CH 222 Winter 2025: **"Kinetics I - 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/8a.htm

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Step One:

Watch the lab video for the "Kinetics I" lab, found here:
http://mhchem.org/y/8.htm
Record the data found at the ***end*** of the lab video on pages Ib-8-4 and Ib-8-5.

Step Two:

Complete pages Ib-8-5 through Ib-8-15 using the "Kinetics I" video and the actual lab instructions on pages Ib-8-2 through Ib-8-3 (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-8-5!

Step Three:

Submit your lab (pages Ib-8-5 through Ib-8-15 *only* to avoid a point penalty**) as a *single* PDF file to the instructor via email (mike.russell@mhcc.edu) on Wednesday, March 5 by 11:59 PM.** I recommend a free program (ex: CamScanner, https://camscanner.com) or a website (ex: CombinePDF, https://combinepdf.com) to convert your work to a PDF file.

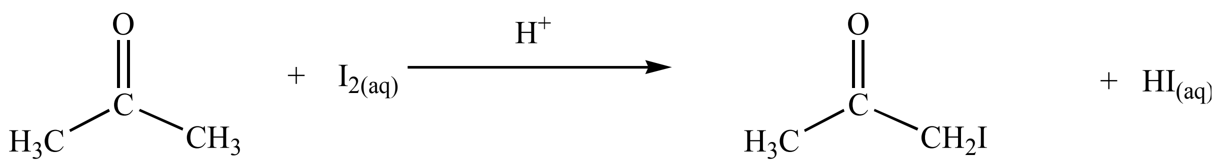
If you have any questions regarding this assignment, please email (mike.russell@mhcc.edu) the instructor! Good luck on this assignment!

Kinetics I - The Iodination Of Acetone

Determining the Rate Constant 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 will determine the **rate law** for a reaction by changing some of the above variables 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 will study the kinetics of the reaction between iodine and acetone:



The rate of this reaction is found to depend on the concentration of the hydrogen ion (acid, HCl) as well as the concentrations of the reactants (acetone and iodine). The **rate law** for this reaction is

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

where **k** is the **rate constant** for the reaction and **m**, **n**, and **p** are the **orders of the reaction** with respect to acetone, hydrogen ions (acid), and iodine, respectively. Although orders of reaction can be any value, for this lab we will be looking only for **integer values** for the orders of reaction (0, 1, 2 are acceptable but not 0.5, 1.3, etc.)

The **rate** of the reaction can also be expressed as the change in the concentration of a reactant divided by the time interval:

$$rate = \frac{-\Delta[I_2]}{\Delta t}$$

The iodination of acetone is easily investigated because iodine (I_2) has a deep yellow/brown color. As the acetone is iodinated and the iodine converted to the iodide anion, this color will disappear, allowing the rate of the reaction to be easily monitored.

We can study the rate of this reaction by simply making I_2 the ***limiting reactant*** in a large excess of acetone and H^+ ion. By measuring the time required for the initial concentration of iodine (I_2) to be used up completely, the rate of the reaction can be determined by the equation

$$rate = \frac{-\Delta[\mathrm{I_2}]}{\Delta \mathrm{t}} = \frac{-([I_2]_{final} - [I_2]_{initial})}{\mathrm{t}_{final} - t_{initial}} = \frac{-(0 - [I_2]_{initial})}{\mathrm{t}_{final} - 0} = \frac{[I_2]_{initial}}{\mathrm{t}_{final}}$$

or simply as

$$rate = \frac{[I_2]}{time}$$

From the rate information, we can determine the **orders** with respect to acetone (**m**), acid (**n**) and iodine (**p**) by varying the amounts of reactants and measuring the effect on the rate. Once the orders of reaction are known, we will be able to calculate the rate constant, k. In Part One of this experiment you will determine the rates of reactions, the orders of the reactants, and finally the rate constant at room temperature.

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. You will perform a series of experiments that will examine the relationship between the concentration of reactants and the time for the iodination of acetone reaction.

For each trial 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.

Start a timer (stopwatch) as you add the iodine to the 125 mL flask with the other chemicals. Swirl the flask **until the yellow color disappears**, **then halt the timer**. It may help to place the flask on a white piece of paper to help discern when the color disappears. Record the time elapsed in seconds.

Repeat this reaction mixture until **two** trials are within 20 seconds of each other.

Repeat this process for each of the four trials listed in the table below. Waste can be placed in the drain or in a waste bottle (probably the better option!) Get a stamp in your lab notebook before leaving lab, then complete the worksheet portions on your own.

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Name:

Kinetics I - The Iodination Of Acetone

Part I: Changing Concentration to Find the Rate Constant

For each mixture listed below, add all of the chemicals but iodine to a 125 mL Erlenmeyer flask. Add the iodine last, starting a stopwatch and measuring how long the reaction takes to turn the solution clear. Time should be recorded in seconds. **Repeat** each reaction mixture until **two** trials are within 20 seconds of each other.... repeat the trial again if the times vary too much. ***NOTE:*** **Record the data found at the end of the video in the places below for Trials #1 - #4.**

Trial #1:

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

Time in seconds for yellow color to disappear, first time: ____________________ seconds

Time in seconds for yellow color to disappear, second time: ____________________ seconds

Time in seconds for yellow color to disappear, third time (if necessary): ____________________ seconds

Trial #2:

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

Time in seconds for yellow color to disappear, first time: ____________________ seconds

Time in seconds for yellow color to disappear, second time: ____________________ seconds

Time in seconds for yellow color to disappear, third time (if necessary): ____________________ seconds

<u>Trial #3</u>:

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

Time in seconds for yellow color to disappear, first time: ____________________ seconds

Time in seconds for yellow color to disappear, second time: ____________________ seconds

Time in seconds for yellow color to disappear, third time (if necessary): ____________________ seconds

<u>Trial #4</u>:

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

Time in seconds for yellow color to disappear, first time: ____________________ seconds

Time in seconds for yellow color to disappear, second time: ____________________ seconds

Time in seconds for yellow color to disappear, third time (if necessary): ____________________ seconds

Before you move on, check your data. Trial #1 should be about half the time of trial #2, and trial #3 and trial #4 should both be about the same time elapsed. If you don't see this trend, contact the instructor.

You are now ready to complete the Kinetics I lab!

Kinetics I - The Iodination Of Acetone – *Worksheet*

The ultimate goal of this section is to find the best value of the rate constant, k, for the iodination of acetone at room temperature. To get there, we need to first find the rates of each reaction, then the order of the reactants (acetone, HCl and iodine), and then finally, the rate constant k.

a. **Find the average time in seconds** for each Trial in Part One.

Example: the first experiment took 230 seconds, the second experiment took 250 seconds.

The *average time* would be: (230 + 250) / 2 = **240 seconds**

Record your reaction times from Part I (which were within 20 seconds of each other) and the final average times here:

	Experiment #1 (s)	Experiment #2 (s)	Average time (s)
Trial #1	____________	____________	____________
Trial #2	____________	____________	____________
Trial #3	____________	____________	____________
Trial #4	____________	____________	____________

Show an example of how you got the average values in seconds here:

b. **Find the concentration of each reactant** (acetone, HCl and iodine) before the reaction started.

Your group took bulk reactants (which were, as a reminder: **1.0 M HCl, 4.0 M acetone**, and **0.0050 M iodine**) then placed them (with water) in an Erlenmeyer flask. The final volume was always 25.00 mL. Mixing chemicals dilutes the concentrations from the "bulk" value to a smaller value.

We can determine these diluted values using: $\mathbf{M_1V_1 = M_2V_2}$

Example: Determine the concentration of iodine in trial #1.

Let $\mathbf{M_1}$ = initial (undiluted) concentration of iodine (0.0050 M), $\mathbf{V_1}$ = 5.00 mL (of undiluted iodine added to the mixture), and $\mathbf{V_2}$ = 25.00 mL (the total volume of the diluted solution once HCl, acetone and water are added). Solving for M_2, the concentration of iodine in the diluted solution, one gets:

> $\mathbf{M_2}$ = 0.0050 M * 5.00 mL / 25.00 mL = **0.0010 M**, which is the concentration of iodine used in the reaction in trial #1.

Example: Determine the concentration of iodine in trial #2.

Solution: Since $\mathbf{M_1}$ and $\mathbf{V_2}$ are the same as in the previous example, we see that only $\mathbf{V_1}$ has changed to 10.00 mL. Rearranging for $\mathbf{M_2}$ as before:

> $\mathbf{M_2}$ = 0.0050 M * 10.00 mL / 25.00 mL = **0.0020 M**, the concentration of iodine used in the reaction in trial #2.

i. **Calculate the concentration of acetone ($\mathbf{M_2}$) used in Trial 1 - Trial 4.**

The **concentration** of acetone was 4.0 M in the "bulk" solution (the "$\mathbf{M_1}$" value.)
The final volume ($\mathbf{V_2}$) is always 25.00 mL.
Trial 1, Trial 2 and Trial 4 used 5.00 mL of acetone from the bulk solution, but Trial 3 used 10.00 mL (your "$\mathbf{V_1}$" values)

	volume acetone (mL)	**Bulk Acetone (M)**	**Total Volume (mL)**	**acetone (M)**
Trial #1	5.00	4.0 M	25.00	__________
Trial #2	5.00	4.0 M	25.00	__________
Trial #3	10.00	4.0 M	25.00	__________
Trial #4	5.00	4.0 M	25.00	__________

ii. **Calculate the concentration of HCl (M_2) used in Trial 1 - Trial 4.**

The **concentration** of HCl was 1.0 M in the "bulk" solution (the "$\mathbf{M_1}$" value.)
The final volume (**V_2**) is always 25.00 mL.
Trial 1, Trial 2 and Trial 3 used 5.00 mL of acetone from the bulk solution, but Trial 4 used 10.00 mL (your "$\mathbf{V_1}$" values)

	volume HCl (mL)	**Bulk HCl (M)**	**Total Volume (mL)**	**HCl (M)**
Trial #1	5.00	1.0 M	25.00	________
Trial #2	5.00	1.0 M	25.00	________
Trial #3	5.00	1.0 M	25.00	________
Trial #4	10.00	1.0 M	25.00	________

iii. **Complete the following table** showing the diluted concentrations of all reactants used in each trial. The I_2 concentrations have been completed for you (see the examples at the beginning of part b, above.)

	acetone (M) *step i, above*	**HCl (M)** *step ii, above*	**I_2 (M)**
Trial #1	________	________	0.0010
Trial #2	________	________	0.0020
Trial #3	________	________	0.0010
Trial #4	________	________	0.0010

Use this area to show relevant calculations and at least one example as to how you found a diluted concentration.

c. **Find the rate of each trial**

In this lab, rate is best described by: **rate = [I_2]/(average time in seconds)** (see first page of lab)
Find the rate for each reaction by using the calculated iodine concentration then dividing by the average time for that trial (Part Three, section a, above).

Example: Trial #1 required an average of 240 seconds to go to completion. Trial #2 required an average of 496 seconds to go to completion. Find the rate of reaction for both trial #1 and trial #2.

Solution: The rate of reaction is equal to the concentration of iodine divided by the average time elapsed for the reaction. (your values are in Part Three, section a.)

In this example, the first two times are 240 s and 496 s. Using the appropriate concentrations, we can calculate rate:

rate (trial #1) = [I_2]/(average time in seconds) = 0.0010 M / 240 s = $\mathbf{4.2 * 10^{-6}\ M\ s^{-1}}$.

rate (trial #2) = [I_2]/(average time in seconds) = 0.0020 M / 496 s = $\mathbf{4.0 * 10^{-6}\ M\ s^{-1}}$.

Now, **calculate your rate values** by filling in the table below:

	I_2 (M)	**average time (s)**	**rate ($M\ s^{-1}$)**
Trial #1	0.0010	__________	__________
Trial #2	0.0020	__________	__________
Trial #3	0.0010	__________	__________
Trial #4	0.0010	__________	__________

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

d. **Find the order** of the reaction with respect to each reactant (**m**, **n**, and **p**)

To find the order of each reactant, we need to localize the effect that each reactant had on the rate. This is expressed in a **rate order** which, for CH 222 and CH 223, **can only be equal to zero, one or two** (no fractions, negative numbers, etc.)

In this lab, it is safe to assume that **water does not affect the rate**, so we can ignore its contribution to any changes that the rate might exhibit.

Notice that **trial 2 has twice as much iodine as trial 1**. Any changes to the rates of trial 1 and trial 2 are a direct result of the iodine (HCl and acetone are constant), and we will use these trials to calculate **p, the reaction order for iodine.**

Also notice that **trial 3 has twice as much acetone as trial 1**, so any changes to the rates of trials 3 and 1 will be the result of acetone only (HCl and iodine are constant), and we will use these trials to calculate **m, the reaction order for acetone.**

Lastly, notice that **trial 4 has twice as much HCl as trial 1**, yet acetone and iodine concentrations remain constant; we will use these trials to calculate **n, the reaction order for HCl.**

Example: Find the order of reaction with respect to iodine (p) if these values are used:

rate (trial #1) = $[I_2]$/(average time in seconds) = 0.0010 M / 240 s = $\mathbf{4.2 * 10^{-6}\ M\ s^{-1}}$.

rate (trial #2) = $[I_2]$/(average time in seconds) = 0.0020 M / 496 s = $\mathbf{4.0 * 10^{-6}\ M\ s^{-1}}$.

Solution: Notice how in trial #2 we doubled the concentration of $[I_2]$ while leaving the other reactants' concentrations (HCl, acetone) unchanged. An effect upon the rate of the reaction will reflect the influence of the iodine only, and this will allow us to determine **p**.

In these sample calculations, doubling the concentration of iodine (to 0.0020 M from 0.0010 M) spawned a negligible change in the rate ($4.0 * 10^{-6}$ M s^{-1} versus $4.2 * 10^{-6}$ M s^{-1}). Because we are only concerned with whole integer values of rate orders, this implies a **zero order reactant**, and **p = 0**.

A more formal approach to determining p would be as follows: divide the expression for rate 2 by the expression for rate 1; this results in the following:

$$\frac{\text{rate 2}}{\text{rate 1}} = \frac{k[\text{acetone}]^m[\text{HCl}]^n[I_2]^p}{k[\text{acetone}]^m[\text{HCl}]^n[I_2]^p}$$

The values of k, [acetone] and [HCl] remain constant between trial 1 and trial 2 (only $[I_2]$ changes), so the expression reduces to

$$\frac{\text{rate 2}}{\text{rate 1}} = \frac{4.0 * 10^{-6}}{4.2 * 10^{-6}} = \frac{[0.0020]^p}{[0.0010]^p} = 2^p$$

$$0.95 = (2)^p$$

Taking the logarithm of both sides leads to

$$\log 0.95 = \log 2^p = p \log 2$$

and solving for p:

$$p = \frac{\log 0.95}{\log 2} = -0.074 \approx 0$$

Therefore, the order with respect to iodine equals **zero**, or **p = 0**.

Now, **find the order of each reactant** by filling in the tables below. ***Note:*** **rate** values from Part 3 section c; **concentration** (M) values from Part 3, section b, subsection iii. Show your work on the following page. ***Remember*** that m, n and p can be **0, 1, or 2 only**! Round your answers as necessary!

For I_2 (p):

	I_2 (M)	**rate (M s^{-1})**	**My value of p is:**
Trial #1	0.0010	__________	__________
Trial #2	0.0020	__________	

For acetone (m):

	acetone (M)	**rate (M s^{-1})**	**My value of m is:**
Trial #1	__________	__________	__________
Trial #3	__________	__________	

For HCl (n):

	HCl (M)	**rate (M s^{-1})**	**My value of n is:**
Trial #1	__________	__________	__________
Trial #4	__________	__________	

Use this space to show **how** you got each of the orders of reaction (m, n and p):

e. **Find the value of k, the rate constant** for the iodination of acetone reaction.

You now have all the necessary information to calculate the rate constant, k, for each trial. For this reaction,

$$\text{rate} = k[\text{acetone}]^m[\text{HCl}]^n[I_2]^p$$

Rate values appear in Part 3 section c
[acetone], [HCl] and [I2] are the concentrations for each trial (Part 3, section b, subsection iii)
m, n and p are the orders of reaction (Part 3, section d)

Example: Find the value of k for trial #1 if the rate = 4.2×10^{-6} M s^{-1} and assuming that the order with respect to acetone (m) and HCl (n) is two and the order with respect to I_2 (p) is zero.

Solution: In trial #1, the diluted concentration of acetone is 0.80 M, the HCl is 0.20 M and I_2 is 0.0010 M. Using the given values, we can calculate k from the following equation:

$$\text{rate} = k[\text{acetone}]^m[\text{HCl}]^n[I_2]^p$$

$$4.2 * 10^{-6} = k[0.80]^2[0.20]^2[0.0010]^0$$

$$k = 4.2 * 10^{-6} / [0.80]^2[0.20]^2$$

and solving for k we get a value of $\mathbf{k = 1.6 * 10^{-4}\ M^{-1}\ s^{-1}}$

Now **find the rate constant k for each trial** by completing the table:

My value of m = __________ *(these can be found in section d, above)*

My value of n = __________

My value of p = __________

	acetone (M)	HCl (M)	I_2 (M)	rate ($M\ s^{-1}$)	value of k
Trial #1	__________	__________	0.0010	__________	__________
Trial #2	__________	__________	0.0020	__________	__________
Trial #3	__________	__________	0.0010	__________	__________
Trial #4	__________	__________	0.0010	__________	__________

Concentrations in section b, subsection iii *Rate values in section c*

Average value of k: ______________________

parts per thousand of your four k values: ______________________

Use this space to show **a sample calculation for k** and also relevant **parts per thousand calculations**:

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

Kinetics I - The Iodination Of Acetone - Postlab Questions:

1. In a reaction involving the iodination of acetone, the following reaction mixture was used: 5.00 mL 4.0 M acetone, 5.00 mL 1.0 M HCl, 5.00 mL 0.0050 M I_2, and 10.0 mL water.

    a. What was the molarity of the acetone in the reaction mixture? (Recall that $M_1V_1 = M_2V_2$)

    b. The color of the above reaction mixture disappeared in 250 seconds. What was the rate of the reaction? (Hint: First determine the initial concentration of the iodine, then use the equation for rate from the lab.)

2. A second reaction mixture was made: 10.00 mL acetone, 5.00 mL HCl, 5.00 mL I_2, and 5.00 mL of H_2O.

    a. What was the molarity of the acetone in this reaction mixture?

    b. The iodine color disappeared in 120 seconds. What was the rate of the reaction?

    c. Determine the order of the reaction (m) with respect to acetone using the information from question 1 and 2. (Round off the value of m to the nearest integer)

3. A third reaction mixture is made: 10.00 mL acetone, 5.00 mL HCl, 10.00 mL I_2. If the reaction is zero order with respect to iodine, how long will it take for the iodine color to disappear? (*Hint:* rate = $[I_2]$/time, use the rate from question 2b and the new $[I_2]$ to solve for the time elapsed.)

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