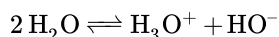


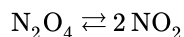
## 10.1: The Concept of Equilibrium Reactions

Pure dinitrogen tetroxide ( $\text{N}_2\text{O}_4$ ) is a colorless gas that is widely used as a rocket fuel. Although  $\text{N}_2\text{O}_4$  is colorless, when a container is filled with pure  $\text{N}_2\text{O}_4$ , 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  $\text{N}_2\text{O}_4$ , but has become a mixture of dinitrogen tetroxide and nitrogen dioxide;  $\text{N}_2\text{O}_4$  is undergoing a *decomposition reaction* to form  $\text{NO}_2$ . If the gaseous mixture is cooled, it again turns colorless and analysis tells us that it is again, almost pure  $\text{N}_2\text{O}_4$ ; this means that the  $\text{NO}_2$  in the mixture can also undergo a *synthesis reaction* to re-form  $\text{N}_2\text{O}_4$ . Initially, only  $\text{N}_2\text{O}_4$  is present. As the reaction proceeds, the concentration of  $\text{N}_2\text{O}_4$  decreases and the concentration of  $\text{NO}_2$  increases. However, if you examine the figure, after some time, the concentrations of  $\text{N}_2\text{O}_4$  and  $\text{NO}_2$  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  $\text{N}_2\text{O}_4$ , forming two moles of  $\text{NO}_2$ , 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.



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:



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  $\text{N}_2\text{O}_4$  varies as the *square* of the partial pressure of  $\text{NO}_2$ . The Ideal Gas Laws tell us that the partial pressure of a gas,  $P_{\text{gas}}$ , 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:

$$\frac{(P_{\text{NO}_2})^2}{P_{\text{N}_2\text{O}_4}} = K$$

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