# *Math Tutor* WRITING RATE LAWS

Factors such as surface area and temperature affect the rate of reactions because they affect the frequency and energy of collisions between particles. The concentrations of reactants can also affect the frequency of collisions. If other factors are kept constant, the rates of most chemical reactions will be determined by the concentrations of reactants. Thus, it is possible to write an equation called a *rate law* that relates the rate of a reaction to the concentrations of reactants.

## SAMPLE

**Fluorine gas reacts with chlorine dioxide according to the following equation.**

$$F_2(g) + 2ClO_2(g) \longrightarrow 2FClO_2(g)$$

**Use the following experimental data to write a rate law for this reaction.**

| Trial | Concentration of $F_2$ | Concentration of $ClO_2$ | Rate (mol/L•s) |
|---|---|---|---|
| 1 | 0.10 *M* | 0.10 *M* | $1.1 \times 10^{-3}$ |
| 2 | 0.20 *M* | 0.10 *M* | $2.2 \times 10^{-3}$ |
| 3 | 0.10 *M* | 0.20 *M* | $2.2 \times 10^{-3}$ |
| 4 | 0.20 *M* | 0.20 *M* | $4.4 \times 10^{-3}$ |

To write the rate law, first examine the data to see how the rate of reaction changes as the concentrations of the reactants change.

- When $[F_2]$ doubles and $[ClO_2]$ remains constant, the rate of reaction doubles from $1.1 \times 10^{-3}$ mol/L•s to $2.2 \times 10^{-3}$ mol/L•s. So, the rate is directly proportional to $[F_2]$, or $R \propto [F_2]$.
- When $[ClO_2]$ doubles and $[F_2]$ remains constant, the rate of reaction also doubles from $1.1 \times 10^{-3}$ mol/L•s to $2.2 \times 10^{-3}$ mol/L•s. So, the rate is directly proportional to $[ClO_2]$, or $R \propto [ClO_2]$.
- Because rate is proportional to both $[F_2]$ and $[ClO_2]$, you can write the rate law $R = k[F_2][ClO_2]$. The data from Trial 4 help confirm the rate law because when both $[F_2]$ and $[ClO_2]$ double, the rate increases by a factor of four, from $1.1 \times 10^{-3}$ mol/L•s to $4.4 \times 10^{-3}$ mol/L•s.

## PRACTICE PROBLEMS

**1.** Nitrogen monoxide and oxygen react to produce nitrogen dioxide according to the following equation:

$$O_2(g) + 2NO(g) \longrightarrow 2NO_2(g)$$

Use the data in the following table to write a rate law for this reaction.

| Trial | $[O_2]$ | $[NO]$ | Reaction Rate (mol/L•s) |
|---|---|---|---|
| 1 | $1.20 \times 10^{-2}$ *M* | $1.40 \times 10^{-2}$ *M* | $3.30 \times 10^{-3}$ |
| 2 | $2.40 \times 10^{-2}$ *M* | $1.40 \times 10^{-2}$ *M* | $6.60 \times 10^{-3}$ |
| 3 | $1.20 \times 10^{-2}$ *M* | $2.80 \times 10^{-2}$ *M* | $1.32 \times 10^{-2}$ |

**2.** Hydrogen reacts with ethyne, $C_2H_2$, to produce ethane, $C_2H_6$, as shown below:

$$2H_2(g) + C_2H_2(g) \longrightarrow C_2H_6(g)$$

Use the data in the following table to write a rate law for this reaction.

| Trial | $[H_2]$ | $[C_2H_2]$ | Reaction Rate (mol/L•min) |
|---|---|---|---|
| 1 | 0.20 *M* | 0.20 *M* | $1.5 \times 10^{-4}$ |
| 2 | 0.40 *M* | 0.20 *M* | $3.0 \times 10^{-4}$ |
| 3 | 0.20 *M* | 0.40 *M* | $1.5 \times 10^{-4}$ |