

## Introduction to Reaction Rates

Introduction & Methods of Initial Rates

## Collision Theory

- In order for reactions to occur, 3 things must happen:
  1. Atoms or molecules must \_\_\_\_\_
  2. They must collide with the proper \_\_\_\_\_
  3. They must collide with sufficient \_\_\_\_\_

## Chemical Kinetics

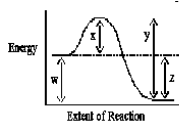
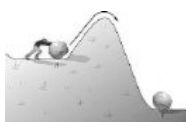
- \_\_\_\_\_ - area of chemistry that is concerned with reaction rates and reaction mechanisms
- The rate of a reaction depends on the collision \_\_\_\_\_ of the reactants and on the collision \_\_\_\_\_.

## Factors that Affect Reaction Rates

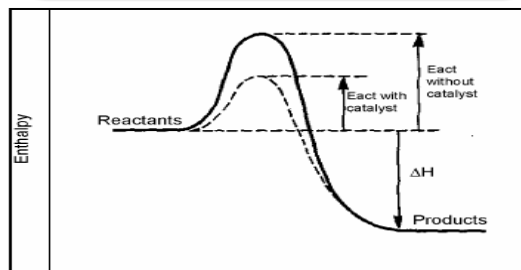
1. Nature of the \_\_\_\_\_
2. \_\_\_\_\_ of the reactants
3. \_\_\_\_\_
4. \_\_\_\_\_
5. \_\_\_\_\_

## Activation Energy

- - the minimum amount of energy that particles must collide with to make a reaction occur



## Activation Energy & Catalyst



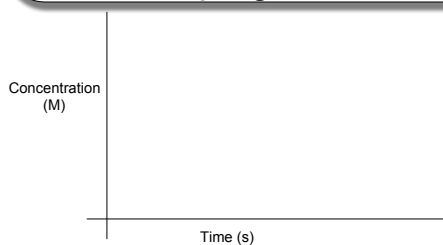
## Intro to Reaction Rates

- – how fast the reaction takes place
- The rate is usually measured in  $\text{Ms}^{-1}$

## Graphing the reaction progress

- For the reaction
- $2\text{NO}_2 \rightarrow 2\text{NO} + \text{O}_2$
- Graph what happens to the concentrations of the reactants & products as the reaction progresses

## Graphing the reaction progress



## Rate Law Types

- Differential Rate Law
  - Usually just called the rate law
  - Comes from calculus (derivative)
  - Finding the slope of the curve
  - Slope rate
- Integrated Rate Law
  - From integrals

## Form of the Rate Law

- Differential Rate Law
- Rate =  $k[A]^x[B]^y$
- Rate =  $M s^{-1}$
- K = rate constant – found only experimentally

## Determining the Order of the Rate Law

- Rate =  $k[A]$  1<sup>st</sup> order
- Rate =  $k[A]^2$  2<sup>nd</sup> order
- Rate =  $k[A][B]$  2<sup>nd</sup> order
- Rate =  $k[A]^2[B]$  3<sup>rd</sup> order
- Rate =  $k[A][B]^2$  3<sup>rd</sup> order
- To get the reaction order, you just add up all of the exponents

## Units of k

- Units of rate are usually Ms<sup>-1</sup>
- Can get units of k from the formula for the rate law
- 1<sup>st</sup> order
  - K = s<sup>-1</sup>
- 2<sup>nd</sup> order
  - K = M<sup>-1</sup>s<sup>-1</sup>
- 3<sup>rd</sup> order
  - K = M<sup>-2</sup>s<sup>-1</sup>

## Integrated Rate Law

- Zero order
  - Rate = k
- 1<sup>st</sup> order
  - $\ln[A] = -kt + \ln[A_0]$
- 2<sup>nd</sup> order
  - $1/A = kt = 1/[A_0]$

## Method of Initial Rates

- Choose 2 experiments in which all but 1 variable changes when divided
- Rate 1 =  $k[A]^x[B]^y$   
Rate 2 =  $k[A]^x[B]^y$
- Cancel out all but 1 of the exponents to solve for x or y
- Solve for the other (x or y)
- Use one experiment to solve for k
- Find units for k

## Example

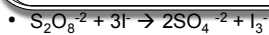
**TABLE 12.4** Initial Rates from Three Experiments for the Reaction  
 $\text{NH}_4^+(\text{aq}) + \text{NO}_2^-(\text{aq}) \rightarrow \text{N}_2(\text{g}) + 2\text{H}_2\text{O}(\text{l})$

Experiment	Initial Concentration of $\text{NH}_4^+$	Initial Concentration of $\text{NO}_2^-$	Initial Rate (mol/L · s)
1	0.100 M	0.0050 M	$1.35 \times 10^{-7}$
2	0.100 M	0.010 M	$2.70 \times 10^{-7}$
3	0.200 M	0.010 M	$5.40 \times 10^{-7}$

What is the rate law?

What is the value for k including units?

## Example 2

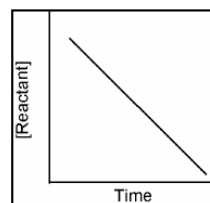


Experiment	$[\text{S}_2\text{O}_8^{2-}]$	$[\text{I}^-]$	Rate
1	0.018	0.036	$2.6 \times 10^{-6}$
2	0.027	0.036	$3.9 \times 10^{-6}$
3	0.036	0.054	$7.4 \times 10^{-6}$
4	0.050	0.072	$1.4 \times 10^{-5}$

- determine the rate law
- Determine the value of the rate constant
- How is the rate of the disappearance of  $\text{S}_2\text{O}_8^{2-}$  related to the disappearance of  $\text{I}^-$ ?
- What is the rate when  $[\text{S}_2\text{O}_8^{2-}] = 0.015\text{M}$  &  $[\text{I}^-] = 0.040\text{M}$ ?

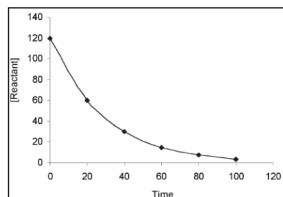
## Graphs and kinetic data

- a straight-line graph shows zero order w.r.t that reactant.



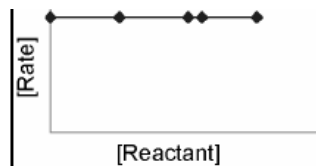
## Graphs and kinetic data

- a constant half-life graph shows first order w.r.t that reactant.



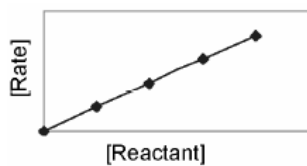
## Plotting initial rates against concentrations of the reactants.

- Zero order – rate is independent of concentration



Plotting initial rates against concentrations of the reactants

- First order reaction – rate is directly proportional to concentration



Plotting initial rates against concentrations of the reactants

- Second order reaction - varies

