FLINN SCIENTIFIC

Exploring Equilibrium

It Works Both Ways

Introduction

The word *equilibrium* has two roots: *aequi*, meaning *equal*, and *libra*, meaning *weight* or *balance*. Our physical sense of equilibrium—in the notion of a stream or the swing of a pendulum—suggests an equal balance of opposing forces. How does this physical sense of equilibrium translate to chemical equilibrium? Explore the nature and consequences of equilibrium in chemical reactions.

Concepts

- Reversible reactions
- Chemical equilibrium
- Complex-ion reaction
- Acid–base indicators

Background

Physical changes, such as melting ice or dissolving sugar, are often recognized by noting that these processes can be easily reversed. Some common examples of chemical change, such as burning wood or spoiling food, generally cannot be reversed. A closer look at chemical change, however, reveals that many chemical reactions are also reversible.

Consider the following example of a reversible chemical reaction (Equation 1). At high pressures and in the presence of a catalyst, nitrogen and hydrogen react to form ammonia. If the temperature is high enough, however, ammonia decomposes to reform its constituent elements. The reaction can go both ways! This reversible reaction is represented symbolically using double arrows.

$$N_2(g) + 3H_2(g) \rightleftharpoons 2NH_3(g) \qquad \textit{Equation 1}$$

What happens when nitrogen and hydrogen are allowed to react? In a closed system, the concentrations of nitrogen and hydrogen will decrease, and the concentration of ammonia will steadily increase as the reaction proceeds in the forward direction. Soon, however, the concentration of ammonia will be large enough that the reverse reaction will begin to take place at a significant rate as well. Eventually, as the reaction occurs in both the forward and the reverse directions, the number of ammonia molecules being formed will become equal to the number of ammonia molecules being consumed. At this point, no further changes will be observed in the overall concentrations of nitrogen, hydrogen and ammonia. This is the point of chemical equilibrium.

Chemical equilibrium is defined as the state in which the rate of the forward reaction equals the rate of the reverse reaction and the concentrations of reactants and products remain constant with time. Note that this definition describes a dynamic picture of equilibrium. The reactions continue, but there is an equal balance of opposing reaction rates.

What happens when the equilibrium is disturbed? Any factor that changes the rate of the forward or the reverse reaction will change the amounts of reactants and products that are present at equilibrium. Reaction conditions that are known to affect the rates of chemical reactions include the concentrations of reactants and the temperature. In this experiment, we will investigate how changes in reaction conditions affect the amounts of reactants and products present at equilibrium.

Experiment Overview

The purpose of this experiment is to explore the nature of chemical equilibrium and to identify conditions that affect the position of equilibrium. Two different reversible reactions will be studied.

Reaction of iron(III) nitrate with potassium thiocyanate will be used to study *complex-ion equilibrium*. Iron(III) ions react with thiocyanate ions to form $FeSCN^{2+}$ complex ions (Equation 2). The effects of changing the concentrations of reactants and of changing the reaction temperature will be investigated.

$$Fe^{3+}(aq) + SCN^-(aq) \rightleftharpoons FeSCN^{2+}(aq) \qquad \textit{Equation 2}$$

The properties of an indicator will be used to study *acid–base equilibrium*. An indicator is a dye that can gain or lose hydrogen ions to form substances that have different colors. Equation 3 summarizes the reversible reaction of the indicator bromcresol green (HIn). HIn represents an uncharged indicator molecule and In^- an indicator anion formed after the molecule has lost a hydrogen ion. The color of the indicator in the presence of either excess H^+ or OH^- ions (see Equation 4) will show how changing the concentration of a product affects the equilibrium shown in Equation 3.

$$\underset{\text{Color A}}{HIn(aq)} \rightleftharpoons \underset{\text{Color B}}{H^+(aq) + In^-(aq)} \qquad \textit{Equation 3}$$

$$H^+(aq) + OH^-(aq) \rightarrow H_2O(l) \qquad \textit{Equation 4}$$

Exploring Equilibrium It Works Both Ways Lab

Clemens Wendtner

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