$$2NO_2(g) + F_2(g) \rightleftarrows 2NO_2F(g)$$

$$R_f = k_f[NO_2][F_2]$$

$$K_c = \frac{[NO_2F]_{eq}^2}{[NO_2]_{eq}^2[F_2]_{eq}}$$

$$\begin{array}{ll} \text{Step 1} & NO_2(g) + F_2(g) \rightleftarrows NO_2F(g) + \cancel{F(g)} \\ \text{Step 2} & NO_2(g) + \cancel{F(g)} \rightleftarrows NO_2F(g) \\ \hline & 2NO_2(g) + F_2(g) \rightleftarrows 2NO_2F(g) \end{array}$$

$$K_{c_1} = \frac{[NO_2F]_{eq}[F]_{eq}}{[NO_2]_{eq}[F_2]_{eq}}$$

$$K_{c_2} = \frac{[NO_2F]_{eq}}{[NO_2]_{eq}[F]_{eq}}$$

$$\underbrace{K_{c_1} \cdot K_{c_2}}_{\downarrow} = \frac{[NO_2F]_{eq}\cancel{[F]_{eq}}}{[NO_2]_{eq}[F_2]_{eq}} \cdot \frac{[NO_2F]_{eq}}{[NO_2]_{eq}\cancel{[F]_{eq}}}$$

$$K_c = \frac{[NO_2F]_{eq}^2}{[NO_2]_{eq}^2[F_2]_{eq}}$$

Manipulating the Equilibrium Constant

1. When reactions are added, their equilibrium constants are multiplied.

2. When you multiply a reaction by a constant, you raise the equilibrium constant to that power.

3. When you reverse the direction of a reaction, you take the reciprocal of its equilibrium constant.

$$2A \rightleftharpoons B + C$$

$$K_c = \frac{[B]_{eq}[C]_{eq}}{[A]_{eq}^{2}}$$

$$4A \rightleftharpoons 2B + 2C$$

$$K_c' = \frac{[B]_{eq}^{2}[C]_{eq}^{2}}{[A]_{eq}^{4}}$$

$$= \left(\frac{[B]_{eq}[C]_{eq}}{[A]_{eq}^{2}}\right)^{2}$$

$$= \left(K_c\right)^{2}$$

$$\sqrt{x} \cdot \sqrt{x} = x$$

$$x^{\frac{1}{2}} \cdot x^{\frac{1}{2}} = x^{\frac{1}{2}+\frac{1}{2}}$$

$$= x^{1} = x$$

$$2A \rightleftarrows B + C$$

$$K_c = \frac{[B]_{eq}[C]_{eq}}{[A]_{eq}^2}$$

$$B + C \rightleftarrows 2A$$

$$K_c'' = \frac{[A]_{eq}^2}{[B]_{eq}[C]_{eq}}$$

$$= \frac{1}{\frac{[B]_{eq}[C]_{eq}}{[A]_{eq}^2}}$$

$$= \frac{1}{K_c}$$

$$2A(g) \rightleftarrows B(g) + C(g)$$

We start with pure A at a concentration of 1.50 mol L^{-1}. At equilibrium, the concentration of substance B is found to be 0.50 mol L^{-1}. What is the numerical value of K_c for this reaction?