Key Questions

1. What are the units of:
   a. \( [A] \) (where \( A \) stands for the formula of some reactant)? \( \text{mol/L (or M)} \)
   b. the rate of a reaction? \( \text{mol/L s} \)
   c. the rate constant of a first-order reaction? \( \text{Hint: Use eq 1 and dimensional analysis using your answers to parts (a) and (b). } \text{Show your work below.} \)
      \[ \text{Rate} = k [A] \]
      \[ \frac{M}{s} = k M \]
      \[ k = \frac{1}{s} \text{ (or s}^{-1}) \]
   d. the rate constant of a second-order reaction? \( \text{Hint: Use eq 2 and dimensional analysis using your answers to parts (a) and (b). } \text{Show your work below} \)
      \[ R = k [A]^2 \]
      \[ \frac{M}{s} = k M^2 \]
      \[ \frac{M}{s} \frac{1}{M^2} = k \]
      \[ k = \frac{1}{M s} = \frac{L}{\text{mol} \cdot \text{s}} \text{ (or M}^{-1} \text{s}^{-1}) \]

2. Consider the following reaction that is first order with respect to \( A \) and first order with respect to \( B \):

   \[ A + B \rightarrow \text{Products} \]

   a. By what factor would the reaction rate increase if \([A]\) were doubled but \([B]\) remains unchanged? \(2\)
   b. By what factor would the reaction rate increase if \([B]\) were tripled but \([A]\) remains unchanged? \(3\)
   c. Write the rate law for the reaction: \( k [A][B] = \text{RATE} \) \( \text{(it's an equation!)} \)
   d. By what factor would the reaction rate increase if \([A]\) were tripled and \([B]\) were doubled? \(6\)

The Model: Overall Rate Orders

For the generic chemical reaction: \( a A + b B \rightarrow d D + e E \)
the rate law is:
\[ \text{rate} = k [A]^x [B]^y \]

where \( k \) is the rate constant, \( x \) is the order with respect to \( A \), and \( y \) is the order with respect to \( B \). The overall order of the reaction is \( x + y \).
3. For each reaction and its corresponding rate law, indicate what the order is with respect to each reactant and the overall order of each reactant.

<table>
<thead>
<tr>
<th>Reaction</th>
<th>Rate Law</th>
<th>Order with respect to</th>
<th>Overall Reaction Order</th>
</tr>
</thead>
<tbody>
<tr>
<td>a. 2 NO₂ + F₂ → 2 NO₂F</td>
<td>rate = (k \left[ \text{NO}_2 \right] \left[ \text{F}_2 \right] )</td>
<td>NO₂: 1</td>
<td>F₂: 1</td>
</tr>
<tr>
<td>b. 2 NO + O₂ → 2 NO₂</td>
<td>rate = (k \left[ \text{O}_2 \right] \left[ \text{NO} \right]^2 )</td>
<td>NO: 2</td>
<td>O₂: 1</td>
</tr>
<tr>
<td>c. NO + N₂O₅ → 3 NO₂</td>
<td>rate = (k \left[ \text{N}_2\text{O}_5 \right] )</td>
<td>NO: 0</td>
<td>N₂O₅: 1</td>
</tr>
</tbody>
</table>

4. True or False (circle your choice): There is a one-to-one relationship between the stoichiometric coefficient of a reactant and the order with respect to that reactant in the rate law. Explain your answer.

The Order has to be determined experimentally, whereas the stoichiometric coefficient can be determined without experimenting. **Use Rxs a” & “c” above in #3 to support this response!**

**The Model: Determination of Rate Orders**

So at this point, one may ask the question: “If stoichiometric coefficients do not determine the orders, then how are rate laws determined?” A very common way of determining the order with respect to a reactant in the rate law is the Method of Initial Rates. The standard strategy to employ is to perform the same reaction several times but at the same temperature. In each of the trials, the concentration of one of the reactants is changed while the concentrations of any other reactants are held constant. Any observed change in the initial rate of the reaction is, therefore, due to the change in concentration of that one reactant. The initial rate of the one trial serves as a “reference” to which the other trials’ initial rates are compared. By determining the ratio of the initial rates of two trials in which \([A]\) is different (but the concentration of any other reactant is the same), the order with respect to \(A\) may be determined.

\[
\begin{align*}
[A]_1 &= \text{molar concentration of A in trial 1} \\
Rate_1 &= \text{initial rate of reaction in trial 1} \\
[A]_2 &= \text{molar concentration of A in trial 2} \\
rate_2 &= \text{initial rate of reaction in trial 2} \\
x &= \text{order with respect to A} \\
Rate_2 &= \left( \frac{[A]_2}{[A]_1} \right)^x \\
Rate_1 &= \left( \frac{[A]_1}{[A]_2} \right)^x
\end{align*}
\]

**Key Questions**

5. Fill in the blanks in the following table. Be able to show your work and/or explain how you arrived at your answers.

<table>
<thead>
<tr>
<th>Factor by which ([A]) is changed in going from trial 1 to 2</th>
<th>Factor by which initial rate of reaction changes in going from trial 1 to 2</th>
<th>(x) (order with respect to (A))</th>
</tr>
</thead>
<tbody>
<tr>
<td>a. ([A]) is doubled</td>
<td>rate doubles</td>
<td>1</td>
</tr>
<tr>
<td>b. ([A]) is halved</td>
<td>rate remains constant</td>
<td>0</td>
</tr>
<tr>
<td>c. ([A]) is tripled</td>
<td>rate increases by factor of 9</td>
<td>2</td>
</tr>
<tr>
<td>d. ([A]) is tripled</td>
<td>rate triples</td>
<td>1</td>
</tr>
<tr>
<td>e. ([A]) is halved</td>
<td>rate is fourth, ??</td>
<td>Unclear</td>
</tr>
</tbody>
</table>
6. The oxidation of thallium (I) to thallium(III) by cerium(IV) is catalyzed by manganese(II). The kinetics of the aqueous reaction was observed several times at the same temperature:

\[
2 \text{ Ce}^{4+}(aq) + \text{Tl}^+(aq) \xrightarrow{\text{Mn}^{2+}} 2 \text{ Ce}^{3+}(aq) + \text{Tl}^3+(aq)
\]

<table>
<thead>
<tr>
<th>Experiment</th>
<th>[Ce(^{4+})](_{\text{initial}})</th>
<th>[Tl(^+)](_{\text{initial}})</th>
<th>[Mn(^{2+})]</th>
<th>Initial rate of formation of Tl(^3+) (M s(^{-1}))</th>
</tr>
</thead>
<tbody>
<tr>
<td>1</td>
<td>0.0050</td>
<td>0.0025</td>
<td>0.0010</td>
<td>2.0 × 10(^{-4})</td>
</tr>
<tr>
<td>2</td>
<td>0.0050</td>
<td>0.0025</td>
<td>0.0010</td>
<td>2.0 × 10(^{-4})</td>
</tr>
<tr>
<td>3</td>
<td>0.0100</td>
<td>0.0025</td>
<td>0.0010</td>
<td>4.0 × 10(^{-4})</td>
</tr>
<tr>
<td>4</td>
<td>0.0050</td>
<td>0.0025</td>
<td>0.0030</td>
<td>6.0 × 10(^{-4})</td>
</tr>
</tbody>
</table>

**a.** Determine the rate orders with respect to Ce\(^{4+}\), to Tl\(^+\), and to Mn\(^{2+}\) and explain or show how you arrived at your answers.

i. The order with respect to Ce\(^{4+}\) is __0__ . **Work/Explanation:**

\[
\frac{\text{[Ce}^{4+}\text{]}}{\text{[Ce}^{4+}\text{]}} = \frac{2.0 \times 10^{-4}}{2.0 \times 10^{-4}} = \frac{0.0010}{0.0050}
\]

\[
X = 1
\]

ii. The order with respect to Tl\(^+\) is __0__ . **Work/Explanation:**

\[
\frac{\text{Tl}^+}{\text{Tl}^+} = \frac{2.0 \times 10^{-4}}{2.0 \times 10^{-4}} = \frac{0.0025}{0.0050}
\]

\[
X = 0
\]

iii. The order with respect to Mn\(^{2+}\) is __1__ . **Work/Explanation:**

\[
\frac{\text{Mn}^{2+}}{\text{Mn}^{2+}} = \frac{6.0 \times 10^{-4}}{2.0 \times 10^{-4}} = \frac{0.0030}{0.0010}
\]

\[
X = 1
\]

**b.** Write the rate law for the reaction.

\[\text{Rate} = k \text{ [Ce}^{4+}\text{] [Mn}^{2+}\text{]}\]

**c.** Calculate the value of the rate constant. **Show your work using correct units and sig. figs. and circle your answer.**

\[
k = \frac{\text{Rate}}{\text{[Ce}^{4+}\text{][Mn}^{2+}\text{]}} = \frac{4.0 \times 10^{-4}}{(0.0100)(0.0010)} = 40.4 \text{ M}^{-1}\text{s}^{-1}
\]

**GOOD, BUT SINCE THE "MATH" IS SO SIMPLE, YOU COULDN'T HAVE MADE A MISTAKE!**

**WRITE THIS DOWN AS AN EXPLANATION**

**Since in Expt 1, 3, 4 \text{[Ti}^+\text{]} \& \text{[Mn}^{2+}\text{]} ARE CONSTANT, BUT \text{[Ce}^{4+}\text{]} DOUBLES, THIS CAUSES THE RXN RATE TO DOUBLE.**

\(\therefore\) The order w.r.t. Ce\(^{4+}\) = 1
d. What would the initial rate of the formation of $\text{TI}^+$ be if $[\text{Ce}^{4+}]_{\text{initial}} = 0.0100 \ M$, $[\text{TI}^+]_{\text{initial}} = 0.00500 \ M$, and $[\text{Mn}^{2+}] = 0.00300 \ M$? Show work using correct units and sig. figs., and circle your answer.

$$\text{Rate} = k \left[\text{Ce}^{4+}\right] \left[\text{TI}^+\right]$$

$$\text{Rate} = 400 \ \text{m}^{-1} \ \text{s}^{-1} \left[0.0100 \ M \ \text{Ce}^{4+}\right] \left[0.00500 \ M \ \text{TI}^+\right]$$

$$\text{Rate} = 0.0812 \ \text{M/s}$$

Model: Kinetics of a Unimolecular Reaction

Under certain conditions ozone in the atmosphere decomposes by dissociation ($\text{O}_3 \rightarrow \text{O}_2 + \text{O}$) with the kinetics given below. The data in the graph (fig. 1) and in table 1 are described by the three equations below.

Equations that describe the decomposition of ozone, $\text{O}_3$:

Equation 1. reaction rate $= -\Delta [\text{O}_3]/\Delta t$

Equation 2. rate law: $\text{Rate} = -k \ [\text{O}_3]$

Equation 3. integrated rate law: $\ln \left([\text{O}_3]/[\text{O}_3]_0\right) = -k \ t$ Where: $[\text{O}_3]_0 =$ conc. of $\text{O}_3$ @ $t = 0$

![Graph showing the decomposition of ozone.](image)

Figure 1. Data describing the decomposition of ozone. Note that the curved line refers to the left scale, and the straight line refers to the right scale.

Table 1. Data on Ozone Decomposition (half-life time = 1270 s)

<table>
<thead>
<tr>
<th>Time (s)</th>
<th>0</th>
<th>100.</th>
<th>900.</th>
<th>1000.</th>
<th>6900.</th>
<th>7000.</th>
</tr>
</thead>
<tbody>
<tr>
<td>$\text{O}_3$ concentration ($x 10^{-5}$ M)</td>
<td>89.63</td>
<td>84.87</td>
<td>54.84</td>
<td>51.92</td>
<td>2.07</td>
<td>1.96</td>
</tr>
<tr>
<td>$\ln([\text{O}_3]/[\text{O}_3]_0)$</td>
<td>0.00</td>
<td>-0.0546</td>
<td>-0.4914</td>
<td>-0.5460</td>
<td>-3.767</td>
<td>-3.822</td>
</tr>
</tbody>
</table>

17. Draw lines on the graph in fig. 1 to mark the half-life time and the concentration at the half-life time and label this time with $t_{1/2}$ and the concentration with $C_{t_{1/2}}$. 

Page 5 of 8
7. Use figure 1 to determine how long it takes for half the ozone to decompose. What is the concentration of ozone after that time?

\[ \frac{8.4 \times 10^{-3} \text{ mmol/L}}{2} = 4.4 \times 10^{-5} \text{ M O}_3 \]

**Using Fig. 1 should give ~1200s**

8. Use figure 1 to determine initial reaction rate of the ozone decomposition. Provide both the magnitude and the units.

\[ -5.2 \times 10^{-6} \text{ M} - 9.0 \times 10^{-5} \text{ M} = 3.8 \times 10^{-8} \text{ M/s} \]

\[ \text{SF16s}!! \]

9. Use figure 1 to determine the approximate rate of the reaction after 1.92 hours have passed.

\[ -2.0 \times 10^{-6} \text{ M} - 6.0 \times 10^{-6} \text{ M} \]

\[ \frac{1000 \text{ s}}{4500 \text{ s}} \]

However, the slope is almost 0, so the rate is almost 0.

10. Referring to figure 1, how does the rate of the reaction change with time? Why does the rate change with time? The rate of reaction decreases. This is because the concentration of the reactant is decreasing over time so fewer collisions of reactive particles are occurring, which means the rate of reaction is slower.

11. The reaction is ______ order of the reaction with respect to ozone.

12. Overall, the decomposition of ozone reaction is ______ order.

\[ \text{From Eqn. 2: } R = -k[\text{O}_3] \]

13. If the concentration of ozone were doubled, the reaction rate would ________.

14. What parameter in the integrated rate law determines the slope of the straight line in the graph?

\[ \text{m = slope = } k \]

15. Why does the concentration data, when plotted in the above graph with time on the x-axis, produce a curved line in one case and a straight line in the other?

The curved line is logarithmic due to the exponential decrease of concentration of O₃ over time. The straight line is straightened out by the use of natural log in the integrated rate law.
Team Responses to ALE 2 - Spring 2014  

Team No. 4  Section B

16. How do you think an increase in temperature would affect the rate constant for the decomposition of ozone? Explain.

THE INCREASE IN TEMPERATURE WILL INCREASE THE RATE OF REACTION. SINCE THE RATE OF REACTION IS DIRECTLY PROPORTIONAL TO THE RATE CONSTANT, IT WILL CAUSE THE RATE CONSTANT TO INCREASE ALONG WITH THE TEMPERATURE.

17. Draw lines on the graph in fig. 1 to mark the half-life time and the concentration at the half-life time and label this time with t_1/2 and the concentration with C_{1/2}.

18. Write the rate law (i.e. the differential rate law, equation 2) for the ozone decomposition reaction.

\[ \text{RATE LAW:} \quad -k \left[ \text{O}_3 \right] = \frac{\Delta [\text{O}_3]}{\Delta t} = \text{RATE} \]

19. Write the integrated rate law (i.e. equation 3) for the ozone decomposition reaction.

\[ \ln \left( \frac{[\text{O}_3]}{[\text{O}_3]_0} \right) = -kt \]

20. Calculate the rate constant (use sig. figs. and units) for the ozone decomposition reaction from the slope of the straight line in fig. 1.

\[ \ln \left( \frac{[\text{O}_3]}{[\text{O}_3]_0} \right) = -kt \rightarrow -k = \frac{-\ln \left( \frac{[\text{O}_3]}{[\text{O}_3]_0} \right)}{t} \]

\[ \text{At} \ t = 1000 \rightarrow k = \frac{-\ln \left( \frac{81.92}{89.63} \right)}{1000 \text{ s}} = 5.460 \times 10^{-4} \text{ s}^{-1} \]

Exercises involving Rate Laws

21. Given the following rate law: Rate = k [CHCl₃] [Cl₂]^{1/2}

a. What is the reaction order with respect to chloroform, CHCl₃? \( \frac{1}{2} \) order

b. What is the reaction order with respect to chlorine? \( \frac{1}{2} \) order

c. What is the overall reaction order? \( \frac{3}{2} \) order

d. If the concentration of chloroform is cut in half, what will happen to the reaction rate? Give a quantitative response.

THE RATE WILL DECREASE BY A FACTOR OF 2.

\[ \frac{\text{RATE}_2}{\text{RATE}_1} = \left( \frac{1}{2} \right)^{1/2} \rightarrow \text{RATE}_2 = \frac{1}{2} \text{ RATE}_1 \]

?? Unclear

e. If the concentration of chlorine is tripled, what will happen to the reaction rate? Give a quantitative response.

THE RATE WILL BE \( \sqrt{3} \) FASTER.

\[ \frac{\text{RATE}_2}{\text{RATE}_1} = \left( \frac{[\text{Cl}_2]_2}{[\text{Cl}_2]_1} \right)^{1/2} = \left( \frac{3}{1} \right)^{1/2} \quad \text{RATE}_2 = \sqrt{3} \text{ faster} \]

\[ = 1.73 \text{ times faster} \]
16. How do you think an increase in temperature would affect the rate constant for the decomposition of ozone? Explain.

An increase in temperature will increase the rate constant. This is because temperature makes more particles have the ability to react, so the rate of reaction increases and the rate constant must increase.

17. Draw lines on the graph in fig. 1 to mark the half-life time and the concentration at the half-life time and label this time with $t_{1/2}$ and the concentration with $C_{1/2}$.

18. Write the rate law (i.e. the differential rate law, equation 2) for the ozone decomposition reaction.

$$\text{Rate} = -k \left[ O_3 \right] = \frac{\Delta [O_3]}{\Delta t}$$

19. Write the integrated rate law (i.e. equation 3) for the ozone decomposition reaction.

$$\ln \left( \frac{[O_3]}{[O_3]_0} \right) = -kt$$

20. Calculate the rate constant (use sig. figs. and units) for the ozone decomposition reaction from the slope of the straight line in fig. 1.

$$-\ln \left( \frac{[O_3]}{[O_3]_0} \right) = k$$

$$-\ln \left( \frac{84.87 \times 10^{-6} \text{ M}}{80.62 \times 10^{-6} \text{ M}} \right) = k$$

$$k = 5.46 \times 10^{-4} \text{ s}^{-1}$$

**Exercises involving Rate Laws**

21. Given the following rate law: Rate = $k [\text{CHCl}_3] [\text{Cl}_2]^{\frac{1}{2}}$

a. What is the reaction order with respect to chloroform, $\text{CHCl}_3$? _____order

b. What is the reaction order with respect to chlorine? _____order

c. What is the overall reaction order? _____order

d. If the concentration of chloroform is cut in half, what will happen to the reaction rate? Give a quantitative response.

The reaction rate is cut in half.

e. If the concentration of chlorine is tripled, what will happen to the reaction rate? Give a quantitative response.

The reaction rate increases by a factor of $\sqrt{3} = 1.73$ times.
22. The rate of a reaction is expressed in terms of changes in concentration of reactants and products. Write the balanced chemical equation for the following the reaction described by the following mathematical equation:

\[
\text{Rate} = \frac{1}{2} \frac{\Delta \text{[N}_2\text{O}_5]}{\Delta t} = \frac{1}{4} \frac{\Delta \text{[NO}_2]}{\Delta t} = \frac{\Delta \text{[O}_2]}{\Delta t}
\]

\[2 \text{N}_2 \text{O}_5 \rightarrow 4 \text{NO}_2 + \text{O}_2\]

23. Give the individual reaction orders for each substance and the overall reaction order for the following rate law:

\[\text{Rate} = k \left[\text{O}_3\right]^2 \div \left[\text{O}_2\right]\]

Reaction Order for: a.) \text{O}_3 = 2 \quad \text{b.) O}_2 = -1 \quad \text{c.) Overall reaction} = 1

24. By what factor does the rate change for the rate in the previous problem change if.....

a. \text{[O}_3\text{]} is doubled? \quad \text{↑} 4 \times \quad \text{c. [O}_2\text{]} is doubled? \quad \text{↓} \frac{1}{2} \quad 

b. \text{[O}_2\text{]} is halved? \quad \text{↑} 2 \times \quad \text{d. [O}_3\text{]} is halved? \quad \text{↓} \frac{1}{4}

25. Find the rate law, calculate the rate constant and determine the overall reaction order for the reaction \(\text{H}_2 + \text{I}_2 \rightarrow 2 \text{HI}\) from the following data at 450°C. Show/explain your work using correct units and sig. figs.

<table>
<thead>
<tr>
<th>Experiment</th>
<th>Initial Rate (mol L(^{-1}) s(^{-1}))</th>
<th>Initial [H(_2)] (mol/L)</th>
<th>Initial [I(_2)] (mol/L)</th>
</tr>
</thead>
<tbody>
<tr>
<td>1</td>
<td>1.9 x 10(^{-23})</td>
<td>0.0113</td>
<td>0.0011</td>
</tr>
<tr>
<td>2</td>
<td>1.1 x 10(^{-22})</td>
<td>0.0220</td>
<td>0.0033</td>
</tr>
<tr>
<td>3</td>
<td>9.3 x 10(^{-23})</td>
<td>0.0550</td>
<td>0.0011</td>
</tr>
<tr>
<td>4</td>
<td>1.9 x 10(^{-22})</td>
<td>0.0220</td>
<td>0.0056</td>
</tr>
</tbody>
</table>

\[\frac{\text{Rate}_3}{\text{Rate}_1} = \left(\frac{[\text{H}_2]\_3}{[\text{H}_2]\_1}\right)^{m} \Rightarrow \frac{9.3 \times 10^{-23} \text{ M/s}}{1.9 \times 10^{-22} \text{ M/s}} = \left(\frac{0.0550 \text{ M}}{0.0113 \text{ M}}\right)^{m} \]

\[4.895 = (4.867)^{m} \Rightarrow m = \frac{\log (4.895)}{\log (4.867)} \approx 1\]

\[\frac{\text{Rate}_4}{\text{Rate}_2} = \left(\frac{1.9 \times 10^{-22} \text{ M/s}}{1.1 \times 10^{-22} \text{ M/s}}\right) = \left(\frac{0.0056 \text{ M}}{0.0033 \text{ M}}\right)^{n} \Rightarrow \frac{1.2273}{(1.697)^{n}} = \frac{\log (1.2273)}{\log (1.697)} \approx 1 \Rightarrow n = 0.967 \approx 1 \]

\[k = \frac{\text{Rate}}{[\text{H}_2][\text{I}_2]} \Rightarrow \frac{1.9 \times 10^{-23} \text{ M/s}}{(0.0113 \text{ M})(0.0011 \text{ M})} = 1.5 \times 10^{-3} \text{ M/s} \]

Overall order = 2
22. The rate of a reaction is expressed in terms of changes in concentration of reactants and products. Write the balanced chemical equation for the following the reaction described by the following mathematical equation:

\[
\text{Rate} = \left( \frac{1}{2} \frac{\Delta [\text{N}_2\text{O}_5]}{\Delta t} \right) = \left( \frac{1}{4} \frac{\Delta [\text{NO}_2]}{\Delta t} \right) = \left( \frac{\Delta (\text{O}_2)}{\Delta t} \right) \Rightarrow 2\text{N}_2\text{O}_5 \rightarrow 4\text{NO}_2 + \text{O}_2
\]

23. Give the individual reaction orders for each substance and the overall reaction order for the following rate law:

\[
\text{Rate} = k \left[ \text{O}_3 \right]^2 \left[ \text{O}_2 \right]^{-1}
\]

Reaction Order for:

- \( \text{O}_3 = 2 \)
- \( \text{O}_2 = -1 \)
- Overall reaction = 1

24. By what factor does the rate change for the rate in the previous problem change if:

- \( \text{[O}_3 \text{]} \) is doubled? \( \times \frac{1}{2} \)
- \( \text{[O}_2 \text{]} \) is doubled? \( \times \frac{1}{2} \)
- \( \text{[O}_3 \text{]} \) is halved? \( \times 2 \)
- \( \text{[O}_2 \text{]} \) is halved? \( \times \frac{1}{4} \)

25. Find the rate law, calculate the rate constant and determine the overall reaction order for the reaction \( \text{H}_2 + \text{I}_2 \rightarrow 2\text{HI} \) from the following data at 450°C. Show/explain your work using correct units and sig. figs.

<table>
<thead>
<tr>
<th>Experiment</th>
<th>Initial Rate (mol L⁻¹ s⁻¹)</th>
<th>Initial [H₂] (mol/L)</th>
<th>Initial [I₂] (mol/L)</th>
</tr>
</thead>
<tbody>
<tr>
<td>1</td>
<td>1.9 x 10⁻²³</td>
<td>0.0113</td>
<td>0.0011</td>
</tr>
<tr>
<td>2</td>
<td>1.1 x 10⁻²²</td>
<td>0.0220</td>
<td>0.0033</td>
</tr>
<tr>
<td>3</td>
<td>9.3 x 10⁻²³</td>
<td>0.0550</td>
<td>0.0011</td>
</tr>
<tr>
<td>4</td>
<td>1.9 x 10⁻²²</td>
<td>0.0220</td>
<td>0.0056</td>
</tr>
</tbody>
</table>

\[
\frac{1.9 \times 10^{-22} \text{ M/s}}{1.9 \times 10^{-22} \text{ M/s}} = \left( \frac{0.056 \text{ M}}{0.033 \text{ M}} \right)^x = 1.7 = (1.7)^x \Rightarrow x = 1
\]

\[
\frac{9.3 \times 10^{-23} \text{ M/s}}{1.9 \times 10^{-23} \text{ M/s}} = \left( \frac{0.056 \text{ M}}{0.0113 \text{ M}} \right)^x = 4.9 = (4.9)^x \Rightarrow x = 1
\]

Using Trial 1:

\[
K(0.0113 \text{ M})(0.0011 \text{ M}) = 1.9 \times 10^{-23} \text{ M/s}
\]

\[
K = 1.5 \times 10^{-18} \text{ M}^{-1} \text{s}^{-1}
\]

\[
1.5 \times 10^{-18} \text{ M}^{-1} \text{s}^{-1} \left[ \text{H}_2 \right] \left[ \text{I}_2 \right] = \text{rate}
\]
22. The rate of a reaction is expressed in terms of changes in concentration of reactants and products. Write the balanced chemical equation for the following the reaction described by the following mathematical equation:
\[ \text{Rate} = \frac{1}{2} \frac{\Delta [N_2O_5]}{\Delta t} = \frac{1}{4} \frac{\Delta [NO_3]}{\Delta t} = \frac{\Delta [O_2]}{\Delta t} \]
\[ 2N_2O_5 \rightarrow 4NO_2 + O_2 \]

23. Give the individual reaction orders for each substance and the overall reaction order for the following rate law:
\[ \text{Rate} = k \frac{[O_3]^2}{[O_2]} \]

Reaction Order for:
- a.) $O_3 = 2$
- b.) $O_2 = 1$
- c.) Overall reaction = 1

24. By what factor does the rate change for the rate in the previous problem change if:
- a. $[O_3]$ is doubled?
- b. $[O_2]$ is halved?
- c. $[O_2]$ is doubled?
- d. $[O_3]$ is halved?

25. Find the rate law, calculate the rate constant and determine the overall reaction order for the reaction $H_2 + I_2 \rightarrow 2 HI$ from the following data at 450°C. Show/explain your work using correct units and sig. figs.

<table>
<thead>
<tr>
<th>Experiment</th>
<th>Initial Rate (mol L(^{-1}) s(^{-1}))</th>
<th>Initial $[H_2]$ (mol/L)</th>
<th>Initial $[I_2]$ (mol/L)</th>
</tr>
</thead>
<tbody>
<tr>
<td>1</td>
<td>$1.9 \times 10^{-23}$</td>
<td>0.0113</td>
<td>0.0011</td>
</tr>
<tr>
<td>2</td>
<td>$1.1 \times 10^{-22}$</td>
<td>0.0220</td>
<td>0.0033</td>
</tr>
<tr>
<td>3</td>
<td>$9.3 \times 10^{-23}$</td>
<td>0.0550</td>
<td>0.0011</td>
</tr>
<tr>
<td>4</td>
<td>$1.9 \times 10^{-22}$</td>
<td>0.0220</td>
<td>0.0056</td>
</tr>
</tbody>
</table>

Rate Law: $k [H_2]^a [I_2]^b$

$\frac{\text{Rate}_3}{\text{Rate}_1} = \left(\frac{[H_2]_3}{[H_2]_1}\right)^a = \left(\frac{0.0056}{0.0113}\right)^a = \left(\frac{0.895}{1.0}\right)^a$

$\frac{\text{Rate}_2}{\text{Rate}_1} = \left(\frac{[I_2]_2}{[I_2]_1}\right)^b = \left(\frac{0.0033}{0.0056}\right)^b = \left(\frac{0.558}{1.0}\right)^b$

$a = 1.0$, $b = 0.967 \times 1$