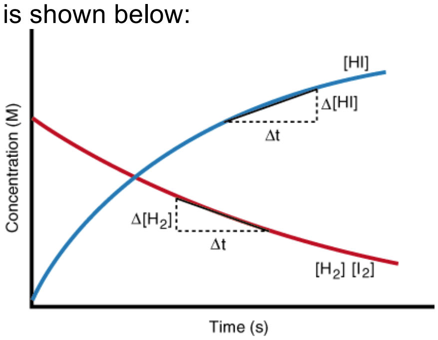
Objective 5. Understand factors that determine reaction rate and describe reaction rate with rate law, order, rate constant, and activation energy.

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.

2. In Chem 1A, we looked at chemical reactions. For example, hydrogen reacts with iodine to form hydrogen iodide:

$H_2 + I_2$ --> 2 HI

HI is a stronger acid than HCl. It is a source of 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 Δ[HI] and Δt. Δ[HI]/Δt = slope. What does this slope represent?
c. The rate of I_2 = - Δ [   ]/Δ___. 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 = - Δ [   ]/Δ____ = ___ [ ]^x [ ]^y

3. Consider the reaction: A + B --> C   rate = k [A]^x [B]^y

a. Rate law does not tell you:	(i) Order	(ii) rate constant	(iii) exothermic
b. 0th order in B means if [B] doubles,	(i) Rate doesn't change	(ii) rate doubles	(iii) rate triples
c. Rate constant, k, changes with	(i) Concentration	(ii) time	(iii) temperature

d. k varies with T (k = A e $^{-Ea/RT}$) means:
(i) As T increases, k decreases   (ii) $E_a > 0$ (iii) E_a changes with Temperature

4. Consider the reaction: X ---> products
a. This reaction is 0th order in reactant X. The rate law is rate = k $[X]^0$. If [X] doubles, what happens to the rate?
b. This reaction is 1st order in reactant X. What is the rate law? If [X] doubles, what happens to the rate?
c. This reaction is 2nd order in reactant X. What is the rate law? If [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 _____.
f. In the experiment in 4e, the concentration of X doubles. What happens to k? What happens to the rate?
g. In the experiment in 4e, the temperature is lowered to 10^{o}C. What happens to k? What happens to the rate?

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:

O_3 (g) + NO (g) ----> NO_2 (g) + O_2 (g).

Experiments with ozone and NO gave the following data:

Experiment	[O_3 (g)]	[NO (g)]	Rate, M/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 [O_3 (g)]^x [NO (g)]^y. Determine x and y.
Hint: Compare Experiments 1 and 2. [O_3 (g)] doubles and [NO (g)] is constant and rate doubles so 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?

6. Bromate ion (BrO_3^-) 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.
$BrO_3^- + 5\ Br^- + 6\ H^+ \text{--->} 3\ Br_2 + 3\ H_2O$

Experiment	[BrO_3^-], M	[Br^-], M	[H^+], M	Rate, M/sec
1	0.10	0.10	0.10	8.0×10^{-4}
2	0.20	0.10	0.10	1.6×10^{-3}
3	0.20	0.20	0.10	3.2×10^{-3}
4	0.10	0.10	0.20	3.2×10^{-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.

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			

8. You did the iodine clock reaction in Lab 3. The activation energy for this reaction is 20 kJ/mole. The rate constant at 20°C is 2.5 M^{-1} sec^{-1}.
a. Calculate the rate constant for this reaction at 60°C. (Should k be larger or smaller at this temperature compared to 20°C? What equation compares k at two different temperatures?)
b. Your lab partner measures k = 1.7 M^{-1} sec^{-1} but forgot to record the temperature. Calculate the temperature that corresponds to this rate constant. (Should temperature be greater than 20°C or less than 20°C? What equation compares temperature at two different k's?)

9. The rate constant for the reaction H_2 (g) + I_2 (g) ---> 2 HI (g) is 5.4×10^{-4} $M^{-1}s^{-1}$ at 326 °C. At 410 °C the rate constant was found to be 2.8×10^{-2} $M^{-1}s^{-1}$. Calculate the activation energy for this reaction. (Answer: between 140 and 180 kJ/mole)