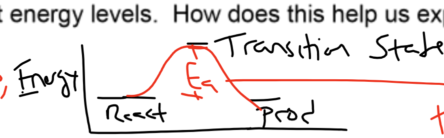
KineticsKey

Monday, October 24, 2016 7:30 AM

1. Draw a reaction coordinate and label it with all important energy levels. How does this help us explain reaction rates?

For reactants to get converted to products, they must overcome the energy barrier (E_a)



relate it to the rate constant (k)

2. What is activation energy and how can it be decreased?

It's the energy "hump" that reactants need to overcome on their path to become products. The only what to change it is adding a catalyst.

3. What are 3 ways to change the rate of a reaction?

Add a catalyst (changes the activation energy and therefore the rate constant (k)), change the temperature (changes k), change the concentration.

4. Consider two reactions with identical rate constants – one of these reactions is 1st order and the other is 2nd order. Explain why the rate of the 2nd order reaction will decrease more quickly than the rate of the 1st order reaction.

The rate in the 2nd order rate constant is dependent on [A]^2. This means that any change in concentration has a more dramatic influence on rate than a first order, which is dependent on [A]^1.

5. Consider the following reactions. Determine the order of the reaction and write the simplest rate law possible (e.g. 3rd order would be rate=$k[A]^3$)

Reaction	k_1	
A	1136 s^{-1}	1st
B	83822 M s^{-1}	0th
C	10.88 $M^{-2}s^{-1}$	3rd

Hint: consider the units of the rate constant

6. For the reaction below, use the method of initial rates to determine the rate constant and rate law. Make sure to use the correct units.

$$2\ NO\ (g) + O_2\ (g) \rightarrow 2\ NO_2$$

Experiment	$[O_2]$ (M)	[NO] (M)	Rate ($M\ s^{-1}$)
1	0.010	0.010	1.0×10^{-6}
2	0.005	0.010	0.5×10^{-6}
3	0.010	0.005	2.5×10^{-7}

Strategy: The rate law will have the form rate = k $[NO]^x [O_2]^y$. Experiments 1 and 2 have equal [NO], so dividing the rate law of 1 by 2 will allow you to solve for y.

$$\frac{1}{2}: \quad \frac{1\times10^{-6}}{0.5\times10^{-6}} = \frac{k}{k}\left(\frac{0.01}{0.01}\right)^x\left(\frac{0.01}{0.005}\right)^y$$

$$2 = 2^y \qquad y = 1$$

$$\frac{1}{3}: \quad \frac{1\times10^{-6}}{2.5\times10^{-7}} = \frac{k}{k}\left(\frac{0.01}{0.005}\right)^x\left(\frac{0.01}{0.01}\right)^y$$

$$4 = (2)^x$$

$$x = 2$$

$$\text{rate} = k[NO]^2[O_2]^1$$

$$1\times10^{-6}\ \frac{M}{s} = k(0.01\ M)^2(0.01\ M)$$

$$k = 1\ M^{-2}s^{-1}$$

$$\text{rate} = 1\ M^{-2}s^{-1}\ [NO]^2[O_2]^1$$

7. For the reaction in problem 6, determine the **rate of the reaction** when the concentration of each reactant is 0.25 mM. *Strategy: this is plug and chug. Plug the 0.25 mM in for [NO] and $[O_2]$. Be careful with units.*

$$\text{rate} = (1\ M^{-2}s^{-1})(2.5\times10^{-4}\ M)^2(2.5\times10^{-4}\ M)$$

$$\text{rate} = 1.5625\times10^{-11}\ \frac{M}{s}$$

8. For the reaction below, use the method of initial rates to determine the rate constant and rate law. Make sure to use the correct units.

$$CO\ (g) + NO_2\ (g) \rightarrow NO\ (g) + CO_2\ (g)$$

Experiment	$[NO_2]$ (mM)	[CO] (mM)	Rate (mM s^{-1})
1	0.15	0.15	0.011
2	0.30	0.15	0.045
3	0.60	0.30	0.18
4	0.60	0.60	0.18

$\frac{1}{2}$ $\qquad \frac{0.011}{0.045} = \frac{k}{k}\left(\frac{0.15}{0.3}\right)^x\left(\frac{0.15}{0.15}\right)^y$

$\frac{1}{4} = \left(\frac{1}{2}\right)^x \qquad \log\frac{1}{4} = x\log\frac{1}{2}$

$x = 2$

$rate = k[NO_2]^2[CO]^{\emptyset}$

$0.011\,\frac{mM}{s} = k(0.15\,mM)^2$

$k = 0.489\ mM^{-1}s^{-1}$

$\frac{3}{4}$ $\qquad \frac{0.18}{0.18} = \frac{k}{k}\left(\frac{0.6}{0.6}\right)^x\left(\frac{0.3}{0.6}\right)^y$

$1 = \left(\frac{0.3}{0.6}\right)^y$

$y = \emptyset$

$rate = 0.489\ mM^{-1}s^{-1}[NO_2]^2$

9. For the reaction in problem 8, determine the **rate of the reaction** when the concentration of each reactant is 0.25 M.

250 mM

$rate = 0.489\ mM^{-1}s^{-1}(250\ mM)^2$

$rate = 30555.6\ \frac{mM}{s}$