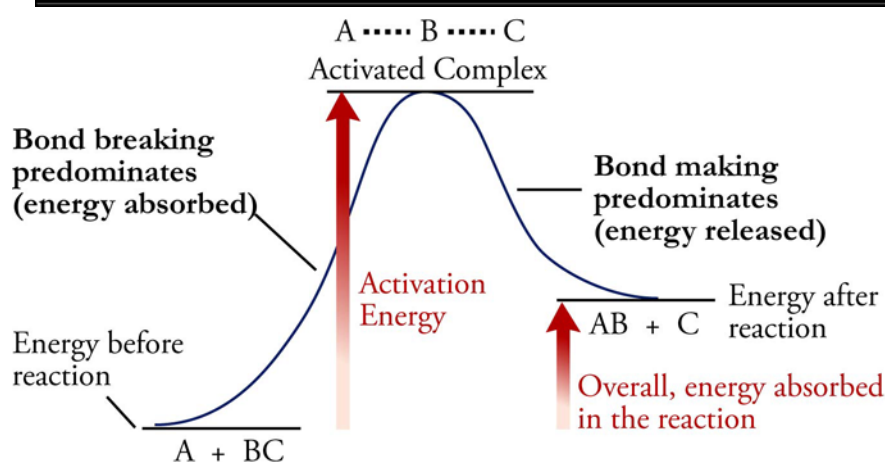


## Chapter 16 - The Process of Chemical Reactions



### ◆ Review Skills

#### 16.1 Collision Theory: A Model for the Reaction Process

- The Basics of Collision Theory
- Endergonic Reactions
- Summary of Collision Theory

#### 16.2 Rates of Chemical Reactions

- Temperature and Rates of Chemical Reactions
- Concentration and Rates of Chemical Reactions
- Catalysts
- Homogeneous and
- Heterogeneous Catalysts

#### *Special Topic 16.1: Green Chemistry - The Development of New and Better Catalysts*

#### 16.3 Reversible Reactions and Chemical Equilibrium

- Reversible Reactions and Dynamic Equilibrium
- Equilibrium Constants
- Determination of Equilibrium Constant Values
- Equilibrium Constants and Extent of Reaction
- Heterogeneous Equilibria
- Equilibrium Constants and Temperature

[Internet: Calculating Concentrations and Gas Pressures](#)

[Internet: pH and pH Calculations](#)

[Internet: Weak Acids and Equilibrium Constants](#)

#### 16.4 Disruption of Equilibrium

- The Effect of Changes in Concentrations on Equilibrium Systems

[Internet: Changing Volume and Gas Phase Equilibrium](#)

- Le Châtelier's Principle
- The Effect of Catalysts on Equilibria

*Special Topic 16.2: The Big Question—How Did We Get Here?*

### ◆ Chapter Glossary

[Internet: Glossary Quiz](#)

### ◆ Chapter Objectives

Review Questions

Key Ideas

Chapter Problems

## Section Goals and Introductions

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### Section 16.1 Collision Theory: A Model for the Reaction Process

#### Goals

- *To describe a model, called collision theory, that helps us to visualize the process of many chemical reactions.*
- *To use collision theory to explain why not all collisions between possible reactants lead to products.*
- *To use collision theory to explain why possible reactants must collide with an energy equal to or above a certain amount to have the possibility of reacting and forming products.*
- *To show how the energy changes in chemical reactions can be described with diagrams.*
- *To use collision theory to explain why possible reactants must collide with a specific orientation to have the possibility of reacting and forming products.*

Once again, this chapter emphasizes that if you develop the ability to visualize changes on the particle level, it will help you understand and explain many different things. This section introduces you to a model for chemical change that is called collision theory, which helps you explain the factors that affect the rates of chemical reactions. These factors are described in Section 16.2.

### Section 16.2 Rates of Chemical Reactions

#### Goals

- *To show how rates of chemical reactions are described.*
- *To explain why increased temperature increases the rates of most chemical reactions.*
- *To explain why increased concentration of reactants increases the rates of chemical reactions.*
- *To describe how catalysts increase the rates of certain chemical reactions.*

This section shows how collision theory helps you explain the factors that affect rates of chemical changes. These factors include amounts of reactants and products, temperature, and catalysts.

### Section 16.3 Reversible Reactions and Chemical Equilibrium

#### Goals

- *To explain why chemical reactions that are reversible come to a dynamic equilibrium with equal forward and reverse rates of reaction.*
- *To show what equilibrium constants are and how they can be determined.*
- *To describe how equilibrium constants can be used to show the relative amounts of reactants and products in the system at equilibrium.*
- *To explain the effect of temperature on equilibrium systems and equilibrium constants.*

This section takes the basic ideas of dynamic equilibrium introduced in Chapter 14 and applies them to reversible chemical changes. This is a very important topic, so plan to spend some extra time on this section, if necessary. You will also learn how equilibrium constants are used to describe the relative amounts of reactants and products for a chemical reaction at

equilibrium, and you will learn how these values can be calculated. Finally, you will learn more about the effect of temperature on chemical changes. See the three related sections on our Web site:

[Internet: Calculating Concentrations and Gas Pressures](#)

[Internet: pH and pH Calculations](#)

[Internet: Weak Acids and Equilibrium Constants](#)

### **Section 16.4 Disruption of Equilibrium**

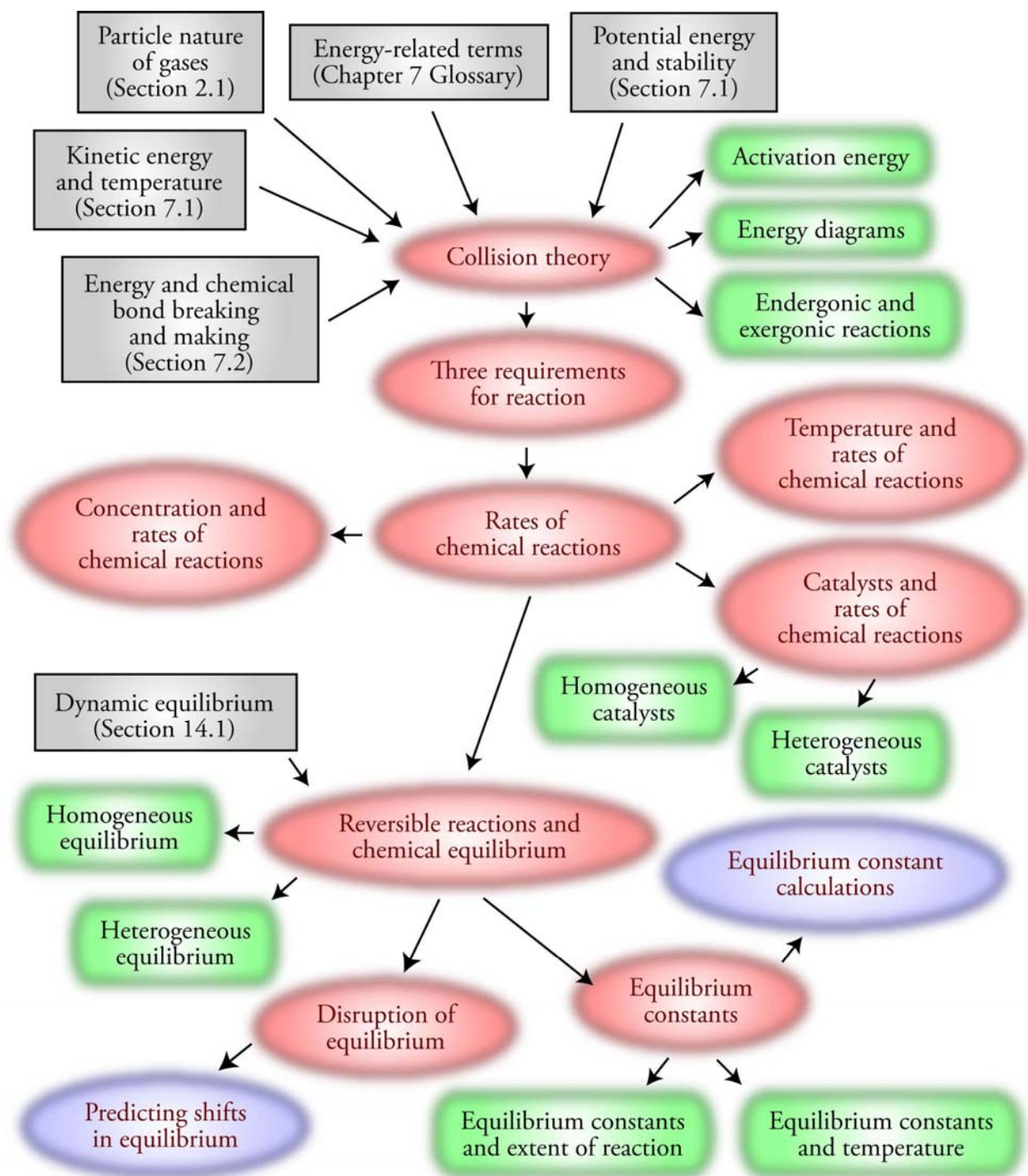
*Goal: To describe how equilibrium systems can be disrupted and show you how to predict whether certain changes on a system at equilibrium will lead to more products, more reactants, or neither.*

Although the concept of chemical equilibrium is very important, many reversible reactions in nature never form equilibrium systems. This section's description of the ways that equilibrium systems can be disrupted will help you to understand why this is true. The ability to predict the effects of changes on equilibrium systems will help you understand the ways that research and industrial chemists create conditions for their chemical reactions that maximize the rate at which desirable reactions move to products and minimize that rate at which undesirable reactions take place.

See the section on our Web site that provides information on *Changing Volumes and Gas Phase Equilibrium*.

[Internet: Changing Volume and Gas Phase Equilibrium](#)

## Chapter 16 Map



## Chapter Checklist

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- Read the Review Skills section. If there is any skill mentioned that you have not yet mastered, review the material on that topic before reading this chapter.
- Read the chapter quickly before the lecture that describes it.
- Attend class meetings, take notes, and participate in class discussions.
- Work the Chapter Exercises, perhaps using the Chapter Examples as guides.
- Study the Chapter Glossary and test yourself on our Web site:

[\*Internet: Glossary Quiz\*](#)

- Study all of the Chapter Objectives. You might want to write a description of how you will meet each objective.
- This chapter has logic sequences in Figures 16.11, 16.13, 16.15, 16.22, and 16.25. Convince yourself that each of the statements in these sequences logically leads to the next statement.
- To get a review of the most important topics in the chapter, fill in the blanks in the Key Ideas section.
- Work all of the selected problems at the end of the chapter, and check your answers with the solutions provided in this chapter of the study guide.
- Ask for help if you need it.

## Web Resources

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[\*Internet: Calculating Concentrations and Gas Pressures\*](#)

[\*Internet: pH and pH Calculations\*](#)

[\*Internet: Weak Acids and Equilibrium Constants\*](#)

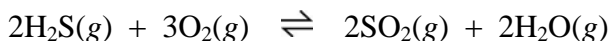
[\*Internet: Changing Volume and Gas Phase Equilibrium\*](#)

[\*Internet: Glossary Quiz\*](#)

## Exercises Key

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**Exercise 16.1 – Writing Equilibrium Constant Expressions:** Sulfur dioxide,  $\text{SO}_2$ , one of the intermediates in the production of sulfuric acid, can be made from the reaction of hydrogen sulfide gas with oxygen gas. Write the equilibrium constant expressions for  $K_C$  and  $K_P$  for the following equation for this reaction. (*Obj's 24 & 25*)



$$K_C = \frac{[\text{SO}_2]^2 [\text{H}_2\text{O}]^2}{[\text{H}_2\text{S}]^2 [\text{O}_2]^3} \quad K_P = \frac{P_{\text{SO}_2}^2 P_{\text{H}_2\text{O}}^2}{P_{\text{H}_2\text{S}}^2 P_{\text{O}_2}^3}$$

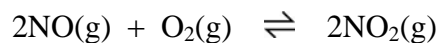
**Exercise 16.2 – Equilibrium Constant Calculation:** Ethanol,  $C_2H_5OH$ , can be made from the reaction of ethylene gas,  $C_2H_4$ , and water vapor. A mixture of  $C_2H_4(g)$  and  $H_2O(g)$  is allowed to come to equilibrium in a container at  $110\text{ }^\circ\text{C}$ , and the partial pressures of the gases are found to be  $0.35\text{ atm}$  for  $C_2H_4(g)$ ,  $0.75\text{ atm}$  for  $H_2O(g)$ , and  $0.11\text{ atm}$  for  $C_2H_5OH(g)$ . What is  $K_P$  for this reaction at  $110\text{ }^\circ\text{C}$ ? (*Obj 26*)

$$C_2H_4(g) + H_2O(g) \rightleftharpoons C_2H_5OH(g)$$

$$K_P = \frac{P_{C_2H_5OH}}{P_{C_2H_4} P_{H_2O}} = \frac{0.11\text{ atm}}{0.35\text{ atm} (0.75\text{ atm})} = \mathbf{0.42\ 1/atm\ or\ 0.42}$$

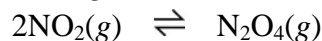
**Exercise 16.3 – Predicting the Extent of Reaction:** Using the information in Table 16.1, predict whether each of the following reversible reactions favors reactants, products, or neither at  $25\text{ }^\circ\text{C}$ . (*Obj 27*)

- a. This reaction is partially responsible for the release of pollutants from automobiles.



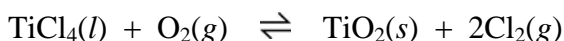
According to Table 16.1, the  $K_P$  for this reaction is  $2.2 \times 10^{12}$ , so it **favors products**.

- b. The  $NO_2(g)$  molecules formed in the reaction in part (a) can combine to form  $N_2O_4$ .



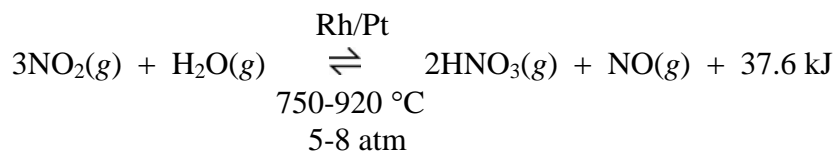
According to Table 16.1, the  $K_P$  for this reaction is  $6.7$ . **Neither reactants nor products are favored.**

**Exercise 16.4 – Writing Equilibrium Constants for Heterogeneous Equilibria:** The following equation describes one of the steps in the purification of titanium dioxide, which is used as a white pigment in paints. Liquid titanium(IV) chloride reacts with oxygen gas to form solid titanium oxide and chlorine gas. Write  $K_C$  and  $K_P$  expressions for this reaction. (*Obj 24 & 25*)



$$K_C = \frac{[Cl_2]^2}{[O_2]} \quad K_P = \frac{P_{Cl_2}^2}{P_{O_2}}$$

**Exercise 16.5 – Predicting the Effect of Disruptions on Equilibrium:** Nitric acid can be made from the exothermic reaction of nitrogen dioxide gas and water vapor in the presence of a rhodium and platinum catalyst at 700-900 °C and 5-8 atm. Predict whether each of the following changes in the equilibrium system will shift the system to more products, to more reactants, or to neither. Explain each answer in two ways: (1) by applying Le Châtelier's principle and (2) by describing the effect of the change on the forward and reverse reaction rates. (Oblis 40- 42 & 44- 46)



- a. The concentration of  $\text{H}_2\text{O}$  is increased by the addition of more  $\text{H}_2\text{O}$ .
  - (1) Using Le Châtelier's Principle, we predict that the system will **shift to more products** to partially counteract the increase in  $\text{H}_2\text{O}$ .
  - (2) The increase in the concentration of water vapor speeds the forward reaction without initially affecting the rate of the reverse reaction. The equilibrium is disrupted, and the system **shifts to more products** because the forward rate is greater than the reverse rate.
- b. The concentration of  $\text{NO}_2$  is decreased.
  - (1) Using Le Châtelier's Principle, we predict that the system will **shift to more reactants** to partially counteract the decrease in  $\text{NO}_2$ .
  - (2) The decrease in the concentration of  $\text{NO}_2(\text{g})$  slows the forward reaction without initially affecting the rate of the reverse reaction. The equilibrium is disrupted, and the system **shifts toward more reactants** because the reverse rate is greater than the forward rate.
- c. The concentration of  $\text{HNO}_3(\text{g})$  is decreased by removing the nitric acid as it forms.
  - (1) Using Le Châtelier's Principle, we predict that the system will **shift to more products** to partially counteract the decrease in  $\text{HNO}_3$ .
  - (2) The decrease in the concentration of  $\text{HNO}_3(\text{g})$  slows the reverse reaction without initially affecting the rate of the forward reaction. The equilibrium is disrupted, and the system **shifts toward more products** because the forward rate is greater than the reverse rate.
- d. The temperature is decreased from 1000 °C to 800 °C.
  - (1) Using Le Châtelier's Principle, we predict that the system shifts in the exothermic direction to partially counteract the decrease in temperature. As the system **shifts toward more products**, energy is released, and the temperature increases.
  - (2) The decreased temperature decreases the rates of both the forward and reverse reactions, but it has a greater effect on the endothermic reaction. Because the forward reaction is exothermic, the reverse reaction must be endothermic. Therefore, the reverse reaction is slowed more than the forward reaction. The system **shifts toward more products** because the forward rate becomes greater than the reverse rate.

- e. The Rh/Pt catalyst is added to the equilibrium system.
- (1) *Le Châtelier's Principle does not apply here.*
  - (2) *The catalyst speeds both the forward and the reverse rates equally. Thus there is no shift in the equilibrium. The purpose of the catalyst is to bring the system to equilibrium faster.*

## Review Questions Key

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1. Describe what you visualize occurring inside a container of oxygen gas, O<sub>2</sub>, at room temperature and pressure.

*The gas is composed of O<sub>2</sub> molecules that are moving constantly in the container. For a typical gas, the average distance between particles is about ten times the diameter of each particle. This leads to the gas particles themselves taking up only about 0.1% of the total volume. The other 99.9% of the total volume is empty space. According to our model, each O<sub>2</sub> molecule moves freely in a straight-line path until it collides with another O<sub>2</sub> molecule or one of the walls of the container. The particles are moving fast enough to break any attraction that might form between them, so after two particles collide, they bounce off each other and continue on alone. Due to collisions, each particle is constantly speeding up and slowing down, but its average velocity stays constant as long as the temperature stays constant.*

2. Write in each blank the word that best fits the definition.
- Energy** is the capacity to do work.
  - Kinetic energy** is the capacity to do work due to the motion of an object.
  - A(n) **endergonic** change is a change that absorbs energy.
  - A(n) **exergonic** change is a change that releases energy.
  - Thermal energy** is the energy associated with the random motion of particles.
  - Heat** is thermal energy that is transferred from a region of higher temperature to a region of lower temperature as a result of the collisions of particles.
  - A(n) **exothermic** change is a change that leads to *heat* energy being evolved from the system to the surroundings.
  - A(n) **endothermic** change is a change that leads the system to absorb *heat* energy from the surroundings.
  - A(n) **catalyst** is a substance that speeds a chemical reaction without being permanently altered itself.
3. When the temperature of the air changes from 62 °C at 4:00 A.M. to 84 °C at noon on a summer day, does the average kinetic energy of the particles in the air increase, decrease, or stay the same?

*Increased temperature means **increased average kinetic energy.***

4. Explain why it takes energy to break an O–O bond in an O<sub>3</sub> molecule.

*Separate atoms are less stable, and therefore, higher potential energy than atoms in a bond. The Law of Conservation of Energy states that energy cannot be created or destroyed, so energy must be added to the system. It always takes energy to break attractions between particles.*

5. Explain why energy is released when two oxygen atoms come together to form an O<sub>2</sub> molecule.

*Atoms in a bond are more stable, and therefore, lower potential energy. The Law of Conservation of Energy states that energy cannot be created or destroyed, so energy is released from the system. Energy is always released when new attractions between particles are formed.*

6. Explain why some chemical reactions *release heat* to their surroundings.

*If the bonds in the products are stronger and lower potential energy than in the reactants, energy will be released from the system. If the energy released is due to the conversion of potential energy to kinetic energy, the temperature of the products will be higher than the original reactants. The higher temperature products are able to transfer heat to the surroundings, and the temperature of the surroundings increases.*

7. Explain why some chemical reactions *absorb heat* from their surroundings.

*If the bonds in the products are weaker and higher potential energy than in the reactants, energy must be absorbed. If the energy absorbed is due to the conversion of kinetic energy to potential energy, the temperature of the products will be lower than the original reactants. The lower temperature products are able to absorb heat from the surroundings, and the temperature of the surroundings decreases.*

8. What are the general characteristics of any dynamic equilibrium system?

*The system must have two opposing changes, from state A to state B and from state B to state A. For a dynamic equilibrium to exist, the rates of the two opposing changes must be equal, so there are constant changes between state A and state B but no net change in the components of the system.*

## Key Ideas Answers

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9. At a certain stage in the progress of a reaction, bond breaking and bond making are of equal importance. In other words, the energy necessary for bond breaking is **balanced** by the energy supplied by bond making. At this turning point, the particles involved in the reaction are joined in a structure known as the activated complex, or **transition state**.
11. In a chemical reaction, the **minimum** energy necessary for reaching the activated complex and proceeding to products is called the activation energy. Only the collisions that provide a net kinetic energy **equal to** or **greater than** the activation energy can lead to products.
13. The energies associated with endergonic (or endothermic) changes are described with **positive** values.
15. Because the formation of the **new bonds** provides some of the energy necessary to break the **old bonds**, the making and breaking of bonds must occur more or less simultaneously. This is possible only when the particles collide in such a way that the bond-forming atoms are **close to each other**.
17. Increased temperature means an increase in the average kinetic energy of the collisions between the particles in a system. This leads to an increase in the **fraction** of the collisions that have enough energy to reach the activated complex (the activation energy).
19. One of the ways in which catalysts accelerate chemical reactions is by providing a(n) **alternative pathway** between reactants and products that has a(n) **lower activation energy**.

21. If the catalyst is not in the same state as the reactants, the catalyst is called a(n) **heterogeneous** catalyst.
23. The extent to which reversible reactions proceed toward products before reaching equilibrium can be described with a(n) **equilibrium constant**, which is derived from the ratio of the concentrations of products to the concentrations of reactants at equilibrium. For homogeneous equilibria, the concentrations of all reactants and products can be described in **moles per liter**, and the concentration of each is raised to a power equal to its **coefficient** in a balanced equation for the reaction.
25. The **larger** a value for  $K$ , the farther the reaction shifts toward products before the rates of the forward and reverse reactions become equal and the concentrations of reactants and products stop changing.
27. Changing temperature always causes a shift in equilibrium systems—sometimes toward more products and sometimes toward more **reactants**.
29. If the forward reaction in a reversible reaction is endergonic, increased temperature will shift the system toward more **products**.
31. Le Châtelier's principle states that if a system at equilibrium is altered in a way that **disrupts** the equilibrium, the system will shift in such a way as to counter the **change**.

## Problems Key

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### Section 16.1 Collision Theory: A Model for the Reaction Process

33. Assume that the following reaction is a single-step reaction in which one of the O–O bonds in  $O_3$  is broken and a new N–O bond is formed. The heat of reaction is  $-226$  kJ/mol.
- $$NO(g) + O_3(g) \rightarrow NO_2(g) + O_2(g) + 226 \text{ kJ}$$
- a. With reference to collision theory, describe the general process that takes place as this reaction moves from reactants to products. (*Obj 2*)
- NO and  $O_3$  molecules are constantly moving in the container, sometimes with a high velocity and sometimes more slowly. The particles are constantly colliding, changing their direction of motion, and speeding up or slowing down. If the molecules collide in a way that puts the nitrogen atom in NO near one of the outer oxygen atoms in  $O_3$ , one of the O–O bonds in the  $O_3$  molecule begins to break, and a new bond between one of the oxygen atoms in the ozone molecule and the nitrogen atom in NO begins to form. If the collision yields enough energy to reach the activated complex, it proceeds on to products. If the molecules do not have the correct orientation, or if they do not have enough energy, they separate without a reaction taking place.*
- b. List the three requirements that must be met before a reaction between  $NO(g)$  and  $O_3(g)$  is likely to take place. (*Obj 11*)
- NO and  $O_3$  molecules must collide, they must collide with the correct orientation to form an N–O bond at the same time that an O–O bond is broken, and they must have the minimum energy necessary to reach the activated complex (the activation energy).*

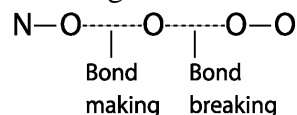
- c. Explain why  $\text{NO}(g)$  and  $\text{O}_3(g)$  must collide before a reaction can take place. (Obj 3)

*The collision brings the atoms that will form the new bonds close, and the net kinetic energy in the collision provides the energy necessary to reach the activated complex and proceed to products.*

- d. Explain why it is usually necessary for the new  $\text{N}-\text{O}$  bonds to form at the same time that the  $\text{O}-\text{O}$  bonds are broken. (Obj 4)

*It takes a significant amount of energy to break  $\text{O}-\text{O}$  bonds, and collisions between particles are not likely to provide enough. As  $\text{N}-\text{O}$  bonds form, they release energy, so the formation of the new bonds can provide energy to supplement the energy provided by the collisions. The sum of the energy of collision and the energy released in bond formation is more likely to provide enough energy for the reaction.*

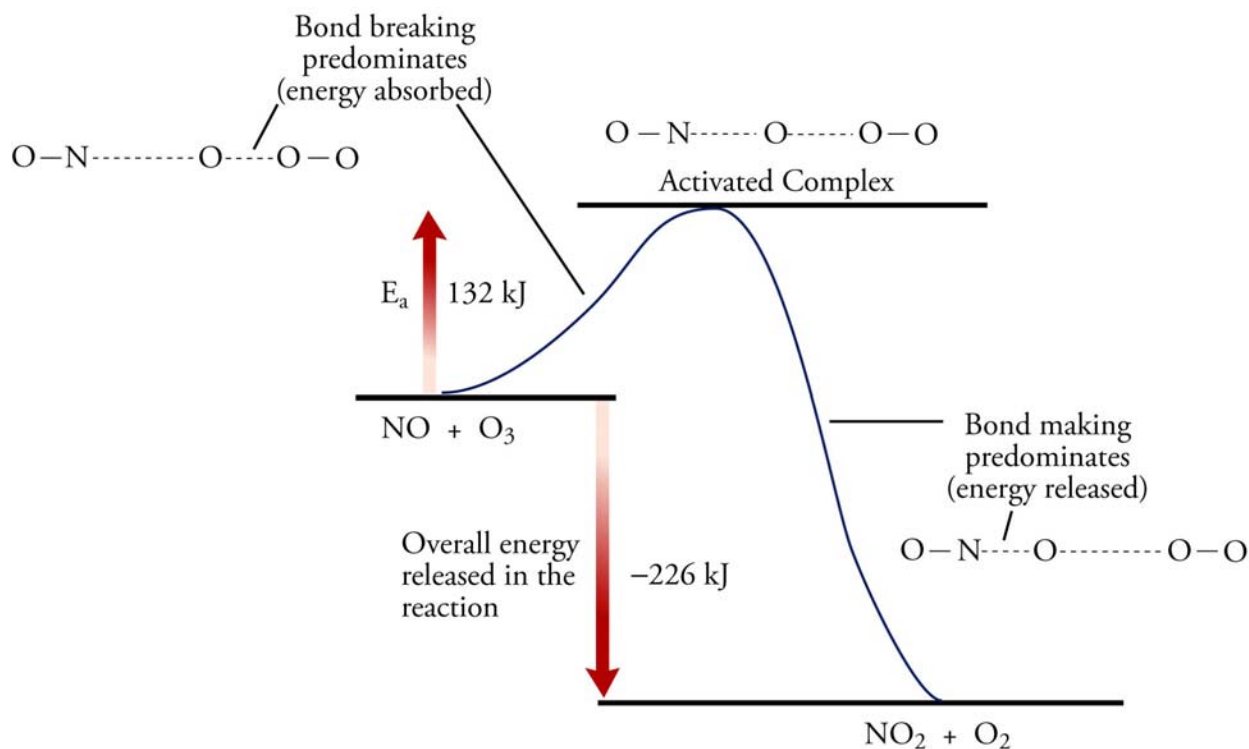
- e. Draw a rough sketch of the activated complex.



- f. Explain why a collision between  $\text{NO}(g)$  and  $\text{O}_3(g)$  must have a certain minimum energy (activation energy) in order to proceed to products. (Obj 5)

*In the initial stage of the reaction, the energy released in bond making is less than the energy absorbed by bond breaking. Therefore, energy must be available from the colliding particles to allow the reaction to proceed. At some point in the change, the energy released in bond formation becomes equal to the energy absorbed in bond breaking. If the colliding particles have enough energy to reach this point (in other words, if they have the activation energy), the reaction proceeds to products.*

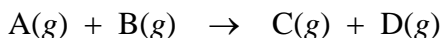
- g. The activation energy for this reaction is 132 kJ/mol. Draw an energy diagram for this reaction, showing the relative energies of the reactants, the activated complex, and the products. Using arrows show the activation energy and heat of reaction. (Obj 7)



- a. Is this reaction exothermic or endothermic? (Obj 6 & 8)  
*The negative sign for the heat of reaction shows that energy is released overall, so the reaction is **exothermic**.*
- b. Explain why  $\text{NO}(g)$  and  $\text{O}_3(g)$  molecules must collide with the correct orientation if a reaction between them is likely to take place. (Obj 10)  
*For a reaction to be likely, new bonds must be made at the same time as other bonds are broken. Therefore, the nitrogen atom in  $\text{NO}$  must collide with one of the outer oxygen atoms in  $\text{O}_3$ .*

### Section 16.2 Rates of Chemical Reactions

35. Consider the following general reaction for which gases A and B are mixed in a constant volume container.



What happens to the rate of this reaction when

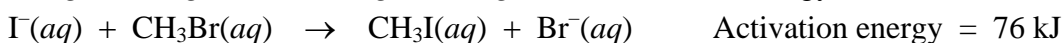
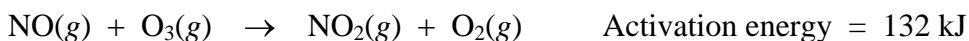
- a. more gas A is added to the container?  
*Increased concentration of reactant A leads to increased rate of collision between A and B and therefore leads to **increased rate of reaction**.*
- b. the temperature is decreased?  
*Decreased temperature leads to decreased average kinetic energy of collisions between A and B. This leads to a decrease in the percentage of collisions with the*

minimum energy necessary for the reaction and therefore leads to **decreased rate of reaction**.

- c. a catalyst is added that lowers the activation energy?

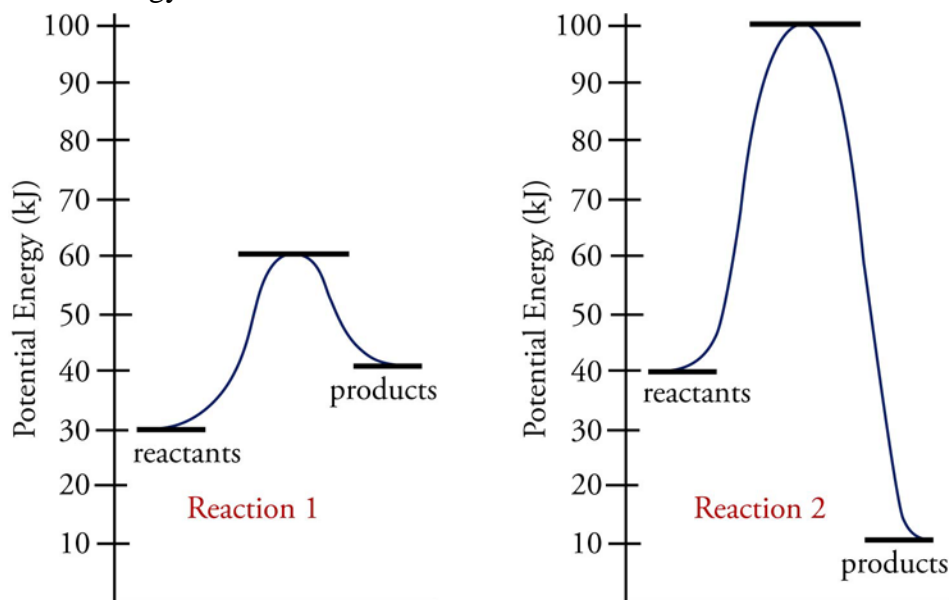
With a lower activation energy, there is a greater percentage of collisions with the minimum energy necessary for the reaction and therefore an **increased rate of reaction**.

37. The reactions listed below are run at the same temperature. The activation energy for the first reaction is 132 kJ/mol. The activation energy for the second reaction is 76 kJ/mol. In which of these reactions would a higher fraction of collisions between reactants have the minimum energy necessary to react (the activation energy)? Explain your answer.



At a particular temperature, the lower the activation energy is, the higher the percentage of collisions with at least that energy or more will be. Thus the **second reaction** would have the higher fraction of collisions with the activation energy.

39. Two reactions can be described by the energy diagrams below. What is the approximate activation energy for each reaction? Which reaction is exothermic and which is endothermic?



The approximate activation energy for reaction 1 is 30 kJ and for reaction 2 is 60 kJ. Reaction 1 is endothermic, and reaction 2 is exothermic.

41. Explain why chlorine atoms speed the conversion of ozone molecules,  $\text{O}_3$ , and oxygen atoms,  $\text{O}$ , into oxygen molecules,  $\text{O}_2$ . (Obj 14)

In part, chlorine atoms are a threat to the ozone layer just because they provide another pathway for the conversion of  $\text{O}_3$  and  $\text{O}$  to  $\text{O}_2$ , but there is another reason. The reaction between  $\text{O}_3$  and  $\text{Cl}$  that forms  $\text{ClO}$  and  $\text{O}_2$  has an activation energy of 2.1 kJ/mole. At 25 °C, about three of every seven collisions (or 43%) have enough energy to reach the activated complex. The reaction between  $\text{O}$  and  $\text{ClO}$  to form  $\text{Cl}$  and  $\text{O}_2$  has an activation energy of only 0.4 kJ/mole. At 25 °C, about 85% of the collisions have at least this energy. The uncatalyzed reaction has an activation energy of about 17 kJ/mole. At 25 °C (298 K), about one of every one thousand collisions (or 0.1%) between  $\text{O}_3$  molecules and

*O atoms has a net kinetic energy large enough to form the activated complex and proceed to products. Thus a much higher fraction of the collisions have the minimum energy necessary to react for the catalyzed reaction than for the direct reaction between  $O_3$  and  $O$ . Thus a much greater fraction of the collisions has the minimum energy necessary for the reaction to proceed for the catalyzed reaction than for the uncatalyzed reaction. Figures 16.14 and 16.15 of the textbook illustrate this.*

43. Using the proposed mechanism for the conversion of  $NO(g)$  into  $N_2(g)$  and  $O_2(g)$  as an example, write a description of the four steps thought to occur in heterogeneous catalysis. (06/16)

**Step 1:** *The reactants ( $NO$  molecules) collide with the surface of the catalyst where they bind to the catalyst. This step is called adsorption. The bonds within the reactant molecules are weakened or even broken as the reactants are adsorbed. ( $N-O$  bonds are broken.)*

**Step 2:** *The adsorbed particles (separate  $N$  and  $O$  atoms) move over the surface of the catalyst.*

**Step 3:** *The adsorbed particles combine to form products ( $N_2$  and  $O_2$ ).*

**Step 4:** *The products ( $N_2$  and  $O_2$ ) leave the catalyst.*

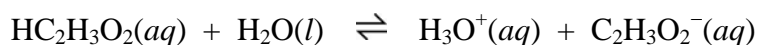
*See Figure 16.16 of the textbook.*

### Section 16.3 Reversible Reactions and Chemical Equilibrium

45. Equilibrium systems have two opposing rates of change that are equal. For each of the following equilibrium systems that were mentioned in earlier chapters, describe what is changing in the two opposing rates.

- a. a solution of the weak acid acetic acid,  $HC_2H_3O_2$  (Chapter 6)

*Acetic acid molecules react with water to form hydronium ions and acetate ions, and at the same time, hydronium ions react with acetate ions to return to acetic acid molecules and water.*



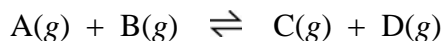
- b. pure liquid in a closed container (Chapter 12)

*Liquids evaporate to form vapor at a rate that is balanced by the return of vapor to liquid.*

- c. a closed bottle of carbonated water with 4 atm of  $CO_2$  in the gas space above the liquid (Chapter 13)

*Carbon dioxide escapes from the solution at a rate that is balanced by the return of  $CO_2$  to the solution.*

47. Two gases, A and B, are added to an empty container. They react in the following reversible reaction.



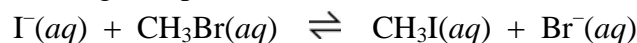
- a. When is the forward reaction rate greatest: (1) when A and B are first mixed, (2) when the reaction reaches equilibrium, or (3) sometime between these two events?

*The forward reaction rate is at its peak when A and B are first mixed. Because A and B concentrations are diminishing as they form C and D, the rate of the forward reaction declines steadily until equilibrium is reached.*

- b. When is the reverse reaction rate greatest: (1) when A and B are first mixed, (2) when the reaction reaches equilibrium, or (3) sometime between these two events?

*The reverse reaction rate is at its peak when the reaction reaches equilibrium. Because C and D concentrations are increasing as they form from A and B, the rate of the reverse reaction increases steadily until equilibrium is reached.*

49. Assume that in the following reversible reaction both the forward and the reverse reactions take place in a single step.



- a. With reference to the changing forward and reverse reaction rates, explain why this reaction moves toward a dynamic equilibrium with equal forward and reverse reaction rates. (Obj 20)

*When  $I^{-}$  ions and  $CH_3Br$  molecules are added to a container, they begin to collide and react. As the reaction proceeds, the concentrations of  $I^{-}$  and  $CH_3Br$  diminish, so the rate of the forward reaction decreases. Initially, there are no  $CH_3I$  molecules or  $Br^{-}$  ions in the container, so the rate of the reverse reaction is initially zero. As the concentrations of  $CH_3I$  and  $Br^{-}$  increase, the rate of the reverse reaction increases.*

*As long as the rate of the forward reaction is greater than the rate of the reverse reaction, the concentrations of the reactants ( $I^{-}$  and  $CH_3Br$ ) will steadily decrease, and the concentrations of products ( $CH_3I$  and  $Br^{-}$ ) will constantly increase. This leads to a decrease in the forward rate of the reaction and an increase in the rate of the reverse reaction. This continues until the two rates become equal. At this point, our system has reached a dynamic equilibrium.*

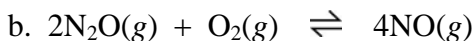
- b. Describe the changes that take place once the reaction reaches an equilibrium state. Are there changes in the concentrations of reactants and products at equilibrium? Explain your answer. (Obj 21)

*In a dynamic equilibrium for reversible chemical reactions, the forward and reverse reaction rates are equal, so although there are constant changes between reactants and products, there is no net change in the amounts of each.  $I^{-}$  and  $CH_3Br$  are constantly reacting to form  $CH_3I$  and  $Br^{-}$ , but  $CH_3I$  and  $Br^{-}$  are reacting to reform  $CH_3Br$  and  $I^{-}$  at the same rate. Thus there is no net change in the amounts of  $I^{-}$ ,  $CH_3Br$ ,  $CH_3I$ , or  $Br^{-}$ .*

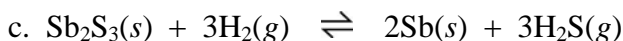
50. Write  $K_C$  and  $K_P$  expressions for each of the following equations. (*Obj's 22 & 23*)



$$K_C = \frac{[\text{C}_2\text{H}_2][\text{H}_2]^3}{[\text{CH}_4]^2} \quad K_P = \frac{P_{\text{C}_2\text{H}_2} P_{\text{H}_2}^3}{P_{\text{CH}_4}^2}$$



$$K_C = \frac{[\text{NO}]^4}{[\text{N}_2\text{O}]^2[\text{O}_2]} \quad K_P = \frac{P_{\text{NO}}^4}{P_{\text{N}_2\text{O}}^2 P_{\text{O}_2}}$$

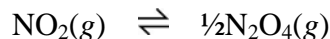


$$K_C = \frac{[\text{H}_2\text{S}]^3}{[\text{H}_2]^3} \quad K_P = \frac{P_{\text{H}_2\text{S}}^3}{P_{\text{H}_2}^3}$$

52. A mixture of nitrogen dioxide and dinitrogen tetroxide is allowed to come to equilibrium at 30 °C, and the gases partial pressures are found to be 1.69 atm  $\text{N}_2\text{O}_4$  and 0.60 atm  $\text{NO}_2$ .

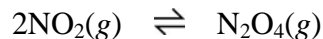
(*Obj 24*)

a. On the basis of these data, what is  $K_P$  for the following equation?



$$K_P = \frac{P_{\text{N}_2\text{O}_4}^{1/2}}{P_{\text{NO}_2}} = \frac{(1.69)^{1/2}}{0.60} = 2.2$$

b. On the basis of these data, what is  $K_P$  for the following equation?



$$K_P = \frac{P_{\text{N}_2\text{O}_4}}{P_{\text{NO}_2}^2} = \frac{1.69}{(0.60)^2} = 4.7$$

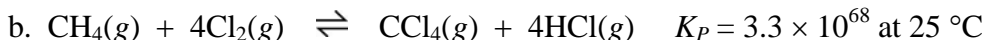
c. Table 14.1 lists the  $K_P$  for the equation in Part (b) as 6.7 at 25 °C. Explain why your answer to Part (b) is not 6.7.

*Changing temperature leads to a change in the value for an equilibrium constant. (Because  $K_P$  for this reaction decreases with increasing temperature, the reaction must be exothermic.)*

54. Predict whether each of the following reactions favors reactants, products, or neither at the temperature for which the equilibrium constant is given. (*Obj 25*)



**$K_P < 10^{-2}$  so reactants favored**

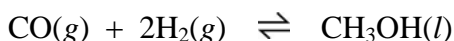


**$K_P > 10^2$  so products favored**



**$10^{-2} < K_P < 10^2$  so neither favored**

56. Write the  $K_c$  expression for the following equation. Explain why the concentration of  $\text{CH}_3\text{OH}$  is left out of the expression. (Obj's 22 & 26)

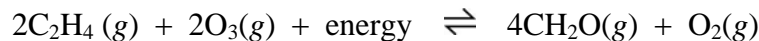


$$K_c = \frac{1}{[\text{CO}][\text{H}_2]^2}$$

If the number of moles of  $\text{CH}_3\text{OH}(l)$  in the container is doubled, its volume doubles too, leaving the concentration (mol/L) of the methanol constant. Increasing or decreasing the total volume of the container will not change the volume occupied by the liquid methanol, so the concentration (mol/L) of the  $\text{CH}_3\text{OH}(l)$  also remains constant with changes in the volume of the container. The constant concentration of methanol can be incorporated into the equilibrium constant itself and left out of the equilibrium constant expression.

$$K' = \frac{[\text{CH}_3\text{OH}]}{[\text{CO}][\text{H}_2]^2} \quad \frac{K'}{[\text{CH}_3\text{OH}]} = \frac{1}{[\text{CO}][\text{H}_2]^2} = K_c$$

58. Ethylene,  $\text{C}_2\text{H}_4$ , is one of the organic substances found in the air we breathe. It reacts with ozone in an endothermic reaction to form formaldehyde,  $\text{CH}_2\text{O}$ , which is one of the substances in smoggy air that cause eye irritation.



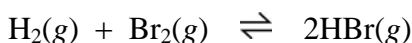
- a. Why does the forward reaction take place more rapidly in Los Angeles than in a wilderness area of Montana with the same air temperature?

*Los Angeles has a much higher ozone concentration than in the Montana wilderness.*

- b. For a variety of reasons, natural systems rarely reach equilibrium, but if this reaction was run in the laboratory, would increased temperature for the reaction at equilibrium shift the reaction to more reactants or more products?

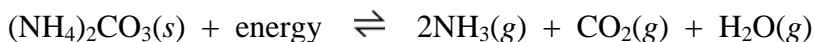
**Toward more products** (*Increased temperature favors the endothermic direction of reversible reactions.*)

60. When the temperature of an equilibrium system for the following reaction is increased, the reaction shifts toward more reactants. Is the reaction endothermic or exothermic?



*Increased temperature favors the endothermic direction of reversible reactions, so this reaction is endothermic in the reverse direction and **exothermic** in the forward direction.*

62. Assume that you are picking up a few extra dollars to pay for textbooks by acting as a trainer's assistant for a heavyweight boxer. One of your jobs is to wave smelling salts under the nose of the fighter to clear his head between rounds. The smelling salts are ammonium carbonate, which decomposes in the following reaction to yield ammonia. The ammonia does the wakeup job. Suppose the fighter gets a particularly nasty punch to the head and needs an extra jolt to be brought back to his senses. How could you shift the following equilibrium to the right to yield more ammonia?



*Increased temperature will drive this endothermic reaction toward products, so warming the smelling salt container in your hands will increase the amount of ammonia released.*

64. Formaldehyde,  $\text{CH}_2\text{O}$ , is one of the components of embalming fluids and has been used to make foam insulation and plywood. It can be made from methanol,  $\text{CH}_3\text{OH}$  (often called wood alcohol). The heat of reaction for the combination of gaseous methanol and oxygen gas to form gaseous formaldehyde and water vapor is  $-199.32$  kJ per mole of  $\text{CH}_2\text{O}$  formed, so the reaction is exothermic.



- a. Increased temperature drives the reaction toward reactants and lowers the value for the equilibrium constant. Explain why this is true. (*Obj's 27 & 28*)

*Increased temperature increases the rate of both the forward and the reverse reactions, but it increases the rate of the endergonic reaction more than it increases the rate of the exergonic reaction. Therefore, changing the temperature of a chemical system at equilibrium will disrupt the balance of the forward and reverse rates of reaction and shift the system in the direction of the endergonic reaction. Because this reaction is exothermic in the forward direction, it must be endothermic in the reverse direction. Increased temperature shifts the system toward more reactants, decreasing the ratio of products to reactants and, therefore, decreasing the equilibrium constant.*

- b. This reaction is run by the chemical industry at  $450\text{-}900$  °C, even though the equilibrium ratio of product to reactant concentrations is lower than at room temperature. Explain why this exothermic chemical reaction is run at high temperature despite this fact. (*Obj 29*)

*To maximize the percent yield at equilibrium, the reaction should be run at as low a temperature as possible, but at low temperature, the rates of the forward and reverse reactions are both very low, so it takes a long time for the system to come to equilibrium. In this case, it is best to run the reaction at high temperature to get to equilibrium quickly. (The unreacted methanol can be recycled back into the original reaction vessel after the formaldehyde has been removed from the product mixture. )*

### Section 16.4 Disruption of Equilibrium

66. Urea,  $\text{NH}_2\text{CONH}_2$ , is an important substance in the production of fertilizers. The equation shown below describes an industrial reaction that produces urea. The heat of reaction is  $-135.7$  kJ per mole of urea formed. Predict whether each of the following changes in the equilibrium system will shift the system to more products, to more reactants, or to neither. Explain each answer in two ways: (1) by applying Le Châtelier's principle and (2) by describing the effect of the change on the forward and reverse reaction rates. (*Obj's 30-33*)

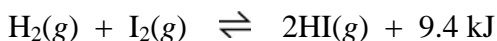


- a. The concentration of  $\text{NH}_3$  is increased by the addition of more  $\text{NH}_3$ . (In the industrial production of urea, an excess of ammonia is added so that the ratio of  $\text{NH}_3$  to  $\text{CO}_2$  is 3:1.)
- *Using Le Châtelier's Principle, we predict that the system will **shift to more products** to partially counteract the increase in  $\text{NH}_3$ .*
  - *The increase in the concentration of ammonia speeds the forward reaction without initially affecting the rate of the reverse reaction. The equilibrium is*

*disrupted, and the system **shifts to more products** because the forward rate is greater than the reverse rate.*

- b. The concentration of  $\text{H}_2\text{O}(g)$  is decreased by removing water vapor.
- *Using Le Châtelier's Principle, we predict that the system will **shift to more products** to partially counteract the decrease in  $\text{H}_2\text{O}$ .*
  - *The decrease in the concentration of  $\text{H}_2\text{O}(g)$  slows the reverse reaction without initially affecting the rate of the forward reaction. The equilibrium is disrupted, and the system **shifts toward more products** because the forward rate is greater than the reverse rate.*
- c. The temperature is increased from  $25\text{ }^\circ\text{C}$  to  $190\text{ }^\circ\text{C}$ . (In the industrial production of urea, ammonia and carbon dioxide are heated to  $190\text{ }^\circ\text{C}$ .)
- *Using Le Châtelier's Principle, we predict that the system shifts in the endothermic direction to partially counteract the increase in temperature. Because the forward reaction is exothermic, the reverse reaction must be endothermic. As the system **shifts toward more reactants**, energy is absorbed, and the temperature decreases.*
  - *The increased temperature increases the rates of both the forward and reverse reactions, but it has a greater effect on the endothermic reaction. Thus the system **shifts toward more reactants** because the reverse rate becomes greater than the forward rate.*

68. Hydriodic acid, which is used to make pharmaceuticals, is made from hydrogen iodide. The hydrogen iodide is made from hydrogen gas and iodine gas in the following exothermic reaction.

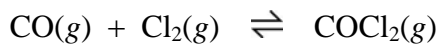


What changes could you make for this reaction at equilibrium to shift the reaction to the right and maximize the concentration of hydrogen iodide in the final product mixture?

*The **addition of either  $\text{H}_2$  or  $\text{I}_2$  (or both)** would increase the concentrations of reactants, increasing the rate of collision between them, increasing the forward rate, and shifting the system toward more product.*

*Lower temperature favors the exothermic direction of the reaction, so **lower temperature** would shift this reaction to a higher percentage of products at equilibrium.*

70. Phosgene gas,  $\text{COCl}_2$ , which is a very toxic substance used to make pesticides and herbicides, is made by passing carbon monoxide gas and chlorine gas over solid carbon, which acts as a catalyst.



If the carbon monoxide concentration is increased by adding CO to an equilibrium system of this reaction, what effect, if any, does it have on the following? (Assume constant temperature.)

- The concentration of  $\text{COCl}_2$  after the system has shifted to come to a new equilibrium.  
*The system will shift toward products, which leads to **increased  $\text{COCl}_2$** .*
- The concentration of  $\text{Cl}_2$  after the system has shifted to come to a new equilibrium.  
*The system will shift toward products, which leads to **decreased  $\text{Cl}_2$** .*
- The equilibrium constant for the reaction.  
*Equilibrium constants are unaffected by reactant and product concentrations, so the **equilibrium constant remains the same**.*