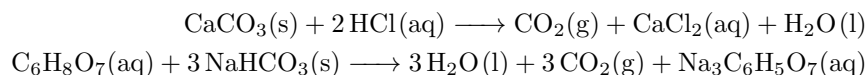


## Lab 1B: Factors Affecting Reaction Rate

### 1 Purpose

To investigate the effect concentration, temperature, and surface area have on the rate of reaction in the following two reactions:



The first reaction will be between Calcium Carbonate and Hydrochloric Acid, varying the concentration of the acid. The second reaction will be between the Citric Acid and Sodium Bicarbonate in a denture tablet, varying the temperature and surface area.

### 2 Procedure

#### 2.1 Concentration Test Procedure

1. Two similar  $\text{CaCO}_3$  rocks were picked and weighed.
2. The weight of a beaker was measured and the scale zeroed.
3. 50.0mL of 1.0M HCl was poured into the beaker.
4. The weight of the HCl and a  $\text{CaCO}_3$  rock combined was measured and recorded.
5. The  $\text{CaCO}_3$  rock was dropped into the beaker.
6. Weight measurements of the system were taken every 30 seconds for 5 minutes.
7. The process was repeated with the other rock in 3.0M HCl.

#### 2.2 Temperature and Surface Area Test Procedure

1. 50.0mL of room temperature water was poured into two beakers, and 50.0mL of boiling water into another beaker.
2. 3 denture tablets were picked and weighed.
3. A beaker with water was zeroed on the scale.
4. A denture tablet was dropped into the water and weight measurements were taken every 30 seconds for 5 minutes, stopping once fully dissolved.
5. This was done 3 times. A control test with room temperature water, a surface area test with a denture tablet crushed into a fine powder, and a heated test with hot water.

### 3 Data/Observations

#### 3.1 Concentration Test

##### 3.1.1 Concentration Test Data

Table 1: Initial Parameters of Open  $\text{CaCO}_3(\text{s})$  and  $\text{HCl}(\text{aq})$  Reaction.

Parameter	Test 1	Test 2
Concentration $\text{HCl}$ (M)	1.0	3.0
Mass $\text{HCl}$ (g)	17.9	32.5
Mass $\text{CaCO}_3$ (g)	10.6	13.7
Mass Total (g)	28.5	46.2

Table 2: Change in Mass Over Time of an Open  $\text{CaCO}_3(\text{s})$  and  $\text{HCl}(\text{aq})$  Reaction.

Time	Test 1 Mass (g)	Test 2 Mass (g)
0:00	28.5	46.1
0:30	28.3	46.0
1:00	28.3	45.8
1:30	28.2	45.6
2:00	28.1	45.4
2:30	28.0	45.3
3:00	28.0	45.1
3:30	27.9	45.0
4:00	27.9	44.9
4:30	27.9	44.9
5:00	27.8	44.8

The mass data collected in Table 2 is the sum of the mass of the  $\text{HCl}(\text{aq})$  solution and the  $\text{CaCO}_3(\text{s})$  rock.

##### 3.1.2 Concentration Test Observations

The two samples of  $\text{CaCO}_3(\text{s})$  were similar in shape and size. Both  $\text{HCl}(\text{aq})$  solutions were clear and colourless. In both tests, fizzing was observed during the reaction. However with the higher concentration acid, a much more vigorous fizzing was observed alongside a pungent smell and the solution becoming darker. The tests were not carried to completion due to time constraints.

#### 3.2 Temperature and Surface Area Test Data

##### 3.2.1 Our Group's Data

Table 3: Initial Parameters of Open Denture Tablet Reactions.

Parameter	Test 1 (Control)	Test 2 (Crushed)	Test 3 (Heated)
Crushed	no	yes	no
Heated	no	no	yes
Mass Beaker-Water (g)	98.0	75.6	80.9
Mass Tablet (g)	2.6	2.5	2.5

Table 4: Change in Mass Over Time of Open Denture Tablet Reactions.

Time	Test 1 Mass (g)	Test 2 Mass (g)	Test 3 Mass (g)
0:00	2.5	2.5	2.5
0:30	2.5	2.4	2.4
1:00	2.5	2.4	2.4
1:30	2.5	2.4	2.4
2:00	2.5	2.4	2.4
2:30	2.5	2.4	—
3:00	2.5	2.3	—
3:30	2.5	2.3	—
4:00	2.5	—	—
4:30	2.5	—	—
5:00	2.4	—	—

The mass data collected here is based on the mass of the denture tablet. The scale was zeroed with the beaker and water. Data was not collected after the tablet was fully dissolved.

### 3.2.2 Abigail & Jersey’s Data (Video)

Table 5: Initial Parameters of Open Denture Tablet Reactions.

Parameter	Test 1 (Control)	Test 2 (Crushed)	Test 3 (Heated)
Crushed	no	yes	no
Heated	no	no	yes
Mass Tablet (g)	2.57	2.53	2.59

Table 6: Change in Mass Over Time of Open Denture Tablet Reactions.

Time	Test 1 Mass (g)	Test 2 Mass (g)	Test 3 Mass (g)
0:00	2.57	2.49	2.59
0:30	2.54	2.53	2.45
1:00	2.47	2.48	2.39
1:30	2.51	2.49	2.27
2:00	2.47	2.53	2.19
2:30	2.46	2.50	2.24
3:00	2.46	2.42	2.20

### 3.2.3 Temperature & Surface Area Test Observations

All three denture tablets were very similar in shape (initially), size, weight, and colour. The crushed tablet was in a fine, even powder. The water appeared clear and odourless prior to the tests. During all three tests, foam formed on the surface of the water and the solution in the beaker turned a green colour. Halfway through the crushed tablet test, the solution turned a turquoise colour similar to seawater. During the heated water test, yellow splotches appeared on the surface of the foam. The crushed tablet and heated water tests were carried until the tablet was visibly dissolved. The control test was not carried to completion due to time constraints.

Minimal weight change was observed in our tests due to the precision of our scale (1 decimal place), so we received the data of another group. However in Abigail and Jersey’s tests (2 decimal places), fluctuating weight change was observed.

## 4 Analysis

### 4.1 Concentration Test Analysis

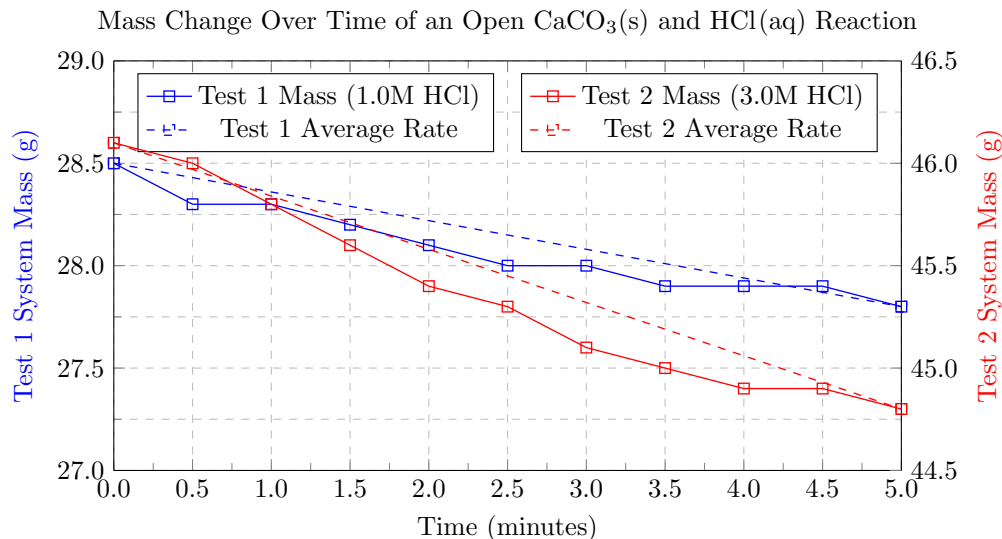


Figure 1: Change in mass of the system with average rates over the first 5 minutes of data.

For the reaction  $\text{CaCO}_3(\text{s}) + 2\text{HCl}(\text{aq}) \longrightarrow \text{CO}_2(\text{g}) + \text{CaCl}_2(\text{aq}) + \text{H}_2\text{O}(\text{l})$ , reduction in weight can be attributed to  $\text{CO}_2$  escaping the system, as it is the only gaseous product. The average rate is therefore in  $\frac{\text{g CO}_2}{\text{min}}$  and is represented in the graph by the slope of the dashed lines. This can be verified by calculating the average rate for both tests from the raw data:

$$\text{Average Rate}_{1.0M} = \frac{\text{Test 1 Mass}_{5:00} - \text{Test 1 Mass}_{0:00}}{5:00 - 0:00} = \frac{27.8 \text{ g} - 28.5 \text{ g}}{5 \text{ min}} = -0.1 \frac{\text{g CO}_2}{\text{min}}$$

$$\text{Average Rate}_{3.0M} = \frac{\text{Test 2 Mass}_{5:00} - \text{Test 2 Mass}_{0:00}}{5:00 - 0:00} = \frac{44.8 \text{ g} - 46.1 \text{ g}}{5 \text{ min}} = -0.26 \frac{\text{g CO}_2}{\text{min}}$$

Based on the calculated average rates, it is evident that the reaction with higher concentration of  $\text{HCl}$  had a faster rate of reaction (its magnitude is greater).  $|-0.26 \frac{\text{g CO}_2}{\text{min}}| > |-0.1 \frac{\text{g CO}_2}{\text{min}}|$

## 4.2 Temperature & Surface Area Test Analysis

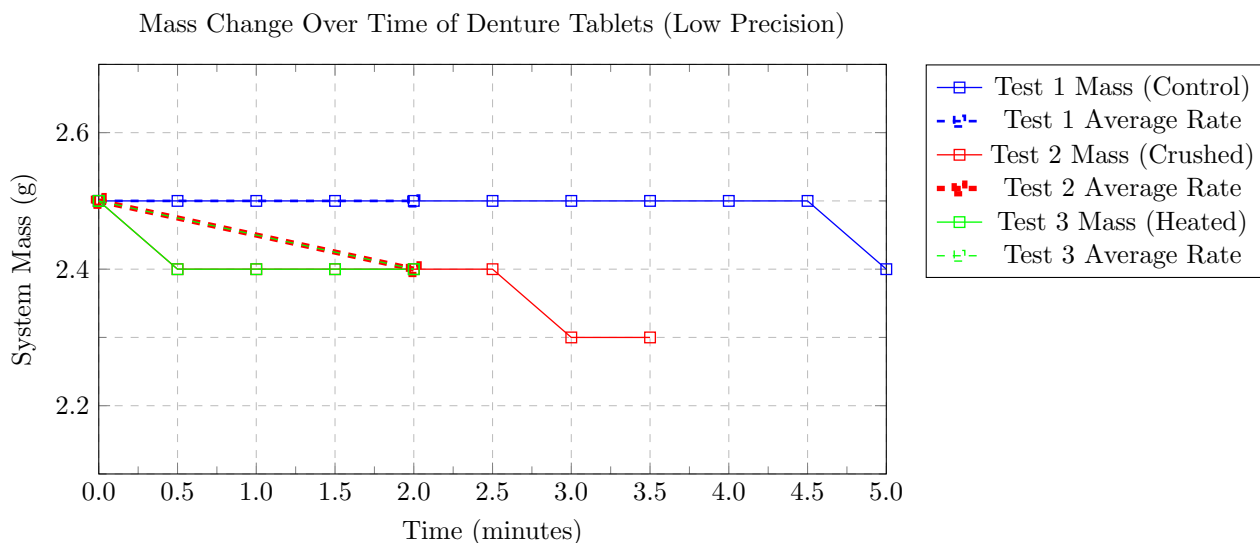


Figure 2: Change in mass of the system with average rates over the first 2 minutes of data.

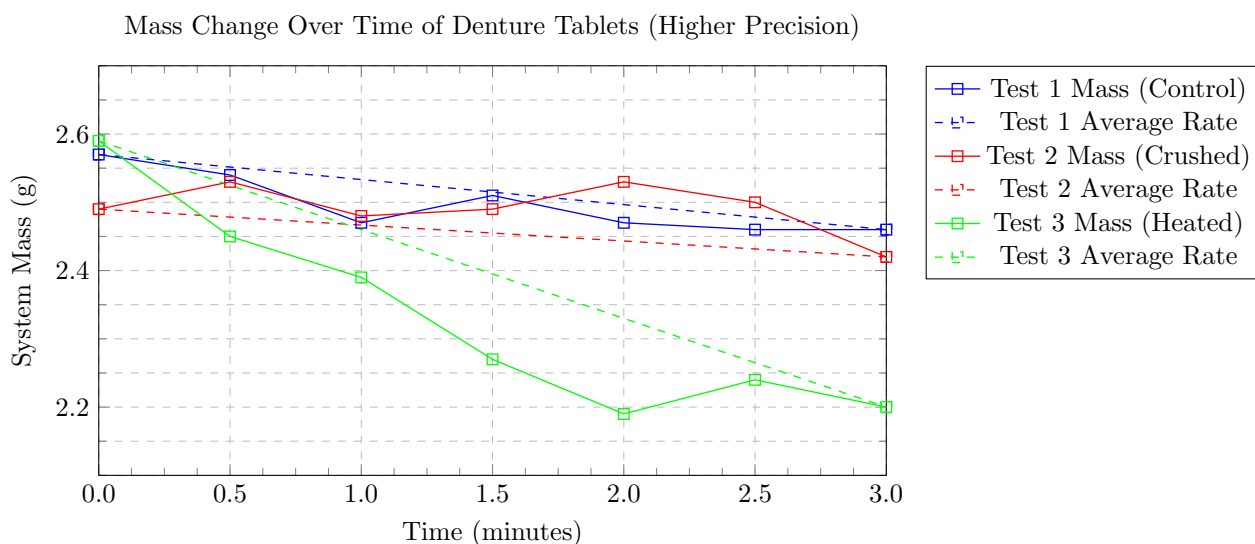


Figure 3: Change in mass of the system with average rates for temperature and surface area test over the first 3 minutes of data with 2 decimal places of precision. (Abigail & Jersey's data.)

As with the first test, in the reaction  $\text{C}_6\text{H}_8\text{O}_7(\text{aq}) + 3\text{NaHCO}_3(\text{s}) \longrightarrow 3\text{H}_2\text{O}(\text{l}) + 3\text{CO}_2(\text{g}) + \text{Na}_3\text{C}_6\text{H}_5\text{O}_7(\text{aq})$ , any reductions in weight can be attributed to  $\text{CO}_2$  escaping the system, being the only gaseous product. The average rate therefore is  $\frac{\text{g CO}_2}{\text{min}}$  and is also represented in the graph by the slope of the dashed lines. We can calculate the average rate over the first two minutes of our data collection and the first three minutes of Abigail & Jersey's data collection to verify this:

$$\text{Average Rate}_{\text{Control}} = \frac{\text{Test 1 Mass}_{2:00} - \text{Test 1 Mass}_{0:00}}{2:00 - 0:00} = \frac{2.5 \text{ g} - 2.5 \text{ g}}{2 \text{ min}} = 0.0 \frac{\text{g CO}_2}{\text{min}}$$

$$\begin{aligned}
\text{Average Rate}_{\text{Crushed}} &= -0.05 \frac{g \text{ CO}_2}{\text{min}} \\
\text{Average Rate}_{\text{Heated}} &= -0.05 \frac{g \text{ CO}_2}{\text{min}} \\
\text{Average Rate}_{\text{Control (AJ)}} &= -0.037 \frac{g \text{ CO}_2}{\text{min}} \\
\text{Average Rate}_{\text{Crushed (AJ)}} &= -0.02 \frac{g \text{ CO}_2}{\text{min}} \\
\text{Average Rate}_{\text{Heated (AJ)}} &= -0.13 \frac{g \text{ CO}_2}{\text{min}}
\end{aligned}$$

$$\begin{aligned}
|\text{Average Rate}_{\text{Heated}}| &> |\text{Average Rate}_{\text{Crushed}}| > |\text{Average Rate}_{\text{Control}}| \\
|\text{Average Rate}_{\text{Heated (AJ)}}| &> |\text{Average Rate}_{\text{Control (AJ)}}| > |\text{Average Rate}_{\text{Crushed (AJ)}}|
\end{aligned}$$

With the data we have, we can conclude that crushing the denture tablets and heating the water both increased the rate of reaction. In Abigail & Jersey's data, the heated test had a much faster rate of reaction than the other two tests. This is further corroborated by our data showing that the heated test fully reacted a full minute before the crushed test. Another interesting observation is that our crushed test had a faster rate of reaction than our control test and fully reacted before the 5 minute time constraint, but in Abigail & Jersey's data, the crushed test had a slightly slower rate of reaction than the control test. This may be due to experimental error ([Section 5.1: Sources of Error](#)).

## 5 Conclusion

In these two experiments, we investigated how the rates of two chemical reactions changed: Calcium carbonate with Hydrochloric acid, and Citric acid with Sodium bicarbonate. We can conclude that increasing concentration, temperature increased the rate of the reactions, while increasing the surface area led to mixed results.

Using the data collected, we could see that the average rate of the first reaction with 3.0 M HCl had a much higher magnitude than with 1.0 M HCl.

$$(|-0.26 \frac{g \text{ CO}_2}{\text{min}}| > |-0.1 \frac{g \text{ CO}_2}{\text{min}}|)$$

The data collected in the second reaction also shows a higher magnitude of average rate when the denture tablets were placed in higher temperature water. The average rate when the denture tablet was crushed diverged in the two results. In our data, the crushed tablet fully reacted faster than the whole tablet, but in Abigail & Jersey's data, the crushed tablet had a lower average rate magnitude than the control test.

$$\begin{aligned}
|\text{Average Rate}_{\text{Heated}}| &> |\text{Average Rate}_{\text{Crushed}}| > |\text{Average Rate}_{\text{Control}}| \\
|\text{Average Rate}_{\text{Heated (AJ)}}| &> |\text{Average Rate}_{\text{Control (AJ)}}| > |\text{Average Rate}_{\text{Crushed (AJ)}}|
\end{aligned}$$

This may be due to the errors discussed in the following section. The increase in rate with increased concentration and temperature can be explained with collision theory. For the first reaction, tripling the concentration of HCl results in more HCl particles in the same volume, increasing the frequency of collisions between HCl and  $\text{CaCO}_3$  particles, thus increasing the rate of reaction. In the second reaction, heating the water increased the kinetic energy of the particles, leading to more particles reaching the activation energy (hitting hard enough), and a faster overall rate. This is coupled with an increase in frequency due to the particles moving more, but this is very minor compared to the former.

In theory, increasing the surface area of a reactant should increase the reaction rate, as the greater area results in more particles being exposed to the other reactant, therefore increasing the frequency of the collisions. The data collected from these experiments, however, is inconclusive on the effect of surface area.

## 5.1 Sources of Error

### 5.1.1 Major Possible Sources of Error

For our reactions, we recorded data using a scale with only one decimal place of precision, which severely limited our ability to analyze the denture tablet reactions. This resulted in one of our calculated average rates being  $0.0 \frac{\text{g CO}_2}{\text{min}}$ , which suggests that the reaction was not happening despite qualitative observations saying otherwise. This also made our calculated average rates for the crushed and heated tests equal, despite one finishing faster than the other.

To supplement our low precision data, we used data collected by Abigail & Jersey, who performed the same experiment but recorded their data with two decimal places of precision. This data was also flawed in that the mass of the system fluctuated throughout the experiment and was observed to increase at a couple points. The system mass in reduction experiments, should not do this, and should instead monotonically decrease.

There are a few possible explanations for this. One is that air currents in the lab, either from movement of people or drafts from windows resulted in various amounts of pressure on the scale. Another major source is that not all of the  $\text{CO}_2$  produced by the reaction escaped the system, due to it being trapped in the large amounts of foam observed in all the tests.

A source of error contributing to actual reaction time could be the mixing of the crushed denture tablet. When poured into the beaker, some portions of the powder rested on top of the foam, preventing reaction as the two compounds needed to be aqueous to react.

### 5.1.2 Minor Possible Sources of Error

Two minor possible sources of error for the weight fluctuations could be miscalibration of the scale, and baseline sensor noise. It is difficult to gauge the effect of these sources of error.

Another minor source of error is inconsistent containers and solution amounts used in the Calcium carbonate reaction. We used a graduated cylinder to measure the reaction rate of the 1.0 M HCl reaction and a beaker for the 3.0 M HCl reaction. This is because we had less 1.0 M HCl solution available, and a thinner container was required to fully cover the  $\text{CaCO}_3$  sample. I expect this to be a negligible source of error, as the  $\text{CaCO}_3$  sample was fully submerged in both cases.

The  $\text{CaCO}_3$  samples not being identical could have also contributed to error, but I expect this to also be negligible when looking at overall reaction rate changes. This could have been mitigated by sanding the rocks samples down to known sizes, or crushing into a powder like the denture tablet test. This was originally planned, but determined it was not necessary to measure overall change and not specific change.

## 5.2 Procedure Reflection

While our procedure was sufficient to investigate the effects of concentration and temperature, conclusive data on surface area was not available. Many of the minor possible sources of error could be mitigated through simple procedural adjustments (consistent equipment, rock crushing, etc.) and the usage of a higher precision scale. For the major sources of error, a draft shield could be placed around the scale to prevent currents, and a proper calibration routine could be done at the start of the experiment to ensure accuracy. Preventing the  $\text{CO}_2$  from being trapped in the foam is more difficult, but could possibly be mitigated through a change in container pressure or the usage of a nonreactive chemical that prevents foam formation.

These changes would allow for more accurate data to be collected and analyzed. Performing the experiment multiple times and averaging the results would also allow for more realistic data to be collected, as outliers would skew the results less.

If more time was given to write the report, the current data could be more thoughtfully analyzed

using linear regressions to provide a trend over the data using every point, instead of an average rate calculated with only the start and end points (only two data points).