# *CH 222 Winter 2025:*
# **"Kinetics II - The Iodination Of Acetone** *(online)*" ***Lab*** **-** *Instructions*

*Note:* **This is the lab for section W1 of CH 222 only.**

- *If you are taking section 01 or section H1 of CH 222*, please use this link: http://mhchem.org/r/9a.htm

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*Step One:*

**Watch the lab video** for the "Kinetics II" lab, found here:

**http://mhchem.org/y/9.htm**

**Record** the data found at the ***end*** of the lab video on page Ib-9-5.

*Step Two:*

**Complete pages Ib-9-5 through Ib-9-11** using the "Kinetics II" video and the actual lab instructions on pages Ib-9-2 through Ib-9-3 (most of this lab is a "tutorial", so you will also find instructions in the lab pages as well.) Include your name on page Ib-9-5!

*Step Three:*

**Submit your lab** (pages Ib-9-5 through Ib-9-11 *only* to avoid a point penalty) **as a *single* PDF file to the instructor via email (mike.russell@mhcc.edu) on Wednesday, March 12 by 11:59 PM. Be sure to include any necessary computer generated graphs as well.** I recommend a free program like CamScanner (https://camscanner.com) to convert your work to a PDF file. Do not include the graph as a separate file.... have all documents in one file to avoid a point penalty.

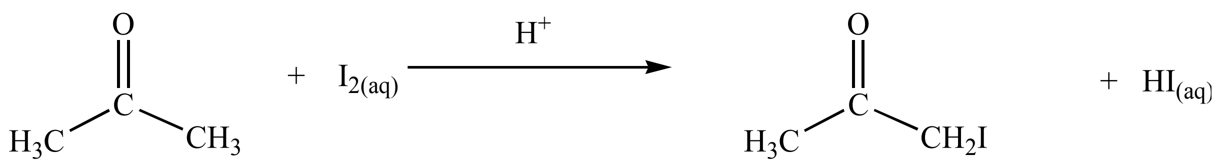
*If you have any questions regarding this assignment, please email (mike.russell@mhcc.edu) the instructor! Good luck on this assignment!*

# Kinetics II - The Iodination Of Acetone

**Determining the Activation Energy for a Chemical Reaction**

The **rate** of a chemical reaction depends on several factors: the nature of the reaction, the concentrations of the reactants, the temperature, and the presence of a possible catalyst. In *Part One* of this experiment we determined the **rate law** for a reaction by changing the concentrations and measuring the rate of the reaction. During *Part Two*, we will explore the relation between the rate constant and temperature to discover the **activation energy** for this reaction.

In this experiment, we continue to study the kinetics of the reaction between iodine and acetone:




In last week's lab, you discovered the **average rate constant, k**, value, as well as the **orders of reaction** (**m, n and p**) which apply to the **rate law** for this reaction:

$$\textbf{rate} = \mathbf{k[acetone]^m[H^+]^n[I_2]^p}$$

In this experiment, you will study the rate of the reaction at different temperatures to find its **activation energy**, $\mathbf{E_a}$. The temperature at which the reaction occurs influences the rate of the reaction. An increase in temperature increases the rate.

As with concentration, there is a quantitative relation between reaction rate and temperature, but here the relation is somewhat more complicated. This relation is based on the idea that to react, the reactant species must have a certain minimum amount of energy present (and the correct geometry, if appropriate) at the time the reactants collide in the reaction step.

This amount of energy, which is typically furnished by the kinetic energy of the species present, is called the **activation energy** ($\mathbf{E_a}$**,** also known as the *energy of activation*) for the reaction. The formula (called the *Arrhenius equation*) relating the rate constant k to absolute Kelvin temperature T and $E_a$ is:

$$\ln k = \frac{-E_a}{RT} + \ln A$$

In this equation, **R** is the gas constant (**8.3145 J/mole K**), and natural logarithms (**ln**) need to be used (*do not use base 10 logs!*) The quantity **A** is referred to as the **collision frequency** and A can be used to determine the fraction of molecules present with sufficient energy and geometry to become products at a given instant in time.

By measuring k at different temperatures, we can graphically determine the activation energy for a reaction. In Part Two of this experiment you will determine the effect of temperature on rate and calculate the activation energy for the reaction.

**DIRECTIONS:**

You will collect data during lab and complete the worksheets at the end. No formal typed lab report is required as long as your writing is legible and easy to read. In this experiment, you will see the effect of temperature upon the reaction.

We shall measure one of the reactions from the "Kinetics I" lab at different temperatures. You do not need to repeat these experiments twice (to be within 20 seconds of each other) as in Part One.

**For each entry in trial #5 listed below**: measure out the appropriate quantities of **1.0 M HCl, 4.0 M acetone** and **water** using a 10.00 mL graduated cylinder and place them in a 125 mL Erlenmeyer flask. Now measure out the appropriate amount of **0.0050 M iodine** in a 10.00 mL graduated cylinder.

Using ice and/or a hot plate, get the solution to a desired **temperature** before adding the iodine. **Record the temperature**, then add the iodine to the Erlenmeyer, start a stopwatch and measure how long the reaction takes to turn the solution clear. Time should be recorded in **seconds**. The iodine does not need to be at the same temperature as the solution in the Erlenmeyer flask.

Temperatures need to be higher than 15 °C (too slow!) and lower than 60 °C (keep the acetone from boiling), and the interval between measurements needs to be at least 5 °C apart.

Waste can be placed in the drain or in a waste bottle (which is probably the better option!) Get a stamp in your lab notebook before leaving lab.

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**Name:**

# Kinetics II - The Iodination Of Acetone

As before, add all the chemicals but iodine to a 125 mL Erlenmeyer flask. Using ice and/or a hot plate, get the solution to a desired **temperature** before adding the iodine. **Record the temperature**, then add the iodine to the Erlenmeyer, start a stopwatch and measure how long the reaction takes to turn the solution clear. Time should be recorded in **seconds**. The iodine does not need to be at the same temperature as the solution in the Erlenmeyer flask.

* Record one trial at **room temperature**. *Hint:* you *may* be able to use some of your data from Part I!
* Record one trial at a **temperature lower than room temperature, but above 15 degrees Celsius**.
* Record **three trials at temperatures higher than room temperature, but under 60 degrees Celsius**.

Temperature differences should be at least 5 degrees Celsius apart (i.e. spread out your temperatures!)

***NOTE:*** **Record the data found at the end of the video in the places below:**

**Trial #5:**

| HCl (mL) | Acetone (mL) | $I_2$ (mL) | Water (mL) | Total Volume (mL) |
|---|---|---|---|---|
| 5 | 5 | 5 | 10 | 25.00 |

Temperature (°C): ____________________ Time (seconds): _______________

Temperature (°C): ____________________ Time (seconds): _______________

Temperature (°C): ____________________ Time (seconds): _______________

Temperature (°C): ____________________ Time (seconds): _______________

Temperature (°C): ____________________ Time (seconds): _______________

*Hint:* As temperature increases, the reaction time should decrease. If you do not see this trend, repeat one or more experiments.

You are now ready to complete the Kinetics II lab!

# Kinetics II - The Iodination Of Acetone – *Worksheet*

Earlier you determined the time elapsed for a given set of concentrations as the temperature was altered. We shall use that information and techniques similar to that of the "Kinetics Part I" lab to determine the **energy of activation** and **collision frequency** for the iodination of acetone reaction using the Arrhenius equation.

a. **Find the inverse Kelvin temperature** for each value in Trial #5.

Convert your temperatures from °C to K, then take the *inverse* of your Kelvin temperatures.

*Example:* Convert 37.5 °C to an inverse Kelvin temperature.

*Solution:* 37.5 °C = 310.7 K. To find the inverse, calculate $(310.7\ K)^{-1}$ = $\mathbf{3.219 * 10^{-3}\ K^{-1}}$

Complete the table below. The first column (Temperature (°C) comes from your data collected in Part Two, Trial #5, while in lab.

| | **Temperature (°C)** | **Temperature (K)** | **Temperature$^{-1}$ ($K^{-1}$)** |
|---|---|---|---|
| Temp #1 | __________ | __________ | __________ |
| Temp #2 | __________ | __________ | __________ |
| Temp #3 | __________ | __________ | __________ |
| Temp #4 | __________ | __________ | __________ |
| Temp #5 | __________ | __________ | __________ |

Use this space to show **a sample calculation** for getting from Temperature (°C) to an inverse Kelvin temperature:

b. **Find the rate** for each temperature value in Trial #5.

Recall from the "Kinetics Part I" lab that, for this experiment:
**rate = $[I_2]$/(time in seconds)**

Use this equation to **find the rate of reaction** (in M $s^{-1}$) for each temperature. Time values come from Trial #5, above:

| | $I_2$ (M) | time (s) | rate (M $s^{-1}$) |
|---|---|---|---|
| Trial #1 | 0.0010 | __________ | __________ |
| Trial #2 | 0.0010 | __________ | __________ |
| Trial #3 | 0.0010 | __________ | __________ |
| Trial #4 | 0.0010 | __________ | __________ |
| Trial #5 | 0.0010 | __________ | __________ |

Use this space to show at least one example of how you calculated the rate of the reaction.

c. **Find the value of the rate constant, k,** for each temperature value in Trial #5.

We will use the process developed in "Kinetics Part I" lab to help us find the values of k for each temperature.

i. First, we need the **diluted concentrations**: *(these can be found in section b of the "Kinetics Part I" lab)*

Concentration (M) of **acetone** when 5.00 mL was used: _______________

Concentration (M) of **HCl** when 5.00 mL was used: _______________

Concentration (M) of **$I_2$** when 5.00 mL was used: _______________

ii. Next, we need the **reaction orders** for each reactant: *("Kinetics Part I", section d)*

My value of m (acetone) = __________

My value of n (HCl) = __________

My value of p ($I_2$) = __________

*If you do not have m = 1, n = 1, p = 0, talk to the instructor!*

iii. Now use the techniques from in Kinetics I", section e, to find the value of k, and then ln k.

Use the **rates from section b** *(above)* and the values for concentration and order (m, n and p) **to find k**. The only variable that will change is the rate; the orders and concentrations remain constant. Take the **natural log (ln) of each k** value as well (i.e. ln ($2.6 * 10^{-5}$) = **-10.56)** (Note: *report your ln k values to the **hundredths** place to satisfy significant figures.*)

| | **rate ($M\ s^{-1}$)** | **k** | **ln k** |
|---|---|---|---|
| Trial #1 | __________ | __________ | __________ |
| Trial #2 | __________ | __________ | __________ |
| Trial #3 | __________ | __________ | __________ |
| Trial #4 | __________ | __________ | __________ |
| Trial #5 | __________ | __________ | __________ |

Show a sample calculation for these steps on the next page.

Use this space to show at least one example of how you calculated the rate constant k and ln k.

d. **Create a graph** of ln k versus inverse Kelvin temperature values

You will be creating a graph of ln k versus inverse temperature to find the energy of activation.

First, collect your inverse temperature (the x-axis) and ln k values (the y-axis) here. Inverse Kelvin temperature values come from section a in this lab. ln k values come from section c, subsection iii, in this lab.

| | **Temperature$^{-1}$ ($K^{-1}$)** | **ln k** |
|---|---|---|
| Trial #1 | __________ | __________ |
| Trial #2 | __________ | __________ |
| Trial #3 | __________ | __________ |
| Trial #4 | __________ | __________ |
| Trial #5 | __________ | __________ |

Using Excel or a similar program, create a **graph** of your ln k values versus the inverse Kelvin temperature values. Make the graph at least as big as half a sheet of paper, and be sure to include unit labels (ln k for the y-axis and (Temperature)$^{-1}$ for the x-axis. *Note* that when using your graphing program, you may need to enter values as decimals, i.e. enter 0.00315 instead of $3.15 \times 10^{-3}$.

**Staple / attach the graph to the end of this lab report packet.**

e. **Find the energy of activation** for the iodination of acetone using the data in Trial #5.

The data points on the graph from the last section should correspond roughly to a straight line with a negative slope. This is the behavior predicted by the Arrhenius equation:

$$\ln k = \frac{-E_a}{RT} + \ln A$$

where **ln k** is the y-axis, **(Temperature in Kelvin)$^{-1}$** is the x-axis, **$-E_a/R$** is the slope, **$R = 8.3145\ J\ K^{-1}\ mol^{-1}$** (the "energy" gas constant), **$E_a$** is the energy of activation, and **A** is the collision frequency.

Perform a linear regression analysis using your calculator or graphing program (inverse Kelvin temperatures will be your x-axis, ln k values will be your y-axis.) Record the values that you collected here:

Slope = ___________ y-intercept = _________ correlation coefficient (r) = ____________

The **energy of activation**, $E_a$, can be determined from the slope. From the value of the slope determined through linear regression, **calculate the activation energy**.

Energy of activation = **–R*slope** = _____________

**Units** for your Energy of activation value = _____________

The collision frequency, A, can be determined from the y-intercept. From the value of the y-intercept determined through linear regression, **calculate the collision frequency**.

Collision frequency = **$e^{y\text{-int}}$** = _____________

***Note that*** **e** is the anti natural logarithm.

f. **You are done!** Finish the postlab question (which is similar to the work you just completed) and you are good to go!

**Kinetics II - The Iodination Of Acetone - Postlab Question:**

The following reaction

$$2\ N_2O_{5(g)} \rightarrow 4\ NO_{2(g)} + O_{2(g)}$$

was studied at several temperatures, and the following values of k were obtained:

| k ($s^{-1}$) | T (˚C) |
|---|---|
| $2.0*10^{-5}$ | 20.0 |
| $7.3*10^{-5}$ | 30.0 |
| $2.7*10^{-4}$ | 40.0 |
| $9.1*10^{-4}$ | 50.0 |
| $2.9*10^{-3}$ | 60.0 |

Using linear regression and the techniques developed in this lab, calculate the **activation energy** and **collision frequency** for this reaction. **Include a computer generated graph** of ln k versus $(T)^{-1}$. *Hint:* make sure you use inverse Kelvin temperatures! Make sure the x-axis lists "0.003" numbers (and not whole integers, etc.)

Slope = ___________ y-intercept = _________ correlation coefficient (r) = ____________

Energy of activation = _____________

**Units** for your Energy of activation value = _____________

Collision frequency = _____________

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