Chapter 19 – Reaction Rates and Equilibrium

Section 19.1 Rates of Reaction

1.1: Explain the collision theory of reactions (q. 38). 19.1

1.2: How is the activation energy of a reaction like a wall or barrier (q. 39)? 19.1

1.3: Where is the formula of a catalyst written in a chemical equation (q. 40)? Why? 19.1

1.4: How is the rate of a reaction influenced by a catalyst? How do catalysts make this possible (q. 41)? 19.1

1.5: When the gas to a stove is turned on, the gas does not burn unless lit by a flame. Once lit, however, the gas burns until turned off. Explain these observations in terms of the effect of temperature on reaction rate (q. 43). 19.1

Section Review 19.1

1. What is meant by the *rate* of a chemical reaction?

2. How does each factor affect the rate of a chemical reaction?

a. temperature –

b. concentration –

c. particle size –

d. an inhibitor –

3. Does every collision between reacting particles lead to products? Explain.

4. Suppose a thin sheet of zinc containing 0.2 mol of the metal is completely covered in air to zinc oxide (ZnO) in one month. How would you express the rate of conversion of the zinc?

5. Refrigerated food stays fresh for long periods. The same food stored at room temperature quickly spoils. Why?

Section 19.2 Reversible Reactions and Equilibrium

2.1: In your own words, define a *reversible reaction* (q. 44). 19.2

2.1: A reversible reaction has reached a state of *dynamic equilibrium*. What does this information tell you (q. 45)? 19.2

2.3: How do the rates of the forward and reverse reactions compare at a state of dynamic chemical equilibrium (q. 46)? 19.2

2.4: What is Le Chatelier's principle? Use it to explain why carbonated drinks go flat when their containers are left open (q. 47). 19.2

Practice Problems

6. How is the equilibrium position of this reaction affected by the following changes?

$$C(s) + H_2O(g) + \text{heat} \rightleftharpoons CO(g) + H_2(g)$$

a. lowering the temperature –

b. increasing the pressure –

c. removing H_2 –

d. adding H_2 –

7. What effect does each change have on the equilibrium position of this reaction?

$N_2(g) + 3\ H_2(g) \rightleftharpoons 2\ NH_3(g) + 92\ kJ$

a. addition of heat –

b. addition of pressure –

c. addition of catalyst –

d. removal of heat –

8. Analysis of an equilibrium mixture of nitrogen, hydrogen, and ammonia contained in a 1-L flask at 300℃ gives the following results: hydrogen, 0.15 mol; nitrogen, 0.25 mol; ammonia, 0.10 mol. Calculate K_{eq} for the reaction.

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

9. Assume you have the mixture described in Problem 8 with the same volume, temperature, and equilibrium concentrations.

a. Calculate K_{eq} for the reaction $2\ NH_3(g) \rightleftharpoons 3\ H_2(g) + N_2(g)$

b. Based on your answers to Problem 8 and Problem 9a, how is the K_{eq} for a forward reaction related to the K_{eq} for a reverse reaction?

10. Suppose the following system reaches equilibrium at a high temperature.

$N_2(g) + O_2(g) \rightleftharpoons 2\ NO(g)$

An analysis of the equilibrium mixture in a 1-L flask gives the following results: nitrogen, 0.50 mol; oxygen, 0.50 mol; nitrogen monoxide, 0.020 mol. Calculate K_{eq} for the reaction.

11. At 750℃ the following reaction reaches equilibrium in a 1-L container.

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

An analysis of the equilibrium mixture provides the following results: hydrogen, 0.053 mol; carbon dioxide, 0.053 mol; water vapor, 0.047 mol; carbon monoxide, 0.047 mol. Calculate K_{eq} for the reaction.

12. The decomposition of hydrogen iodide at 450℃ in a 1-L container produces an equilibrium mixture that contains 0.50 mol of hydrogen. The equilibrium constant is 0.020 for the reaction. How many moles of iodine and hydrogen iodide are present in the equilibrium mixture?

$2\ HI(g) \rightleftharpoons H_2(g) + I_2(g)$

13. Write the expression for the equilibrium constant for each:

a. $2\ HBr(g) \rightleftharpoons H_2(g) + Br_2(g)$

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

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

d. $4\ NH_3(g) + 5\ O_2(g) \rightleftharpoons 6\ H_2O(g) + 4\ NO(g)$

Section Review 19.2

14. How can changes in the equilibrium position be predicted from changes in concentration, temperature, and pressure?

15. How can a balanced chemical equation and experimental data be used to write an equilibrium-constant expression and to calculate its value?

16. What is the significance of double arrows in an equation?

17. How do the amounts of reactants and products change once a reaction has achieved chemical equilibrium?

18. Can a pressure change shift the equilibrium position in every reversible reaction? Explain your answer.

19. Imagine you have determined the following equilibrium constants for several reactions. In which of these reactions are the products favored over the reactants? Why?

a. $K_{eq} = 1 \times 10^2$ b. $K_{eq} = 0.003$ c. $K_{eq} = 3.5$ d. $K_{eq} = 6 \times 10^{-4}$

Section 19.3 Determining Whether a Reaction Will Occur

Section Review 19.3

20 Explain what is meant by the following:

a. entropy

b. free energy

c. spontaneous reaction

d. nonspontaneous reaction

21. What two factors determine whether a reaction is spontaneous?

a.

b.

22. Where does lost free energy typically end up? Does free energy lost as heat ever serve a useful function? Explain.

23. What can change a reaction from nonspontaneous to spontaneous?

24. Suppose the products in a spontaneous process are more ordered than the reactants. Is the entropy change favorable or unfavorable? Explain.

Section 19.4 Calculating Entropy and Free Energy

Practice Problem

25. The standard entropies for some substances at 25℃ are

$KBrO_3(s)$ S^0 = 149.2 J/K · mol
$KBr(s)$ S^0 = 96.4 J/ K · mol
$O_2(g)$ S^0 = 205.0 J/ K · mol

Calculate the ΔS^0 for the reaction

$KBrO_3(s) \rightarrow KBr(s) + 3/2\ O_2(g)$