APPENDICES

TOPIC 3: CHEMICAL KINETICS

Appendix 1: Graphical Determination of the Rate of a Chemical Reaction

The numerical value for the rate of reaction can be determined by studying the change in quantity of a substance at different times. The substance being studied can be either product or reactant. The average rate of reaction can be determined by:

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

or

$$\text{average rate} = \frac{\text{final quantity} - \text{initial quantity}}{\text{final time} - \text{initial time}}$$

The instantaneous rate of a reaction at any time, t, can be found by drawing the tangent to the curve at time = t and then determining the slope of the tangent line.

The reaction you will be studying is

$$2\ HCl(aq) + CaCO_3(s) \rightarrow CaCl_2(aq) + H_2O(l) + CO_2(g)$$

You will be observing the loss in mass of the system as the carbon dioxide produced escapes into the atmosphere from an open container.

Procedure:

1. Place 100 mL of a 3 mol / L HCl solution into a 600 mL beaker. Find the mass of the beaker with the acid and 10 large $CaCO_3$ crystals. Do not add the $CaCO_3$ to the acid at this point.

2. Leave the beaker on the balance and add the $CaCO_3$ to the acid solution. Record the mass of the beaker with acid and $CaCO_3$ at 30 second intervals for 20 minutes.

Questions:

1. Determine the mass of CO_2 produced at each interval. Note: the mass of CO_2 is equal to the mass loss for that interval.

   Mass CO_2 (t = 30 s) = Initial mass – mass (t = 30 s)
   Mass CO_2 (t = 5 min) = Initial mass – mass (t = 5 min)

2. Determine the average rate for the following intervals.

   a. Entire 20 minutes
   b. First 5 minutes
   c. 5 minutes to 15 minutes
   d. Last 5 minutes

3. Construct a graph of mass of CO_2 produced versus reaction time.

4. Use the tangent method to determine the instantaneous rate at 1 minute, 5 minutes, 15 minutes and 20 minutes.

5. Explain why the rate changes as it does over time.

Appendix 2: Kinetics Assignment #1

The formation of nitrogen dioxide from nitrogen dioxide and oxygen gas was studied. The balanced equation for the reaction is:

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

The chemist measured the concentration of the three gases at various time intervals. The data is in the table below. Construct a graph to represent this data. Plot gas concentration along the y-axis and time on the x-axis.

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

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

Questions

1. 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.

2. What is the average rate of the consumption of NO and O_2 and the production of NO_2 over the first 10 minutes, and over the last 10 minutes?

3. Find the instantaneous rate of consumption of O_2 and NO and the instantaneous rate of formation of NO_2 at 4 minutes and 41 minutes into the experiment. Show your work on the graph. Explain why the rate changes.

4. What do you notice about the ratios of the rates oxygen and nitrogen monoxide consumption to the production of nitrogen dioxide?

Answers to Problems: Kinetics Assignment #1

1. average rate for NO = $\frac{0.000115914 \text{ mol / L} - 0.000514 \text{ mol / L}}{71 \text{ min} - 0 \text{ min}}$

$= 5.61 \times 10^{-6}$ mol / L·min

average rate for O_2 = $\frac{0.000144 \text{ mol / L} - 0.000343 \text{ mol / L}}{71 \text{ min} - 0 \text{ min}}$

$= 2.80 \times 10^{-6}$ mol / L·min

average rate for NO_2 = $\frac{0.000398086 \text{ mol / L} - 0 \text{ mol / L}}{71 \text{ min} - 0 \text{ min}}$

$= 5.61 \times 10^{-6}$ mol / L·min

2. Over the first 10 minutes

average rate for NO = $\frac{0.00031192 \text{ mol / L} - 0.000514 \text{ mol / L}}{10 \text{ min} - 0 \text{ min}}$

$= 2.02 \times 10^{-5}$ mol / L·min

average rate for O_2 = $\frac{0.000242 \text{ mol / L} - 0.000343 \text{ mol / L}}{10 \text{ min} - 0 \text{ min}}$

$= 1.01 \times 10^{-5}$ mol / L·min

average rate for NO_2 = $\frac{0.00020208 \text{ mol / L} - 0 \text{ mol / L}}{10 \text{ min} - 0 \text{ min}}$

$= 2.02 \times 10^{-5}$ mol / L·min

Over the last 10 minutes

average rate for NO = $\frac{0.000115914 \text{ mol / L} - 0.000127914 \text{ mol / L}}{71 \text{ min} - 61 \text{ min}}$

$= 1.20 \times 10^{-6}$ mol / L·min

$$\text{average rate for } O_2 = \frac{0.000144 \text{ mol / L} - 0.000150 \text{ mol / L}}{71 \text{ min} - 61 \text{ min}}$$

$$= 6.00 \times 10^{-7} \text{ mol / L·min}$$

$$\text{average rate for } NO_2 = \frac{0.000398086 \text{ mol / L} - 0.000386086 \text{ mol / L}}{71 \text{ min} - 61 \text{ min}}$$

$$= 1.20 \times 10^{-6} \text{ mol / L·min}$$

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

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

$$\text{Instantaneous rate, } O_2 = \frac{0.00022 \text{ mol / L} - 0.00033 \text{ mol / L}}{10 \text{ min} - 0 \text{ min}}$$
$$= 1.10 \times 10^{-5} \text{ mol / L·min}$$

$$\text{Instantaneous rate, } NO_2 = \frac{0.000242 \text{ mol / L} - 0.00003 \text{ mol / L}}{10 \text{ min} - 0 \text{ min}}$$
$$= 2.12 \times 10^{-5} \text{ mol / L·min}$$

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

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

$$\text{Instantaneous rate, } O_2 = \frac{0.00014 \text{ mol / L} - 0.00019 \text{ mol / L}}{70 \text{ min} - 20 \text{ min}}$$
$$= 1.00 \times 10^{-6} \text{ mol / L·min}$$

$$\text{Instantaneous rate, } NO_2 = \frac{0.000365 \text{ mol / L} - 0.000325 \text{ mol / L}}{50 \text{ min} - 30 \text{ min}}$$
$$= 2.00 \times 10^{-6} \text{ mol / L·min}$$

4. The ratio between O_2 and NO_2 is 1:2. The rate consumption of O_2 is one-half the rate of formation of NO_2.

The ratio between NO and NO_2 is 2:2, or 1:1. The rate consumption of NO is equal to the rate of formation of NO_2.

Appendix 3: Kinetics Assignment #2

A chemist is studying the decomposition of di-nitrogen pentoxide at 45ºC. The balanced equation is:

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

The chemist measured the concentration of dinitrogen pentoxide at ten minute intervals for 100 minutes using colorimetry (Spectrophotometer – Spec 20). The data is in the table below. Use the data 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.

Time (min)	Concentration (mol/L)		
	$[N_2O_5]$	$[NO_2]$	$[O_2]$
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		

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

Questions

1. What is the average rate of decomposition of di-nitrogen pentoxide and the formation of nitrogen dioxide and oxygen over the entire 100 minute interval? Determine the rate for each.

2. What is the average rate for the first 20 minutes of the decomposition of N_2O_5 and for the last 20 minutes?

3. 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 80 minutes into the experiment. Show your work on the graph. Explain why the rate changes.

4. Explain why the rate changes between 10 and 80 minutes.

Answers to Problems: Kinetics Assignment #2

1. average rate for $N_2O_5 = \dfrac{0.0006 \text{ mol / L} - 0.0124 \text{ mol / L}}{100 \text{ min} - 0 \text{ min}}$

$= 1.18 \times 10^{-4}$ mol / L·min

average rate for $NO_2 = \dfrac{0.0236 \text{ mol / L} - 0 \text{ mol / L}}{100 \text{ min} - 0 \text{ min}}$

$= 2.36 \times 10^{-4}$ mol / L·min

average rate for $O_2 = \dfrac{0.00590 \text{ mol / L} - 0 \text{ mol / L}}{100 \text{ min} - 0 \text{ min}}$

$= 5.90 \times 10^{-5}$ mol / L·min

2. Over the first 20 minutes

average rate for $N_2O_5 = \dfrac{0.0068 \text{ mol / L} - 0.0124 \text{ mol / L}}{20 \text{ min} - 0 \text{ min}}$

$= 2.80 \times 10^{-4}$ mol / L·min

Over the last 20 minutes

average rate for $N_2O_5 = \dfrac{0.0006 \text{ mol / L} - 0.0011 \text{ mol / L}}{100 \text{ min} - 80 \text{ min}}$

$= 2.50 \times 10^{-5}$ mol / L·min

3. Answers will vary slightly due to the drawing of the tangent line to the point at 10 min.

Instantaneous rate, $N_2O_5 = \dfrac{0.00370 \text{ mol / L} - 0.0118 \text{ mol / L}}{30 \text{ min} - 0 \text{ min}}$
$= 2.70 \times 10^{-4}$ mol / L·min

Instantaneous rate, $NO_2 = \dfrac{0.0165 \text{ mol / L} - 0.0012 \text{ mol / L}}{30 \text{ min} - 0 \text{ min}}$
$= 5.10 \times 10^{-4}$ mol / L·min

Instantaneous rate, $O_2 = \dfrac{0.0025 \text{ mol / L} - 0.0004 \text{ mol / L}}{20 \text{ min} - 0 \text{ min}}$
$= 1.05 \times 10^{-4}$ mol / L·min

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

Instantaneous rate, $N_2O_5 = \underline{0 \text{ mol / L} - 0.002 \text{ mol / L}}$

115 min – 50 min

= 3.08×10^{-5} mol / L·min

Instantaneous rate, NO_2 = $\frac{0.0236 \text{ mol / L} - 0.02 \text{ mol / L}}{100 \text{ min} - 40 \text{ min}}$

= 6.00×10^{-5} mol / L·min

Instantaneous rate, O_2 = $\frac{0.00590 \text{ mol / L} - 0.00520 \text{ mol / L}}{100 \text{ min} - 60 \text{ min}}$

= 1.75×10^{-5} mol / L·min

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

Appendix 4: Chemical Kinetics Problems (BLM)

1. State 3 examples of properties, directly related to reactants or products, that could be used to measure a reaction rate. (van Kessel 365)

2. What property would be appropriate to measure rate in each of the following reactions?

   a. $MnO_4^- + 5\ Fe^{2+} + 8\ H^+ \rightarrow Mn^{2+} + 5\ Fe^{3+} + 4\ H_2O$

   b. $Zn + H_2SO_4 \rightarrow H_2 + ZnSO_4$ (van Kessel 365)

3. What units are used to express reaction rate?

4. In the reaction $3\ H_2 + N_2 \rightarrow 2\ NH_3$, how does the rate of disappearance of hydrogen compare to the rate of disappearance of nitrogen?
   How does the rate of production of NH_3 compare to the rate of disappearance of nitrogen?

5. For the reaction $2\ A + B \rightarrow 3\ C$, it was found that the rate of consumption of B was 0.30 mol / L·s. What was the rate of consumption of A and the rate of formation of C?

6. At a certain temperature, the rate of consumption of N_2O_5 is 2.5×10^{-6} mol / L·s. How fast are NO_2 and O_2 being formed? $2\ N_2O_5 \rightarrow 4\ NO_2 + O_2$

7. Write the rate expressions for the following reactions in terms of the disappearance of the reactants
   and the appearance of the products:

   a. $I^-_{(aq)} + OCl^-_{(aq)} \rightarrow Cl^-_{(aq)} + OI^-_{(aq)}$
   b. $3\ O_{2(g)} \rightarrow 2\ O_{3(g)}$
   c. $4\ NH_{3(g)} + 5\ O_{2(g)} \rightarrow 4\ NO_{(g)} + 6\ H_2O_{(g)}$
   d. $CH_{4(g)} + 2\ O_{2(g)} \rightarrow CO_{2(g)} + 2\ H_2O_{(g)}$ (Chang 538)

8. In a combustion reaction, 8.0 mol of methane gas reacts completely in a 2.00 L container containing excess oxygen gas in 3.2 s.

$$CH_{4(g)} + 2\ O_{2(g)} \rightarrow CO_{2(g)} + 2\ H_2O_{(g)}$$

   a. Calculate the averate rate of consumption of methane gas in mol / L·s.
   b. Calculate the averate rate of consumption of oxygen gas in mol / L·s.
   c. Calculate the averate rate of production of carbon dioxide gas in mol / L·s.
   d. Calculate the averate rate of production of water vapour in mol / L·s. (van Kessel 366)

9. Hydrogen iodide and oxygen react to form iodine gas and water vapour. If oxygen gas reacts at a rate of 0.0042 mol / L·s,

$$4\ HI_{(g)} + O_{2\,(g)} \rightarrow 2\ I_{2(g)} + 2\ H_2O_{(g)}$$

   a. what is the rate of formation of iodine gas in mol / L·s?
   b. what is the rate of formation of water vapour in mol / L·s?
   c. what is the rate of consumption of hydrogen iodide gas in mol / L·s?

10. Consider the reaction, $4\ NO_{2(g)} + O_{2\,(g)} \rightarrow 2\ N_2O_{5(g)}$. Suppose that at a particular moment during the reaction, oxygen is reacting at the rate of 0.024 mol / L·s. Calculate the rate at which N_2O_5 is being formed and calculate the rate at which NO_2 is being consumed. (Chang 538)

Appendix 5: Chemical Kinetics Problems - Answer Key

1. Reactions the produce a gas (measure volume/pressure); reaction that involve the ion as a product (conductivity); reactions that produce a colour change (spectrometer-measure colour intensity)

2. (a) colour change ($Fe^{2+} \rightarrow Fe^{3+}$)
   (b) volume/pressure (H_2(g) produced)

3. mol / L·s

4. $3\ H_2 + N_2 \rightarrow 2\ NH_3$

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

   $3\ H_2 + N_2 \rightarrow 2\ NH_3$

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

5. Rate of consumption of A is

   $2\ A + B \rightarrow 3\ C$

   is twice (2x) the rate of consumption of B (0.30 mol / L·s)

   = 2 x 0.30 mol / L·s

   = 0.60 mol / L·s

   Rate of formation of A is

   $2\ A + B \rightarrow 3\ C$

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

   = 3 x 0.30 mol / L·s

   = 0.90 mol / L·s

6. $2\ N_2O_5 \rightarrow 4\ NO_2 + O_2$

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

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

   rate of formation of NO_2 = 2 x 2.5×10^{-6} mol / L·s = 5.0×10^{-6} mol / L·s

$2\ N_2O_5 \rightarrow 4\ NO_2 + O_2$

The rate of formation of O_2 is 1/2 times as fast as compared to the rate of disappearance of N_2O_5.

rate of formation of O_2 = 1/2 x 2.5 x 10^{-6} mol / L·s = 1.25 x 10^{-6} mol / L·s

7.(a) $I^-_{(aq)} + OCl^-_{(aq)} \rightarrow Cl^-_{(aq)} + OI^-_{(aq)}$

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

(b) $3\ O_{2(g)} \rightarrow 2\ O_{3(g)}$

$$\text{Rate} = -\frac{1}{3}\frac{\Delta[O_2]}{\Delta t} = \frac{1}{2}\frac{\Delta[O_3]}{\Delta t}$$

(c) $4\ NH_{3(g)} + 5\ O_{2(g)} \rightarrow 4\ NO_{(g)} + 6\ H_2O_{(g)}$

$$\text{Rate} = -\frac{1}{4}\frac{\Delta[NH_3]}{\Delta t} = -\frac{1}{5}\frac{\Delta[O_2]}{\Delta t} = \frac{1}{4}\frac{\Delta[NO]}{\Delta t} = \frac{1}{6}\frac{\Delta[H_2O]}{\Delta t}$$

(d) $CH_{4(g)} + 2\ O_{2(g)} \rightarrow CO_{2(g)} + 2\ H_2O_{(g)}$

$$\text{Rate} = -\frac{\Delta[CH_4]}{\Delta t} = -\frac{1\Delta[O_2]}{2\ \Delta t} = \frac{\Delta[CO_2]}{\Delta t} = \frac{1\Delta[H_2O]}{2\Delta t}$$

8. (a) $[CH_4] = \frac{\text{mol}}{\text{L}} = \frac{8.0\ \text{mol}}{2.00\ \text{L}} = 4.0\ \text{mol / L}$

Rate of consumption of CH_4 = $\frac{\text{concentration}}{\text{time}} = \frac{4.0\ \text{mol / L}}{3.2\ \text{s}}$ = 1.25 mol / L·s

(b) Rate of consumption of O_2

$CH_{4(g)} + 2\ O_{2(g)} \rightarrow CO_{2(g)} + 2\ H_2O_{(g)}$

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

2 x 1.25 mol / L·s = 2.50 mol / L·s

(c) Rate of production of CO_2

$$CH_{4(g)} + 2\ O_{2(g)} \rightarrow CO_{2(g)} + 2\ H_2O_{(g)}$$

The rate of production of CO_2 is the same as compared to the rate of consumption of CH_4 is

1 x 1.25 mol / L·s = 1.25 mol / L·s

(d) Rate of production of H_2O

$$CH_{4(g)} + 2\ O_{2(g)} \rightarrow CO_{2(g)} + 2\ H_2O_{(g)}$$

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

2 x 1.25 mol / L·s = 2.50 mol / L·s

9. (a) $4\ HI_{(g)} + O_{2\ (g)} \rightarrow 2\ I_{2(g)} + 2\ H_2O_{(g)}$

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

= 2 x 0.0042 mol / L·s = 0.0084 mol / L·s

(b) $4\ HI_{(g)} + O_{2\ (g)} \rightarrow 2\ I_{2(g)} + 2\ H_2O_{(g)}$

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

= 2 x 0.0042 mol / L·s = 0.0084 mol / L·s

(c) $4\ HI_{(g)} + O_{2\ (g)} \rightarrow 2\ I_{2(g)} + 2\ H_2O_{(g)}$

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

= 4 x 0.0042 mol / L·s = 0.0168 mol / L·s

10. (a) $4\ NO_{2(g)} + O_{2\ (g)} \rightarrow 2\ N_2O_5$

The rate of formation of N_2O_5 is 2 times as fast as compared to the rate of consumption of O_2.

= 2 x 0.024 mol / L·s = 0.048 mol / L·s

(b) $4\ NO_{2(g)} + O_{2\,(g)} \rightarrow 2\ N_2O_5$

The rate of consumption of NO_2 is 4 times as fast as compared to the rate of consumption of O_2.

$= 4 \times 0.024$ mol / L·s = 0.096 mol / L·s

Appendix 6: Lab - Factors Affecting Reaction Rates

It is very important for a chemist to understand the conditions that affect the rate of a chemical reaction. In chemical manufacturing processes, controlling the rate of a given reaction can make all the difference between an economical process and an uneconomical one. For instance, using catalysts is one method for getting reactions to speed up.

In this experiment you will carry out different reactions. You will examine the effect of the nature of the reactants, changing temperature, effect of a catalyst and changing particle size.

Hypothesis:
Predict how each of the factors – temperature change, catalyst, particle size and nature of reaction, will affect the rate of reaction and explain why you made that prediction.

Procedure:

Part A

1. Set up five test tubes as follows:

Reagent	Test Tube				
	#1	#2	#3	#4	#5
0.10 mol/L $Na_2C_2O_4$	3 mL	-	3 mL	3 mL	3 mL
0.10 mol/L $FeSO_4$	-	3 mL	-	-	-
1.0 mol/L H_2SO_4	1 mL	1 mL	2 mL	1 mL	1 mL
Water	1 mL	1 mL	-	1 mL	1 mL
0.10 mol/L $MnSO_4$	-	-	-	-	3 drops
Temperature	20ºC	20ºC	20ºC	50ºC	20ºC
0.020 mol/L $KMnO_4$ (drops)	3	3	3	3	3
Reaction Time (s)					

2. Add all reagents specified except the 3 drops of 0.020 mol/L $KMnO_4$, potassium permanganate, and MIX.

3. Test tube #1 is the control test tube, all others will be compared to it.

4. Add three drops of 0.020 mol/L $KMnO_4$ to test tube #1 and MIX. Record the time for the pink-purple colour to disappear.

5. Repeat the process for each of the other test tubes. Note: Test tube #4 must be at 50ºC BEFORE adding the $KMnO_4$.

Part B

1. To test tube #1 from Part A, add 3 more drops of $KMnO_4$. Record the reaction time.

2. When the reaction is complete, repeat this about 4 more times.

Addition Sets of Drops	Reaction Time (s)
1	
2	
3	
4	
5	

Part C

1. Place a few crystals of calcium carbonate, $CaCO_3$(s), in beaker #1 and an equal amount of powder in beaker #2.

2. Add about 20 mL of 3.0 mol/L HCl to each beaker. Compare the reactions in each beaker.

	Contents	Observations
Beaker #1		
Beaker #2		

Questions:

The net equations for the reactions that took place in test tubes #1 and #2:

#1 $5C_2O_4^{2-}{}_{(aq)} + 2MnO_4^{-}{}_{(aq)} + 16H^{+}{}_{(aq)} \rightarrow 10CO_{2(g)} + 2Mn^{2+}{}_{(aq)} + 8H_2O_{(l)}$

#2 $5Fe^{2+}{}_{(aq)} + MnO_4^{-}{}_{(aq)} + 8H^{+}{}_{(aq)} \rightarrow 5Fe^{3+}{}_{(aq)} + Mn^{2+}{}_{(aq)} + 4H_2O_{(l)}$

1. Compare the data obtained with test tubes #1 and #2, together with the equations above. Explain why you obtained the data that you did.

2. Compare the data obtained with test tubes #1 and #3. Make a statement about how the rate of the reaction and concentration are related. Explain this relationship.

3. Compare the data obtained with test tubes #1 and #4. Make a statement about how temperature and reaction rate are related. Explain why they are related in this way.

4. Compare the data obtained with test tubes #1 and #5. Explain why the $MnSO_4$ affected the reaction in this way.

5. If the complete balanced equation for test tube #1 is

   $5H_2C_2O_{4(aq)} + 2KMnO_{4(aq)} + 3H_2SO_{4(aq)} \rightarrow 10CO_{2(g)} + 2MnSO_{4(aq)} + 8H_2O_{(l)} + K_2SO_{4(aq)}$

   Make a statement describing your results from Part B. Explain the peculiar results you obtained. If you cannot see why they are peculiar, ask the teacher.

6. What did you notice about the reactions in the two beakers in Part C? How were they similar and how were they different? Explain your observations.

7. Explain why an iron nail, when heated in a Bunsen burner will only glow red, whereas steel wool will burn away.

8. Why is milk kept in the refrigerator?

Conclusion:
Explain how each of the factors studied during this experiment affects the rate of a chemical reaction and provide evidence for your conclusion. Did your hypothesis predict the same results?

Instructors Notes:

Part A

Test tube #1 tends to vary in reaction times from 5 – 12 minutes. Students need only to record times to the nearest second, as this is not a purely quantitative investigation.

The permanganate ion may not convert entirely to the Mn (II) ion, which is colourless, in the test tube with the added catalyst. The disappearance of the purple to the brown/yellow of the Mn (IV) is acceptable as an endpoint.

Part B

This is an optional portion that can be done if time permits and poses a challenge for the students. Some students will need more guidance in answering question #5. If Time is short, 3 or 4 trials will establish a trend.

Students will find that the reaction times actually decrease, rather than increase, because the reaction itself produces $MnSO_4$, which is a catalyst for the reaction.. They found in Part A that a decreased concentration of reactants increases reaction time. The students will tend to conclude that they are increasing the amount of $KMnO_4$, which increases the rate. They will need to be reminded that every time they perform the reaction the other reactant concentrations are reduced.

In this reaction, the Mn (IV) tends to persist, resulting ion a yellowish solution. As long as the students are consistent with the colour they choose as the endpoint the reaction times should decrease for each successive trial.

Part C

Students should use 150 mL beakers for this part of the lab.

Appendix 2.1.d **Student Experiment**

Factors Affecting the Rate of a Reaction

(Nature of Reactants, Surface Area, Temperature, and Catalyst)

Part 1: Nature of the Reactants

1. Add 20 drops of 3.0 mol L^{-1} 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 all 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^{-1} 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^{-1} 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^{-1} aqueous zinc(II) nitrate solution to a third well, 20 drops of 1.0 mol L^{-1} iron(III) nitrate solution to a fourth well, and 20 drops of 1.0 mol L^{-1} 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 test tube? 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?

Part 2: Surface Area (degree of subdivision of solid)

1. Add 30 drops of 3.0 mol L^{-1} 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.

9. Observe and record your observations!

Questions:

- What happened in each test tube? Identify the products in each case.

Appendix 2.1.d Factors Affecting the Rate of a Reaction (Contd.) Student Experiment

- Write a balanced equation to represent each reaction.
- How much time do the solids take to react in each solution?

Part 3: 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^{-1} aqueous potassium permanganate, $KMnO_4$, solution (made acidic with sulfuric acid) to each of two 13x100 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^{-1} 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^{-1} oxalic acid solution to the test tube in the hot water bath. Stir.

   The reaction is: $2\ MnO_4^-(aq) + 5\ H_2C_2O_4(aq) + 6\ H_3O^+(aq) \rightarrow 2\ Mn^{2+}(aq) + 10\ CO_2(g) + 14\ H_2O(l)$
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 13x100 mm test tubes. Place one of the test tubes in the hot water bath, one in the cold water bath, and leave one at room temperature. Wait about 2 min. for the solutions to come to temperature.
8. Cut three pieces of magnesium ribbon 0.5 cm long. Add one piece to each of the three test tubes. Observe the time required for each piece to completely disappear.
9. Does the reaction take the same time at each temperature? Explain.

Part 4: Catalyst

1. Add 2 mL of 0.01 mol L^{-1} aqueous potassium permanganate (KMnO4) solution (made acidic with sulfuric acid) to each of two 13x100 mm test tubes.
2. To one of the test tubes, add 5 drops of 0.01 mol L^{-1} manganese(II) sulfate solution.
3. Add 5 mL of 0.02 mol L^{-1} oxalic acid solution to each of the test tubes, stopper and shake.
4. Observe and record your observations!

Questions:

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

Appendix 2.1.e **Teacher Support Material**

Factors Affecting the Rate of a Reaction

(Nature of Reactants, Surface Area, Temperature, and Catalyst)

These 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 (also shown in Appendix 2.1.h)

Appendix 2.1.b and Appendix 2.1.f can be used to demonstrate the effect of changing concentration.

Solutions:

- 3.0 mol L^{-1} hydrochloric acid solution. Dilute 258 mL of concentrated (11.6 mol L^{-1} or 36%) hydrochloric acid to 1.0 L with distilled water.
- 1.0 mol L^{-1} zinc(II) nitrate solution. Dissolve 29.7 g of zinc(II) nitrate hexahydrate, $Zn(NO3)_2 \cdot 6H_2O$, in 100 mL water
- 1.0 mol L^{-1} iron(III) nitrate solution. Dissolve 4.04 g of iron(III) nitrate nonahydrate, $Fe(NO_3)_3 \cdot 9H_2O$, in 100 mL water.
- 1.0 mol L^{-1} copper(II) nitrate solution. Dissolve 29.6 g of copper(II) nitrate hexahydrate, $Cu(NO_3)_2 \cdot 6H_2O$, in 100 mL water.
- 0.01 mol L^{-1} potassium permanganate. Add 1 mL of concentrated (17.8 mol L^{-1} or 95%) sulfuric acid to 75 mL water and add enough water to make the total volume up to 100 mL. Dissolve 1.58 g of potassium permanganate ($KMnO_4$) in this solution.
- 0.02 mol L^{-1} oxalic acid solution. Dissolve 2.52 g of oxalic acid (ethanedioic acid, $HOOCCOOH \cdot 2H_2O$) in 100 mL water.
- 0.01 mol L^{-1} manganese(II) sulfate solution. Dissolve 2.23 g of manganese(II) sulfate tetrahydrate (manganous sulfate, $MnSO_4 \cdot 4H_2O$) in 100 mL water.

NOTE:

1. Sulfate 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.