Name _____ Date _____ Class _____

CHAPTER 18 STUDY GUIDE FOR CONTENT MASTERY

Chemical Equilibrium

Section 18.1 Equilibrium: A State of Dynamic Balance

In your textbook, read about chemical equilibrium.

Complete each statement.

1. When a reaction results in almost complete conversion of reactants to products, chemists say the reaction goes to _____.
2. A reaction that can occur in both the forward and the reverse directions is called a(n) _____.
3. _____ is a state in which the forward and reverse reactions balance each other because they take place at equal rates.
4. At equilibrium, the concentrations of reactants and products are _____, but that does not mean that the amounts or concentrations are _____.
5. Equilibrium is a state of _____, not one of _____.

In your textbook, read about equilibrium expressions and constants.

For each statement below, write *true* or *false*.

_____ 6. The law of chemical equilibrium states that at a given pressure, a chemical system may reach a state in which a particular ratio of reactant to product concentrations has a constant value.

_____ 7. The equation $H_2(g) + I_2(g) \rightleftharpoons 2HI(g)$ is an example of a homogeneous equilibrium.

_____ 8. If an equilibrium constant has a value less than one, the reactants are favored at equilibrium.

_____ 9. The value for K_{eq} is constant only at a specific volume.

_____ 10. If the equilibrium constant for a reaction at 300 K is 49.7, the concentration of the reactants will be greater than the concentration of the products.

_____ 11. A heterogeneous equilibrium means that reactants and products are present in more than one state.

_____ 12. The product of the forward chemical reaction is HI, for the equilibrium expression:

$$K_{eq} = \frac{[HI]^2}{[H_2][I_2]}$$

Section 18.1 *continued*

In your textbook, read about determining equilibrium constants.

A chemist did two experiments to determine the equilibrium constant for the reaction of sulfur dioxide with oxygen to form sulfur trioxide. Use the table showing the results of the experiments to answer the following questions.

$2SO_2(g) + O_2(g) \rightleftharpoons 2SO_3(g)$ at 873 K			
Experiment 1		**Experiment 2**	
Initial concentrations	Equilibrium concentrations	Initial concentrations	Equilibrium concentration
$[SO_2] = 2.00M$	$[SO_2] = 1.50M$	$[SO_2] = 0.500M$	$[SO_2] = 0.590M$
$[O_2] = 1.50M$	$[O_2] = 1.26M$	$[O_2] = 0M$	$[O_2] = 0.0450M$
$[SO_3] = 3.00M$	$[SO_3] = 3.50M$	$[SO_3] = 0.350M$	$[SO_3] = 0.260M$

13. Write the equation to calculate the equilibrium constant for the reaction.

_____

14. Is this reaction an example of a homogeneous or heterogeneous equilibrium?

_____

15. Calculate the equilibrium constant from the data obtained in experiment 1.

_____

_____

16. What is the equilibrium constant for the reaction in experiment 2?

_____

_____

17. Was it necessary to calculate both equilibrium constants? Why or why not?

_____

_____

18. What does this experiment show about the initial concentrations of products and reactants in a reversible reaction?

_____

_____

_____

Section 18.2 Factors Affecting Chemical Equilibrium

In your textbook, read about Le Châtelier's Principle.

Answer the following questions.

1. What does Le Châtelier's Principle say?

_____

_____

2. What are three kinds of stresses that can be placed on a system?

_____

_____

For each reaction below, state the direction, left or right, in which the equilibrium will shift when the indicated substance is added. Identify one other way in which the reaction could be shifted in the same direction you indicated. (Hint: There may be more than one way to do this.)

3. Reaction: $N_2(g) + 3H_2(g) \rightleftharpoons 2NH_3(g)$; NH_3 added

_____

4. Reaction: $H_2(g) + I_2(g) \rightleftharpoons 2HI(g)$; H_2 added

_____

5. Reaction: $CO(g) + H_2O \rightleftharpoons CO_2(g) + H_2(g)$; H_2O added

_____

6. Reaction: $2SO_2(g) + O_2(g) \rightleftharpoons 2SO_3(g)$; SO_3 added

_____

7. Reaction: $2SO_2(g) + O_2(g) \rightleftharpoons 2SO_3(g)$; SO_2 added

_____

8. Reaction: $2NCl_3(g) \rightleftharpoons N_2(g) + 3Cl_2(g)$; NCl_3 added

_____

Section 18.2 *continued*

In your textbook, read about factors affecting chemical equilibrium.

Use each of the terms below just once to complete the passage.

right	exothermic	increase	stress	catalyst	energy
smallest	change	reverse	constant	forward	

When you decrease the volume of a reaction vessel, you **(9)**_____ the pressure. This causes a reaction at equilibrium to shift to the side with the **(10)**_____ number of moles. If the reaction has an equal number of moles of reactants and products, changing the volume of the reaction vessel causes no **(11)**_____ in the equilibrium.

Changing the temperature of a reaction at equilibrium alters both the equilibrium **(12)**_____ and the equilibrium position. When a reaction is **(13)**_____, which means it releases energy, lowering the temperature shifts the equilibrium to the **(14)**_____ because the forward reaction liberates heat and removes the **(15)**_____.

A **(16)**_____ speeds up a reaction by lowering the **(17)**_____ requirements for the reaction, but it does so equally in both the **(18)**_____ and the **(19)**_____ directions. The reaction will reach equilibrium more quickly, but with no change in the amount of product formed.

For each reaction below, indicate in which direction the equilibrium shifts when the stated stress is applied to the system. Write *R* if the reaction shifts to the right, *L* if it shifts to the left, or *NC* if there is no change.

	Reaction	Stress
_____ **20.**	$PCl_5(g) \rightleftharpoons PCl_3(g) + Cl_2(g) + \text{heat}$	temperature increase
_____ **21.**	$CO(g) + Fe_3O_4(s) \rightleftharpoons CO_2(g) + 3FeO(s)$	volume increase
_____ **22.**	$C_2H_2(g) + H_2O(g) \rightleftharpoons CH_3CHO(g) + \text{heat}$	temperature decrease
_____ **23.**	$2NO(g) + H_2(g) \rightleftharpoons N_2O(g) + H_2O(g) + \text{heat}$	volume decrease
_____ **24.**	$\text{Heat} + H_2(g) + I_2(g) \rightleftharpoons 2HI(g)$	temperature decrease
_____ **25.**	$H_2(g) + Cl_2(g) \rightleftharpoons 2HCl(g) + \text{heat}$	volume decrease

Section 18.3 Using Equilibrium Constants

In your textbook, read about calculating equilibrium concentrations.

Answer the following questions.

1. What can you use the equilibrium constant to do?

2. Given the reaction: $N_2 + O_2 \rightleftharpoons 2NO$ for which the K_{eq} at 2273 K is 1.2×10^{-4}

a. Write the equilibrium constant expression for the reaction.

b. Write the equation that would allow you solve for the concentration of NO.

c. What is the concentration of NO if $[N_2] = 0.166M$ and $[O_2] = 0.145M$?

3. What is the solubility product constant?

4. What is the solubility product constant expression for the reaction: $Mg_3(PO_4)_2(s) \rightleftharpoons 3Mg^{2+}(aq) + 2PO_4^{3-}(aq)$

5. Given the equilibrium $BaSO_4(s) \rightleftharpoons Ba^{2+}(aq) + SO_4^{2-}(aq)$, what is the solubility product constant expression?

6. The solubility product constant for $BaSO_4$ at 298 K is 1.1×10^{-10}. Calculate the solubility of $BaSO_4$ in mol/L at 298 K.

Section 18.3 *continued*

In your textbook, read about predicting precipitates.

The solubility product constant can be used to determine if a precipitate will form when two aqueous solutions are mixed together. First, calculate the concentrations of the ions in the final solution. Use the solubility product constant expression to calculate the ion product (Q_{sp}) for the substance that might precipitate. Compare the result with the K_{sp} of the substance.

7. What can you say about a solution when

a. Q_{sp} is greater than K_{sp}?

b. Q_{sp} is equal to K_{sp}?

c. Q_{sp} is less than K_{sp}?

8. Predict whether a precipitate of AgBr will form if 100 mL of 0.0025*M* $AgNO_3$ and 100 mL of 0.0020*M* NaBr are mixed.

9. Explain briefly why Ag_3PO_4 might be more soluble in water than in the same volume of a solution containing Na_3PO_4.