

7.1: The Concept of Dynamic Equilibrium

Skills to Develop

- To understand what is meant by chemical equilibrium.

Two Methods to Describe the Likely Result of a Chemical Reaction

In the previous chapter, we discussed the principles of chemical kinetics, which deal with the *rate of change*, or how quickly a given chemical reaction occurs. We now turn our attention to the *extent* to which a reaction occurs and how reaction conditions affect the final concentrations of reactants and products. We will learn two "languages" to describe the extent of a reaction. First we will describe the final, likely state of a chemical reaction in terms of the ratio of the amount of products to the amount of reactants at equilibrium. Then we will apply the principles of **thermodynamics** to show how energy flow determines the likely amount of product made by a reaction when it reaches dynamic equilibrium.

Equilibrium

For most of the reactions that we have discussed so far, you may have assumed that once reactants are converted to products, they are likely to remain that way. In fact, however, virtually all chemical reactions are *reversible* to some extent. That is, an opposing reaction occurs in which the products react, to a greater or lesser degree, to re-form the reactants. Eventually, the forward and reverse reaction rates become the same, and the system reaches **chemical equilibrium**, the point at which the composition of the system no longer changes with time.



Figure 7.1.1: Dinitrogen tetroxide is a powerful oxidizer that reacts spontaneously upon contact with various forms of hydrazine, which makes the pair a popular propellant combination for rockets. Nitrogen dioxide at $-196\text{ }^{\circ}\text{C}$, $0\text{ }^{\circ}\text{C}$, $23\text{ }^{\circ}\text{C}$, $35\text{ }^{\circ}\text{C}$, and $50\text{ }^{\circ}\text{C}$. (NO_2) converts to the colorless dinitrogen tetroxide (N_2O_4) at low temperatures, and reverts to NO_2 at higher temperatures. Figure used with permission from Wikipedia (CC BY-SA 3.0; [Eframgoldberg](#)).

Chemical equilibrium is a dynamic process that consists of a forward reaction, in which reactants are converted to products, and a reverse reaction, in which products are converted to reactants. At equilibrium, the forward and reverse reactions proceed at equal rates. Consider, for example, a simple system that contains only one reactant and one product, the reversible dissociation of dinitrogen tetroxide (N_2O_4) to nitrogen dioxide (NO_2). You may recall that NO_2 is responsible for the brown color we associate with smog. When a sealed tube containing solid N_2O_4 (mp = -9.3°C ; bp = 21.2°C) is heated from -78.4°C to 25°C , the red-brown color of NO_2 appears (Figure 7.1.1). The reaction can be followed visually because the product (NO_2) is colored, whereas the reactant (N_2O_4) is colorless:



The double arrow indicates that both the forward reaction



and reverse reaction



occurring simultaneously (i.e, the reaction is reversible). However, this does not necessarily mean the system is equilibrium, as section 7. 4 demonstrates.

Figure 7.1.2 shows how the composition of this system would vary as a function of time at a constant temperature. If the initial concentration of NO_2 were zero, then it increases as the concentration of N_2O_4 decreases. Eventually the composition of the system stops changing with time, and chemical equilibrium is achieved. Conversely, if we start with a sample that contains no N_2O_4 but an initial NO_2 concentration twice the initial concentration of N_2O_4 (Figure 7.1.2a), in accordance with the stoichiometry of the reaction, we reach exactly the same equilibrium composition (Figure 7.1.2b). Thus equilibrium can be approached from either direction in a chemical reaction.

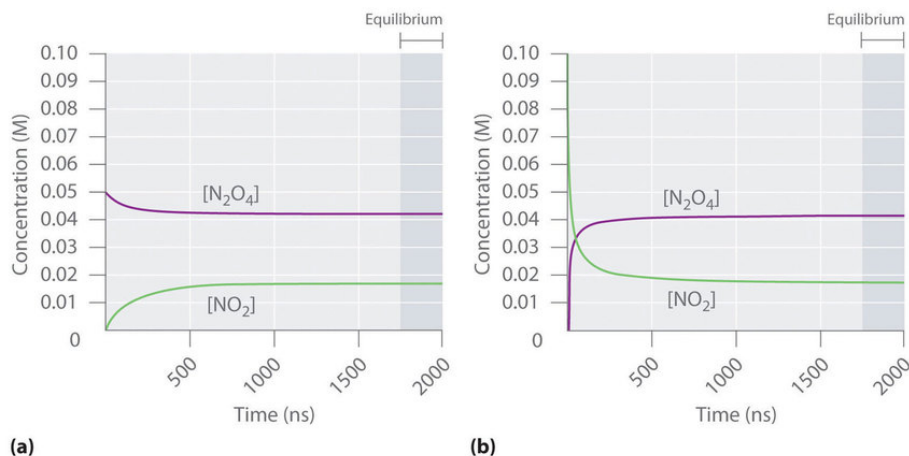


Figure 7.1.2: The Composition of $\text{N}_2\text{O}_4/\text{NO}_2$ Mixtures as a Function of Time at Room Temperature. (a) Initially, this idealized system contains 0.0500 M gaseous N_2O_4 and no gaseous NO_2 . The concentration of N_2O_4 decreases with time as the concentration of NO_2 increases. (b) Initially, this system contains 0.1000 M NO_2 and no N_2O_4 . The concentration of NO_2 decreases with time as the concentration of N_2O_4 increases. In both cases, the final concentrations of the substances are the same: $[\text{N}_2\text{O}_4] = 0.0422 \text{ M}$ and $[\text{NO}_2] = 0.0156 \text{ M}$ at equilibrium.

Figure 7.1.3 shows the forward and reverse reaction rates for a sample that initially contains pure NO_2 . Because the initial concentration of N_2O_4 is zero, the forward reaction rate (dissociation of N_2O_4) is initially zero as well. In contrast, the reverse reaction rate (dimerization of NO_2) is initially very high ($2.0 \times 10^6 \text{ M/s}$), but it decreases rapidly as the concentration of NO_2 decreases. As the concentration of N_2O_4 increases, the rate of dissociation of N_2O_4 increases. Eventually, the forward and reverse reaction rates become identical, and the system has reached chemical equilibrium. If the forward and reverse reactions occur at different rates, then the system is not at equilibrium.

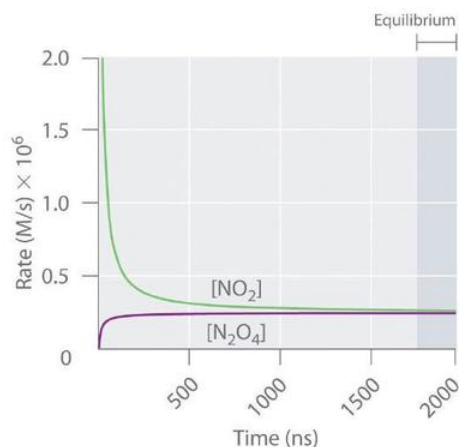


Figure 7.1.3: The Forward and Reverse Reaction Rates as a Function of Time for the $N_2O_{4(g)} \rightleftharpoons 2NO_{2(g)}$ System Shown in Part (b) in Figure 7.1.2

Initially, the rate of dimerization of NO_2 (reverse reaction) is quite rapid, but this rate decreases over time as the concentration of NO_2 decreases. Because the initial concentration of N_2O_4 is zero, the rate of the dissociation reaction (forward reaction) at $t = 0$ is also zero. As the dimerization reaction proceeds, the N_2O_4 concentration increases, and its rate of dissociation also increases. Eventually the rates of the two reactions are equal: chemical equilibrium has been reached, and the concentrations of N_2O_4 and NO_2 no longer change.

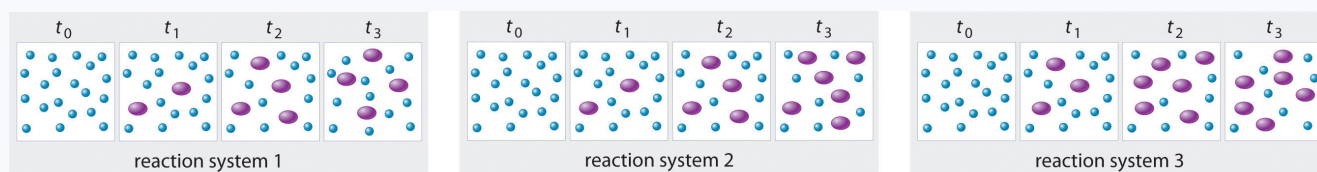
At equilibrium, the forward reaction rate is equal to the reverse reaction rate.

Example 7.1.1

The three reaction systems (1, 2, and 3) depicted in the accompanying illustration can all be described by the equation:



where the blue circles are A and the purple ovals are B . Each set of panels shows the changing composition of one of the three reaction mixtures as a function of time. Which system took the longest to reach chemical equilibrium?



Given: three reaction systems

Asked for: relative time to reach chemical equilibrium

Strategy:

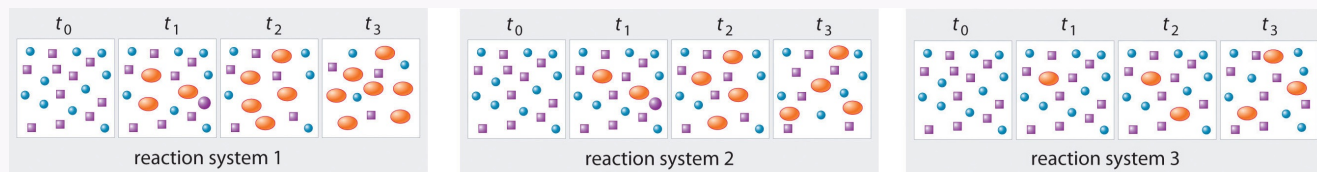
Compare the concentrations of A and B at different times. The system whose composition takes the longest to stabilize took the longest to reach chemical equilibrium.

Solution:

In systems 1 and 3, the concentration of A decreases from t_0 through t_2 but is the same at both t_2 and t_3 . Thus systems 1 and 3 are at equilibrium by t_3 . In system 2, the concentrations of A and B are still changing between t_2 and t_3 , so system 2 may not yet have reached equilibrium by t_3 . Thus system 2 took the longest to reach chemical equilibrium.

Exercise 7.1.1

In the following illustration, A is represented by blue circles, B by purple squares, and C by orange ovals; the equation for the reaction is $A + B \rightleftharpoons C$. The sets of panels represent the compositions of three reaction mixtures as a function of time. Which, if any, of the systems shown has reached equilibrium?



Answer

system 2

Summary

At equilibrium, the forward and reverse reactions of a system proceed at equal rates. Chemical equilibrium is a dynamic process consisting of forward and reverse reactions that proceed at equal rates. At equilibrium, the composition of the system no longer

changes with time. The composition of an equilibrium mixture is independent of the direction from which equilibrium is approached.

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