

Half-life <-- these are just little chemistry/math problems. Each rate order has a specific half-life form. Often you'll be asked to find k , the rate constant, or, if you are given the rate constant and the initial concentration, to find the half-life.

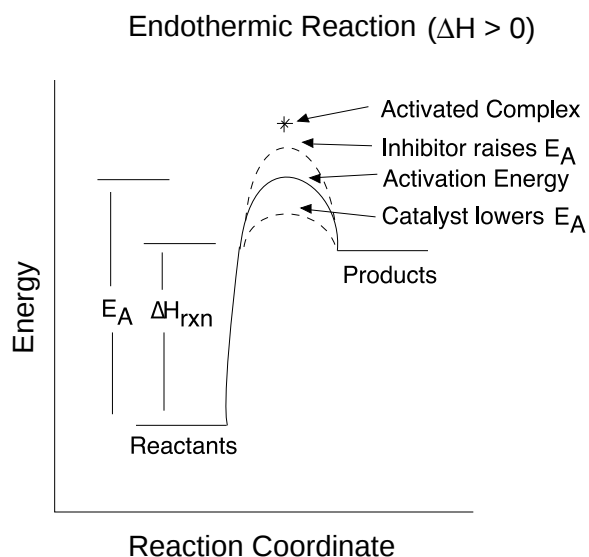
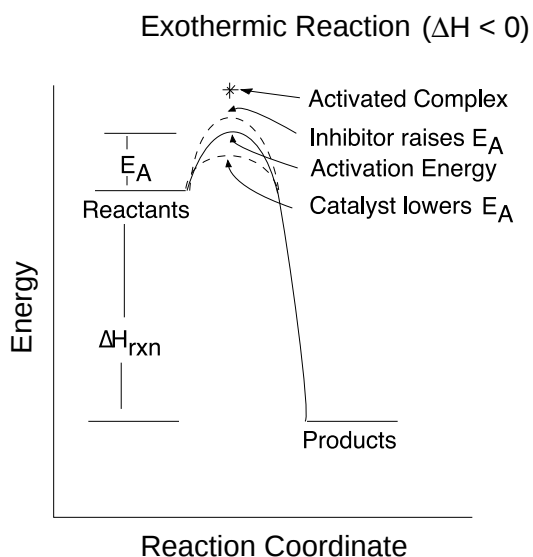
- What you want to do is look at the linear-forms of the concentration time expressions for each rate order.
- ★ Always remember that in half-life problems you don't have to be given numbers for concentrations. What you know is that at the half-life time the concentration of the reacting species is equal to half of its initial concentration ($[A] = 1/2[A]_0$).

Here's a summary for zero, first, and second order reactions:

Order	Linear $[A]$ vs time form	1/2-life expression	Dependence upon concentration
0	$[A] - [A]_0 = -kt$	$t_{1/2} = \frac{[A]_0}{2k}$	yes
1	$\ln[A] - \ln[A]_0 = -kt$	$t_{1/2} = \frac{.69}{k}$	no
2	$\frac{1}{[A]} - \frac{1}{[A]_0} = kt$	$t_{1/2} = \frac{1}{[A]_0 k}$	yes

Collision Theory, Activation Energy and Activated Complex

- This is just about trying to explain how the actual reaction occurs. Like what makes two reactant molecules actually react with each other when they collide. Sometimes they react, sometimes they don't. It's explained in terms of an activation energy that is required to trigger the reaction and an intermediate molecule, called the activated complex that is the unstable intermediate species occurring when the molecules collide. It's diagrammed out below for both an exothermic and an endothermic reaction (know each of these very well):



Activation Energy Problems

- ★ With the idea of activation energy, temperature really comes into play. The higher the temperature, the more energy that's in the system and the closer the molecules are to getting over the activation energy hump.
- ★ The important equation here is the Arrhenius equation that relates the rate constant to temperature.

Arrhenius Equation: $k = A e^{-\frac{E_A}{RT}}$

Linearized Arrhenius Equation: $\ln k = \ln A - \frac{E_A}{RT}$

k = rate constant

A = frequency factor

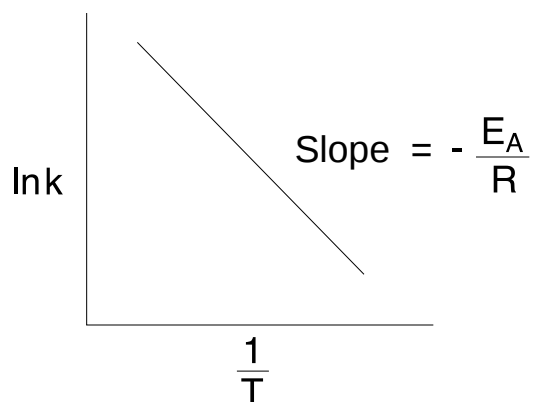
E_A = Activation Energy (Joules)

R = Universal gas constant (8.31 J/mol•K)

T = Temperature (Kelvin)

So, $\ln k = \ln A - \frac{E_A}{RT}$ is a linear equation that can be used to find a number of things.

- A. Plot the natural log of the rate constant, $\ln(k)$ vs. inverse temperature, $\frac{1}{T}$ and the slope of the line will give you the activation energy



- B. Or, if you know E_A and k at one temperature (Activation energy doesn't change with temperature) then you can find k at a different temperature using the following equation:

$$\ln \frac{k_2}{k_1} = \frac{E_A}{R} \left[\frac{1}{T_1} - \frac{1}{T_2} \right]$$

Also, if you know the rate constant and temperature at two points you can use the above equation to solve for the activation energy.

Student Packet: 4.4: Collision Theory: AssessMe questions

- 1) The rate constant of a reaction is generally expected to
 - A) be independent of temperature.
 - B) decrease with increasing temperature.
 - C) increase with increasing temperature.
 - D) increase with increasing temperature, only if the reaction is endothermic.
 - E) I'm clueless

- 2) An approximate rule is that an increase in temperature of 10°C will double the reaction rate. Which of the following is doubled on the molecular level? The:
 - A) activation energy.
 - B) average kinetic energy.
 - C) average velocity of molecules.
 - D) number of collisions.
 - E) I'm clueless

- 3) The decomposition of NaClO_3 is a first order reaction. In an experiment, 90% of a sample of NaClO_3 decomposed in 48.0 minutes. How long would it take for 50% of a sample to decompose?
A) 14.4 min B) 21.6 min C) 24.0 min D) 43.2 min E) I'm clueless

- 4) The half-life of ^{210}Po is 140 days. What mass of a 2.00 g sample will be left after 700 days?
A) 1.00 g B) 0.50 g C) 0.25 g D) 0.06 g E) I'm clueless

- 5) Which rate law is first order in B and second order in A?
 - A) $d[A]/dt = k_{\text{obs}}[A][B]$
 - B) $d[A]/dt = k_{\text{obs}}[A]^2[B]$
 - C) $d[A]/dt = k_{\text{obs}}[A]^4[B]^{-1}$
 - D) $d[A]/dt = k_{\text{obs}}[A]^5/[B]$
 - E) I'm clueless

- 6) Which rate law is second order overall?
 - A) $d[A]/dt = k_{\text{obs}}[A][B]$
 - B) $d[A]/dt = k_{\text{obs}}[A]^2[B]$
 - C) $d[A]/dt = k_{\text{obs}}[A]^4[B]^{-1}$
 - D) $d[A]/dt = k_{\text{obs}}[A]^5/[B]$
 - E) I'm clueless

Student Packet: 4.4: Collision Theory: ShowMe questions

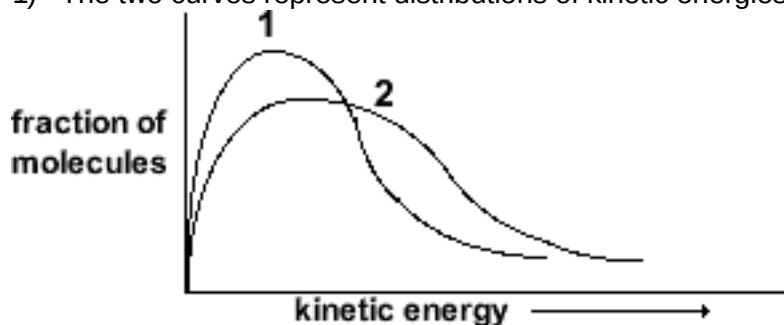
- 1) The following table shows the rate constants for a typical reaction. From these data calculate the activation energy for the reaction.

<u>T (K)</u>	<u>k (sec⁻¹)</u>
573.9	2.50×10^{-5}
583.8	5.32×10^{-5}
618.1	6.32×10^{-4}
643.2	3.20×10^{-3}

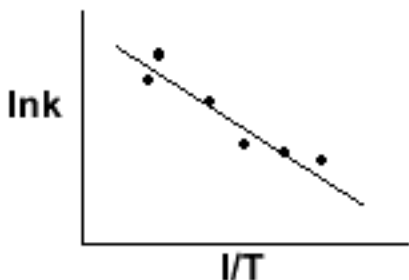
- 2) The denaturation of a certain virus is a first order process with an E_a of 277 kJ/mol. At 300 K the rate constant is $8.09 \times 10^{-8} \text{ sec}^{-1}$. What is the rate constant at 127°C?
- 3) Sketch the energy possible for the decomposition of HI into H_2 and I_2 . $E_a = 182 \text{ kJ/mol}$ assuming $\text{DE} = 20 \text{ kJ/mol}$.
- 4) The rate constant for a specific reaction is $3.2 \times 10^{-4} \text{ L/(mol}\cdot\text{sec)}$ at 610 K and $4.1 \times 10^{-3} \text{ L/(mol}\cdot\text{sec)}$ at 670 K. Find the activation energy and then calculate the rate constant at 700 K.

Student Packet: 4.4: Collision Theory: TestMe questions

- 1) The two curves represent distributions of kinetic energies in a sample of gas molecules.



- A) Curve 1 is at a higher temperature
 B) Curve 2 is at a higher temperature
 C) Curve 2 is at equilibrium
 D) Curve 1 is at equilibrium
 E) Both II and III
- 2) If the rate constant for a reaction is $2.50 \times 10^{-5} \text{ sec}^{-1}$ at 573.9 K, what is the magnitude of the rate constant at 530 K given that the energy of activation is 216 kJ/mol?
- A) $-4.08 \times 10^{-5} \text{ sec}^{-1}$ B) $5.88 \times 10^{-7} \text{ sec}^{-1}$ C) $4.08 \times 10^{-5} \text{ sec}^{-1}$
 D) $7.62 \times 10^{-4} \text{ sec}^{-1}$ E) None of the above
- 3) Which of the following reactions would be the fastest assuming that the reactions are of the same order, involve the same starting concentrations of reactants and have the same magnitude for the frequency factor, A, in the Arrhenius equation? The reactions have activation energies of:
- A) 75 kJ/mol B) 45 kJ/mol C) 172 kJ/mol D) 200 J/mol E) 45 J/mol
- 4) In order for the gases NO and NO₃ to react:
- I. the reactant particles, NO and NO₃ must bump into each other
 II. the molecules collision energy must exceed the activation energy of the reaction
 III. the molecular orientation must be correct for reaction
 IV. the molecules activation energy must exceed their collision energy
- A) I, II B) I, II, III, IV C) I, II, III D) I, III, IV E) I, III
- 5) The rate of a reaction was measured at several temperatures. A graph of $\ln k$ versus $1/T$ was plotted:



The line has the formula $y = -9.522 \times 10^3(1/T) + 334$

The activation energy of this reaction is:

- A) 182 J/mol B) $9.522 \times 10^3 \text{ J/mol}$ C) 79.17 kJ/mol D) 9.522 kJ/mol E) 277 kJ/mol