

Chemistry 032 Exam #2 - Chapter 13 - Chemical Kinetics:

Chemical Kinetics - is the study of how fast reactions take place.

Reaction Rates:

- ❖ Variable - instantaneously vs millions of years
- ❖ Being able to manipulate reactions - specifically increase reaction rate - is vital to industrial processes.
- ❖ Most reactions occur faster at higher temperatures

Collision Theory of Chemical Reactions:

- ❖ Reactions generally occur as a result of collisions between reacting molecules
- ❖ Collision theory - the reaction rate is **directly** proportional to the number of molecular collisions per second. (Eq. 1)
- ❖ Effective collisions = collisions resulting in a chemical reaction



- ❖ Ineffective collisions = collisions that do not result in chemical reactions.



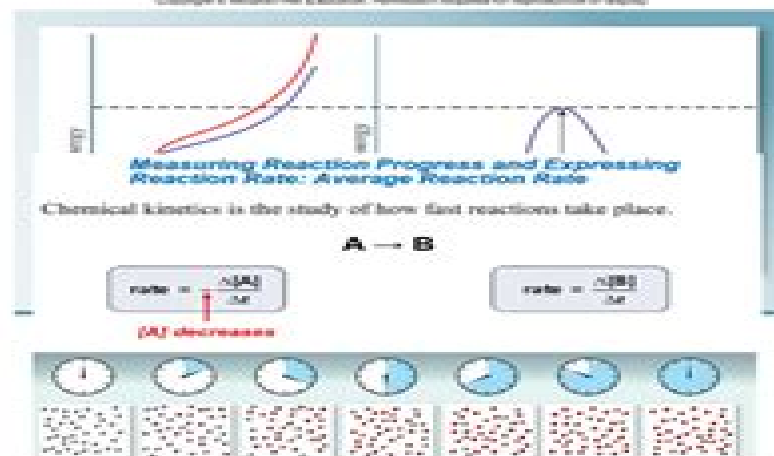
- ❖ Activation Energy (E_a) = minimum energy needed to initiate a chemical reaction

When molecules collide in an effective collision, they form an *activated complex* (also called the *transition state*).



- ❖ Temperature difference:
 - Red line = Higher temperature
 - Blue line = Lower Temperature

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Average Rate -

- ❖ $\text{Br}_2 (\text{aq}) + \text{HCOOH} (\text{aq}) \rightarrow 2 \text{Br}^- (\text{aq}) + 2 \text{H}^+ (\text{aq})$
- ❖ Average Rate = $-\Delta[\text{Br}_2]/\Delta t = -[\text{Br}_2]_{\text{final}} - [\text{Br}_2]_{\text{initial}} / (t_{\text{final}} - t_{\text{initial}})$
- ❖ First 50 seconds: Average Rate = $(.0101 - .0120) \text{ M} / (50.0 - 0.0) \text{ s} = 3.8 \times 10^{-5} \text{ M/s}$
- ❖ First 100 seconds: Average Rate = $(.00846 - .0120) \text{ M} / (100.0 - 0.0) \text{ s} = 3.54 \times 10^{-5} \text{ M/s}$

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