

g. The rate law for the reaction of NO and Cl<sub>2</sub> at 250 K:  $2NO + Cl_2 \rightarrow 2NOCl$

Reactant Concentrations (mol/L)		Rate of Disappearance of NO <sub>2</sub> (mol/L·s)
[NO]	[Cl <sub>2</sub> ]	
0.000	0.000	$2.5 \times 10^{-3}$
0.000	0.000	$5.0 \times 10^{-3}$
0.000	0.000	$5.0 \times 10^{-3}$

- Determine the order of the reaction for each reactant.  
 NO is second order, because doubling its concentration quadruples the rate (page 1 of 2)  
 Cl<sub>2</sub> is first order, because doubling its concentration doubles the rate (page 1 of 2)
- Write the rate law for the reaction.

$$\text{rate} = k[NO]^2[Cl_2]$$

- Calculate the rate constant with units.  
 $\text{rate} = k[NO]^2[Cl_2]$   
 $2.5 \times 10^{-3} = k(0.000)^2(0.000)$   
 $k = 2.5 \times 10^6 \text{ L}^2/\text{mol}^2\text{s}$
- Calculate the rate at the instant when [NO] = 0.000 mol/L, and [Cl<sub>2</sub>] = 0.000 mol/L.  
 $\text{rate} = (2.5 \times 10^6) (0.000)^2 (0.000)$   
 $\text{rate} = 2.5 \times 10^{-3} \text{ s}^{-1} \rightarrow 2.5 \times 10^{-3} \text{ M/s}$
- Calculate the [NO] when [Cl<sub>2</sub>] = 0.000 mol/L, during the second trial.  
 Find the [NO] using  $[NO] = 0.000 - 0.000 = 0.000$   
 Now, use the MOLE BALANCE to find the CHANGE in [NO] = 0.000M Cl<sub>2</sub> = (same amount) Cl<sub>2</sub> = 0.000M NO  
 Find the [NO] concentration:  $0.000 - 0.000 = 0.000\text{M}$
- Calculate the rate under the conditions given in "a."  
 $\text{rate} = (2.5 \times 10^6) (0.000)^2 (0.000)$   
 $\text{rate} = 2.5 \times 10^{-3} \text{ s}^{-1} \rightarrow 2.5 \times 10^{-3} \text{ M/s}$
- If the first trial is repeated at 700K with the exact same concentrations, would the initial rate still be  $2.5 \times 10^{-3}$ ? Explain your answer.  
 No. It was proved the temperature is a different suggestion, the rate constant will change. At a higher temperature, the rate constant will be larger, changing the rate of the reaction.

28. Carbon monoxide reacts with O<sub>2</sub> to form CO<sub>2</sub>.



[CO] (mol/L)	[O <sub>2</sub> ] (mol/L)	Rate (mol/L·s)
0.01	0.01	$5.00 \times 10^{-3}$
0.02	0.01	$1.00 \times 10^{-2}$
0.01	0.02	$1.00 \times 10^{-2}$

- What is the rate law for this reaction?  
 $\text{rate} = k[CO]^2[O_2]$
- What is the order of the reaction with respect to CO?  
 second order
- What is the order with respect to O<sub>2</sub>?  
 first order
- What is the overall order of the reaction?  
 third order
- What is the value of the rate constant, k, with units?  
 $\text{rate} = k[CO]^2[O_2]$   
 $5.00 \times 10^{-3} = k(0.01)^2(0.01)$   
 $k = 4.0 \times 10^6 \text{ L}^3/\text{mol}^3\text{s}$
- During trial 2, at the exact instant that [CO] = 0.000 M, what is the [O<sub>2</sub>?  
 Find the [O<sub>2</sub>] using  $[O_2] = 0.01 - 0.000 = 0.000$   
 Now, use the MOLE BALANCE to find the CHANGE in [O<sub>2</sub>] = 0.000M O<sub>2</sub> = (same amount) O<sub>2</sub> = 0.000M CO  
 Find the [O<sub>2</sub>] concentration:  $0.000 - 0.000 = 0.000\text{M}$
- Calculate the rate at the exact instant described in "f."  
 $\text{rate} = (4.0 \times 10^6) (0.000)^2 (0.000)$   
 $\text{rate} = 0.000 \text{ mol/L}^3/\text{s}$   
 $\text{rate} = 0.000 \text{ s}^{-1} \rightarrow 0.0 \times 10^{-3} \text{ M/s}$

# Ap Chemistry Kinetics Lab Answers

**Lei Shi**



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