1. A chemical reaction occurs when reactant atoms or molecules collide with sufficient energy and orientation for bonds to break and form. The three factors that determines how fast or slow a reaction occurs are temperature, concentration, and a catalyst.
a. (i) As temperature increases, do molecules move faster or slower?
(ii) When molecules collide at higher temperature, do they collide with more energy or less energy?
(iii) When molecules collide at higher temperature, will molecules bounce off each other or will molecules break apart (bonds break)?
b. (i) When concentration increases, will there be more or less reactant molecules in the same volume of solution?
(ii) When concentration increases, will the frequency of collisions increase or decrease? In other words, when
concentration increases, will there be more collisions or less collisions in a certain period of time?
(iii) What is the probability of a molecule breaking apart if collisions occur more frequently?
c. We will look at the effect of an catalyst in Objective 6 Reaction mechanisms.

Answers:
a (i). As temperature increases, molecules move faster. (Temperature is a measure of the average kinetic energy of a substance.)
(ii) When molecules collide at higher temperature, they collide with more energy.
(iii) When molecules collide at higher temperature, molecules will break apart (bonds break).
b. (i) When concentration increases, there be more reactant molecules in the same volume of solution.
(ii) When concentration increases, the frequency of collisions increase. In other words, when concentration increases, there will be more collisions in a certain period of time.
(iii) The probability of a molecule breaking apart if collisions occur more frequently is higher.
c. We will look at the effect of an catalyst in Objective 6 Reaction mechanisms.
2. In Chem 1A, we looked at chemical reactions. For example, hydrogen reacts with iodine to form hydrogen iodide:
$\mathrm{H}_{2}+\mathrm{I}_{2}$--> 2 HI
HI is a stronger acid than HCl . It is a source of $\mathrm{I}_{2}$ when oxidized in air.
The concentration vs. time graph (http://www.sparknotes.com/chemistry/kinetics/ratelaws/section1.rhtml) for this reaction is shown below:

a. Which curve represents what happens to HI with time?
b. The blue curve shows $\Delta[\mathrm{HI}]$ and $\Delta \mathrm{t} . \Delta[\mathrm{HI}] / \Delta \mathrm{t}=$ slope. What does this slope represent?
c. The rate of $\mathrm{I}_{2}=-\Delta[\quad] / \Delta$ $\qquad$ Why is the sign negative?
d. What are the units for the rate of $I_{2}$ ?
e. Write a rate law for this reaction. Fill in the blanks: rate $=-\Delta[] / \Delta]_{工}=\ldots \quad[]^{x}[]^{y}$

Answers:
a. The blue curve represents $\mathrm{HI} . \mathrm{HI}$ is a product and the concentration of the HI product increases with reaction time.
b. Slope $=\Delta[\mathrm{HI}] / \Delta \mathrm{t}=$ rate of HI production with time.

Note the slope is steeper (rate is faster) at the start of the reaction and the slope decreases (rate is slower) as the reaction time increases.
c. The rate of $\mathrm{I}_{2}=-\Delta\left[\mathrm{I}_{2}\right] / \Delta \mathrm{t}$. The sign is negative because $\mathrm{I}_{2}$ is consumed (disappears) as the reaction proceeds.
d. The units for the rate of $I_{2}$ is M/time where $M=$ Molarity.
e. rate law: rate $=-\Delta\left[\mathrm{H}_{2}\right] / \Delta \mathrm{t}=-\Delta\left[\mathrm{I}_{2}\right] / \Delta \mathrm{t}=\mathrm{k}\left[\mathrm{H}_{2}\right]^{\mathrm{x}}\left[\mathrm{I}_{2}\right]^{\mathrm{y}}$ where $\mathrm{k}=$ rate constant, x and $\mathrm{y}=$ order of reaction with respect to each reactant.
3. Consider the reaction: $\quad A+B$--> $C \quad$ rate $=k[A]^{x}[B]^{y}$
a. Rate law does not tell you:
(i) Order
(ii) rate constant
(iii) exothermic
(i) Rate doesn't change
(ii) rate doubles
(iii) rate triples
(i) Concentration
(ii) time
(iii) temperature
b. Oth order in $B$ means if $[B]$ doubles,
c. Rate constant, $k$, changes with
d. k varies with $\mathrm{T}\left(\mathrm{k}=\mathrm{A} \mathrm{e}^{-\mathrm{Ea} / \mathrm{RT}}\right)$ means:
(i) As T increases, k decreases
(ii) $\mathrm{E}_{\mathrm{a}}>0$
(iii) $\mathrm{E}_{\mathrm{a}}$ changes with Temperature
Answers:
a. Rate law does not tell you: (iii) exothermic
b. Oth order in B means if [B] doubles, (i) Rate doesn't change
c. Rate constant, k , changes with (iii) temperature
d. $k$ varies with $T\left(k=A e^{-E a / R T}\right)$ means: (ii) $E_{a}>0$
4. Consider the reaction: $X$---> products
a. This reaction is 0 th order in reactant $X$. The rate law is rate $=k[X]^{0}$. If $[X]$ doubles, what happens to the rate?
b. This reaction is 1 st order in reactant $X$. What is the rate law? If $[X]$ doubles, what happens to the rate?
c. This reaction is 2 nd order in reactant $X$. What is the rate law? If $[\mathrm{X}]$ triples, what happens to the rate?
$d$. The reaction rate stays constant as $[X]$ doubles. What is the rate law? What is the reaction order with respect to $X$ ?
e. In an experiment done at room temperature, $[X]=1$. The rate is equal to $\qquad$ .
f. In the experiment in 4 e , the concentration of X doubles. What happens to $\overline{\mathrm{k}}$ ? What happens to the rate?
g. In the experiment in 4 e , the temperature is lowered to $10^{\circ} \mathrm{C}$. What happens to k ? What happens to the rate?

Answers:
a. This reaction is 0 th order in reactant $X$. The rate law is rate $=k[X]^{0}$. If $[X]$ doubles, the rate does not change (stays the same).

$$
\frac{\text { rate of Experiment } 2}{\text { rate of Experiment } 1}=\frac{k[X]^{0}}{k[X]^{0}}=\frac{k[2]^{0}}{k[1]^{0}}=1
$$

b. This reaction is 1 st order in reactant $X$. The rate law is rate $=k[X]$. If $[X]$ doubles, the rate doubles.

$$
\frac{\text { rate of Experiment } 2}{\text { rate of Experiment } 1}=\frac{k[X]^{1}}{k[X]^{1}}=\frac{k[2]^{1}}{k[1]^{1}}=2
$$

c. This reaction is $2 n d$ order in reactant $X$. The rate law is rate $=k[X]^{2}$. If $[X]$ triples, the rate increases by a factor of 9 .

$$
\frac{\text { rate of Experiment } 2}{\text { rate of Experiment } 1}=\frac{k[X]^{2}}{k[X]^{2}}=\frac{k[3]^{2}}{k[1]^{2}}=9
$$

d. The reaction rate stays constant as $[X]$ doubles. The rate law is rate $=k[X]^{0}$. The reaction order with respect to $X$ is 0 . e. In an experiment done at room temperature, $[X]=1$. The rate $=k[X]^{x}=k[1]^{x}=k$.
f. In the experiment in 4 e , the concentration of $X$ doubles. $k$ stays the same. Can't say what happens to the rate because we do not know the order of the reaction with respect to $X$.
g. In the experiment in 4 e , the temperature is lowered to $10^{\circ} \mathrm{C} . \mathrm{k}$ decreases. Rate decreases.
5. Studies show that CFC's destroy the ozone layer. Another way that ozone in the upper atmosphere is destroyed is by high flying aircraft that produce NO :

$$
\mathrm{O}_{3}(\mathrm{~g})+\mathrm{NO}(\mathrm{~g})--->\mathrm{NO}_{2}(\mathrm{~g})+\mathrm{O}_{2}(\mathrm{~g})
$$

Experiments with ozone and NO gave the following data:

| Experiment | $\left[\mathrm{O}_{3}(\mathrm{~g})\right]$ | $[\mathrm{NO}(\mathrm{g})]$ | Rate, $\mathrm{M} / \mathrm{min}$ |
| :--- | :--- | :--- | :--- |
| 1 | 0.22 | 0.51 | 3.77 |
| 2 | 0.45 | 0.50 | 7.50 |
| 3 | 0.23 | 1.53 | 11.5 |
| 4 | 0.44 | 1.50 |  |
| 5 | 0.90 |  | 23.0 |
| 6 |  | 3.01 | 15.0 |

It is colder in the upper atmosphere where this reaction occurs than on the surface of the Earth.
a. Determine the rate law for this reaction, e.g., rate $=k\left[O_{3}(g)\right]^{x}[N O(g)]^{y}$. Determine $x$ and $y$.

Hint: Compare Experiments 1 and 2. $\left[\mathrm{O}_{3}(\mathrm{~g})\right]$ doubles and $[\mathrm{NO}(\mathrm{g})]$ is constant and rate doubles so $\mathrm{x}=1$.
b. Calculate the rate constant, k.
c. Fill in the blanks for Experiments 4, 5, and 6.

Hint: To determine the rate for Experiment 4, compare Experiments 2 and 4 OR Experiments 3 and 4. Which concentration stays constant and which concentration changes? OR Use the rate law and substitute the concentrations and k and solve for rate.
d. What happens to the rate if this reaction occurred on the surface of the Earth?
e. What happens to the rate constant if this reaction occurred on the surface of the Earth?

Answers:
a. Compare Experiments 1 and 2. $\left[\mathrm{O}_{3}(\mathrm{~g})\right]$ doubles and $[\mathrm{NO}(\mathrm{g})]$ is constant and rate doubles so $\mathrm{x}=1$, i.e., $1^{\text {st }}$ order with respect to $\mathrm{O}_{3}$.
Compare Experiments 1 and 3. $\left[\mathrm{O}_{3}(\mathrm{~g})\right]$ is constant and $[\mathrm{NO}(\mathrm{g})]$ triples and rate triples so $\mathrm{y}=1$, i.e., $1^{\text {st }}$ order with respect to NO.
Rate law: rate $=\mathrm{k}\left[\mathrm{O}_{3}(\mathrm{~g})\right]^{1}[\mathrm{NO}(\mathrm{g})]^{1}$.
b. To calculate k, need to know rate law, rate, and concentrations of each reactant. Use either Experiments 1, 2 , or 3 to calculate k .
Use Experiment 1: 3.77 M/min $=k(0.22 M)(0.51 M)$. Solve for $k=3.77 \mathrm{M} / \mathrm{min} /(0.22 \mathrm{M})(0.51 \mathrm{M})=33.6 \mathrm{M}^{-1} \mathrm{~min}^{-1}$.
Use Experiment 2: $7.50 \mathrm{M} / \mathrm{min}=k(0.45 \mathrm{M})(0.50 \mathrm{M})$. Solve for $k=7.50 \mathrm{M} / \mathrm{min} /(0.45 \mathrm{M})(0.50 \mathrm{M})=33.3 \mathrm{M}^{-1} \mathrm{~min}^{-1}$.
Use Experiment 3: $11.5 \mathrm{M} / \mathrm{min}=\mathrm{k}(0.23 \mathrm{M})(1.53 \mathrm{M})$. Solve for $\mathrm{k}=11.5 \mathrm{M} / \mathrm{min} /(0.23 \mathrm{M})(1.53 \mathrm{M})=32.7 \mathrm{M}^{-1} \mathrm{~min}^{-1}$.
The rate constant, $k$, should be the same for each Experiment. It is within experimental error.
c. Experiment 4: rate $=33 \mathrm{M}^{-1} \mathrm{~min}^{-1}(0.44)(1.50)=\mathrm{M} / \mathrm{min}$

Experiment 5: $23.0=33 \mathrm{M}^{-1} \mathrm{~min}^{-1}(0.90)([\mathrm{NO}(\mathrm{g})])==>$ Solve for $[\mathrm{NO}(\mathrm{g})]=0.77 \mathrm{M}$
Experiment 6: $15.0=33 \mathrm{M}^{-1} \mathrm{~min}^{-1}\left(\left[\mathrm{O}_{3}(\mathrm{~g})\right]\right)(3.01)==>$ Solve for $\left[\mathrm{O}_{3}(\mathrm{~g})\right]=0.15 \mathrm{M}$
d. This reaction occurs in the upper atmosphere. It is warmer on the surface of the earth so the rate increases.
$e$. It is warmer on the surface of the earth so the rate increases and rate constant increases.
6. Bromate ion $\left(\mathrm{BrO}_{3}{ }^{-}\right)$from potassium bromate is a flour improver that strengthens dough and allows for higher rising in the oven. Problem is, bromate is a suspected carcinogen.
$\mathrm{BrO}_{3}^{-}+5 \mathrm{Br}^{-}+6 \mathrm{H}^{+}-->3 \mathrm{Br}_{2}+3 \mathrm{H}_{2} \mathrm{O}$

| Experiment | $\left[\mathrm{BrO}_{3}{ }^{-}\right], \mathrm{M}$ | $[\mathrm{Br}], \mathrm{M}$ | $\left[\mathrm{H}^{+}\right], \mathrm{M}$ | Rate, $\mathrm{M} / \mathrm{sec}$ |
| :--- | :--- | :--- | :--- | :--- |
| 1 | 0.10 | 0.10 | 0.10 | $8.0 \times 10^{-4}$ |
| 2 | 0.20 | 0.10 | 0.10 | $1.6 \times 10^{-3}$ |
| 3 | 0.20 | 0.20 | 0.10 | $3.2 \times 10^{-3}$ |
| 4 | 0.10 | 0.10 | 0.20 | $3.2 \times 10^{-3}$ |
| 5 | 0.50 | 0.20 | 0.20 |  |
| 6 | 1.0 | 0.20 |  | 0.256 |

a. Determine the rate law for this reaction.
b. Calculate the rate constant.
c. Fill in the blanks for Experiments 5 and 6.

Answers:
Rate law: rate $=\mathrm{k}\left[\mathrm{BrO}_{3}^{-}\right]^{\mathrm{x}}\left[\mathrm{Br}^{-}\right]^{\mathrm{y}}\left[\mathrm{H}^{+}\right]^{\mathrm{z}}$.
a. Compare Experiments 1 and 2. $\left[\mathrm{BrO}_{3}{ }^{-}\right]$doubles and $[\mathrm{Br}]$ and $\left[\mathrm{H}^{+}\right]$are constant and rate doubles so $x=1$, i.e., $1^{\text {st }}$ order with respect to $\mathrm{BrO}_{3}{ }^{-}$.
Compare Experiments 2 and $3 .\left[\mathrm{BrO}_{3}{ }^{-}\right]$is constant, $[\mathrm{Br}]$ doubles, and $\left[\mathrm{H}^{+}\right]$is constant and rate doubles so $y=1$, i.e., $1^{\text {st }}$ order with respect to $\mathrm{Br}^{-}$.
Compare Experiments 1 and $4 .\left[\mathrm{BrO}_{3}{ }^{-}\right]$is constant, $[\mathrm{Br}]$ is constant, and $\left[\mathrm{H}^{+}\right]$doubles and rate quadruples so $z=2$, i.e., $2^{\text {nd }}$ order with respect to $\mathrm{H}^{+}$.
Rate law: rate $=\mathrm{k}\left[\mathrm{BrO}_{3}^{-}\right]^{1}\left[\mathrm{Br}^{-}\right]^{1}\left[\mathrm{H}^{+}\right]^{2}$.
b. To calculate $k$, need to know rate law, rate, and concentrations of each reactant. Use either Experiments 1 , 2, 3, or 4 to calculate k .
Use Experiment 1: $8.0 \times 10^{-4} \mathrm{M} / \mathrm{min}=\mathrm{k}(0.10 \mathrm{M})(0.10 \mathrm{M})(0.10)^{2}$.
Solve for $k=8.0 \times 10^{-4} \mathrm{M} / \mathrm{min} /\left((0.10 \mathrm{M})(0.10 \mathrm{M})(0.10 \mathrm{M})^{2}\right)=8 \mathrm{M}^{-3} \mathrm{~min}^{-1}$. (Note units of $k$ depend on reaction order.)
Use Experiment 2: $1.6 \times 10^{-3} \mathrm{M} / \mathrm{min}=\mathrm{k}(0.20 \mathrm{M})(0.10 \mathrm{M})(0.10)^{2}$.
Solve for $k=1.6 \times 10^{-3} \mathrm{M} / \mathrm{min} /\left((0.20 \mathrm{M})(0.10 \mathrm{M})(0.10 \mathrm{M})^{2}\right)=8 \mathrm{M}^{-3} \mathrm{~min}^{-1}$.
The rate constant, $k$, should be the same for each Experiment. It is within experimental error.
c. Experiment 5: rate $=8 \mathrm{M}^{-3} \mathrm{~min}^{-1}(0.50 \mathrm{M})(0.20 \mathrm{M})(0.20)^{2}=0.032 \mathrm{M}\left(10 x\right.$ faster rate than Experiment $4==>\left[\mathrm{BrO}_{3}{ }^{-}\right]$is quintupled, $\left[\mathrm{Br}^{-}\right]$is doubled, and $\left[\mathrm{H}^{+}\right]$is constant. $5 \times 2 \times 1=10$ so concentration increases by 10 for reactants that are $1^{\text {st }}$ order and constant for reactant that is $2^{\text {nd }}$ order so rate increases by factor of 10.)
Experiment 6: $0.256=8 \mathrm{M}^{-3} \mathrm{~min}^{-1}(1.0 \mathrm{M})(0.20 \mathrm{M})\left(\left[\mathrm{H}^{+}\right]\right)^{2}==>$ Solve for $\left[\mathrm{H}^{+}\right]=0.4 \mathrm{M}(8 x$ faster rate than Experiment $5==>$
$\left[\mathrm{BrO}_{3}{ }^{-}\right]$is doubled, $[\mathrm{Br}]$ is constant, and $\left[\mathrm{H}^{+}\right]$is doubled. $2 \times 1 \times 2^{2}=8$ so concentration increases by 8.)
7. In Lab 1, you heated $95 \%$ ethanol with vinegar in the presence of a sulfuric acid catalyst to make ethyl acetate (finger nail polish remover).
For each change in reaction condition, determine whether each quantity increases, decreases, or stays the same.

| Change in Reaction Condition | Reaction Rate | Rate Constant | Activation Energy |
| :--- | :--- | :--- | :--- |
| Raise the temperature |  |  |  |


| Use diluted vinegar solution |  |  |  |
| :--- | :--- | :--- | :--- |
| Use $100 \%$ ethanol |  |  |  |

Answers:
Reaction rate changes with change in temperature or concentration.
Rate constant changes with change in temperature.
Activation energy does not change with temperature or concentration.

| Change in Reaction Condition | Reaction Rate | Rate Constant | Activation Energy |
| :--- | :--- | :--- | :--- |
| Raise the temperature | increases | increases | stays the same |
| Use diluted vinegar solution | decreases | stays the same | stays the same |
| Use 100\% ethanol | increases | stays the same | stays the same |

8. You did the iodine clock reaction in Lab 3. The activation energy for this reaction is $20 \mathrm{~kJ} / \mathrm{mole}$. The rate constant at $20^{\circ} \mathrm{C}$ is $2.5 \mathrm{M}^{-1} \mathrm{sec}^{-1}$.
a. Calculate the rate constant for this reaction at $60^{\circ} \mathrm{C}$. (Should k be larger or smaller at this temperature compared to $20^{\circ} \mathrm{C}$ ? What equation compares k at two different temperatures?)
b. Your lab partner measures $\mathrm{k}=1.7 \mathrm{M}^{-1} \mathrm{sec}^{-1}$ but forgot to record the temperature. Calculate the temperature that corresponds to this rate constant. (Should temperature be greater than $20^{\circ} \mathrm{C}$ or less than $20^{\circ} \mathrm{C}$ ? What equation compares temperature at two different k's?)
Answers:
a. Use Arrhenius equation: $k=A e^{-E a / R T}$

In $\left(k_{1} / k_{2}\right)=-\left(E_{a} / R\right)\left(1 / T_{1}-1 / T_{2}\right)$
$\mathrm{k}_{1}=$ ?
$\mathrm{k}_{2}=2.5 \mathrm{M}^{-1} \mathrm{sec}^{-1}$
$\mathrm{E}_{\mathrm{a}}=20 \mathrm{~kJ} /$ mole $=20,000 \mathrm{~J} / \mathrm{mole}$
$\mathrm{R}=8.31 \mathrm{~J} / \mathrm{mole} \mathrm{K}$
$\mathrm{T}_{1}=60^{\circ} \mathrm{C}=333 \mathrm{~K}$
$\mathrm{T}_{2}=20^{\circ} \mathrm{C}=293 \mathrm{~K}$
In $\left(\mathrm{k}_{1} / 2.5 \mathrm{M}^{-1} \mathrm{sec}^{-1}\right)=-(20,000 \mathrm{~J} /$ mole $/ 8.31 \mathrm{~J} /$ mole K $)(1 / 333 \mathrm{~K}-1 / 293 \mathrm{~K})$
$\ln \left(k_{1} / 2.5 \mathrm{M}^{-1} \mathrm{sec}^{-1}\right)=0.987$
$\left(\mathrm{k}_{1} / 2.5 \mathrm{M}^{-1} \mathrm{sec}^{-1}\right)=2.68$
Solve for $k_{1}=6.7 \mathrm{M}^{-1} \mathrm{sec}^{-1}$. $\mathrm{k}_{1}$ should be larger at $60^{\circ} \mathrm{C}$ compared to $20^{\circ} \mathrm{C}$.
b. a. Use Arrhenius equation: $k=A e^{-E a / R T}$

In $\left(k_{1} / k_{2}\right)=-\left(E_{a} / R\right)\left(1 / T_{1}-1 / T_{2}\right)$
$\mathrm{k}_{1}=1.7 \mathrm{M}^{-1} \mathrm{sec}^{-1}$
$\mathrm{k}_{2}=2.5 \mathrm{M}^{-1} \mathrm{sec}^{-1}$
$\mathrm{E}_{\mathrm{a}}=20 \mathrm{~kJ} / \mathrm{mole}=20,000 \mathrm{~J} / \mathrm{mole}$
$R=8.31 \mathrm{~J} /$ mole K
$\mathrm{T}_{1}=$ ?
$\mathrm{T}_{2}=20^{\circ} \mathrm{C}=293 \mathrm{~K}$
$\ln \left(1.7 \mathrm{M}^{-1} \mathrm{sec}^{-1} / 2.5 \mathrm{M}^{-1} \mathrm{sec}^{-1}\right)=-(20,000 \mathrm{~J} /$ mole $/ 8.31 \mathrm{~J} /$ mole K$)\left(1 / \mathrm{T}_{1}-1 / 293 \mathrm{~K}\right)$
$-0.386=-(20,000 \mathrm{~J} /$ mole $/ 8.31 \mathrm{~J} /$ mole K$)\left(1 / \mathrm{T}_{1}-1 / 293 \mathrm{~K}\right)$
$0.000160=\left(1 / T_{1}-1 / 293 K\right)$
$0.00357=1 / \mathrm{T}_{1}$
Solve for $T_{1}=280 \mathrm{~K} . \mathrm{T}_{1}$ should be smaller than 293 K because k is smaller than k at $20^{\circ} \mathrm{C}$.
9. The rate constant for the reaction $\mathrm{H}_{2}(\mathrm{~g})+\mathrm{I}_{2}(\mathrm{~g})-->2 \mathrm{HI}(\mathrm{g})$ is $5.4 \times 10^{-4} \mathrm{M}^{-1} \mathrm{~s}^{-1}$ at $326{ }^{\circ} \mathrm{C}$. At $410^{\circ} \mathrm{C}$ the rate constant was found to be $2.8 \times 10^{-2} \mathrm{M}^{-1} \mathrm{~s}^{-1}$. Calculate the activation energy for this reaction. (Answer: between 140 and 180 kJ/mole)
Answers:
Use Arrhenius equation: $\mathrm{k}=\mathrm{Ae}-\mathrm{Ea/RT}$
$\ln \left(k_{1} / k_{2}\right)=-\left(E_{a} / R\right)\left(1 / T_{1}-1 / T_{2}\right)$
$\mathrm{k}_{1}=5.4 \times 10^{-4} \mathrm{M}^{-1} \mathrm{sec}^{-1}$
$\mathrm{k}_{2}=2.8 \times 10^{-2} \mathrm{M}^{-1} \mathrm{sec}^{-1}$
$\mathrm{E}_{\mathrm{a}}=$ ? $\mathrm{J} / \mathrm{mole}$
$R=8.31 \mathrm{~J} / \mathrm{mole} \mathrm{K}$
$\mathrm{T}_{1}=326^{\circ} \mathrm{C}=599 \mathrm{~K}$
$\mathrm{T}_{2}=410^{\circ} \mathrm{C}=683 \mathrm{~K}$
In $\left(5.4 \times 10^{-4} \mathrm{M}^{-1} \mathrm{sec}^{-1} / 2.8 \times 10^{-2} \mathrm{M}^{-1} \mathrm{sec}^{-1}\right)=-\left(\mathrm{E}_{\mathrm{a}} / 8.31 \mathrm{~J} / \mathrm{mole} \mathrm{K}\right)(1 / 599-1 / 683 \mathrm{~K})$
$-3.95=-\left(E_{a} / 8.31 \mathrm{~J} /\right.$ mole K $)(1 / 599-1 / 683 \mathrm{~K})$
$19230=\mathrm{E}_{\mathrm{a}} / 8.31 \mathrm{~J} /$ mole K
Solve for $\mathrm{E}_{\mathrm{a}}=159800 \mathrm{~J} / \mathrm{mole}=159.8 \mathrm{~kJ} / \mathrm{mole}$.

