

Topics:

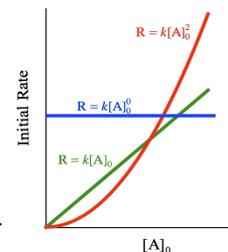
1) **Reaction Rates** (*The change in the concentration of substance per unit time*)

- Determine the average rate of a reaction based on the change in conc. of a substance.
- Use stoichiometry to convert from the average rate of appearance/disappearance of one substance in a reaction to the average rate of appearance/disappearance of another substance in the reaction.
- The instantaneous rate of disappearance of a reactant is faster in the beginning of a reaction.

$$Rate_A = \frac{\Delta[A]}{\Delta t}$$

2) **Rate Laws:** For the reaction  $A + B \rightarrow C$  the general rate law is  $Rate = k[A]^x [B]^y$

- Use experimental data of concentrations and rates to determine the order with respect to any reactant.
  - zero order—the rate does not change with change in concentration of substance.
  - first order—the rate doubles when the concentration of substance doubles
  - second order—the rate quadruples when the concentration of the substance doubles.
- Analyze graphs of Rate vs  $[A]_0$  to determine the order of that substance. See graph →



3) **Integrated Rate Laws:** (Rate Laws consisting of concentration, k and time)

- First order:  $\ln[A]_t - \ln[A]_0 = -kt$       Graph: **ln[A] vs time** is linear       $Half-life = t_{1/2} = \frac{0.693}{k}$
- Second order:  $\frac{1}{[A]_t} - \frac{1}{[A]_0} = kt$       Graph of  $\frac{1}{[A]}$  vs. **time** is linear       $\frac{1}{2}$ -life depends on  $[A]$ , t
- Zero order:  $[A] = [A]_0 - kt$       Graph: **[A] vs time** is linear       $\frac{1}{2}$ -life depends on  $[A]$ , t

4) **Activation Energy, Collision Theory and the Arrhenius Equation**

- Use collision theory to explain why the rate of a reaction increases when one increases concentration of a reactant, increases surface area, increases temperature or uses a catalyst.
- Label energy diagrams: activation energy, activated complex, and  $\Delta H$
- Use Arrhenius equation to determine how temperature and  $E_A$  affects k (rate constant)  
When T increases, k increases (rate also increases); When  $E_A$  is larger, k is lower (rate is also slower)

$$\ln k = -\left(\frac{E_a}{R}\right)\left(\frac{1}{T}\right) + \ln A; \text{ Graph: } \ln k \text{ versus } 1/T \text{ is linear where } m = -\left(\frac{E_a}{R}\right) \rightarrow \ln \frac{k_2}{k_1} = -\frac{E_a}{R} \left(\frac{1}{T_2} - \frac{1}{T_1}\right)$$

5) **Reaction Mechanisms:** (Rates of reactions can be used as evidence to support a reaction mechanism.)

- The rate determining step is the step in the mechanism with the highest activation energy. (slowest)
- The coefficients of the reactants in the rate determining step are the exponents in the rate law.
  - If  $aA + bB \rightarrow cC$  is the rate determining step then the rate law is  $Rate = k [A]^a [B]^b$
  - This is because the rate of a rxn is dependent on the molecules colliding in the rate determining step.
  - If a reactant is not involved in the rate determining step, the rxn is zero order w/ respect to that reactant.

6) **Catalysts:** chemical substances that speed up the rate of a reaction without being consumed.

- Catalysts allow the reaction to proceed by a pathway (mechanism) that is more kinetically favorable, so the overall rate increases.
- Catalysts often lower the activation energy of the reaction and can also increase the frequency of collision.
- Catalysts are not consumed because they react and then they are regenerated.