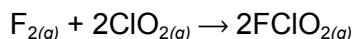


1. B

Explanation: Before you work with any chemical reaction, always make sure to balance the reaction first. Note that the presented reaction above is not balanced as written. After balancing, it will become:



Now that it is balanced, we can say that option A is incorrect because the rate of formation of FCIO_2 is twice the rate of disappearance of F_2 . Option C is also incorrect since the rate of disappearance of ClO_2 is twice the rate of disappearance of F_2 . Option B is the correct answer because it is true that the rate of disappearance of ClO_2 is equal to the rate of formation of FCIO_2 . You can try to solve this problem mathematically by using $\frac{\Delta F_2}{\Delta t} = -0.28 \text{ M/s}$. However, bear in mind that the negative sign only implies that the species is being consumed as the reaction progresses and does not indicate a negative rate. Therefore, we only compare the magnitude of the calculated rate and not the sign.

2. C

Explanation: We can use the method of initial rates to determine the rate law expression. To determine reaction order with respect to H_2 , we can use experiments 1 and 2, while experiments 2 and 3 can be used to determine reaction order with respect to NO. The general rate expression is $\text{rate} = k[\text{NO}]^x[\text{H}_2]^y$.

Reaction order with respect to NO	Reaction order with respect to H_2
$\frac{\text{rate}_2}{\text{rate}_3} = \frac{k [\text{NO}]_2^x [\text{H}_2]_2^y}{k [\text{NO}]_3^x [\text{H}_2]_3^y}$	$\frac{\text{rate}_2}{\text{rate}_1} = \frac{k [\text{NO}]_2^x [\text{H}_2]_2^y}{k [\text{NO}]_1^x [\text{H}_2]_1^y}$
<p>Since k and $[\text{H}_2]$ are constant, we can cancel these terms, leaving us with:</p>	<p>Since k and $[\text{NO}]$ are constant, we can cancel these terms, leaving us with:</p>

$\frac{rate_2}{rate_3} = \frac{[NO]_2^x}{[NO]_3^x}$	$\frac{rate_2}{rate_1} = \frac{[H_2]_2^y}{[H_2]_1^y}$
Substitution of known values yield:	Substitution of known values yield:
$\frac{4.68 \times 10^{-6}}{1.17 \times 10^{-6}} = \left(\frac{0.100}{0.050}\right)^x$ $4 = 2^x$ $x = 2$	$\frac{4.68 \times 10^{-6}}{2.34 \times 10^{-6}} = \left(\frac{0.032}{0.016}\right)^y$ $2 = 2^y$ $y = 1$

From these calculations, the final form of the rate law expression can be written as rate = $k[NO]^2[H_2]$ and the overall reaction order is 3.

3. B

Explanation: From calculation in number 2, we knew that the rate law expression is:

$$rate = k[NO]^2[H_2]$$

Let us set two scenarios here. We will double the concentration of NO while holding all the other parameters constant. For convenience, let us arbitrarily use 1 M and 2 M as the concentration of NO.

<p>[NO] = 1; [H₂] = 1; k is constant</p> $rate = k[NO]^2[H_2]$ $rate = k(1)^2(1)$ $rate = k$	<p>[NO] = 2; [H₂] = 1; k is constant</p> $rate = k[NO]^2[H_2]$ $rate = k(2)^2(1)$ $rate = 4k$
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Dividing the rate on the right by the rate on the left gives us 4, which means that the rate will be quadrupled if the concentration of NO is doubled. This makes sense because the reaction is second order with respect to the concentration of NO.

Take note that you will get the same result if you use other arbitrary concentrations of NO (if you want to confirm this, try using 2 and 4!). However, for simplicity of calculations, I recommend you use the lowest possible integer.

4. C

Explanation: We can use the same analysis utilized in number 3, although this time, we should use 2 and 1 as arbitrary values of $[H_2]$ (while keeping all the other variables constant) since the concentration of H_2 is halved.

$[H_2] = 2; [NO] = 1; k$ is constant	$[H_2] = 1; [NO] = 1; k$ is constant
rate = $k[NO]^2[H_2]$ rate = $k(1)^2(2)$ rate = $2k$	rate = $k[NO]^2[H_2]$ rate = $k(1)^2(1)$ rate = k

Dividing the rate in the right and the rate in the left gives us $\frac{1}{2}$, which implies that reducing the concentration of H_2 to half of its original concentration will cause the rate to be reduced to half of its original concentration.

5. C

Explanation: Option A is incorrect because the rate law expression stated is for a first-order reaction. For zeroth-order reaction, the rate law is simply rate = k . Option B seems correct, but it is actually false because as stated, t is plotted against $[A]_t$. If this is done, time will be the dependent variable, while $[A]_t$ will be the independent variable. To be correct for zeroth-order



Chemical Kinetics

Answer Key

reactions, the graph must plot the variable $[A]_t$ against t , and it must have the highest r^2 . Option C is correct because it states the correct characteristics of a zeroth-order reaction.



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To God be the glory!