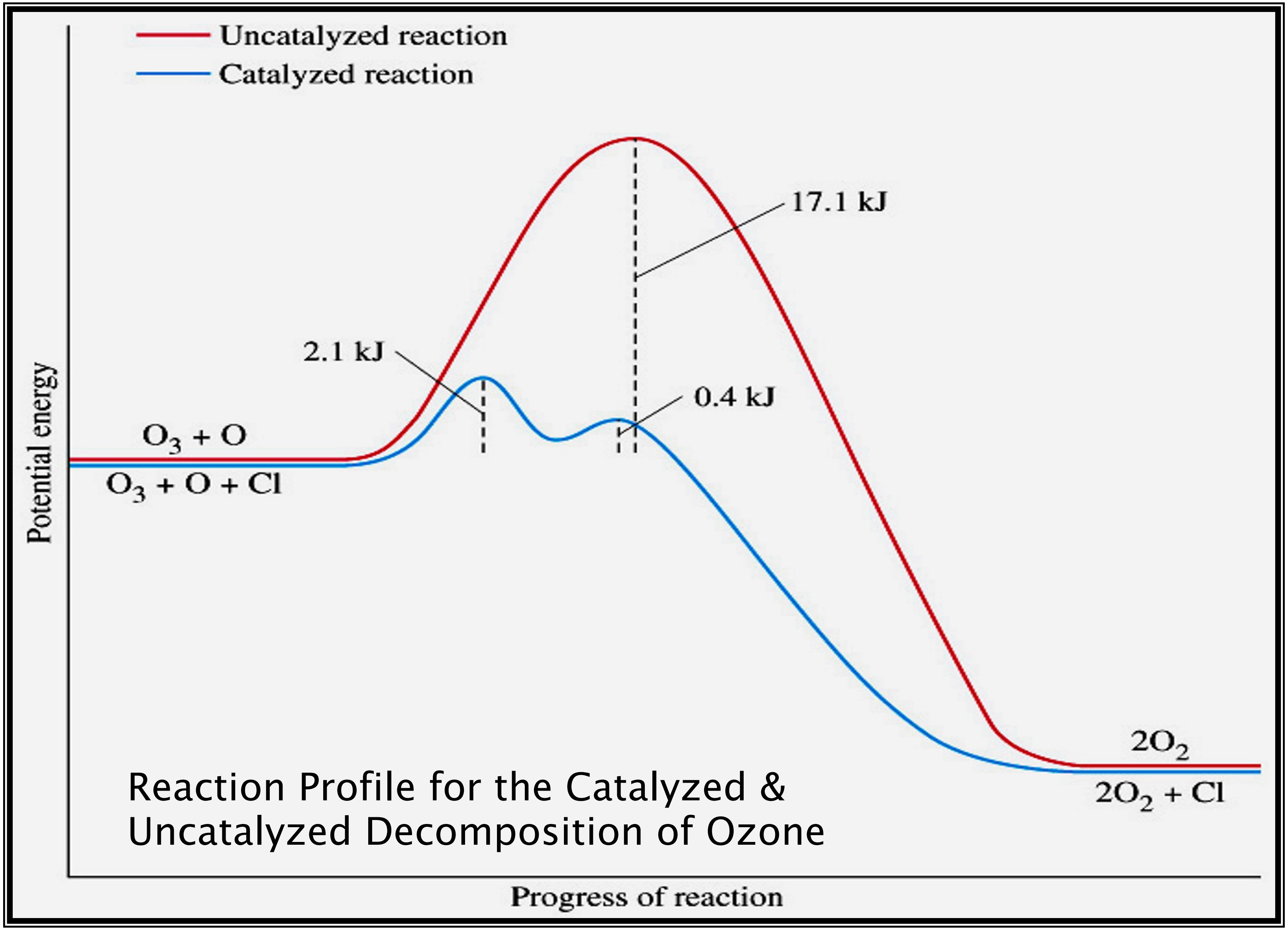
Homogeneous catalysts

- In the same phase as the reacting medium
- May have larger impact on rate (may be able to create a lower energy activated complex than heterogeneous)

Effect of a Catalyst on Reaction Profile & Activation Energy



Effect of a Catalyst on Reaction Profile & Activation Energy

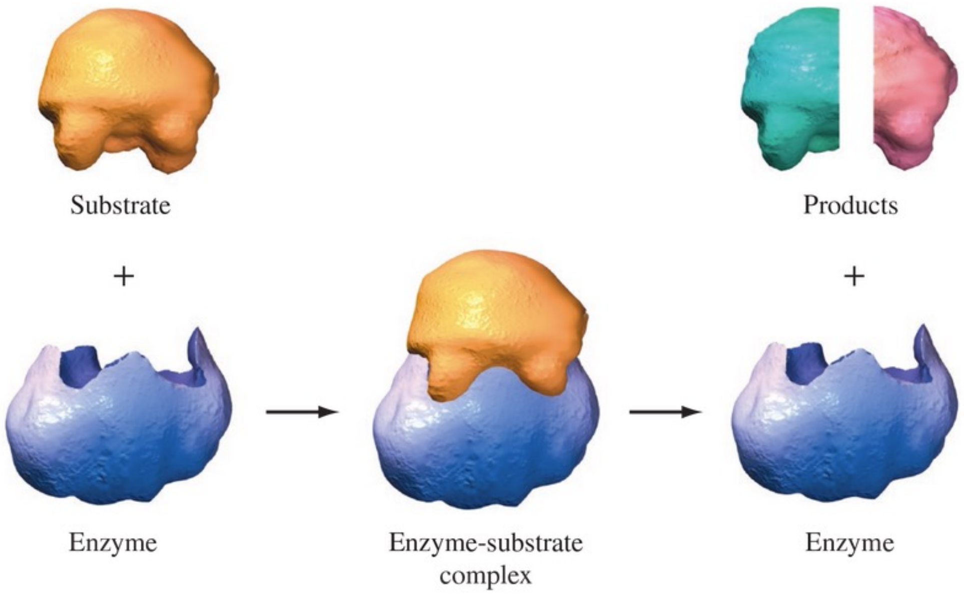


Reaction Profile for the Catalyzed & Uncatalyzed Decomposition of Ozone

Enzymes

Enzymes

- Control the rate of biological reactions
- Very specific
  - Only act on specific molecules called **substrates**
- Often proteins – shape determines functionality
  - Partially due to hydrogen bonding (CHM101!)
- Some enzymes also act as **inhibitors**



Catalysts & Reaction Mechanisms

When considering mechanisms:

- Catalysts control the rate of the reaction.
- The rate of the elementary step that involves the catalyst will control the observed rate.
- Rate constant is different from the rate constant without the catalyst – called the catalytic rate constant.