

TOPIC 3: CHEMICAL KINETICS

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Appendix 3.1: Graphical Determination of Reaction Rate: Lab Activity

Introduction

The rate of a reaction can be calculated by studying the change in the amount of a product or a reactant at different times.

The **average rate of reaction** can be calculated using the following formula:

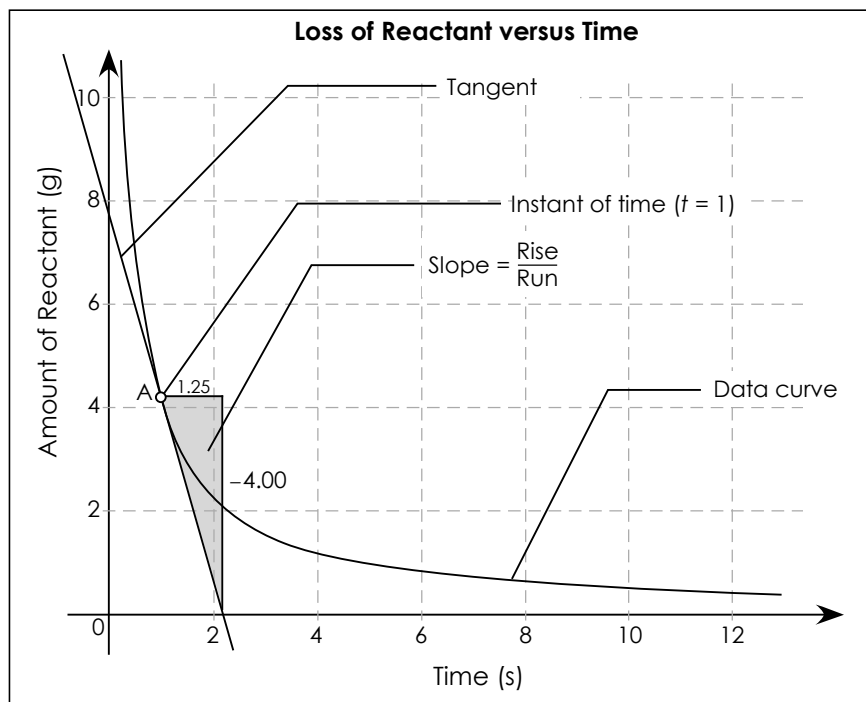
$$\text{Average rate} = \frac{\Delta \text{ amount of substance}}{\Delta \text{ time}}$$

Or

$$\text{Average rate} = \frac{\text{amount of substance at the end} - \text{amount of substance initially}}{\text{final time} - \text{initial time}}$$

This calculation, however, is only the average rate of reaction over a time period. It would be more useful to know the rate of the reaction at a specific time during the reaction. This rate, called the *instantaneous rate of reaction*, can be determined by measuring the amount of change in a product or a reactant at several times during a reaction. Using this data, a graph can be created and the instantaneous rate of reaction can be determined by drawing a tangent to the graph at any time and finding the slope of that tangent.

In reality, calculus is needed to find this slope, but an approximation can be determined by drawing a tangent line and finding the slope (as shown in the figure below). To do this, select two points on the tangent and calculate the slope using rise over run.

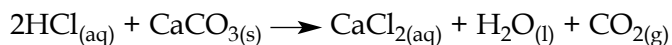


Appendix 3.1: Graphical Determination of Reaction Rate: Lab Activity (continued)

Purpose

In this lab activity, you will measure the loss of mass of a reactant at several times during a chemical reaction. Using the previous graph of the data, you will calculate the average and instantaneous rates of reaction.

The reaction involved is



You will measure the loss of mass in this reaction as the carbon dioxide is released

Caution:

HCl is an acid. Gloves, goggles, and clothing protection must be worn.

Procedure

1. Place 10 to 12 large pieces of CaCO_3 into a paper cup or on filter paper on a scale. Pour 100 mL of 3.0 mol/L HCl solution into a 500 mL beaker. Place the beaker on the scale beside the CaCO_3 . Record the total mass of everything.
2. With a stopwatch ready and the beaker on the scale, the person timing the lab activity should indicate when to pour the CaCO_3 chips into the acid and start the timer. Be sure to put the cup or filter paper back on the scale—it must remain there until the end of the experiment.
3. Record the mass every 30 seconds for 20 minutes.

Questions

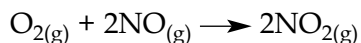
1. The loss in mass in this reaction equals the amount of CO_2 produced. Calculate the mass of CO_2 produced for each 30-second time interval.
2. Calculate the average reaction rate. Using the average rate of reaction formula (provided at the start of this lab activity), determine the average rate of this reaction for the following time intervals:
 - a) First 5 minutes
 - b) First 10 minutes
 - c) Last 5 minutes
 - d) Last 10 minutes
 - e) From 5 to 15 minutes
 - f) For the entire 20 minutes
3. Construct a graph of mass of CO_2 produced versus time.

Appendix 3.1: Graphical Determination of Reaction Rate: Lab Activity *(continued)*

4. Calculate the instantaneous rate of reaction. On your graph, mark the point, draw an approximate tangent line, and calculate the slope of the tangent for the following instants of time:
 - a) 30 seconds
 - b) 60 seconds
 - c) 5 minutes
 - d) 10 minutes
 - e) 15 minutes
 - f) 20 minutes
5. What did you observe in the rate of this reaction from beginning to end? Why does the reaction rate change over time?
6. Explain when it would be useful to know the average rate of reaction and when you would need to know the instantaneous rate of reaction.

Appendix 3.2A: Chemical Kinetics: Assignment 1

A chemist is studying the formation of nitrogen dioxide from nitrogen monoxide and oxygen gas. The balanced equation for the reaction is:



The chemist measured the concentration of the three gases at various time intervals and recorded the data in the table below.

Change in Concentration of Reactants and Products Over Time			
Time (min)	Concentration (mol/L)		
	[O ₂]	[NO]	[NO ₂]
0	0.000343	0.000514	0
2	0.000317	0.000461	0.000053
4	0.000289	0.000406	0.000108
6	0.000271	0.000368	0.000146
10	0.000242	0.000311	0.000204
16	0.000216	0.000259	0.000256
26	0.000189	0.000206	0.000308
41	0.000167	0.000162	0.000353
51	0.000158	0.000143	0.000372
61	0.000150	0.000127	0.000387
71	0.000144	0.000116	0.000399

Questions

- Construct a graph to represent the data provided in the table above. Plot gas concentration along the y -axis and time on the x -axis.

Average rates over a period of time can be calculated by connecting two points on your curve with a straight line and determining the slope.

Instantaneous rates are determined by drawing a tangent line to the curve at the point of interest and determining the slope of the tangent line.

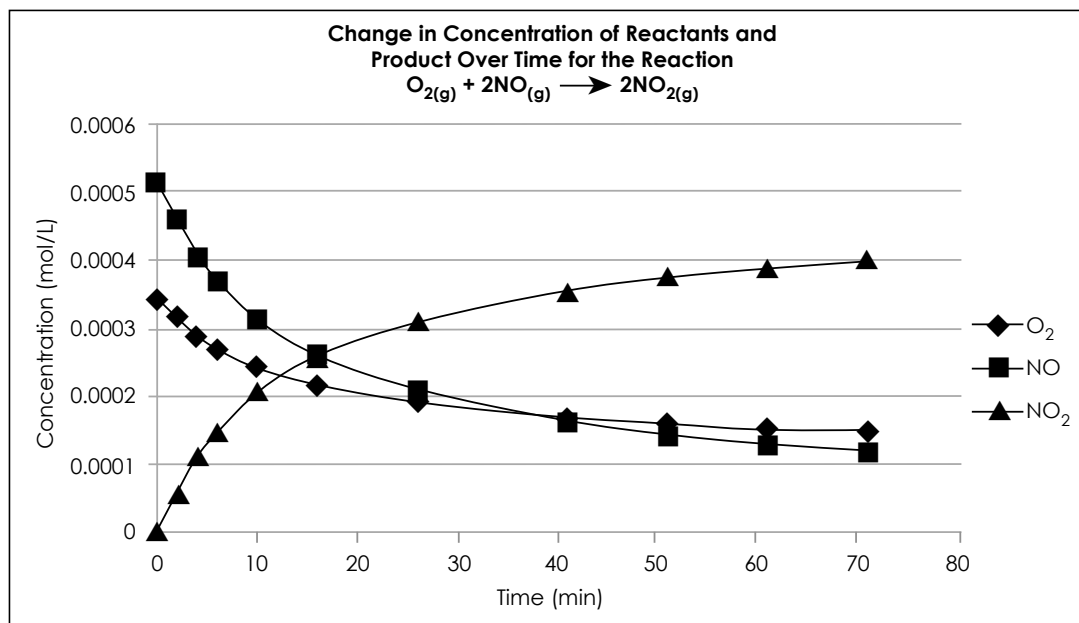
Appendix 3.2A: Chemical Kinetics: Assignment 1 *(continued)*

2. What is the average rate of reaction for nitrogen oxide and oxygen and the formation of nitrogen dioxide over the entire 71-minute interval? Determine the rate for each.
3. What is the average rate of the consumption of NO and O₂ and the production of NO₂ over the first 10 minutes and over the last 10 minutes?
4. Find the instantaneous rate of consumption of O₂ and NO and the instantaneous rate of formation of NO₂ at 4 minutes and at 41 minutes into the experiment. Show your work on the graph. Explain why the rate changes.
5. What do you notice about the ratios of the rates of oxygen and nitrogen monoxide consumption to the production of nitrogen dioxide?

Appendix 3.2B: Chemical Kinetics: Assignment 1 (Answer Key)

Answers to Questions

1. Completed graph:



2. Over the 71-minute interval:

$$\begin{aligned} \text{Average rate for NO} &= \frac{0.000116 \text{ mol/L} - 0.000514 \text{ mol/L}}{71 \text{ min} - 0 \text{ min}} \\ &= 5.61 \times 10^{-6} \text{ mol/L} \cdot \text{min} \end{aligned}$$

$$\begin{aligned} \text{Average rate for O}_2 &= \frac{0.000144 \text{ mol/L} - 0.000343 \text{ mol/L}}{71 \text{ min} - 0 \text{ min}} \\ &= 2.80 \times 10^{-6} \text{ mol/L} \cdot \text{min} \end{aligned}$$

$$\begin{aligned} \text{Average rate for NO}_2 &= \frac{0.000399 \text{ mol/L} - 0 \text{ mol/L}}{71 \text{ min} - 0 \text{ min}} \\ &= 5.61 \times 10^{-6} \text{ mol/L} \cdot \text{min} \end{aligned}$$

Appendix 3.2B: Chemical Kinetics: Assignment 1 (Answer Key) (continued)

3. Over the first 10 minutes:

$$\begin{aligned}\text{Average rate for NO} &= \frac{0.000311 \text{ mol/L} - 0.000514 \text{ mol/L}}{10 \text{ min} - 0 \text{ min}} \\ &= 2.03 \times 10^{-5} \text{ mol/L} \cdot \text{min}\end{aligned}$$

$$\begin{aligned}\text{Average rate for O}_2 &= \frac{0.000242 \text{ mol/L} - 0.000343 \text{ mol/L}}{10 \text{ min} - 0 \text{ min}} \\ &= 1.01 \times 10^{-5} \text{ mol/L} \cdot \text{min}\end{aligned}$$

$$\begin{aligned}\text{Average rate for NO}_2 &= \frac{0.000204 \text{ mol/L} - 0 \text{ mol/L}}{10 \text{ min} - 0 \text{ min}} \\ &= 2.04 \times 10^{-5} \text{ mol/L} \cdot \text{min}\end{aligned}$$

Over the last 10 minutes:

$$\begin{aligned}\text{Average rate for NO} &= \frac{0.000116 \text{ mol/L} - 0.000127 \text{ mol/L}}{71 \text{ min} - 61 \text{ min}} \\ &= 1.10 \times 10^{-6} \text{ mol/L} \cdot \text{min}\end{aligned}$$

$$\begin{aligned}\text{Average rate for O}_2 &= \frac{0.000144 \text{ mol/L} - 0.000150 \text{ mol/L}}{71 \text{ min} - 61 \text{ min}} \\ &= 6.00 \times 10^{-7} \text{ mol/L} \cdot \text{min}\end{aligned}$$

$$\begin{aligned}\text{Average rate for NO}_2 &= \frac{0.000399 \text{ mol/L} - 0.000387 \text{ mol/L}}{71 \text{ min} - 61 \text{ min}} \\ &= 1.20 \times 10^{-6} \text{ mol/L} \cdot \text{min}\end{aligned}$$

4. Answers will vary slightly due to the drawing of the tangent line to the point at 4 minutes.

$$\begin{aligned}\text{Instantaneous rate, NO} &= \frac{0.00028 \text{ mol/L} - 0.000485 \text{ mol/L}}{10 \text{ min} - 0 \text{ min}} \\ &= 2.05 \times 10^{-5} \text{ mol/L} \cdot \text{min}\end{aligned}$$

$$\begin{aligned}\text{Instantaneous rate, O}_2 &= \frac{0.00022 \text{ mol/L} - 0.00033 \text{ mol/L}}{10 \text{ min} - 0 \text{ min}} \\ &= 1.10 \times 10^{-5} \text{ mol/L} \cdot \text{min}\end{aligned}$$

Appendix 3.2B: Chemical Kinetics: Assignment 1 (Answer Key) (continued)

$$\begin{aligned}\text{Instantaneous rate, NO}_2 &= \frac{0.000242 \text{ mol/L} - 0.00003 \text{ mol/L}}{10 \text{ min} - 0 \text{ min}} \\ &= 2.12 \times 10^{-5} \text{ mol/L} \cdot \text{min}\end{aligned}$$

Answers will vary slightly due to the drawing of the tangent line to the point at 41 minutes.

$$\begin{aligned}\text{Instantaneous rate, NO} &= \frac{0.000125 \text{ mol/L} - 0.00018 \text{ mol/L}}{60 \text{ min} - 30 \text{ min}} \\ &= 1.83 \times 10^{-6} \text{ mol/L} \cdot \text{min}\end{aligned}$$

$$\begin{aligned}\text{Instantaneous rate, O}_2 &= \frac{0.00014 \text{ mol/L} - 0.00019 \text{ mol/L}}{70 \text{ min} - 20 \text{ min}} \\ &= 1.00 \times 10^{-6} \text{ mol/L} \cdot \text{min}\end{aligned}$$

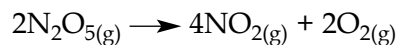
$$\begin{aligned}\text{Instantaneous rate, NO}_2 &= \frac{0.000365 \text{ mol/L} - 0.000325 \text{ mol/L}}{50 \text{ min} - 30 \text{ min}} \\ &= 2.00 \times 10^{-6} \text{ mol/L} \cdot \text{min}\end{aligned}$$

5. The ratio between O₂ and NO₂ is 1:2. The rate of consumption of O₂ is one-half the rate of formation of NO₂.

The ratio between NO and NO₂ is 2:2 or 1:1. The rate of consumption of NO is equal to the rate of formation of NO₂.

Appendix 3.3A: Chemical Kinetics: Assignment 2

A chemist is studying the decomposition of dinitrogen pentoxide at 45°C. The balanced equation is:



The chemist measured the concentration of dinitrogen pentoxide at 10-minute intervals for 100 minutes, using colorimetry (spectrophotometry), and recorded the data in the table below.

Decomposition of Dinitrogen Pentoxide at 45°C			
Time (min)	Concentration (mol/L)		
	[N ₂ O ₅]	[NO ₂]	[O ₂]
0	0.0124	0	0
10	0.0092		
20	0.0068		
30	0.0050		
40	0.0037		
50	0.0027		
60	0.0020		
70	0.0014		
80	0.0011		
90	0.0008		
100	0.0006		

Appendix 3.3A: Chemical Kinetics: Assignment 2 (continued)

Questions

1. Use the data from the table above and the balanced equation to calculate the concentration of nitrogen dioxide and oxygen gas at each interval. Then construct a graph to represent this data. Plot gas concentration along the y -axis and time on the x -axis.

Average rates over a period of time can be calculated by connecting two points on your curve with a straight line and determining the slope.

Instantaneous rates are determined by drawing a tangent line to the curve at the point of interest and determining the slope of the tangent line.

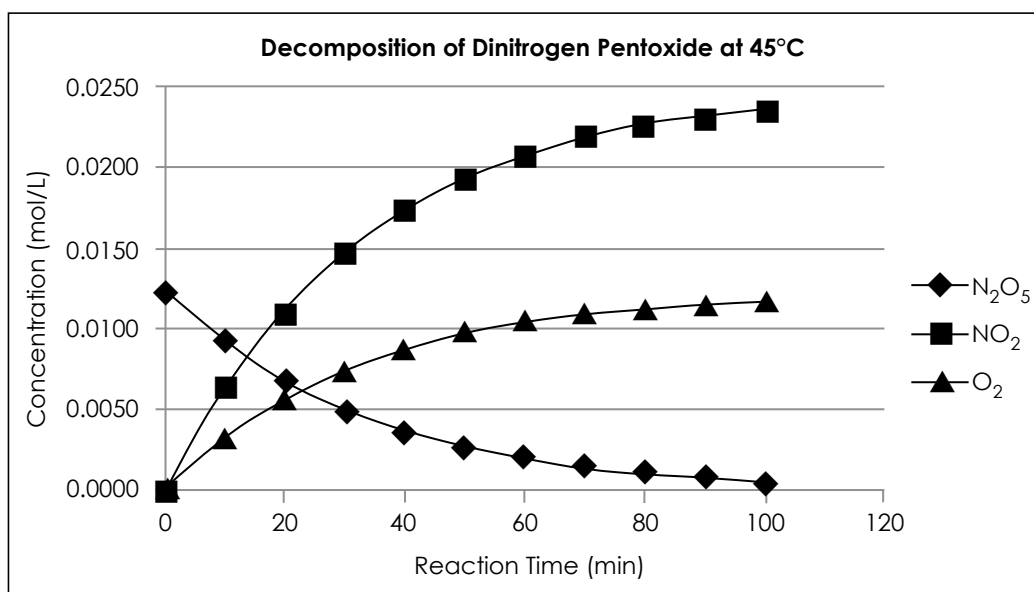
2. What is the average rate of decomposition of dinitrogen pentoxide and the formation of nitrogen dioxide and oxygen over the entire 100-minute interval? Determine the rate for each.
3. What is the average rate for the first 20 minutes of the decomposition of N_2O_5 and for the last 20 minutes?
4. Find the instantaneous rate of decomposition of N_2O_5 and the instantaneous rate of formation of NO_2 and O_2 at 10 minutes and at 80 minutes into the experiment. Show your work on the graph. Explain why the rate changes.
5. Explain why the rate changes between 10 and 80 minutes.

Appendix 3.3B: Chemical Kinetics: Assignment 2 (Answer Key)

Answers to Questions

1. Completed chart and graph:

Decomposition of Dinitrogen Pentoxide at 45°C			
Time (min)	Concentration (mol/L)		
	[N ₂ O ₅]	[NO ₂]	[O ₂]
0	0.0124	0	0
10	0.0092	0.0064	0.0032
20	0.0068	0.0112	0.0056
30	0.0050	0.0148	0.0074
40	0.0037	0.0174	0.0087
50	0.0027	0.0194	0.0097
60	0.0020	0.0208	0.0104
70	0.0014	0.0220	0.0110
80	0.0011	0.0226	0.0113
90	0.0008	0.0232	0.0116
100	0.0006	0.0236	0.0118



Appendix 3.3B: Chemical Kinetics: Assignment 2 (Answer Key) (continued)

2. Over the 100-minute interval:

$$\begin{aligned}\text{Average rate for N}_2\text{O}_5 &= \frac{0.0006 \text{ mol/L} - 0.0124 \text{ mol/L}}{100 \text{ min} - 0 \text{ min}} \\ &= 1.18 \times 10^{-4} \text{ mol/L} \cdot \text{min}\end{aligned}$$

$$\begin{aligned}\text{Average rate for NO}_2 &= \frac{0.0236 \text{ mol/L} - 0 \text{ mol/L}}{100 \text{ min} - 0 \text{ min}} \\ &= 2.36 \times 10^{-4} \text{ mol/L} \cdot \text{min}\end{aligned}$$

$$\begin{aligned}\text{Average rate for O}_2 &= \frac{0.00590 \text{ mol/L} - 0 \text{ mol/L}}{100 \text{ min} - 0 \text{ min}} \\ &= 5.90 \times 10^{-5} \text{ mol/L} \cdot \text{min}\end{aligned}$$

3. Over the first 20 minutes:

$$\begin{aligned}\text{Average rate for N}_2\text{O}_5 &= \frac{0.0068 \text{ mol/L} - 0.0124 \text{ mol/L}}{20 \text{ min.} - 0 \text{ min}} \\ &= 2.80 \times 10^{-4} \text{ mol/L} \cdot \text{min}\end{aligned}$$

Over the last 20 minutes:

$$\begin{aligned}\text{Average rate for N}_2\text{O}_5 &= \frac{0.0006 \text{ mol/L} - 0.0011 \text{ mol/L}}{100 \text{ min} - 80 \text{ min}} \\ &= 2.50 \times 10^{-5} \text{ mol/L} \cdot \text{min}\end{aligned}$$

4. Answers will vary slightly due to the drawing of the tangent line to the point at 10 minutes.

$$\begin{aligned}\text{Instantaneous rate for N}_2\text{O}_5 &= \frac{0.00370 \text{ mol/L} - 0.0118 \text{ mol/L}}{30 \text{ min} - 0 \text{ min}} \\ &= 2.70 \times 10^{-4} \text{ mol/L} \cdot \text{min}\end{aligned}$$

$$\begin{aligned}\text{Instantaneous rate for NO}_2 &= \frac{0.0165 \text{ mol/L} - 0.0012 \text{ mol/L}}{30 \text{ min} - 0 \text{ min}} \\ &= 5.10 \times 10^{-4} \text{ mol/L} \cdot \text{min}\end{aligned}$$

Appendix 3.3B: Chemical Kinetics: Assignment 2 (Answer Key) (continued)

$$\begin{aligned}\text{Instantaneous rate for O}_2 &= \frac{0.0025 \text{ mol/L} - 0.0004 \text{ mol/L}}{20 \text{ min} - 0 \text{ min}} \\ &= 1.05 \times 10^{-4} \text{ mol/L} \cdot \text{min}\end{aligned}$$

Answers will vary slightly due to the drawing of the tangent line to the point at 80 minutes.

$$\begin{aligned}\text{Instantaneous rate for N}_2\text{O}_5 &= \frac{0 \text{ mol/L} - 0.002 \text{ mol/L}}{115 \text{ min} - 50 \text{ min}} \\ &= 3.08 \times 10^{-5} \text{ mol/L} \cdot \text{min}\end{aligned}$$

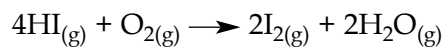
$$\begin{aligned}\text{Instantaneous rate for NO}_2 &= \frac{0.0236 \text{ mol/L} - 0.02 \text{ mol/L}}{100 \text{ min} - 40 \text{ min}} \\ &= 6.00 \times 10^{-5} \text{ mol/L} \cdot \text{min}\end{aligned}$$

$$\begin{aligned}\text{Instantaneous rate for O}_2 &= \frac{0.00590 \text{ mol/L} - 0.00520 \text{ mol/L}}{100 \text{ min} - 60 \text{ min}} \\ &= 1.75 \times 10^{-5} \text{ mol/L} \cdot \text{min}\end{aligned}$$

5. There are fewer reactant particles available over time, so the rate gets slower as the reaction proceeds.

Appendix 3.4A: Chemical Kinetics Problems *(continued)*

9. Hydrogen iodide and oxygen react to form iodine gas and water vapour. If oxygen gas reacts at a rate of $0.0042 \text{ mol/L}\cdot\text{s}$,

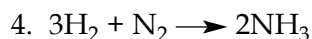


- What is the rate of formation of iodine gas in $\text{mol/L}\cdot\text{s}$?
- What is the rate of formation of water vapour in $\text{mol/L}\cdot\text{s}$?
- What is the rate of consumption of hydrogen iodide gas in $\text{mol/L}\cdot\text{s}$?

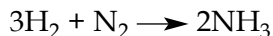
Appendix 3.4B: Chemical Kinetics Problems (Answer Key)

- Three examples of properties that could be used to measure reaction rate are
 - reactions that produce a gas (measure volume/pressure)
 - reactions that involve the ion as a product (conductivity)
 - reactions that produce a colour change (spectrometer – measure colour intensity)
- Observing the colour change would be the best indicator of reaction rate. When the permanganate ion (MnO_4^-) disappears, the pink-purple colour of the solution will disappear. You may see a red-brown colour appear as the Fe^{3+} forms.
 - Because hydrogen gas is produced, you could collect the gas and measure the volume produced as time passes. Or, if the container is sealed, pressure increase could be measured.

3. $\text{mol/L}\cdot\text{s}$

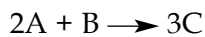


The rate of disappearance of H_2 is 3 times as fast as the rate of disappearance of N_2 .



The rate of production of NH_3 is 2 times as fast as the rate of disappearance of N_2 .

5. The rate of consumption of A is

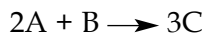


This is twice ($2\times$) the rate of consumption of B ($0.30 \text{ mol/L}\cdot\text{s}$)

$$= 2 \times 0.30 \text{ mol/L}\cdot\text{s}$$

$$= 0.60 \text{ mol/L}\cdot\text{s}$$

The rate of formation of C is

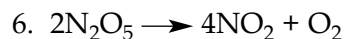


Appendix 3.4B: Chemical Kinetics Problems (Answer Key) (continued)

This is three times (3×) the rate of consumption of B(0.30 mol/L·s)

$$= 3 \times 0.30 \text{ mol/L}\cdot\text{s}$$

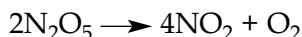
$$= 0.90 \text{ mol/L}\cdot\text{s}$$



2:4 ratio, which simplifies to a 1:2 ratio.

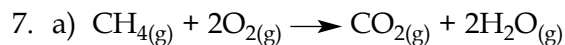
The rate of formation of NO_2 is 2 times as fast as the rate of disappearance of N_2O_5 .

The rate of formation of $\text{NO}_2 = 2 \times 2.5 \times 10^{-6} \text{ mol/L}\cdot\text{s} = 5.0 \times 10^{-6} \text{ mol/L}\cdot\text{s}$.

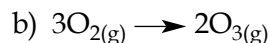


The rate of formation of O_2 is half as fast as the rate of disappearance of N_2O_5 .

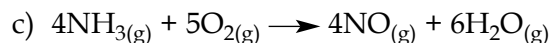
The rate of formation of $\text{O}_2 = 1/2 \times 2.5 \times 10^{-6} \text{ mol/L}\cdot\text{s} = 1.25 \times 10^{-6} \text{ mol/L}\cdot\text{s}$.



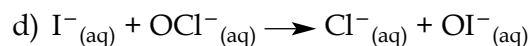
$$\text{Rate} = -\frac{\Delta[\text{CH}_4]}{\Delta t} = -\frac{1}{2} \frac{\Delta[\text{O}_2]}{\Delta t} = \frac{\Delta[\text{CO}_2]}{\Delta t} = \frac{1}{2} \frac{\Delta[\text{H}_2\text{O}]}{\Delta t}$$



$$\text{Rate} = \frac{1}{3} \frac{\Delta[\text{O}_2]}{\Delta t} = \frac{1}{2} \frac{\Delta[\text{O}_3]}{\Delta t}$$



$$\text{Rate} = -\frac{1}{4} \frac{\Delta[\text{NH}_3]}{\Delta t} = -\frac{1}{5} \frac{\Delta[\text{O}_2]}{\Delta t} = \frac{1}{4} \frac{\Delta[\text{NO}]}{\Delta t} = \frac{1}{6} \frac{\Delta[\text{H}_2\text{O}]}{\Delta t}$$

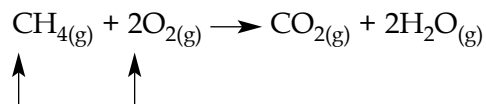


$$\text{Rate} = -\frac{\Delta[\text{I}^-]}{\Delta t} = -\frac{\Delta[\text{OCl}^-]}{\Delta t} = \frac{\Delta[\text{Cl}^-]}{\Delta t} = \frac{\Delta[\text{OI}^-]}{\Delta t}$$

Appendix 3.4B: Chemical Kinetics Problems (Answer Key) (continued)

8. Rate of consumption of CH_4 = concentration/time = $4.0 \text{ mol/L} / 3.2 \text{ s} = 1.25 \text{ mol/L}\cdot\text{s}$

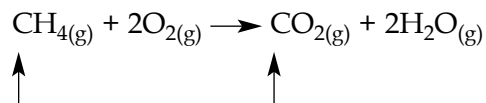
- a) Rate of consumption of O_2



The rate of consumption of O_2 is 2 times as fast as the rate of consumption of CH_4 .

$$2 \times 1.25 \text{ mol/L}\cdot\text{s} = 2.50 \text{ mol/L}\cdot\text{s}$$

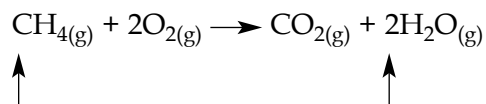
- b) Rate of production of CO_2



The rate of production of CO_2 is the same as the rate of consumption of CH_4 .

$$1 \times 1.25 \text{ mol/L}\cdot\text{s} = 1.25 \text{ mol/L}\cdot\text{s}$$

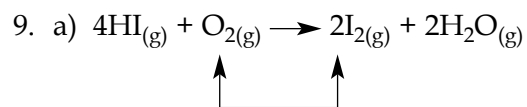
- c) Rate of production of H_2O



The rate of production of H_2O is 2 times as fast as the rate of consumption of CH_4 .

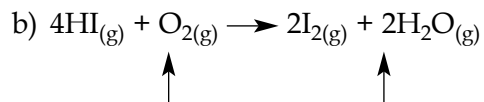
$$2 \times 1.25 \text{ mol/L}\cdot\text{s} = 2.50 \text{ mol/L}\cdot\text{s}$$

Appendix 3.4B: Chemical Kinetics Problems (Answer Key) (continued)



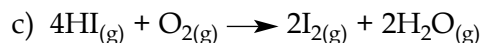
The rate of formation of I_2 is 2 times as fast as the rate of consumption of O_2 .

$$2 \times 0.0042 \text{ mol/L}\cdot\text{s} = 0.0084 \text{ mol/L}\cdot\text{s}$$



The rate of formation of H_2O is 2 times as fast as the rate of consumption of O_2 .

$$2 \times 0.0042 \text{ mol/L}\cdot\text{s} = 0.0084 \text{ mol/L}\cdot\text{s}$$



The rate of consumption of HI is 4 times as fast as the rate of consumption of O_2 .

$$4 \times 0.0042 \text{ mol/L}\cdot\text{s} = 0.0168 \text{ mol/L}\cdot\text{s}$$

Appendix 3.5A: Factors Affecting the Rate of Reactions: Lab Activity

Teacher Notes

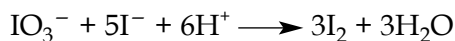
This is a well-known lab activity sometimes called the Iodine Clock Reaction or the Harcourt-Esson Reaction. The sudden, dramatic change in colour from a clear or cloudy, white solution to a dark, blue-black solution helps indicate the end of the reaction. By varying concentration and temperature, students can measure the time required for the colour change to occur, and thus determine the effects of these factors on the rate of the reaction.

This lab activity uses a three-step reaction that produces iodine. The iodine will then form a dark blue complex with starch.

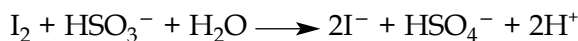
The first and rate-determining step between iodate and metabisulphite ions generates iodide ions:



The excess iodate oxidizes the iodide to form iodine:



However, the iodine is reduced immediately by the metabisulphite ions back to iodide:



Only when the metabisulphite is fully consumed will the elemental iodine remain and react with the starch (not shown in these reactions). The solution then quickly becomes blue.

Pre-lab Preparation

Two solutions need to be prepared for this lab activity:

- **Solution A:** Saturated potassium iodate (KIO_3)
 - You will need approximately 80 mL per lab group.
 - This solution can be prepared well in advance.

Caution:

- Print out and review with students the Material Safety Data Sheets (MSDS) for potassium iodate, sodium metabisulphite, and sulphuric acid.
- All lab participants must wear personal safety equipment for protection of eyes, hands, and clothes.
- Spills must be properly cleaned immediately.
- Review safety procedures for using a hot plate and handling hot solutions.

Appendix 3.5A: Factors Affecting the Rate of Reactions: Lab Activity *(continued)*

- **Solution B:** Sodium metabisulphite ($\text{Na}_2\text{S}_2\text{O}_5$), sulphuric acid (H_2SO_4), and soluble starch
 - You will need the following per litre of water:
 - 25.0 g $\text{Na}_2\text{S}_2\text{O}_5$
 - 5.0 mL concentrated H_2SO_4
 - 40 g soluble starch
 - You will need approximately 100 mL per lab group.
 - This solution does not store well and should be mixed no earlier than the day before the lab activity.
 - To dissolve starch, start by mixing it thoroughly in a few mL of distilled water, and then pour it into boiling water (less than one litre). Once dissolved, let the water cool, add the H_2SO_4 , add the $\text{Na}_2\text{S}_2\text{O}_5$, and stir until dissolved and bring up to 1.0 L.
 - Test the speed of the reaction beforehand. If the reaction happens too quickly to measure the time accurately, dilute the KIO_3 stock solution.

Materials (per group)

two 100 mL beakers
250 mL beaker
two 10 mL graduated cylinders
two test tubes (18 × 250 mm)
temperature probe
timer/stopwatch
personal safety equipment
hot plate
ice
distilled water
paper towel
Solution A (10 mL each trial)
Solution B (10 mL each trial)

Appendix 3.5A: Factors Affecting the Rate of Reactions: Lab Activity (continued)

Procedure**Part A: Concentration of Reactants and Rate of Reaction**

1. Label or mark one graduated cylinder as **cylinder A** and one beaker as **beaker A**. Measure out exactly 10.0 mL of Solution A into this graduated cylinder and pour it into a 100 mL beaker.
2. Measure out exactly 10.0 mL of Solution B using a second graduated cylinder and pour it into another 100 mL beaker.
3. With a stopwatch ready, have the person timing the reaction give the signal to mix the solutions and start timing. Quickly pour Solution A into Solution B, and immediately swirl the solutions several times. Put the beaker on the paper towel (for contrast). When a colour change occurs, stop the watch and record the time.
4. Rinse and dry out the beakers.

Repeat steps 1 to 4, changing the concentration of Solution A by mixing the following amounts in beaker A for each trial.

Trial	mL of Solution A	mL of Distilled Water
2	9.0	1.0
3	8.0	2.0
4	7.0	3.0
5	6.0	4.0
6	5.0	5.0

Optional: Obtain data from other groups for their trials.

Part B: Temperature and Reaction Rate

1. Prepare an ice-water bath by filling a 250 mL beaker half way with cold water and ice.
2. Measure 10.0 mL of Solution A into a test tube labelled **test tube A**, and 10.0 mL of Solution B in another test tube.
3. Place the test tubes into the ice bath and leave them there until the temperature of the test tubes is equal to the ice bath. Record the temperature.
4. With a stopwatch ready, have the person timing the reaction give the signal to mix the solutions and start timing. Quickly pour Solution A into Solution B, and immediately swirl the test tube several times. Return the mixture immediately to the ice bath.
5. Stop the time when the colour change occurs. Record the time.
6. Rinse and dry the test tubes.

Appendix 3.5A: Factors Affecting the Rate of Reactions: Lab Activity *(continued)*

Repeat steps 2 to 5 with a couple of warm tap-water baths (at different temperatures) and then with an 80°C to 90°C bath using a hot plate.

Share your data with other students and obtain data from other groups to obtain a range of temperatures.

Analysis

1. If you have shared data from Part A with other students, calculate the average reaction rate for each trial.
2. Plot the data from the two trials (Part A and Part B) in two different graphs.

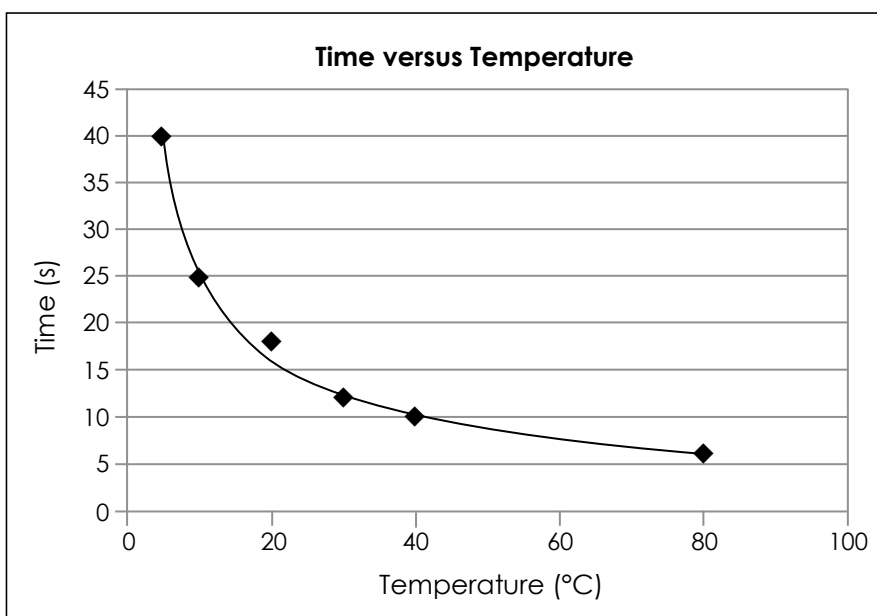
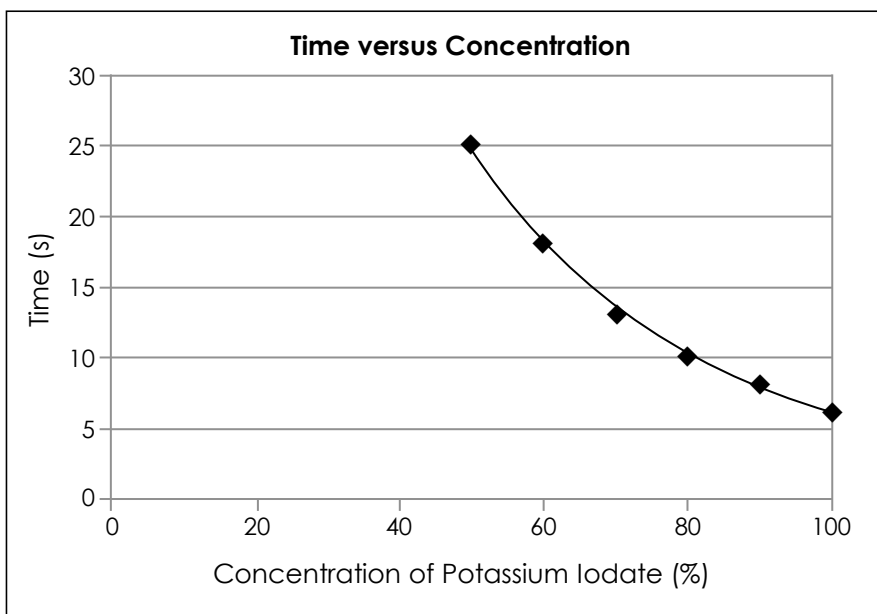
Conclusions

1. Make general statements about the effects of concentration and temperature on the rate of reaction.
2. Use the collision theory to explain your observations.

Appendix 3.5B: Factors Affecting the Rate of Reactions: Lab Activity (Answer Key)

Analysis

The graphs should look similar to the following:



Appendix 3.5B: Factors Affecting the Rate of Reactions: Lab Activity (Answer Key)
(continued)

Conclusions

1. From the graph of **Time versus Concentration**, one can conclude that there is an inverse relationship between the time it takes for a reaction to finish and the concentration of one of the reactants. This relationship is not linear – in other words, the greater the concentration, the faster the reaction.

From the graph of **Time versus Temperature**, one can conclude that there is an inverse, non-linear relationship. As the temperature increases, the speed of the reaction also increases.

2. According to the collision theory, the rate of reaction should increase (speed up) when the concentration of a reactant increases because there are more particles of the reactant present, increasing the chances of a collision with enough energy to start a reaction.

When heated, particles increase in kinetic energy, meaning that there is a higher likelihood of a collision between particles that have enough energy for the reaction to occur. Therefore, reactions proceed faster when the temperature is increased.

Appendix 3.6A: Factors Affecting the Rate of a Reaction: Lab Activity

Part A: Nature of the Reactants

1. Add 20 drops of 3.0 mol/L hydrochloric acid solution to each of five wells of a 24-well test plate.
2. Place a small piece of magnesium in the first well, a small piece of aluminum in the second, a small piece of zinc in the third, a piece of iron in the fourth, and a piece of copper in the fifth.
3. Observe and record your observations.

Questions

- What gas is produced? How do you know?
 - Write a balanced equation to represent the reaction.
 - Do all the metals take the same time to react?
 - Rank the metals in order of reactivity.
4. Add 13 drops of water and 7 drops of 3.0 mol/L hydrochloric acid solution to one well of a 24-well test plate. Stir with a glass capillary tube (sealed at one end) to mix the solution.
 5. Add 13 drops of water and 7 drops of 3.0 mol/L acetic acid solution to a second well of a 24-well test plate. Stir with a glass capillary tube to mix the solution.
 6. Add 20 drops of 1.0 mol/L aqueous zinc(II) nitrate solution to a third well, 20 drops of 1.0 mol/L iron(III) nitrate solution to a fourth well, and 20 drops of 1.0 mol/L copper(II) nitrate solution to a fifth well of the 24-well test plate.
 7. Place a small piece of magnesium in each of the five solutions.
 8. Observe and record your observations.

Questions

- What happened in each well? Identify the products in each case.
- Write a balanced equation to represent each reaction.
- How much time does the magnesium take to react in each solution?

Appendix 3.6A: Factors Affecting the Rate of a Reaction: Lab Activity *(continued)*

Part B: Surface Area (degree of subdivision of a solid)

1. Add 30 drops of 3.0 mol/L hydrochloric acid solution to each of four wells of a 24-well test plate.
2. To the first well, add a piece (a marble chip is suitable) of calcium carbonate (CaCO_3). To the second well, add a similar amount of finely ground (powdered) calcium carbonate.
3. To the third well, add a piece of “mossy” zinc. To the fourth well, add a similar amount of finely divided zinc (20-mesh) or powdered zinc.
4. Observe and record your observations.

Questions

- What happened in each well? Identify the products in each case.
- Write a balanced equation to represent each reaction.
- How much time do the solids take to react in each solution?

Part C: Temperature

1. Prepare a hot water bath by heating about 150 mL of water in a 250 mL beaker to boiling. Set aside.
2. Add 2 mL of 0.01 mol/L aqueous potassium permanganate (KMnO_4) solution (made acidic with sulphuric acid) to each of two 13 × 100 mm test tubes.
3. Place one of the test tubes of potassium permanganate solution into the hot water bath. While it is coming up to temperature, proceed to the next step.
4. Add 5 mL of 0.02 mol/L oxalic acid solution to the second test tube (at room temperature). Stir with a stirring rod.
5. Add 5 mL of 0.02 mol/L oxalic acid solution to the test tube in the hot water bath. Stir.
6. Reheat the water in your water bath to boiling, and set it aside again. Prepare a cold water bath by adding ice cubes to 50 mL of water in a 250 mL beaker.
7. Add 3 mL of water and 1 mL of 3.0 mol/L hydrochloric acid solution to each of three 13 × 100 mm test tubes. Place one of the test tubes in the hot water bath, place one in the cold water bath, and leave one at room temperature. Wait about 2 minutes for the solutions to come to temperature.

Appendix 3.6A: Factors Affecting the Rate of a Reaction: Lab Activity *(continued)*

- Cut three 0.5 cm long pieces of magnesium ribbon. Add one piece to each of the three test tubes. Observe the time required for each piece to disappear completely.

Question

- Does the reaction take the same time at each temperature? Explain.

Part D: Catalyst

- Add 2 mL of 0.01 mol/L aqueous potassium permanganate (KMnO_4) solution (made acidic with sulphuric acid) to each of two 13×100 mm test tubes.
- To one of the test tubes, add 5 drops of 0.01 mol/L manganese(II) sulphate solution.
- Add 5 mL of 0.02 mol/L oxalic acid solution to each of the test tubes, stopper the tubes, and shake.
- Observe and record your observations.

Question

- Does the reaction take the same time in each test tube? Explain.

Appendix 3.6B: Factors Affecting the Rate of a Reaction: Lab Activity (Teacher Notes)

The four experiments demonstrate factors affecting the rate of a reaction, including

- nature (identity) of reactants
- surface area (degree of subdivision of a solid for heterogeneous reactants)
- temperature
- catalyst

Solutions

- 3.0 mol/L hydrochloric acid solution. Dilute 258 mL of concentrated (11.6 mol/L or 36%) hydrochloric acid to 1.0 L with distilled water.
- 1.0 mol/L zinc(II) nitrate solution. Dissolve 29.7 g of zinc(II) nitrate hexahydrate, $\text{Zn}(\text{NO}_3)_2 \cdot 6\text{H}_2\text{O}$, in 100 mL water.
- 1.0 mol/L iron(III) nitrate solution. Dissolve 4.04 g of iron(III) nitrate nonahydrate, $\text{Fe}(\text{NO}_3)_3 \cdot 9\text{H}_2\text{O}$, in 100 mL water.
- 1.0 mol/L copper(II) nitrate solution. Dissolve 29.6 g of copper(II) nitrate hexahydrate, $\text{Cu}(\text{NO}_3)_2 \cdot 6\text{H}_2\text{O}$, in 100 mL water.
- 0.01 mol/L potassium permanganate. Add 1 mL of concentrated (17.8 mol/L or 95%) sulphuric acid to 75 mL water and add enough water to take the total volume up to 100 mL. Dissolve 1.58 g of potassium permanganate (KMnO_4) in this solution.
- 0.02 mol/L oxalic acid solution. Dissolve 2.52 g of oxalic acid (ethanedioic acid, $\text{HOCCOOH} \cdot 2\text{H}_2\text{O}$) in 100 mL water.
- 0.01 mol/L manganese(II) sulphate solution. Dissolve 2.23 g of manganese(II) sulphate tetrahydrate (manganous sulphate, $\text{MnSO}_4 \cdot 4\text{H}_2\text{O}$) in 100 mL water.

Notes

1. Sulphate compounds can be used instead of nitrate compounds. Be sure to adjust the masses used for the difference in molar mass.
2. Students may be more successful if the oxide layer is removed from the magnesium ribbon by rubbing the ribbon's surface gently with emery paper before the magnesium samples are distributed.

