For the reaction \( \text{J} \rightarrow \text{Products} \), the following initial rates were obtained for the respective initial concentrations of J:

<table>
<thead>
<tr>
<th>([\text{J}]_0 ) (M)</th>
<th>0.005</th>
<th>0.0082</th>
<th>0.017</th>
<th>0.03</th>
</tr>
</thead>
<tbody>
<tr>
<td>( V_0 ) (M s(^{-1}))</td>
<td>3.6E-7</td>
<td>9.6E-7</td>
<td>4.1E-6</td>
<td>1.3E-5</td>
</tr>
</tbody>
</table>

Determine the reaction’s order with respect to concentrations of J.
Solution: Create a log-log plot and determine the slope.

\[
\begin{array}{c|cccc}
\ln[\text{J}]_0 & -5.298 & -4.8036 & -4.0745 & -3.5066 \\
\ln V_0 & -14.837 & -13.856 & -12.405 & -11.251 \\
\end{array}
\]

The slope of this log-log plot is exactly 2.00; and so the order of the reaction with respect to J concentrations is 2: \( \text{Rate} = \frac{\text{d}[\text{J}]}{\text{d}t} = k[\text{J}]^2 \)

Experimental data are rarely this clean; and in general experimentally determined reaction orders rarely deliver whole integers. One generally rounds off to the nearest integer.
First Order Kinetics (semi-log plots).

For the reaction: \( \text{N}_2\text{O}_5(aq) + \text{Br}_2(l) \rightarrow \text{Products} \); the following data are obtained when dinitrogen pentoxide concentrations are followed with respect to time:

<table>
<thead>
<tr>
<th>Time (s)</th>
<th>0</th>
<th>200</th>
<th>400</th>
<th>600</th>
<th>1000</th>
</tr>
</thead>
<tbody>
<tr>
<td>([\text{N}_2\text{O}_5]) (M)</td>
<td>0.110</td>
<td>0.073</td>
<td>0.048</td>
<td>0.032</td>
<td>0.014</td>
</tr>
<tr>
<td>(\ln[\text{N}_2\text{O}_5])</td>
<td>-2.207</td>
<td>-2.617</td>
<td>-3.037</td>
<td>-3.440</td>
<td>-4.269</td>
</tr>
</tbody>
</table>

The slope of the plot is -0.00206 which equals \(-k\).
The rate constant for this reaction is 0.00206 s\(^{-1}\).

The half-life for the reaction is 0.693/k = 336 s.

**Apparently, this reaction does not depend on \(\text{Br}_2(l)\) conc\(^{ns}\).**
Homework example on initial rates.

The initial rate for the reaction $2\text{Fe}^{2+} + \text{Br}_2(l) \rightarrow 2\text{Fe}^{3+} + 2\text{Br}^-(l)$ was monitored at constant $[\text{Br}_2(l)]_0$ while varying $[\text{Fe}^{2+}]_0$. The following initial rate data was obtained:

<table>
<thead>
<tr>
<th>$[\text{Fe}^{2+}]_0$</th>
<th>0.030</th>
<th>0.040</th>
<th>0.050</th>
<th>0.060</th>
<th>0.080</th>
</tr>
</thead>
<tbody>
<tr>
<td>Initial Rate, M s$^{-1}$</td>
<td>0.0122</td>
<td>0.0161</td>
<td>0.0211</td>
<td>0.0240</td>
<td>0.0315</td>
</tr>
</tbody>
</table>

Evaluate reaction order with respect to ferrous ion concentrations.

To start on this problem, assume Initial Rate, $r_0 = k_0[\text{Fe}^{2+}]_0^x[\text{Br}_2]_0^y$

Then $\log r_0 = \log k_0 + x\log [\text{Fe}^{2+}]_0 + y\log [\text{Br}_2]_0$

Next, if one is told that $y = 1$; evaluate $k_0$. What are the units of $k_0$?

Another Method is to GUESS the order with respect to $[\text{Fe}^{2+}]_0$. Yes, it works!!!!

If, say the order (in this case, $x$) is 1.00; then there is a linear relationship between initial rate and initial ferrous ion concentrations, and a plot of Initial rate vs $[\text{Fe}^{2+}]_0^1$ will give a straight line.

If, on the other hand, $x = 2$ (second order in $[\text{Fe}^{2+}]_0$’s; then a plot of Initial rate vs $[\text{Fe}^{2+}]_0^2$ will give a straight line.