SO$_2$Cl$_2$ decomposes by first order kinetics and $k = 2.81 \times 10^{-3}$ min$^{-1}$ at a given temperature. The initial concentration of SO$_2$Cl$_2$ = 0.015 M. Determine the half-life of the reaction.

$$t_{1/2} = \frac{0.6931}{k} \quad t_{1/2} = \frac{1}{(k[A]_0)}$$

1. $t_{1/2} = \frac{0.6931}{2.81 \times 10^{-3} \text{ min}^{-1}} = 246.6$ min
2. $t_{1/2} = \frac{0.6931}{2.81 \times 10^{-3} \text{ min}^{-1}} = 247$ min
3. $t_{1/2} = \frac{1}{(2.81 \times 10^{-3} \text{ min}^{-1} (0.015))} = 2.37 \times 10^4$ min
4. $t_{1/2} = \frac{1}{(2.81 \times 10^{-3} \text{ min}^{-1} (0.015))} = 2.4 \times 10^4$ min
$\text{SO}_2\text{Cl}_2$ decomposes by first order kinetics and $k = 2.81 \times 10^{-3} \text{ min}^{-1}$ at a given temperature. The initial concentration of $\text{SO}_2\text{Cl}_2 = 0.015 \text{ M}$. Determine the half-life of the reaction.

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Select the correct rate law for step 2.

Step 2: \( \text{O}_2 + \text{N}_2\text{O}_2 \overset{k_2}{\rightarrow} \text{NO}_2 + \text{NO}_2 \)

1. \( \text{rate} = k_2[\text{O}_2][\text{N}_2\text{O}_2] \)
2. \( \text{rate} = \frac{k_2[\text{O}_2][\text{N}_2\text{O}_2]}{[\text{NO}_2]^2} \)
3. \( \text{rate} = k_2[\text{O}_2][\text{N}_2\text{O}_2] k^{-2}_2[\text{NO}_2]^2 \)
4. \( \text{rate} = \frac{k_2[\text{O}_2][\text{N}_2\text{O}_2]}{k^{-2}_2[\text{NO}_2]^2} \)
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Step 2: \[ \text{O}_2 + \text{N}_2\text{O}_2 \xrightarrow{k_2} \text{NO}_2 + \text{NO}_2 \]

1. rate = \( k_2[\text{O}_2][\text{N}_2\text{O}_2] \)
2. rate = \( k_2[\text{O}_2][\text{N}_2\text{O}_2]/[\text{NO}_2]^2 \)
3. rate = \( k_2[\text{O}_2][\text{N}_2\text{O}_2] k_{-2}[\text{NO}_2]^2 \)
4. rate = \( k_2[\text{O}_2][\text{N}_2\text{O}_2]/k_{-2}[\text{NO}_2]^2 \)
If the first step is slow and the second step is fast, $k_2[NO] \gg k_{-1}$.

1. $rate = k_1[NO][Br_2]$
2. $rate = (k_1 k_2 k_{-1})[NO]^2[Br_2]$
3. $rate = 2k_1[NO][Br_2]$
4. $rate = (2k_1 k_2 k_{-1})[NO]^2[Br_2]$
If the first step is slow and the second step is fast, \( k_2[\text{NO}] \gg k_{-1} \).

1. \( \text{rate} = k_1[\text{NO}][\text{Br}_2] \)

2. \( \text{rate} = (k_1 k_2 k_{-1})[\text{NO}]^2[\text{Br}_2] \)

3. \( \text{rate} = 2k_1[\text{NO}][\text{Br}_2] \)

4. \( \text{rate} = (2k_1 k_2 k_{-1})[\text{NO}]^2[\text{Br}_2] \)
rate = $k_{obs}([O_3]/[O_2])$

<table>
<thead>
<tr>
<th>Order in O$_2$?</th>
<th>If [O$_2$] is doubled/effect on rate?</th>
</tr>
</thead>
<tbody>
<tr>
<td>1. 0</td>
<td>no effect</td>
</tr>
<tr>
<td>2. 0</td>
<td>double</td>
</tr>
<tr>
<td>3. 1</td>
<td>double</td>
</tr>
<tr>
<td>4. 1</td>
<td>multiply by $\frac{1}{2}$</td>
</tr>
<tr>
<td>5. -1</td>
<td>double</td>
</tr>
<tr>
<td>6. -1</td>
<td>multiply by $\frac{1}{2}$</td>
</tr>
<tr>
<td>7. -1</td>
<td>multiply by -1</td>
</tr>
</tbody>
</table>
rate = \( k_{obs}(\frac{[O_3]}{[O_2]}) \)

<table>
<thead>
<tr>
<th>Order in ( O_2 )?</th>
<th>If ([O_2]) is doubled/effect on rate?</th>
</tr>
</thead>
<tbody>
<tr>
<td>2%</td>
<td>1. 0</td>
</tr>
<tr>
<td>2%</td>
<td>2. 0</td>
</tr>
<tr>
<td>10%</td>
<td>3. 1</td>
</tr>
<tr>
<td>6%</td>
<td>4. 1</td>
</tr>
<tr>
<td>3%</td>
<td>5. -1</td>
</tr>
<tr>
<td>73%</td>
<td>6. -1</td>
</tr>
<tr>
<td>4%</td>
<td>7. -1</td>
</tr>
</tbody>
</table>
If you double both $[O_3]$ and $[O_2]$, the rate will

1. not change.
2. decrease by half.
3. double.
4. triple.
5. quadruple.
If you double both \([O_3]\) and \([O_2]\), the rate will

1. not change.
2. decrease by half.
3. double.
4. triple.
5. quadruple.

![Bar chart with percentages]

- 1: 13%
- 2: 4%
- 3: 81%
- 4: 0%
- 5: 2%
5.111 Principles of Chemical Science
Fall 2014

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