

g. The data in the table are for the reaction of NO and Cl₂ at 260 K: $2\text{NO} + \text{Cl}_2 \rightarrow 2\text{NOCl}$

Reactant Concentrations (mol/L)		Rate of Disappearance of NO ₂ (mol/L·s)
[NO]	[Cl ₂]	
0.000	0.000	2.5×10^{-3}
0.000	0.000	5.0×10^{-3}
0.000	0.000	5.0×10^{-3}

- a. Determine the order of the reaction for each reactant.

NO is second order, because doubling its concentration quadruples the rate (page 1 & 2)

Cl₂ is first order, because doubling its concentration doubles the rate (page 1 & 2)

- b. Write the rate law for the reaction.

$$\text{rate} = k[\text{NO}]^2[\text{Cl}_2]$$

- c. Calculate the rate constant with units.

$$\text{rate} = k[\text{NO}]^2[\text{Cl}_2]$$

$$2.5 \times 10^{-3} = k(0.000)^2(0.000)$$

$$k = 2.5 \times 10^6 \text{ s}^{-1}$$

- d. Calculate the rate at the instant when [NO] = 0.000 mol/L, and [Cl₂] = 0.000 mol/L.

$$\text{rate} = (2.5 \times 10^6 \text{ s}^{-1})(0.000)^2(0.000)$$

$$\text{rate} = 2.5 \times 10^{-3} \text{ s}^{-1} \rightarrow 2.5 \times 10^6 \text{ M/s}$$

- e. Calculate the [NO] when [Cl₂] = 0.000 mol/L, during the second trial.

$$\text{Find the CHANGING: } [\text{Cl}_2] = 0.000 - 0.000 = 0.000$$

$$\text{Now, use the INITIAL-FINAL to find the CHANGING: } [\text{NO}] = 0.000 \text{ M } \text{Cl}_2 \text{ is (used 100\% and } [\text{Cl}_2] = 0.000 \text{ M NO}$$

$$\text{Find the FINAL concentration: } 0.000 - 0.000 = 0.000 \text{ M}$$

- f. Calculate the rate under the conditions given in "e."

$$\text{rate} = (2.5 \times 10^6 \text{ s}^{-1})(0.000)^2(0.000)$$

$$\text{rate} = 2.5 \times 10^{-3} \text{ s}^{-1} \rightarrow 2.5 \times 10^6 \text{ M/s}$$

- g. If the first trial is repeated at 700K, with the exact same concentrations, would the initial rate still be 2.5×10^6 ? Explain your answer.

No. It was repeat the experiment at a different temperature, the rate constant will change. At a higher temperature, the rate constant will be larger, causing the rate to be faster.

28. Carbon monoxide reacts with O₂ to form CO₂.



[CO] (mol/L)	[O ₂] (mol/L)	Rate (mol/L·s)
0.00	0.00	5.00×10^{-3}
0.04	0.00	1.00×10^{-3}
0.00	0.04	1.00×10^{-3}

- a. What is the rate law for this reaction?

$$\text{rate} = k[\text{CO}]^2[\text{O}_2]$$

- b. What is the order of the reaction with respect to O₂?

second order

- c. What is the order with respect to CO?

first order

- d. What is the overall order of the reaction?

third order

- e. What is the value of the rate constant, k, with units?

$$\text{rate} = k[\text{CO}]^2[\text{O}_2]$$

$$5.00 \times 10^{-3} = k(0.00)^2(0.00)$$

$$k = 4.0 \times 10^6 \text{ s}^{-1}$$

- f. During trial 3, at the exact instant that [O₂] = 0.000 M, what is the [CO]?

$$\text{Find the CHANGING: } [\text{O}_2] = 0.000 - 0.000 = 0.000$$

$$\text{Now, use the INITIAL-FINAL to find the CHANGING: } [\text{CO}] = 0.000 \text{ M } \text{O}_2 \text{ is (used 100\% and } [\text{O}_2] = 0.000 \text{ M CO}$$

$$\text{Find the FINAL concentration: } 0.000 - 0.000 = 0.000 \text{ M}$$

- g. Calculate the rate at the exact instant described in "f."

$$\text{rate} = k[\text{CO}]^2[\text{O}_2]$$

$$\text{rate} = (4.0 \times 10^6 \text{ s}^{-1})(0.000)^2(0.000)$$

$$\text{rate} = 4.0 \times 10^{-3} \text{ s}^{-1} \rightarrow 4.0 \times 10^6 \text{ M/s}$$

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A red circular graphic with a gradient, appearing as a partial circle or a stylized arrow pointing to the right, located to the right of the Carolina Biological Supply Company text.

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