

## FRONT OF CARD INFO

## BACK OF CARD INFO

**Factors that Affect rate**

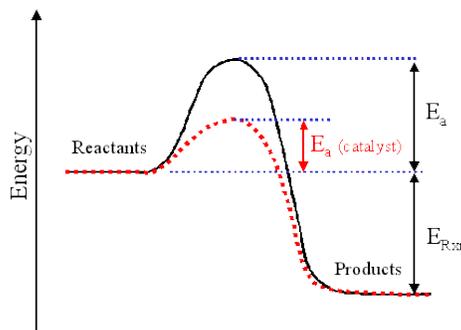
There are five things that affect reaction rates (it is important to not only know what they are but how they affect rates)

1. Nature of reactants (states of matter as well as chemical identity)
2. Concentration of reactant
3. Temperature
4. Surface area of solid reactants
5. Catalysts

Trick alert: remember that adding an inert gas will have NO AFFECT on reaction rates

**Catalysis**

Will speed up the rate of a reaction by either lowering the activation energy or allowing for a different reaction mechanism. In the diagram below the  $E_a$  with a catalysts is lower than the  $E_a$  without. If the reaction require less energy it will happen more often... hence the rate increase ( $E_{rxn}$  here is really  $\Delta H$ ).



A catalyst will lower the  $E_a$  of both forward and reverse reaction equally, therefore equilibrium does not change!

Note: biological catalysts are called enzymes

**Collision Theory of Reaction rates**

In order for reactions to occur the reactants have to collide with each other with enough energy to overcome activation energy ( $E_a$ ) AND the proper 3-D orientation for a reaction to occur.

Most collisions DO NOT result in a reaction

If a collision does result in a reaction we call it an effective collision

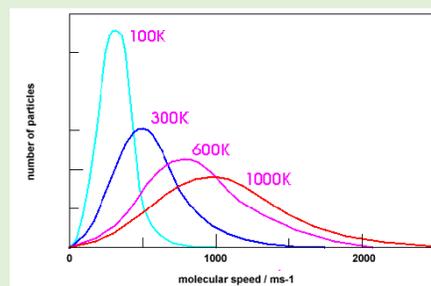
**Activated complex**

Also known as a transition state

Occur at the peak of the energy diagrams  
The easiest way to think about this is to envision a hybrid between reactant molecules and products. This is when reactant bonds are broken and product bonds are formed.

### Maxwell-Boltzmann Distribution

Relates temperature to the average kinetic energy of the molecules  
Since the KE is an average that means not all molecules are moving at the same speed (some faster, some slower).



Notice the curves flatten out when temperature is increased  
The far right of the curve represents the molecules that have enough KE to overcome the  $E_a$  and react

Trick alert: the curves also flatten out as the mass of the molecule decreases since smaller molecules will generally be moving faster

### Reaction Rate

Generally speaking rate is  $\frac{\text{change in concentration}}{\text{change in time}}$

When discussing an individual molecule in a reaction you will use general rates to focus directly on its rate!

There are two methods for solving for this (know both):

1. Instantaneous rate (tangent)
2. Relative rate (average)

### Rate Law

**Analyzes CONCENTRATION vs RATE data**

**Usually presented in table forms**

May also be called the Differential Rate law

The rate law equation is written as  $\text{Rate} = [A]^m[B]^n$   
Where  $m$  and  $n$  are the order with respect to each reactant

$m + n = \text{overall order of reaction}$

### Method of Initial Rates

Reaction rates are measured at the beginning of reactions before there is an accumulation of products

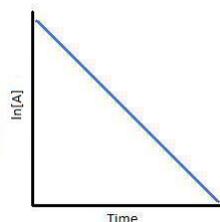
Once products appear it is plausible that the reverse reaction may take place and distort the rate data

<p><b>Zero Order Reactions (Rate Law)</b></p>	<p>The change in concentration of reactant has no effect on the rate.</p> <p>These are not very common.</p> <p>General form of rate equation: Rate = k</p> <p>Trick alert: you do not have to write the reactant in brackets if it is zero order!</p>
<p><b>First Order Reactions (Rate Law)</b></p>	<p>Rate is directly proportional to the reactant concentrations.</p> <p><u>doubling reactant concentration will double rate!</u></p> <p>These are very common!</p> <p>General form of rate equation: Rate = k [A]<sup>1</sup> ... may also be written as Rate = k [A]</p>
<p><b>Second Order Reactions (Rate law)</b></p>	<p>Rate is quadrupled when reactant concentrations are doubled (2<sup>x</sup>) increases by a factor of 9 when reactant concentrations are tripled (3<sup>x</sup>)</p> <p>These are common, particularly in gas-phase reactions.</p> <p>General form of rate equation: Rate = k [A]<sup>2</sup></p>
<p><b>Integrated Rate Law</b></p>	<p><b>Analyzes CONCENTRATION vs TIME data</b></p> <p><b>Can be presented in table form or graphs</b></p> <p>Calculus folks love this because it is simply derivations from graphs!</p> <p>All equations derived will be in the form of y = mx + b</p>
<p><b>Zero Order Reactions (Integrated Rate Law)</b></p>	<p>Relates concentration to time</p> <p>[A] versus time will give straight line</p> <div data-bbox="967 1671 1214 1871" data-label="Figure"> </div> <p><math>[A] = -kt + [A_0]</math></p>

**First Order Reactions  
(Integrated Rate Law)**

Relates Natural log of concentration to time

$\ln[A]$  versus time will give straight line

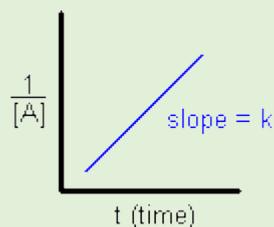


$$\ln[A] = -kt + \ln[A_0]$$

**Second Order Reactions  
(Integrated Rate law)**

Relates the inverse of concentration ( $\frac{1}{[A]}$ ) to time

$\frac{1}{[A]}$  versus time will give a straight line



$$\frac{1}{[A]} = kt + \frac{1}{[A_0]}$$

**Half-Life**

The time it takes for half of the original reactant to disappear (react)

First Order Reactions have CONSTANT Half-lives and are concentration independent

$$t_{\frac{1}{2}} = \frac{0.693}{k}$$

Second order and zero order reactions are concentration dependent and do not have constant half lives

**Reaction Mechanisms**

Series of steps that occur during the conversion of reactants and products (we have called them sub-reactions at times)

Remember that these elementary reactions can be used to determine the rate law for the overall reaction (2 different ways)

1. Slow initial step
2. Fast initial step

<b>Homogeneous Catalyst</b>	exists in the same phase as the reacting molecules
<b>Heterogeneous Catalyst</b>	exists in a different phase than reactants, usually involves gaseous reactants adsorbed on the surface of a solid catalyst