10.1: The Concept of Equilibrium Reactions

Pure dinitrogen tetroxide (N_2O_4) is a colorless gas that is widely used as a rocket fuel. Although N_2O_4 is colorless, when a container is filled with pure N_2O_4, the gas rapidly begins to turn a dark brown. A chemical reaction is clearly occurring, and indeed, chemical analysis tells us that the gas in the container is no longer pure N_2O_4, but has become a mixture of dinitrogen tetroxide and nitrogen dioxide; N_2O_4 is undergoing a *decomposition reaction* to form NO_2. If the gaseous mixture is cooled, it again turns colorless and analysis tells us that it is again, almost pure N_2O_4; this means that the NO_2 in the mixture can also undergo a *synthesis reaction* to re-form N_2O_4. Initially, only N_2O_4 is present. As the reaction proceeds, the concentration of N_2O_4 decreases and the concentration of NO_2 increases. However, if you examine the figure, after some time, the concentrations of N_2O_4 and NO_2 have *stabilized* and, as long as the temperature is not changed, the relative concentrations of the two gasses remain constant.

The reversible reaction of one mole of N_2O_4, forming two moles of NO_2, is a classic example of a **chemical equilibrium**. We encountered the concept of equilibrium in Chapter 9 when we dealt with the autoprotolysis of water to form the hydronium and hydroxide ions, and with the dissociation of weak acids in aqueous solution.

$$2\,H_2O \rightleftharpoons H_3O^+ + HO^-$$

When we wrote these chemical equations, we used a **double arrow** to signify that the reaction proceeded in both directions. Using this convention, the dissociation of dinitrogen tetroxide to form two molecules of nitrogen dioxide can be shown as:

$$N_2O_4 \rightleftarrows 2\,NO_2$$

If the temperature of our gas mixture is again held constant and the total pressure of the gas in the container is varied, analysis shows that the partial pressure of N_2O_4 varies as the *square* of the partial pressure of NO_2. The Ideal Gas Laws tell us that the partial pressure of a gas, P_{gas}, is directly proportional to the *concentration* of that gas in the container). Mathematically, the relationship between the partial pressures of the two gasses can be expressed by the equation below:

$$\frac{(P_{NO_2})^2}{P_{N_2O_4}} = K$$