1

Kinetics

Ch 14

2

Chemical kinetics

- Chemical kinetics – area of chemistry concerned with speeds (or rates) of reactions
- Reaction rate – the speed at which a reaction occurs
- Rate at which amount of reactants is converted to products per unit time

3

Alka-Seltzer activity

- Determine an average rate for the reaction of alka-seltzer with water
- What factors can change this rate? How do they change it? Why do they change it?

4

Factors that affect reaction rates

- Reactant concentration – most reactions react faster if the concentration(s) are increased
- Temperature – most reactions proceed faster if the temperature is increased
- Physical state of reactants (like surface area) – conditions that allow more collisions tend to have faster reactions
- Catalysts – increase reaction rates without being used up

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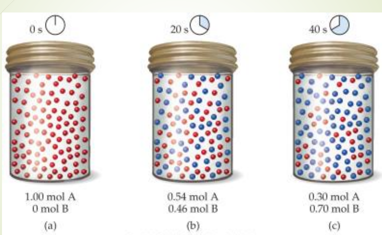
What is happening with time?



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What is happening with time?

- [A] decreases as [B] increases
- Since reaction rate is the change of concentration of a reactant or product with time,

$$Average\ Rate = -\frac{\Delta[A]}{\Delta T}$$

$$Average\ Rate = \frac{\Delta[B]}{\Delta T}$$

- Why is the rate for A negative?

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Average reaction rate

- By convention, rates are always expressed as positive quantities

$$\overline{Rate} = -\frac{\Delta[A]}{\Delta T} = \frac{0.54M - 1.00M}{20s - 0s} = 2.3x10^{-2}M/s$$

$$\overline{Rate} = \frac{\Delta[B]}{\Delta T} = \frac{0.46M - 0.00M}{20s - 0s} = 2.3x10^{-2}M/s$$

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How do you determine the rate?

- Determine the rate by monitoring concentration during a reaction
- $C_4H_9Cl_{(aq)} + H_2O_{(l)} \rightarrow C_4H_9OH_{(aq)} + HCl_{(aq)}$

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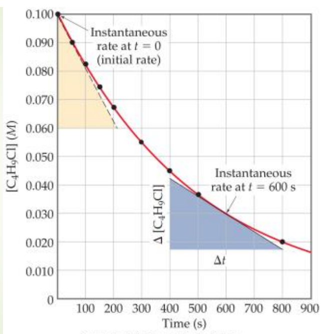

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Average rate of reaction Example

A(g) → B(g)

Time (s)	0.00	10.0	20.0	30.0	40.0
Moles of A	0.124	0.110	0.088	0.073	0.054

The average rate of disappearance of A between 10 s and 20 s is ____.

A. $2.2x10^{-3}$ mol/s
B. $1.1x10^{-3}$ mol/s
C. $4.4x10^{-3}$ mol/s
D. 454 mol/s
E. $9.9x10^{-3}$ mol/s

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Average rate of reaction Example

A(g) → B(g)

Time (s)	0.00	10.0	20.0	30.0	40.0
Moles of A	0.124	0.110	0.088	0.073	0.054

The average rate of disappearance of A between 20 s and 30 s is ____.

A. $-1.5x10^{-3}$ mol/s
B. $-5.0x10^{-4}$ mol/s
C. $1.5x10^{-3}$ mol/s
D. $7.3x10^{-3}$ mol/s
E. $9.3x10^{-3}$ mol/s

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Average vs instantaneous rates

- Average rate measures concentration change over a time period $\lim_{x \to 0} F(x)$
- Instantaneous rate
  - Use tangent line
- Assume "rate" means "instantaneous rate"

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Reaction rates and stoichiometry

- $C_4H_9Cl(aq) + H_2O(l) \rightarrow C_4H_9OH(aq) + HCl(aq)$
- Rate of appearance of C_4H_9OH is equal to rate of disappearance of C_4H_9Cl
- For every 1 mole of C_4H_9Cl consumed, 1 mole of C_4H_9OH is produced
- $Rate = -\frac{\Delta[C_4H_9Cl]}{\Delta t} = \frac{\Delta[C_4H_9OH]}{\Delta t}$

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Reaction rates and stoichiometry

- When relationships are not 1:1?
- $2\ HI(g) \rightarrow H_2(g) + I_2(g)$
- $Rate = -\frac{1}{2}\frac{\Delta[HI]}{\Delta t} = \frac{\Delta[H_2]}{\Delta t} = \frac{\Delta[I_2]}{\Delta t}$

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Reaction rates and stoichiometry

- Generic rates:
- aA + bB → cC + dD
- $Rate = -\frac{1}{a}\frac{\Delta[A]}{\Delta t} = -\frac{1}{b}\frac{\Delta[B]}{\Delta t} = \frac{1}{c}\frac{\Delta[C]}{\Delta t} = \frac{1}{d}\frac{\Delta[D]}{\Delta t}$

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Rate laws and concentration

- How does reaction rate depend on initial concentration?
- Rate law – shows the relationship
- aA + bB → cC + dD
- Rate = $k[A]^m[B]^n$

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Rate laws and concentration

- Rate constant (k)
- k changes with temperature
- Units depend on the rate law for each reaction

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Rate laws and concentration

- Exponents are determined by measuring effect of concentration change on reaction rate
  - Determine experimentally (or use data)
  - NOT THE SAME as stoich coefficients
  - Usually whole numbers (but can be fractions or negative also)

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*Rate law example
One reactant

- Determine rate law and rate constant

Experiment number	Initial A concentration (M)	Observed initial rate (M/s)
1	0.0100	5.4×10^{-7}
2	0.0200	1.08×10^{-6}
3	0.0400	2.15×10^{-6}

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*Rate law example
One reactant

- Determine rate law and rate constant

Experiment number	Initial A concentration (M)	Observed initial rate (M/s)
1	0.0100	5.4×10^{-7}
2	0.0200	2.16×10^{-6}
3	0.0400	8.64×10^{-6}

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*Practice 0.5

- Determine the rate law and rate constant for the reaction using the following data:

Experiment number	Initial B concentration (M)	Observed initial rate (M/s)
1	0.0050	9.3×10^{-3}
2	0.0100	3.7×10^{-2}
3	0.0025	2.3×10^{-3}

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Rate law example
Two reactants

- Determine the rate law and rate constant using initial reaction rate for the following reaction:
- $2\ NO + 2\ H_2 \rightarrow N_2 + 2\ H_2O$

Experiment Number	[NO] (M)	$[H_2]$ (M)	Initial Rate (M/s)
1	0.10	0.10	1.23×10^{-3}
2	0.10	0.20	2.46×10^{-3}
3	0.20	0.10	4.92×10^{-3}

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Practice 1

1. Determine the rate law and rate constant using initial reaction rate for the following reaction:

   $OCl^- + I^- \rightarrow OI^- + Cl^-$

$[OCl^-]$ (M)	$[I^-]$ (M)	Rate (M/s)
1.5×10^{-3}	1.5×10^{-3}	1.36×10^{-4}
3.0×10^{-3}	1.5×10^{-3}	2.72×10^{-4}
1.5×10^{-3}	3.0×10^{-3}	2.72×10^{-4}

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Practice 1

2. Determine the rate law and rate constant using initial reaction rate for the following reaction:

   $2\ NO + 2\ H_2 \rightarrow N_2 + 2\ H_2O$

Experiment Number	[NO] (M)	$[H_2]$ (M)	Initial Rate (M/s)
1	0.10	0.10	1.23×10^{-3}
2	0.10	0.20	2.46×10^{-3}
3	0.20	0.10	4.92×10^{-3}

- Calculate rate when [NO] = 0.050 M and $[H_2]$ = 0.150 M

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*Practice 1.5

- Determine the rate law for the following reaction:
- $BrO_3^-(aq)+5Br^-(aq)+6H^+(aq) \rightarrow 3Br_2(aq)+3H_2O(l)$

	Initial concentrations			Rate in
Mixture	$[BrO_3^-]$ in M	$[Br^-]$ in M	$[H^+]$ in M	M per unit time
A	0.0050	0.025	0.030	10
B	0.010	0.025	0.030	20
C	0.010	0.050	0.030	40
D	0.010	0.050	0.060	160

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Another example with different method More complicated rate law calc

- Determine the rate law and rate constant for the reaction represented below:
- $2A + B \rightarrow C + D$

Experiment	Initial [A] (mol L^{-1})	Initial [B] (mol L^{-1})	Initial rate of formation of C (mol L^{-1} min^{-1})
1	0.125	0.375	2.2×10^{-4}
2	0.375	0.375	6.5×10^{-4}
3	0.750	0.750	2.7×10^{-3}

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Practice 2

Experiment	Initial [A] (mol L^{-1})	Initial [B] (mol L^{-1})	Initial rate of formation of C (mol L^{-1} min^{-1})
1	0.25	0.75	4.3×10^{-4}
2	0.75	0.75	1.3×10^{-3}
3	1.50	1.50	5.3×10^{-3}
4	1.75	?	8.0×10^{-3}

- Rate law, rate constant, ?
- Initial rate of change of [A] in expt 3
- Initial value of [B] in expt 4

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*Challenge problem

- Determine the rate law and value of k
- $2NO(g) + Cl_2(g) \longrightarrow 2NOCl(g)$

Trial	[NO] (mol/L)	$[Cl_2]$ (mol/L)	$-\frac{\Delta[NO]}{\Delta t}$ $(mol\ L^{-1}\ s^{-1})$
1	0.10	0.10	0.00300
2	0.10	0.15	0.00450
3	0.15	0.10	0.00675

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Overall reaction order

- Sum of exponents in rate law
- Ex: $2\ NO + Br_2 \rightarrow 2\ NOBr$

  Rate = $k[NO]^2[Br_2]$

  Order with respect to NO:

  Order with respect to Br_2:

  Overall reaction order:

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Practice 3

Reaction	Rate law	Order with respect to		Overall reaction order
$2\ NO_2 + F_2 \rightarrow 2\ NO_2F$	Rate = $k[NO_2][F_2]$	NO_2	F_2	
$2\ NO + O_2 \rightarrow 2\ NO_2$	Rate = $k[O_2][NO]^2$	NO	O_2	
$NO + N_2O_5 \rightarrow 3\ NO_2$	Rate = $k[N_2O_5]$	NO	N_2O_5	

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Rate laws tell us…

- Rate or speed of a reaction
- How the reaction rate changes when reactant concentrations are changed
- How rate changes due to temperature changes (because k is dependent on temperature)
- The reactant concentration at any time during the reaction

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Usefulness of rate laws…

- Rate law can be converted into equations to relate concentrations to time
- Involves calculus magic (you need to know the results, not the process)

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Reaction orders

- First order
  - Rate depends on concentration of single reactant raised to 1^{st} power
  - Rate = k [A]
- Second order
  - Rate depends on concentration of single reactant raised to 2^{nd} power
  - Rate = k $[A]^2$
  - Or two reactants each raised to 1^{st} power
  - Rate = k [A][B]

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Reaction orders

- Zero order
  - Rate does NOT depend on concentration of A
  - Rate = k

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First order reactions

- A → Products
- $Ave\ rate = -\frac{\Delta[A]}{\Delta T}$
- Rate = k[A]

……………………………fancy math………………………

- $ln\frac{[A]_t}{[A]_o} = -kt \quad OR \quad \ln[A]_t - ln[A]_o = -kt$

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First order reactions

- $ln\frac{[A]_t}{[A]_o} = -kt \quad OR \quad \ln[A]_t - ln[A]_o = -kt$
- Graphing ln[A] vs t will give you a straight line with slope = k

$$\ln[A] = -kt + \ln[A]_o$$
$$y = mx + b$$

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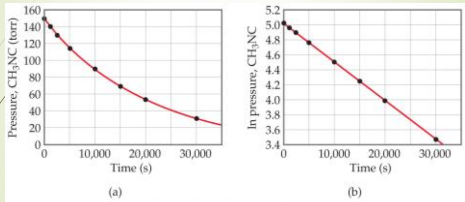
First order reactions
Graphs



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First order reaction
Example

Time (s)	P (mmHg)
0	284
100	220
150	193
200	170
250	150

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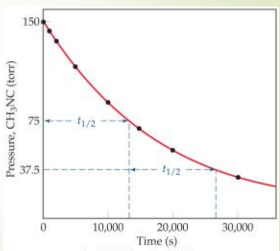
First order reactions
Half-life

- Half-life – time required for reactant concentration to reach half of its initial value ($[A]_o$)
- $\ln[A]_t - \ln[A]_o = -kt$
- Use easy numbers for $[A]_o$ and $[A]_t$
- $t_{1/2} = \frac{0.693}{k}$



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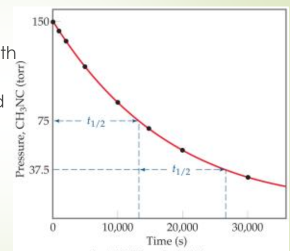
First order reactions
Half-life

- Stays constant with constant T
- Does not depend on initial concentration



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First order reactions
Example

- $CH_3-N\equiv C \rightarrow CH_3-C\equiv N$
- The reaction is a first order reaction. At 230.3˚C, k = 6.29×10^{-4} s^{-1}.
- If [$CH_3-N\equiv C$] is 1.00×10^{-3} M initially, what is [$CH_3-C\equiv N$] after 1.000×10^{3} s?
  - A 5.33×10^{-4} M
  - B 2.34×10^{-4} M
  - C 1.88×10^{-3} M
  - D 4.27×10^{-4} M
  - E 1.00×10^{-6} M

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Summary of first order reactions

- Differential rate law: $Rate = -\frac{\Delta[A]}{\Delta T} = k[A]$
- Integrated rate law: $\ln[A] - \ln[A]_o = -kt$
- Straight line plot: ln[A] vs t
- Slope: slope = k
- Half-life: $t_{1/2} = \frac{0.693}{k}$

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Second order reactions

- A → Products OR A + B → Products
- $Ave\ rate = -\frac{\Delta[A]}{\Delta T} = -\frac{\Delta[B]}{\Delta t}$
- Rate = k[A]^2 OR Rate = k[A][B]

……………………………fancy math……………………

- $\frac{1}{[A]_t} = kt + \frac{1}{[A]_o}$ OR $\frac{1}{[A]_t} - \frac{1}{[A]_o} = kt$

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Second order reactions

- $\frac{1}{[A]_t} = kt + \frac{1}{[A]_o}$ OR $\frac{1}{[A]_t} - \frac{1}{[A]_o} = kt$
- Graphing $\frac{1}{[A]}$ vs t will give you a straight line with slope = k

$$\frac{1}{[A]} = kt + \frac{1}{[A]_o}$$
$$y = mx + b$$

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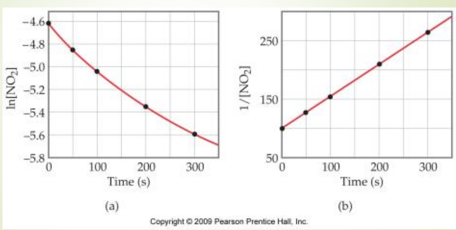
Second order reactions Graphs



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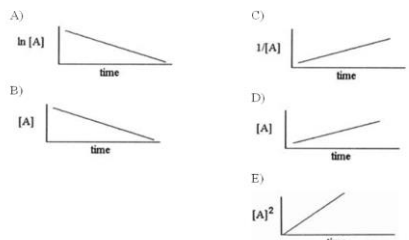
Second order reactions Example 1

Which one of the following graphs shows the correct relationship between concentration and time for a reaction that is second order in [A]?



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Second order reactions Example 2

The following reaction is second order in [A] and the rate constant is 0.039 $M^{-1}s^{-1}$

A → B

The concentration of A was 0.30 M at 23 s. The initial concentration of A was _________ M.

A) 2.4
B) 0.27
C) 0.41
D) 3.7
E) 1.2×10^{-2}

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Summary of second order reactions

- Differential rate law: $Rate = -\frac{\Delta[A]}{\Delta T} = k[A]^2$
- Integrated rate law: $\frac{1}{[A]_t} - \frac{1}{[A]_o} = -kt$
- Straight line plot: 1/[A] vs t
- Slope: slope = k
- Half-life: Calculate

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Zero order reactions

- A → Products
- $Ave\ rate = -\frac{\Delta[A]}{\Delta T} = k$
- Rate = k

.................................*fancy math*...................................

- $[A]_t - [A]_o = -kt \quad OR \quad [A]_t = -kt + [A]_o$

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Zero order reactions

- $[A]_t - [A]_o = -kt \quad OR \quad [A]_t = -kt + [A]_o$
- Graphing [A] vs t will give you a straight line with slope = k

$$[A]_t = -kt + [A]_o$$
$$y = mx + b$$

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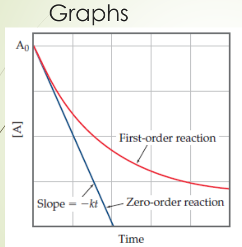
Zero order reactions
Graphs



- When would it be difficult to distinguish between first order and zero order reactions?

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Summary of zero order reactions

- Differential rate law: $Rate = -\frac{\Delta[A]}{\Delta T} = k$
- Integrated rate law: $[A]_t - [A]_o = -kt$
- Straight line plot: [A] vs t
- Slope: slope = k
- Half-life: Calculate

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Practice 4

- $NO_2 \rightarrow NO + \frac{1}{2} O_2$
- Determine if reaction is zero, first, or second order in NO_2
- Write rate law and determine k

Time (s)	$[NO_2]$ (M)
0.0	0.01000
5.0	0.00787
10.0	0.00649
20.0	0.00481
30.0	0.00380

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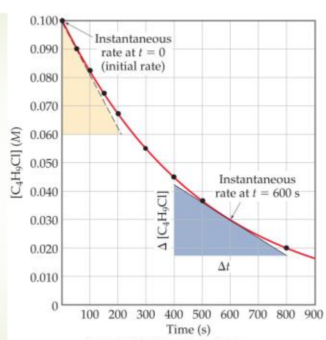
Practice 5

- The graph shows a first order reaction.
- Determine the half-life ($t_{1/2}$)and reaction constant (k).



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Elementary reactions

- Reactions take place because of collisions between reactant molecules
- Elementary reactions take place in single step
- Molecularity – total number of molecules reacting in one step

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Elementary reactions
Molecularity

- Unimolecular involves one molecule; rate is 1st order
- Bimolecular involves two molecules, rate is 2nd order
- Termolecular requires three molecules, rate is 3rd order
  - Almost never heard of because the chances of three molecules coming into contact at same time with correct orientation and energy are VERY small

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Elementary reactions
Rate law

- Elementary reactions have rate laws based on molecularity

TABLE 14.3 • Elementary Reactions and Their Rate Laws

Molecularity	Elementary Reaction	Rate Law
*Uni*molecular	A $\longrightarrow$ products	Rate = $k[A]$
*Bi*molecular	A + A $\longrightarrow$ products	Rate = $k[A]^2$
*Bi*molecular	A + B $\longrightarrow$ products	Rate = $k[A][B]$
*Ter*molecular	A + A + A $\longrightarrow$ products	Rate = $k[A]^3$
*Ter*molecular	A + A + B $\longrightarrow$ products	Rate = $k[A]^2[B]$
*Ter*molecular	A + B + C $\longrightarrow$ products	Rate = $k[A][B][C]$

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Practice 6

1. $H_2 + Br_2 \rightarrow 2\ HBr$
   a. If the reaction occurs in a single elementary reaction, predict its rate law.
   b. If experimental studies show a different rate law, what can you conclude?
2. $2\ NO + Br_2 \rightarrow 2\ NOBr$
   a. Write the rate law for this reaction, assuming it involves a single elementary reaction
   b. Is a single elementary reaction likely for this reaction? Why/why not?

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Collision model

- Molecules must collide to react
- Not all collisions result in reactions (actually, only a small fraction do)
  - Molecules must have proper orientation
  - Molecules must have enough energy to react

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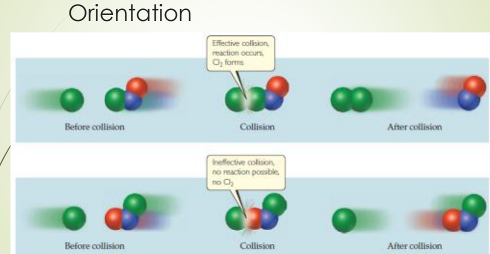
Collision model
Orientation



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Collision model
Activation energy

- Activation energy – minimum energy to initiate chemical reaction
  - Like an energy barrier
  - Colliding molecules must have enough energy to break reactant bonds and create product bonds

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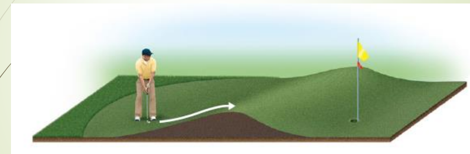
Collision model
Activation energy



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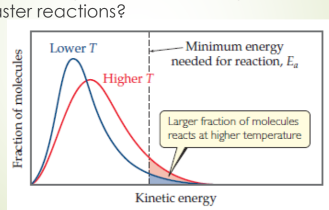
Collision model
Higher temperatures

- Why do higher temperatures generally result in faster reactions?



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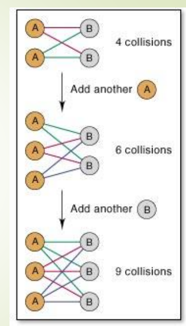
Collision theory
Increasing concentration

- Why does increasing the concentration of a reactant or reactants generally increase the rate?



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Collision theory
Surface area

- Why does changing the surface area affect the rate of a reaction?

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Reaction energy profile

- Show energy of reactants, transition state ("activation complex"), and energy of products
- Use to calculate activation energy (E_a) and reaction enthalpy (ΔH)
- Determine if reaction is endothermic ($+ \Delta H$) or exothermic ($- \Delta H$)

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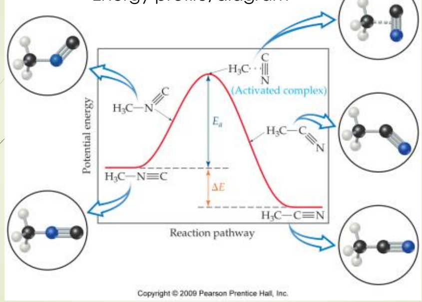
Energy profile/diagram



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Reaction energy profile (energy diagram)

- $H_3C - N \equiv C \rightarrow H_3C - C \equiv N$
- Shows energy of reactants (energy in bonds and phases)
- Additional energy needed to overcome E_a
- Top of hill is the activated complex, or the energy of the transition state
- Products have energy

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Reaction energy profile

- $E_a = E_{top\ of\ hill} - E_{start}$
- $E_a = E_{transition\ state} - E_{reactants}$
- $\Delta E = E_{products} - E_{reactants}$

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Reaction energy profile

- Exothermic – Products have lower energy than reactants
  - $-\Delta E$ or $-\Delta H$
  - Gives off/releases energy
  - Feels hot
- Endothermic – Products have higher energy than reactants
  - $+\Delta E$ or $+\Delta H$
  - Absorbs energy
  - Feels cold

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Arrhenius equation *Not on AP exam*

$$k = Ae^{-E_a/RT}$$

- Relates the factors that affect the reaction rate
  - **k** = rate constant
  - **A** = frequency factor (indicates collision frequency)
  - **e** = base of the natural log
  - **E_a** = Activation Energy **(kJ/mol)**
  - **R** = Gas constant **(8.314 J/K mol)**
  - **T** = Temperature (K)

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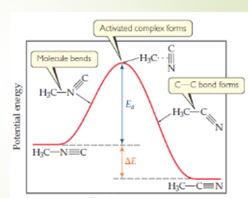
Energy profile

- WHY does increasing T increase reaction rate?
- WHY does decreasing E_a increase rate?
- WHY does increasing [reactants] increase rate?



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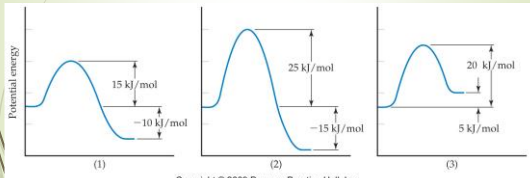
Energy profile example

- Rank the reactions from slowest to fastest



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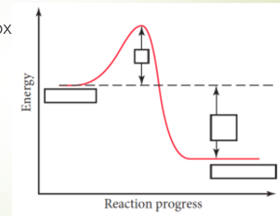
Practice 7

1. Label each box



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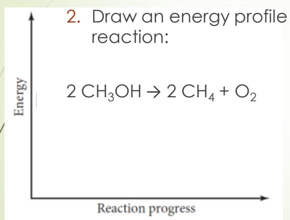
Practice 7

2. Draw an energy profile for the following reaction:

$2\ CH_3OH \rightarrow 2\ CH_4 + O_2$ ΔH = + 253 kJ/mol



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Reaction mechanisms

- Overall chemical equation does not provide information on how a reaction takes place
- Need reaction mechanisms
  - Mechanisms show elementary steps

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Reaction mechanisms
Review vocab

- Molecularity – number of molecules reacting in step
  - Uni-, bi-, ter-

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Reaction mechanisms
New vocab

- Intermediates – formed in one step and consumed in another
  - Not shown in overall equation
  - Can be detected experimentally during rxn
- Catalyst – can react in one step and then reproduced in another
- Rate determining step – slowest step that limits the speed of the reaction

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Reaction mechanisms
Slow first step example

- Write the rate law for the following reaction

$2\,A + B_2 \rightarrow 2\,AB$

Step 1: $A + B_2 \rightarrow AB + B$ *(slow)*

Step 2: $A + B \rightarrow AB$ *(fast)*

- Use coefficients of reactants from SLOW step as exponents in rate law:

Rate = k [A] [B_2]

- Check that rate law is only written in terms of reactants of OVERALL reaction

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Practice 8
Slow first step

1. Write the rate law for:

$NO_2 + CO \rightarrow NO + CO_2$ **(overall)**

Step 1: $NO_2 + NO_2 \rightarrow NO_3 + NO$ (slow)

Step 2: $NO_3 + CO \rightarrow NO_2 + CO_2$ (fast)

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Practice 8
Slow first step

2. Write the rate law for:

$F_2 + 2\,NO_2 \rightarrow 2\,NO_2F$ **(overall)**

Step 1: $F_2 + 2\,NO_2 \rightarrow NO_2F + F + NO_2$ *(slow)*

Step 2: $F + NO_2 \rightarrow NO_2F$ *(fast)*

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Practice 8
Slow first step

3. Write the rate law and overall reaction for:

Step 1: $N_2O \rightarrow N_2 + O$ (slow)

Step 2: $N_2O + O \rightarrow N_2 + O_2$ (fast)

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Practice 8
Slow first step

4. Write the rate law and overall reaction for:

Step 1: $N_2O \rightarrow N_2 + O$ (slow)

Step 2: $N_2O + O \rightarrow N_2 + O_2$ (fast)

Then identify the intermediate

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Reaction mechanisms

- How could you see if the proposed mechanisms are supported?

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Reaction mechanisms
Fast first step example – explained

$2\ NO + O_2 \rightarrow 2\ NO_2$

Step 1: $NO + NO \overset{k_1}{\underset{k_{-1}}{\Leftrightarrow}} N_2O_2$ *(fast)*

Step 2: $N_2O_2 + O_2 \overset{k_2}{\rightarrow} 2\ NO_2$ *(slow)*

- Write rate law using slow step: Rate = $k[N_2O_2][O_2]$
- N_2O_2 is an intermediate and can not be included in the rate law for the overall reaction. Use the fast equilibrium step.

Rate = $k[NO]^2[O_2]$

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Reaction mechanisms
Fast first step example – for MC

$A_2 + B_2 \rightarrow 2\ AB$

Step 1: $A_2 \rightarrow 2\ A$ *(fast)*

Step 2: $A + B_2 \rightarrow AB + B$ *(slow)*

Step 3: $B + A_2 \rightarrow AB + B$ *(fast)*

- Write rate law using slow step: Rate = $k[A][B_2]$
- A is an intermediate. Rearrange 1st step.

$A_2 \rightarrow 2\ A$ and $\frac{1}{2} A_2 \rightarrow A$

Rate = $k[A_2]^{1/2}[B_2]$

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Practice 9
Fast first step

1. Write a rate law that does not contain an intermediate. (Use MC way)

$2\ NO + Br_2 \rightarrow 2\ NOBr$ (overall)

Step 1: $NO + Br_2 \leftrightarrow NOBr_2$ (fast)

Step 2: $NOBr_2 + NO \rightarrow 2\ NOBr$ (slow)

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Practice 9
Fast first step

2. Write a rate law that does not contain an intermediate. (Use non-MC way) What is the overall reaction? What are the intermediates?

Step 1: $Cl_2 \leftrightarrow 2\ Cl$ (fast)

Step 2: $Cl + CHCl_3 \rightarrow HCl + CCl_3$ (slow)

Step 3: $Cl + CCl_3 \rightarrow CCl_4$ (fast)

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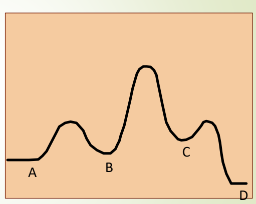
Multistep reaction energy profiles
A → D

1. How many intermediates are formed in the reaction?
2. How many transition states are there?
3. Which step is slowest? Fastest?
4. Is the reaction endo- or exothermic?



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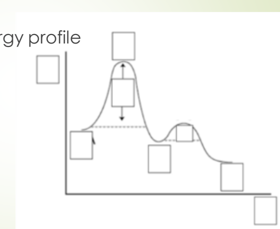
Practice 10
Multistep reaction energy profile

1. Label the energy profile



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Practice 10
Multistep reaction energy profile

2. Draw an energy profile that matches the following steps:

Step 1: $NO_2 + F_2 \rightarrow NO_2F + F$ (slow)
Step 2: $NO_2 + F \rightarrow NO_2F$ (fast)

(the overall reaction is exothermic)

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Practice 10
Multistep reaction energy profile

3. Draw an energy profile that matches the following steps:

Step 1: $A + B_2 \leftrightarrow C + 2\,D$ (fast)
Step 2: $C \rightarrow E$ (slow)
Step 3: $2\,D + E \rightarrow F$ (fast)
(the overall reaction is endothermic)

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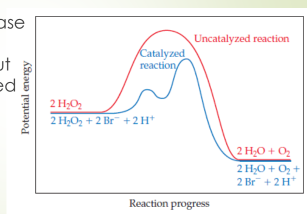
Catalysis

- Catalysts increase the rate of a reaction without being consumed



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Catalysis

- Catalyst may react to form intermediate but is regenerated in subsequent step

$2\,H_2O_2 \rightarrow 2\,H_2O + O_2$ (catalyzed by Br^- in acid)

$H_2O_2 + Br^- + H^+ \rightarrow HOBr + H_2O$
$HOBr + H_2O_2 \rightarrow H_2O + O_2 + Br^- + H^+$

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Catalysis
Identify the intermediates and the catalysts

$2\,H_2O_2 \rightarrow 2\,H_2O + O_2$ (catalyzed by Br^- in acid)

$H_2O_2 + Br^- + H^+ \rightarrow HOBr + H_2O$
$HOBr + H_2O_2 \rightarrow H_2O + O_2 + Br^- + H^+$

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Catalysis

- Catalysts allow reactions to proceed by different mechanism (new pathway)
- New pathway has lower E_a
- More molecules have enough energy (to meet Ea) so more are able to react
- Increases # of effective collisions
- Catalysts DO NOT change ΔE or ΔH

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Catalysis
3 types of catalysts

- Heterogeneous – in different phase as reactants
- Homogeneous – same phase as reactants
- Biological - enzymes

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Catalysis
Homogeneous catalysts

- Reactants and catalysts are in a single phase
  - Liguid
  - Gas
- Elephant toothpaste

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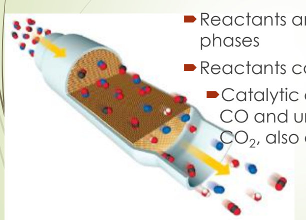
Catalysis
Heterogeneous catalysts



- Reactants and catalysts are in different phases
- Reactants can bind/bond to surface
  - Catalytic converters in cars convert CO and unburned hydrocarbons to CO_2, also convert NO and NO_2 to N_2

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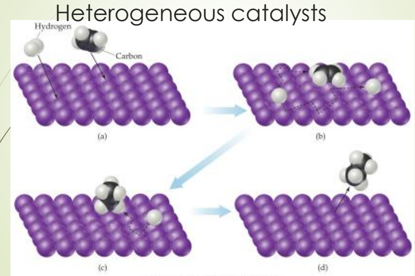
Catalysis
Heterogeneous catalysts



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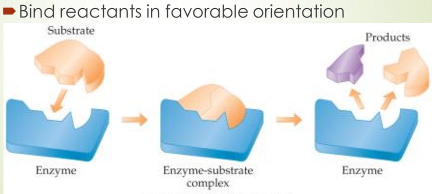
Catalysis
Biological catalysts

- Enzymes are proteins that catalyze reactions
- Bind reactants in favorable orientation



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Catalysis
Energy diagram