Lab Report
Rate of Chemical Reaction: The Iodination of Acetone

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Abstract
The only way to determine the rate of chemical reaction was to conduct and observe data from the experiment. By adjusting the concentration of reactants in each mixture and recording times of chemical reaction, we would be able to find the orders and the rate constant of the reaction, which we then used to calculate the reaction rate. Also, iodine was added to indicate the time of the reaction, as when the color changed.

Purpose
To identify rate of reactions by finding the orders of reaction and rate constant through conducting the experiments. As well as, to compare rates of reaction for each experiments after varying in volume of each reactants.

Introduction
In any kind of reaction, time involves in how fast the reaction takes place. If the reaction occurs in a very short time, it is said to have a fast rate of reaction while it has a slower rate of reaction if it takes longer time. To specify, reaction rate measures how quickly reactants disappear to become products. Concentration of reactants is the fundamental parameter observed. Thus, reaction rate is found by observing the change in concentration of reactants or products over time; the unit is molarity per time (M/t), in which time can be seconds or minutes depending on how fast the reaction occurs. Two equations describe the same reaction rate which are \(- \Delta [\text{Reactants}]/ \Delta t\) for the disappearance of reactants and \(\Delta [\text{Products}]/ \Delta t\) for the formation of products. From figure 1, the changing rate of reactants is negative because the initial concentration reduces as the reaction proceeds. The products, on the other hand, increase throughout the reaction; thus, a sign is positive. Nevertheless, both equations can use to solve for the same reaction rate. Since both of equations are linked and can be expressed as \(- \frac{1}{a} \frac{\Delta [A]}{\Delta t} = - \frac{1}{b} \frac{\Delta [B]}{\Delta t} = \frac{1}{c} \frac{\Delta [C]}{\Delta t} = \frac{1}{d} \frac{\Delta [D]}{\Delta t}\), according to the balanced equation of \(aA+bB \rightarrow cC+dD\).
According to the Rate Law, it is an equation describing the relationship between the rate of reaction, rate constant (k), and the concentration of reactants raised to a certain power.\(^3\) The equation is denoted as \( \text{rate}= k[A]^x[B]^y \) from \( \text{aA+bB} \rightarrow \text{cC+dD} \) where \( x \) is the order of A and \( y \) is the order of B. These orders of reaction are only determined experimentally. They define how rate depends on the power of each concentration.\(^4\) Each specific order can be found from let the fraction of rate\(_1\)/rate\(_2\) equals to the fraction of initial rate\(_1\)/initial rate\(_2\). As soon as the orders are solved, rate constant is determined. To simplify, considered only one reactant, the rate of reaction is calculated using Table 1.\(^5\)

However, rate of the reaction does also depend on addition of catalyst and temperature. Raising the temperature increases kinetic energy which will increase the chance of molecules collision. The products will form faster. Therefore, increasing in temperature raises the rate of reaction. In accordance with UCDAVIS, “If the temperature of a reaction were to reach a certain point where the reactant will begin to degrade, it will decrease the rate of the reaction.”\(^5\) Catalyst is another factor that assists the rate of reaction. Because it reduces the activation of energy which causes the reaction to proceed faster, the reaction rate increases.

<p>| Table 1: equations for solving the reaction rate for each order.(^5) |
|---|---|---|
| <strong>Zero-Order</strong> | <strong>First-Order</strong> | <strong>Second-Order</strong> |
| Rate Law | Rate= k | Rate= k[A] |
|          | Rate= k | Rate= k[A] |
|          | Rate= k[A](^2) | Rate= k[A](^2) |</p>
<table>
<thead>
<tr>
<th><strong>Integrated Rate Law</strong></th>
<th>([A]t = -kt + [A]0)</th>
<th>(\ln[A]t = -kt + \ln[A]0)</th>
<th>(1[A]t = -kt + 1[A]0)</th>
</tr>
</thead>
</table>

**Units of Rate Constant (k):**
- \(\text{mol L}^{-1} \text{s}^{-1}\)
- \(\text{mol L}^{-1} \text{s}^{-1}\)
- \(\text{s}^{-1}\)
- \(\text{Lmol}^{-1} \text{s}^{-1}\)
- \(\text{Lmol}^{-1} \text{s}^{-1}\)

**Linear Plot to Determine (k):**
- \([A][A] \text{ versus time}\)
- \(\ln[A]\ln[A] \text{ versus time}\)
- \(1[A]1[A] \text{ versus time}\)

**Relationship of Rate Constant to the Slope of Straight Line:**
- slope = \(-k\)
- slope = \(-k\)
- slope = \(-k\)  slope = \(-k\)
- slope = \(k\)

**Materials**
- 100 mL Flask
- 10 or 25 mL Cylinder
- Watch clock
- 4M Acetone
- 1M Sulfuric Acid
- 0.005M Iodine in KI

**Method**
1. Fill distilled water in 25-mL flask
2. Draw 5 mL acetone and 5 mL sulfuric acid and 10 mL water and pour into the second 25-mL flask
3. Use dry and clean cylinder to measure 5 mL of \(I_2\) solution. Be careful not to spill the solution on your hands or clothes.
4. Noting the time on the watch clock, pour \(I_2\) solution into the second flask and quickly swirl the flask to thoroughly mix the reagents.
5. Observe the color of the reaction between the first and the second flask. Record time when the color is disappear.
6. Repeat the experiment for the second run (step2-5)  
7. Change the concentration of the reactants according to Table 1 in each mixture, and repeat the experiment (step 2-5)  
8. Record the time of each experiment  
9. Calculate rate of the reaction and the rate constant  

**Table 1: Preparing Each Mixture**

<table>
<thead>
<tr>
<th>Mixture</th>
<th>Acetone (mL)</th>
<th>H₂SO₄ (mL)</th>
<th>H₂O (mL)</th>
<th>I₂ (mL)</th>
<th>Total Volume (mL)</th>
</tr>
</thead>
<tbody>
<tr>
<td>1</td>
<td>5</td>
<td>5</td>
<td>10</td>
<td>5</td>
<td>25</td>
</tr>
<tr>
<td>2</td>
<td>10</td>
<td>5</td>
<td>5</td>
<td>5</td>
<td>25</td>
</tr>
<tr>
<td>3</td>
<td>5</td>
<td>10</td>
<td>5</td>
<td>5</td>
<td>25</td>
</tr>
<tr>
<td>4</td>
<td>5</td>
<td>5</td>
<td>5</td>
<td>10</td>
<td>25</td>
</tr>
</tbody>
</table>

**Result**

**Table 2: Data table reporting time of chemical reaction in each mixture**

<table>
<thead>
<tr>
<th>Mixture</th>
<th>Acetone (mL)</th>
<th>H₂SO₄ (mL)</th>
<th>H₂O (mL)</th>
<th>I₂ (mL)</th>
<th>Time (s)</th>
</tr>
</thead>
<tbody>
<tr>
<td></td>
<td></td>
<td></td>
<td></td>
<td></td>
<td>Run 1</td>
</tr>
<tr>
<td>1</td>
<td>5</td>
<td>5</td>
<td>10</td>
<td>5</td>
<td>63</td>
</tr>
<tr>
<td>2</td>
<td>10</td>
<td>5</td>
<td>5</td>
<td>5</td>
<td>22.8</td>
</tr>
<tr>
<td>3</td>
<td>5</td>
<td>10</td>
<td>5</td>
<td>5</td>
<td>21</td>
</tr>
<tr>
<td>4</td>
<td>5</td>
<td>5</td>
<td>5</td>
<td>10</td>
<td>91.2</td>
</tr>
</tbody>
</table>
Calculation

<table>
<thead>
<tr>
<th>Mixture</th>
<th>Initial Concentration (M)</th>
<th>Rate (r) = \frac{[I_2]}{r}</th>
<th>(r_{1} = 1.58 \times 10^{-5})</th>
</tr>
</thead>
<tbody>
<tr>
<td></td>
<td>Acetone</td>
<td>H(^+)</td>
<td>I(_2)</td>
</tr>
<tr>
<td>1</td>
<td>0.8</td>
<td>0.2</td>
<td>1 \times 10^{-3}</td>
</tr>
<tr>
<td>2</td>
<td>1.6</td>
<td>0.2</td>
<td>1 \times 10^{-3}</td>
</tr>
<tr>
<td>3</td>
<td>0.8</td>
<td>0.4</td>
<td>1 \times 10^{-3}</td>
</tr>
<tr>
<td>4</td>
<td>0.8</td>
<td>0.2</td>
<td>2 \times 10^{-3}</td>
</tr>
</tbody>
</table>

\[ \frac{r_2}{r_1} = \frac{4.42 \times 10^{-3}}{1.58 \times 10^{-5}} = \left( \frac{1.6}{0.8} \right)^m, \quad m = 1 \]

\[ \frac{r_4}{r_1} = \frac{1.09 \times 10^{-3}}{1.58 \times 10^{-5}} = \left( \frac{2 \times 10^{-3}}{1 \times 10^{-3}} \right)^n, \quad n = 0 \]

\[ \frac{r_3}{r_1} = \frac{4.69 \times 10^{-5}}{1.58 \times 10^{-7}} = \left( \frac{0.4}{0.2} \right)^p, \quad p = 1 \]

*** Note: The number of the calculated reaction orders are not an integer, therefore some of them were adjusted and approximated into counting numbers.

Discussion

1. How are time and rate related? How are \(1/\text{time}\) and rate related?

   The rate of reaction is how fast the reaction occurs while time of reaction is how long it takes for the reaction to complete, in which they are inversely proportional (rate = \(\frac{1}{\text{time}}\)). If reaction rate increases, the reaction is completed faster; the reaction time decreases. It also means that rate is equal to \(1/\text{time}\).

2. What does it mean when someone says a reaction is “first order”?

   The first-order reaction is a reaction depending only on the concentration of one compound (reactant) according to the equation of the rate of the first-order reaction, rate = \(k[A]\).

3. In the reaction, A+B ----> C it is found that the reaction is first order in terms of A and B. What happens to the rate of the concentration of A and B are doubled?
According to the reaction $A + B \rightarrow C$, if $A$ and $B$ both are the first order reaction, we will be able to express the relationship of the rate of reaction based on the rate law as rate = $k[A][B]$. Therefore, if the rate of the concentration of $A$ and $B$ are doubled, the rate of chemical reaction will be increased by 4 since the reaction will still have the same rate constant, however, the rate expression will be changed into ‘rate = $k[2*A][2*B]$’ or ‘rate = $4*k[A][B]$’.

4. A second reaction mixture was made up in the following way: 10 mL 4M acetone + 5 mL 1M HCl + 5 mL 0.005M $I_2$ + 5 mL water

   a. What were the initial concentration of acetone, $H^+$ ion and $I_2$ in the reaction mixture?

   b. It took 120 seconds for $I_2$ to disappear from the reaction mixture when it occurred at the same temperature: What was the rate of the reaction?

Write rate of the reaction mixture.

Based on the data from the experiment, we could calculate the initial concentration of acetone, $H^+$ ion and $I_2$ by finding the mole in the mixture of each reactant and divided them by the total volume of the solution which was 0.025 liter. Therefore, we would be able to find the initial concentration of each reactants in the second mixture as Acetone = 1.6 M, $H^+$ ion = 0.2 M, and $I_2 = 1 \times 10^{-3}$ M. However, in order to find the rate of chemical reaction, based on the formula $Rate (r) = \frac{[I_2]_0}{t}$, we also needed the time the reaction had taken, so we could observe the time from the disappearance of Iodine solution due to its brown color. Hence, we applied the formula as $Rate = \frac{1 \times 10^{-3}}{120}$ and found that rate of the reaction is $8.3 \times 10^{-6}$ M/s.

**Conclusion**

By experimenting, we could determine the reaction rate. With different concentration of reactants in each mixture, we recorded times of the reaction when completed—when blue color disappears, and then we calculated the reaction rate using initial concentration of $I_2$ calculated and time recorded.
References


