7.1 Equilibrium Systems

When doing stoichiometry calculations, we generally assume that chemical reactions proceed until one of the reactants runs out. At this point, we say that the reaction has gone to completion. This is true for most chemical reactions. However, some chemical reactions do not proceed to completion. **Figure 1** shows the changes that take place when nitrogen dioxide gas, $NO_2(g)$, is placed in a sealed, evacuated chamber at 25 °C. Nitrogen dioxide gas reacts to form dinitrogen tetroxide gas, $N_2O_4(g)$, as is represented by the equation

$$2\ NO_2(g) \rightarrow N_2O_4(g)$$

The sealed chamber ensures that the reaction takes place in a closed chemical reaction system. In a closed system, energy may transfer from the surroundings to the chemical system, or from the chemical system to the surroundings, but matter cannot. (You encountered this term in Chapter 5.) Nitrogen dioxide is a dark brown gas, whereas dinitrogen tetroxide gas is colourless, so the colour in the chamber provides visual evidence of their relative quantities. The initial dark brown colour of the gas decreases in intensity as it is converted to colourless dinitrogen tetroxide gas.

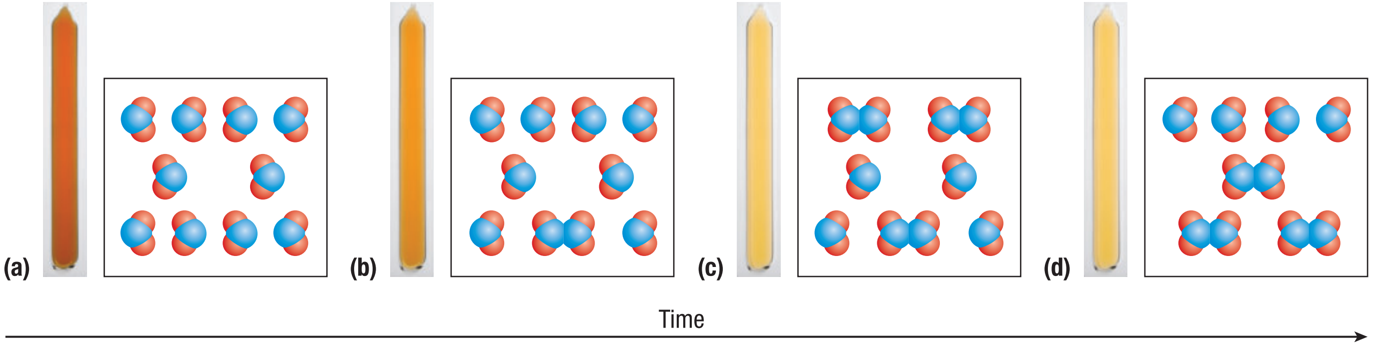

Figure 1 (a) Initially, the vial contains only molecules of brown nitrogen dioxide gas. In the vial shown in (b) some of the nitrogen dioxide gas has been converted to dinitrogen tetroxide gas, which is colourless. Eventually, in (c) and (d), an equilibrium is established, so the gas remains the same colour.

However, even over a long time, the intensity of the brown colour eventually remains constant. The gas never becomes completely colourless. This indicates that the concentration of nitrogen dioxide is no longer changing. The reaction has clearly stopped short of completion. When the concentrations of all reactants and products of a chemical reaction remain constant over time, we say that this system has reached **chemical equilibrium**. Any chemical reaction carried out in a closed system will eventually reach chemical equilibrium. All of the reactions explored in this chapter take place in closed systems.

chemical equilibrium the state of a reaction in which all reactants and products have reached constant concentrations in a closed system

The Equilibrium Condition

All chemical equilibria are dynamic equilibria. A **dynamic equilibrium** is an equilibrium in which the rates of forward and reverse processes are equal. To help picture a dynamic equilibrium, imagine 18 000 people attending a sporting event. Most of them are in their seats for most of the time, but there are always about 400 people at the concession stands buying snacks. This means that on average 17 600 people are in their seats. Throughout the event, people continuously move back and forth from the seats to the concessions but the total number in each area remains constant. Since it is a closed system, we can therefore describe this as a closed system in dynamic equilibrium.

dynamic equilibrium a balance between forward and reverse processes that are occurring simultaneously at equal rates

Why is a chemical equilibrium always a dynamic equilibrium? To answer this, consider the chemical reaction between steam, $H_2O(g)$, and carbon monoxide gas, $CO(g)$, in a closed vessel at a high temperature. The balanced equation for this reaction is

$$H_2O(g) + CO(g) \rightarrow H_2(g) + CO_2(g)$$

Assume that we start with equal amounts of carbon monoxide gas and water vapour. From the balanced equation, we know these substances react in a 1:1 molar ratio. Therefore, if we start off with equal concentrations, the concentrations of these two gases will always be equal. Similarly, since hydrogen gas, $H_2(g)$, and carbon dioxide gas, $CO_2(g)$, are formed in equal amounts, the concentrations of these gases will also always be equal.

Under the conditions of a closed vessel at a high temperature, the chemical reaction begins immediately and progresses very quickly. **Figure 2** shows the changes in the concentrations of reactants and products of this reaction over time. Notice how, as the concentrations of the products increase, the concentrations of the reactants decrease. At the time indicated by the dashed vertical line in Figure 2, the concentrations of reactants and products remain stable. This is the point at which chemical equilibrium has been reached. No further changes in the concentrations of reactants or products will occur (unless the chemical system is changed in some way, as you will see in Section 7.2). WEB LINK

LEARNING TIP

Concentration Notation

By convention, when the formula of a substance is placed between square brackets, we are referring to the concentration of that substance. For example, the concentration of carbon dioxide gas is labelled $[CO_2(g)]$ in Figure 2. Concentration also may be denoted by the symbol *c*.

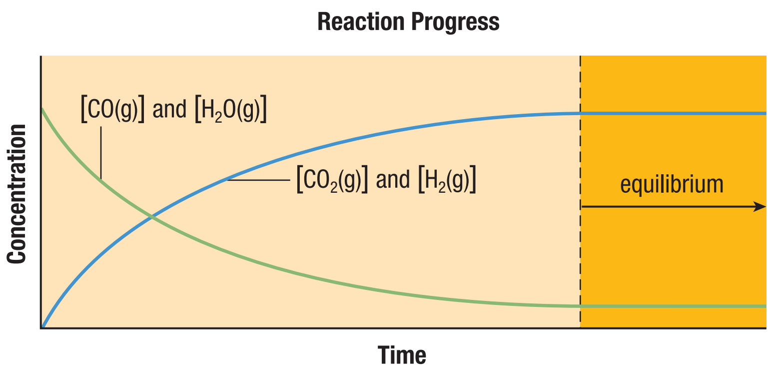

Figure 2 Changes in concentrations over time, when equal amounts of carbon monoxide gas and water vapour are allowed to react in a closed vessel

Scientists call the relative concentration of reactants and products in a chemical reaction system at equilibrium the **equilibrium position**. In the reaction depicted in Figure 2, the equilibrium position is the point at which the relative concentrations of the reactants and products stop changing. By convention, the equilibrium position is communicated in reference to the left-hand side (the reactant side) or the right-hand side (the product side).

equilibrium position the relative concentrations of reactants and products in a system in dynamic equilibrium

The equilibrium position of the reaction between water vapour and carbon monoxide gas lies far to the right. In other words, almost all the reactants are converted to products. However, the concentrations of the reactants never reach zero. If all chemical equilibria are dynamic equilibria, what chemical reactions are occurring?

Reversible reactions are chemical reactions that proceed in both the forward and reverse directions. In a dynamic equilibrium, the reaction rates in the forward and reverse directions of a reversible reaction are equal and so a constant concentration of reactants and products will always be present. Reversible reactions are identified by a double arrow symbol ($\rightleftharpoons$) in a chemical equation. The balanced equation for our example is therefore

reversible reaction a chemical reaction that proceeds in both the forward and reverse directions, setting up an equilibrium in a closed system

$$H_2O(g) + CO(g) \rightleftharpoons H_2(g) + CO_2(g)$$

When the water vapour and carbon monoxide gas were first added to the closed vessel, no hydrogen gas or carbon dioxide gas were present (**Figure 3**). Since there were no reactants available for the reverse reaction, it did not occur. As the forward reaction proceeded, the concentrations of hydrogen gas and carbon dioxide gas increased, providing reactants for the reverse reaction. Consequently, the rate of the reverse reaction increased. The system reached equilibrium once the rate of the forward reaction became equal to the rate of the reverse reaction.

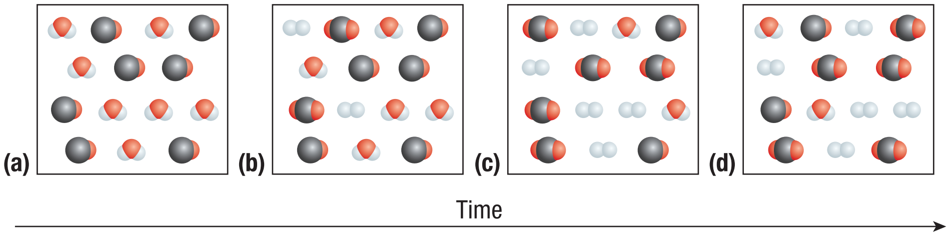

Figure 3 (a) Water vapour and carbon monoxide gas are mixed in equal amounts and (b) begin to react to form gaseous carbon dioxide and hydrogen. After some time, (c) equilibrium is reached and from that point on (d) the numbers of reactant and product molecules then remain constant over time.

Mini Investigation

The Water Exchange

Skills: Predicting, Performing, Observing, Evaluating

SKILLS HANDBOOK A2.1, A3.2, A6.5

When a reversible chemical reaction occurs in a closed system, the system eventually reaches dynamic equilibrium. In this investigation, you and a partner will use volumes of water to model the development of a system at equilibrium.

Equipment and Materials: marker; two 50 mL graduated cylinders; 2 large straws; water

1. Work in pairs. With the marker, label one 50 mL graduated cylinder “A” and the other “B.”
2. Measure 50 mL of water into cylinder A only. Then, place one large straw in cylinder A and another in cylinder B.
3. Create a table to record the volume of water in cylinders A and B. You will need room to record about 20 volumes for each.
4. Working simultaneously, one partner must trap the water in the straw in cylinder A and transfer it to cylinder B while the other partner must trap any water in the straw in cylinder B and transfer it to cylinder A. To trap and transfer the water, first hold the straw flat against the cylinder bottom with one hand. Then, wet one finger on the other hand and use that finger to cover the top of the straw. Lift the straw with both hands and transfer it to the other cylinder.
5. Once you and your partner have transferred your straw to the other cylinder, release the straw and let the water flow out. Then, measure and record the volume in each cylinder.
6. Return each empty straw to its original cylinder.
7. Repeat Steps 4 to 6 until the volumes stop changing.
8. Clean up your area.

A. Using the data in your table, create a graph of the data by plotting volume over the number of transfers. Place the data from both cylinder A and cylinder B on the same graph. T/I C

B. How did the volumes of water in each of the cylinders change over the course of this activity? T/I

C. Describe how you can tell from your graph when equilibrium was established. T/I

D. Consider that the straws, cylinders, and water comprise a system. Is the system open or closed? Why? K/U T/I

E. Explain how this investigation modelled a dynamic equilibrium. T/I A

Forward and Reverse Reactions

In an equilibrium system, an equilibrium position can be reached starting from the forward reaction or from the reverse reaction. Is the equilibrium position the same starting from either side? Consider the chemical reaction system between dinitrogen tetroxide gas, $N_2O_4(g)$, and nitrogen dioxide gas, $NO_2(g)$, that was depicted in Figure 1. In a sealed vessel at SATP, the reaction will progress to chemical equilibrium, according to the balanced chemical equation

$$N_2O_4(g) \rightleftharpoons 2\ NO_2(g)$$

Table 1 shows the initial concentrations and the final concentrations reached at equilibrium for two different experiments. In the first, a sealed 1 L vessel was filled with only nitrogen dioxide gas. In the second, the vessel initially contained only dinitrogen tetroxide.

Table 1 Changes in Concentrations of $NO_2(g)$ and $N_2O_4(g)$ by the Forward or Reverse Reactions

	Initial concentrations (mol/L)		Final concentrations (mol/L)	
	$N_2O_4(g)$	$NO_2(g)$	$N_2O_4(g)$	$NO_2(g)$
Experiment 1	0.750	0	0.721	0.0580
Experiment 2	0	1.50	0.721	0.0580

UNIT TASK BOOKMARK

As you work on the Unit Task described on page 582, you might apply what you learned about forward and reverse reactions in equilibrium systems.

Notice that the initial concentration of dinitrogen tetroxide gas in experiment 1 is half that of the nitrogen dioxide gas in experiment 2. As shown in **Figure 4**, when one molecule of dinitrogen tetroxide gas reacts, two molecules of nitrogen dioxide gas form. The initial concentrations of gases were chosen to reflect this 1:2 ratio. Regardless of which substance was the reactant and which was the product, the final concentrations of the gases at equilibrium are the same (**Figure 5**). WEB LINK

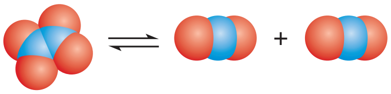

Figure 4 Molecular model of decomposition of dinitrogen tetroxide gas to nitrogen dioxide gas

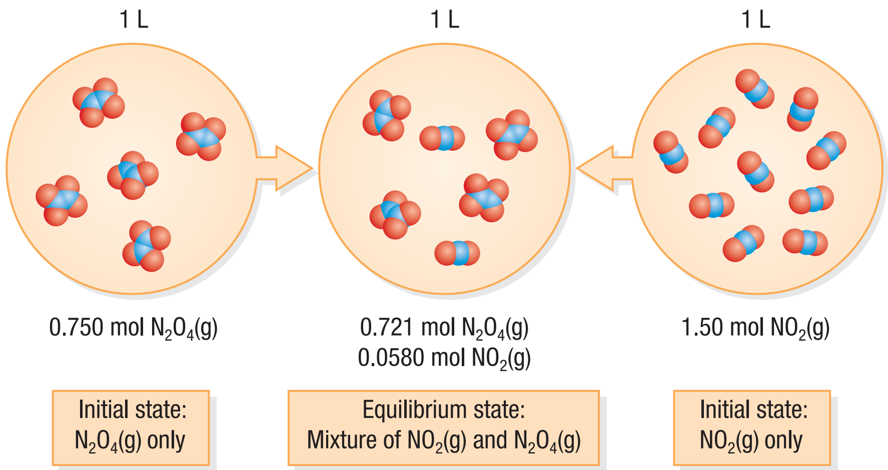

Figure 5 The composition of the mixture at dynamic equilibrium is the same regardless of whether the system started with gaseous nitrogen dioxide or with gaseous dinitrogen tetroxide.

From experiments like this, chemists have made the following generalization:

> For a closed chemical equilibrium system in constant environmental conditions, the same equilibrium concentrations are reached regardless of the direction by which equilibrium was reached.

Investigation 7.1.1

The Extent of a Chemical Reaction (page 472)

In this investigation you will combine various quantities of two solutions that react to form a precipitate. You will then test the precipitate to discover whether any or both of the reactants remain in solution to see if the reaction went to completion or formed an equilibrium.

Stoichiometry and Chemical Equilibria

When the concentrations of reactants and products do not change, is the reaction always at equilibrium? To answer this, consider the synthesis of ammonia gas, $NH_3(g)$. Ammonia gas is an important compound used in the production of plant fertilizers. Canada alone produces over 4000 million t of ammonia every year. Ammonia gas can be synthesized from nitrogen gas and hydrogen gas by the chemical reaction system given by the equation

$$N_2(g) + 3\,H_2(g) \rightleftharpoons 2\,NH_3(g)$$

If you were a manufacturer of ammonia, you would want to maximize the quantity of ammonia gas produced under particular conditions. Manufacturing needs such as yield are one reason that drives chemists to explore the characteristics of chemical equilibria. CAREER LINK

In experiments, when gaseous nitrogen, hydrogen, and ammonia are mixed in a closed vessel at 25 °C, the reaction is very slow. The concentrations of these gases change very little over time, regardless of the initial amounts. Chemists think that lack of reactivity is due to the strength of the bonds in the reactants—substances with high bond energies are less reactive. The bond energy of the N–N bond is 941 kJ/mol and the bond energy of the H–H bond is 432 kJ/mol. Both of these bonds are very strong and so both the reactants in ammonia gas synthesis are very unreactive. Manufacturers of ammonia gas use an iron catalyst to increase the rate of ammonia synthesis. In turn, the rate of the reverse reaction increases with the quantity of ammonia produced until equilibrium is reached.

You can predict the changes in concentration of reactants and products as a system approaches equilibrium from the coefficients of a balanced chemical equation (the stoichiometry). For the ammonia reaction system, the molar ratio of nitrogen gas to hydrogen gas to ammonia gas is 1:3:2. From this ratio, you can predict that, for every 1 mol of nitrogen gas that reacts, 3 mol of hydrogen gas will be consumed and 2 mol of ammonia gas will form. Experiments have confirmed these predictions (**Figure 6**). The decrease in hydrogen gas concentration is 3 times the decrease in nitrogen gas concentration. The increase of ammonia concentration is 2 times the decrease of the nitrogen concentration. You can also use the 1:3:2 molar ratio to predict concentration changes in reactants and products when equilibrium is reached from the opposite direction.

In fact, for many years chemists were unable to find conditions under which this reaction system would reach equilibrium at a rate that converted the reactants to ammonia gas quickly enough to be useful. As you will see in Section 7.3, the scientist who eventually solved this puzzle was given a Nobel prize to honour the achievement.

LEARNING TIP

Concentrations of Gases

Concentrations provided to you in this chapter are given in moles per litre. A concentration of 1.00 mol/L of hydrogen gas, $H_2(g)$, means there is one mole of $H_2(g)$ per litre of space occupied.

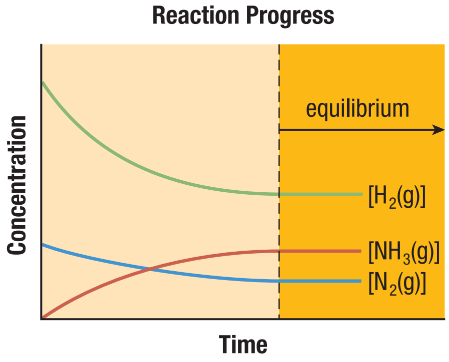

Figure 6 Concentration changes over time for the reaction $N_2(g) + 3\,H_2(g) \rightleftharpoons 2\,NH_3(g)$ when only nitrogen gas and hydrogen gas are mixed initially.

Determining Concentrations for Chemical Equilibria

For chemical equilibrium systems composed of aqueous solutions or gases, one strategy to perform stoichiometric calculations is to use an ICE table. In an ICE table, I stands for "initial" concentrations of reactants and products before the reaction, C for "change" in the concentrations of reactants and products from the start of the reaction to when equilibrium is achieved, and E for "equilibrium" concentrations of reactants and products.

Tutorial 1 Calculating Equilibrium Concentrations

In this tutorial, you will use an ICE table to find concentrations in chemical equilibrium systems. Although stoichiometric calculations use either concentrations or the amount in moles because the volume of the system remains the same, we will always use concentration values in ICE tables and in calculations involving equilibrium. Keep in mind that the values in ICE tables must be in mol/L.

Sample Problem 1: Calculating Equilibrium Concentrations from Initial Reactant Concentrations

Hydrogen fluoride gas, HF(g), is used in the production of many important substances, such as medicines. Hydrogen fluoride may be synthesized from gaseous hydrogen, H_2(g), and fluorine, F_2(g). The balanced chemical equation for this reaction is

$$H_2(g) + F_2(g) \rightleftharpoons 2\,HF(g)$$

When a chemist starts this chemical reaction in a sealed container at SATP, the initial concentration of gaseous hydrogen and of gaseous fluorine is 2.00 mol/L. No hydrogen fluoride gas is present initially. What are the equilibrium concentrations of hydrogen gas and hydrogen fluorine gas, if the equilibrium concentration of floride gas is 0.48 mol/L?

Given: $[H_2(g)]_{initial} = 2.00$ mol/L; $[F_2(g)]_{initial} = 2.00$ mol/L; $[HF(g)]_{initial} = 0$ mol/L; $[F_2(g)]_{equilibrium} = 0.48$ mol/L

Required: $[H_2(g)]_{equilibrium}$; $[HF(g)]_{equilibrium}$

Analysis: Set up an ICE table. Your table should have two columns and three rows. Put the letters I, C, and E in the cells of the first column. Then, write the balanced chemical equation above the second column. Add the initial concentrations to the I row, under the corresponding symbol in the chemical equation.

Use a variable, such as x, to represent the changes in the concentrations of reactants and products. Then add coefficients to x that correspond to the coefficients in the balanced equation.

From the balanced equation, you know that H_2(g) and F_2(g) are converted to HF(g) in a 1:1:2 molar ratio. During the reaction, the concentrations of H_2(g) and F_2(g) decrease as the concentration of HF(g) increases. Since H_2(g) and F_2(g) are consumed in a 1:1 molar ratio, the decrease in their concentrations is $-x$ mol/L. Since 2 mol of HF(g) are produced per 1 mol of H_2(g) and of F_2(g), the increase in concentration of HF(g) is $+2x$. The equilibrium concentrations of H_2(g) and F_2(g) will be their initial concentrations, 2.00 mol/L, minus the decrease in their concentrations (x mol/L).

In the E row of your ICE table, place the expression $2.00 - x$ under both H_2(g) and F_2(g). Since there was initially no HF(g), the equilibrium concentration of HF(g) will be $0 + 2x$, or $2x$, (the increase in its concentration). Your ICE table should now look similar to **Table 2**. Remember that the units of all values are mol/L.

Table 2 ICE Table for Calculating Equilibrium Concentrations

	$H_2(g)$	+ $F_2(g)$	$\rightleftharpoons$ $2\,HF(g)$
I	2.00	2.00	0
C	$-x$	$-x$	$+2x$
E	$2.00 - x$	$2.00 - x$	$2x$

Solution

Step 1. Write and solve the equation for x, using the values in the ICE table.

$[F_2(g)]_{equilibrium}$ was given as 0.48 mol/L. By your analysis, you also know that $[F_2(g)]_{equilibrium}$ is represented by the expression $2.00 - x$.

$$\begin{aligned} 2.00 \text{ mol/L} - x &= 0.48 \text{ mol/L} \\ -x &= 0.48 \text{ mol/L} - 2.00 \text{ mol/L} \\ &= -1.52 \text{ mol/L} \\ x &= 1.52 \text{ mol/L} \end{aligned}$$

Step 2. Use the value of x to calculate the equilibrium concentrations of the other two entities.

$$\begin{aligned} [H_2(g)] &= 2.00 \text{ mol/L} - x \\ &= 2.00 \text{ mol/L} - 1.52 \text{ mol/L} \\ [H_2(g)] &= 0.48 \text{ mol/L} \\ [HF(g)] &= 2x \\ &= 2(1.52 \text{ mol/L}) \\ [HF(g)] &= 3.04 \text{ mol/L} \end{aligned}$$

Statement: The equilibrium concentrations of hydrogen gas and hydrogen fluoride gas are 0.48 mol/L and 3.04 mol/L, respectively.

Sample Problem 2: Equilibrium Concentrations from a Graph of Reaction Progress

When ammonia gas, $NH_3(g)$, is heated, it decomposes to form nitrogen gas, $N_2(g)$, and hydrogen gas, $H_2(g)$. A chemist adds 4.0 mol of ammonia gas to a 2.0 L sealed, rigid container and heats it. **Figure 7** shows the changes in the amount of ammonia gas she observed over time.

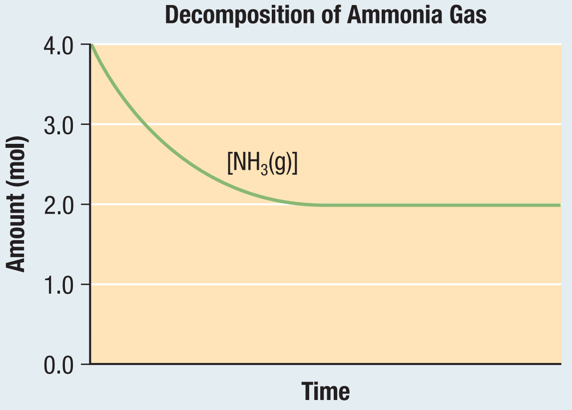

Figure 7

The balanced equation for this chemical reaction system is

$$2\ NH_3(g) \rightleftharpoons N_2(g) + 3\ H_2(g)$$

Determine the equilibrium concentrations of $N_2(g)$ and $H_2(g)$.

Given: initial quantity of NH_3 = 4.0 mol; volume = 2.0 L

Required: $[N_2(g)]_{equilibrium}$; $[H_2(g)]_{equilibrium}$

Solution:

Step 1. Calculate the initial concentration of ammonia gas.

$$[NH_3(g)]_{initial} = \frac{4.0\ mol}{2.0\ L}$$

$$[NH_3(g)]_{initial} = 2.0\ mol/L$$

Step 2. Using Figure 7, determine the amount of ammonia gas at equilibrium, once the amount remains stable

Since the amount is stable at the last time point, this is the amount of ammonia gas at equilibrium.

$$[NH_3(g)]_{equilibrium} = \frac{2.0\ mol}{2.0\ L}$$

$$[NH_3(g)]_{equilibrium} = 1.0\ mol/L$$

Step 3. Use an ICE table to determine the equilibrium concentrations of $N_2(g)$ and $H_2(g)$.

First, list the initial concentrations of reactants and products in the I row. Then, indicate changes in concentrations in the C row by letting multiples of x represent changes in the concentrations of reactants and products. Remember that the multiples of x always correspond to the coefficients in the balanced equation. Then, where appropriate in the E row, use positive x expressions to indicate that a reactant or product increases in concentration as the reaction proceeds, and negative x expressions to indicate that a reactant or product decreases in concentration as the reaction proceeds. Your ICE table should be similar to **Table 3**.

Table 3 ICE Table for Calculating Equilibrium Concentrations

	$2\ NH_3(g)$	$\rightleftharpoons$	$N_2(g)$	+	$3\ H_2(g)$
I	2.0		0		0
C	$-2x$		$+x$		$+3x$
E	$2.0 - 2x$		x		$3x$

Step 4. Calculate the value of x, using the initial, change, and equilibrium concentrations of ammonia gas.

$$\begin{aligned} 2.0\ mol/L - 2x &= 1.0\ mol/L \\ -2x &= -1.0\ mol/L \\ 2x &= 1.0\ mol/L \\ x &= 0.50\ mol/L \end{aligned}$$

Step 5. Use the value of x to calculate the equilibrium concentrations of nitrogen gas and hydrogen gas.

$$\begin{aligned} [N_2(g)]_{equilibrium} &= x \\ [N_2(g)]_{equilibrium} &= 0.50\ mol/L \\ [H_2(g)]_{equilibrium} &= 3x \\ &= 3(0.50\ mol/L) \\ [H_2(g)]_{equilibrium} &= 1.5\ mol/L \end{aligned}$$

Statement: The equilibrium concentration of nitrogen gas is 0.50 mol/L and the equilibrium concentration of hydrogen gas is 1.5 mol/L.

Practice

1. When dinitrogen tetroxide gas, $N_2O_4(g)$, is placed in a sealed container at 100 °C, it decomposes into nitrogen dioxide gas, $NO_2(g)$, according to the following balanced equation:

   $$N_2O_4(g) \rightleftharpoons 2\ NO_2(g)$$

   A laboratory technician places 0.25 mol of dinitrogen tetroxide gas in a 1.0 L closed container. At equilibrium, the concentration of nitrogen dioxide gas is 0.25 mol/L. Use an ICE table to determine the equilibrium concentration of dinitrogen tetroxide gas. T/I [ans: $[N_2O_4(g)]_{\text{equilibrium}} = 0.12$ mol/L]

2. At 35 °C, 3.00 mol of pure nitrosyl chloride gas, NOCl(g), is contained in a sealed 3.00 L flask. The nitrosyl chloride gas decomposes to nitric oxide gas, NO(g), and chlorine gas, $Cl_2(g)$ until equilibrium is reached.

   $$2\ NOCl(g) \rightleftharpoons 2\ NO(g) + Cl_2(g)$$

   At equilibrium, the concentration of nitric oxide gas is 0.043 mol/L. Use an ICE table to determine the equilibrium concentrations of nitrosyl chloride gas and chlorine gas under the conditions used. T/I [ans: $[NOCl(g)]_{\text{equilibrium}} = 0.96$ mol/L; $[Cl_2(g)]_{\text{equilibrium}} = 0.022$ mol/L]

3. A chemist places 2.00 mol of ethene gas, $C_2H_4(g)$, and 1.25 mol of bromine gas, $Br_2(g)$, in a sealed 0.500 L container. **Figure 8** shows the concentration of ethene gas over time. The temperature of the vessel and its contents are kept constant. The balanced chemical equation for the reaction is

   $$C_2H_4(g) + Br_2(g) \rightleftharpoons C_2H_4Br_2(g)$$

   Determine the equilibrium concentration of ethene, bromine, and 1,2-dibromoethane gases. T/I

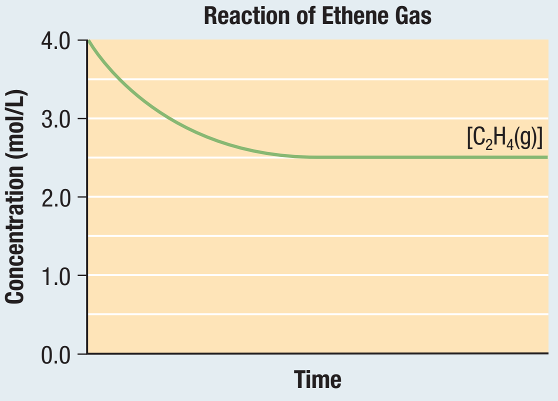

Figure 8

[ans: $[C_2H_4(g)]_{\text{equilibrium}} = 2.5$ mol/L; $[Br_2(g)]_{\text{equilibrium}} = 1.0$ mol/L; $[C_2H_4Br_2(g)]_{\text{equilibrium}} = 1.5$ mol/L]

7.1 Review

Summary

- Some chemical reactions do not go to completion, but to a chemical equilibrium, which is a dynamic equilibrium in a closed system in which the forward and reverse reactions occur at equal rates.
- In an equilibrium system, an equilibrium position can be reached starting from the forward reaction or from the reverse reaction.

Questions

1. In carbonated pop, the bubbles come from carbonic acid, $H_2CO_3(aq)$, in the liquid decomposing to carbon dioxide gas, $CO_2(g)$, and water. K/U A
   (a) Explain how the sealed bottle of pop is an example of a dynamic equilibrium.
   (b) What happens when you open the pop bottle?
2. The following balanced chemical equation represents a chemical system at equilibrium: K/U

$$\underset{\text{colourless}}{C_2H_4(g)} + \underset{\text{brown}}{Br_2(g)} \rightleftharpoons \underset{\text{colourless}}{C_2H_4Br_2(g)}$$

   (a) What visual changes will you see at equilibrium?
   (b) What is occurring at the molecular level?
   (c) Is the equilibrium static or dynamic?
3. When combined in a closed vessel, hydrogen gas and iodine gas will form hydrogen iodide gas until an equilibrium position is reached. K/U T/I C
   (a) Write a balanced equation for this chemical reaction system.
   (b) Suppose you carry out an investigation starting with a 2.0 L flask containing 0.45 mol of hydrogen iodide. Predict how the concentrations of the gases will change as the system reaches equilibrium.
   (c) In a second experiment, the concentration of each of the gases was monitored (**Figure 9**). Complete an ICE table for the reaction.

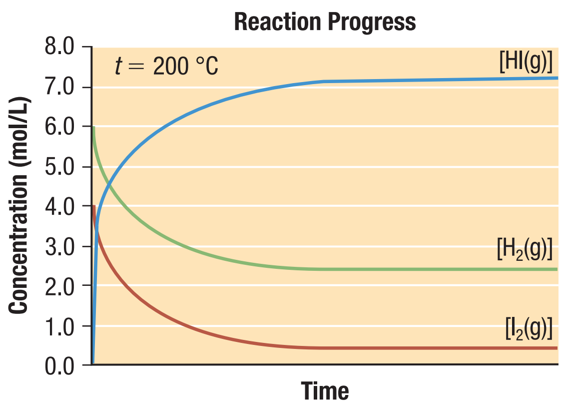

Figure 9

4. Sulfuric acid is an important industrial chemical that is usually produced by a series of reactions. One of these involves an equilibrium between gaseous sulfur dioxide, oxygen, and sulfur trioxide.

$$2\ SO_2(g) + O_2(g) \rightleftharpoons 2\ SO_3(g)$$

   If 2.5 mol of sulfur dioxide gas and 2.0 mol of oxygen gas are placed in a sealed 1.0 L container and allowed to reach equilibrium, 0.75 mol of sulfur dioxide remains. Use an ICE table to determine the concentration of the other gases at equilibrium. T/I
5. Phosphorus pentachloride gas, $PCl_5(g)$, will decompose to phosphorous trichloride, $PCl_3(g)$, and chlorine, $Cl_2(g)$, at 160 °C. In a sealed vessel, the reaction will proceed to equilibrium:

$$PCl_5(g) \rightleftharpoons PCl_3(g) + Cl_2(g)$$

   A chemist places 3.00 mol of phosphorous pentachloride gas into a sealed 1.50 L flask at 160 °C. At equilibrium, he observes there is 0.300 mol of phosphorous trichloride gas and some chlorine gas. Calculate the equilibrium concentrations of gaseous phosphorous pentachloride and chlorine. T/I
6. Fructose is a carbohydrate found in many foods, especially sweet fruits. If you eat a meal that is rich in fructose, the concentration of fructose in solution inside the intestinal tract of the small intestine (i.e., the concentration outside the cells) will be higher than the fructose concentration inside the cells making up the intestine. This causes fructose to diffuse into the cells. A
   (a) Will diffusion result in the absorption of all fructose molecules from the digesting food travelling through the intestine? Explain.
   (b) Conduct research to determine why cells are able to absorb the maximum possible amount of nutrients from their surroundings.

WEB LINK