Pearson

Chapter Summary and Review

Key Learning Outcomes

Chapter Objectives	Assessment
Express Equilibrium Constants for Chemical Equations (15.3)	• Example 15.1 For Practice 15.1 Exercises 21, 22
Manipulate the Equilibrium Constant to Reflect Changes in the Chemical Equation (15.3)	• Example 15.2 For Practice 15.2 For More Practice 15.2 Exercises 27, 28, 29, 30
Relate K_p and K_c (15.4)	• Example 15.3 For Practice 15.3 Exercises 31, 32
Write Equilibrium Expressions for Reactions Involving a Solid or a Liquid (15.5)	• Example 15.4 For Practice 15.4 Exercises 33, 34
Find Equilibrium Constants from Experimental Concentration Measurements (15.6)	• Examples 15.5, 15.6 For Practice 15.5, 15.6 Exercises 35, 36, 43, 44
Predict the Direction of a Reaction by Comparing *Q* and *K* (15.7)	• Example 15.7 For Practice 15.7 Exercises 47, 48, 49, 50
Calculate Equilibrium Concentrations from the Equilibrium Constant and One or More Equilibrium Concentrations (15.8)	• Example 15.8 For Practice 15.8 Exercises 37, 38, 39, 40, 41, 42, 43, 44, 45, 46
Find Equilibrium Concentrations from Initial Concentrations and the Equilibrium Constant (15.8)	• Examples 15.9, 15.10 For Practice 15.9, 15.10 Exercises 53, 54, 55, 56, 57, 58
Calculate Equilibrium Partial Pressures from the Equilibrium Constant and Initial Partial Pressures (15.8)	• Example 15.11 For Practice 15.11 Exercises 59, 60
Find Equilibrium Concentrations from Initial Concentrations in Cases with a Small Equilibrium Constant (15.8)	• Examples 15.12, 15.13 For Practice 15.12, 15.13 Exercises 61, 62
Determine the Effect of a Concentration Change on	• Example 15.14 For Practice 15.14

Determine the Effect of a Concentration Change on Equilibrium (15.9)	• Example 15.14 For Practice 15.14 Exercises 63, 64, 65, 66
Determine the Effect of a Volume Change on Equilibrium (15.9)	• Example 15.15 For Practice 15.15 Exercises 67, 68
Determine the Effect of a Temperature Change on Equilibrium (15.9)	• Example 15.16 For Practice 15.16 Exercises 69, 70

Key Terms

Section 15.2

reversible

dynamic equilibrium

Section 15.3

equilibrium constant (K)

law of mass action

Section 15.7

reaction quotient (Q)

Section 15.9

Le Châtelier's principle

Key Concepts

The Equilibrium Constant (15.1)

- The equilibrium constant, K expresses the relative concentrations of the reactants and the products at equilibrium.
- The equilibrium constant measures how far a reaction proceeds toward products: a large K (much greater than 1) indicates a high concentration of products at equilibrium, and a small K (much less than 1) indicates a low concentration of products at equilibrium.

Dynamic Equilibrium (15.2)

- Most chemical reactions are reversible; they can proceed in either the forward or the reverse direction.
- When a chemical reaction is in dynamic equilibrium, the rate of the forward reaction equals the rate of the reverse reaction, so the net concentrations of reactants and products do not change. However, this does *not* imply that the concentrations of the reactants and the products are equal at equilibrium.

The Equilibrium Constant Expression (15.3)

- The equilibrium constant expression is given by the law of mass action and is equal to the concentrations of the products, raised to their stoichiometric coefficients, divided by the concentrations of the reactants, raised to their stoichiometric coefficients.

- When the equation for a chemical reaction is reversed, multiplied, or added to another equation, *K* must be modified accordingly.

The Equilibrium Constant, *K* (15.4)

- The equilibrium constant can be expressed in terms of concentrations (K_c) or in terms of partial pressures (K_p). These two constants are related by Equation 15.2. Concentration must always be expressed in units of molarity for K_c. Partial pressures must always be expressed in units of atmospheres for K_p.

States of Matter and the Equilibrium Constant (15.5)

- The equilibrium constant expression contains only partial pressures or concentrations of reactants and products that exist as gases or solutes dissolved in solution. Pure liquids and solids are not included in the expression for the equilibrium constant.

Calculating *K* (15.6)

- We can calculate the equilibrium constant from equilibrium concentrations or partial pressures by substituting measured values into the expression for the equilibrium constant (as obtained from the law of mass action).
- In most cases, we can calculate the equilibrium concentrations of the reactants and products—and therefore the value of the equilibrium constant—from the initial concentrations of the reactants and products and the equilibrium concentration of *just one* reactant or product.

The Reaction Quotient, *Q* (15.7)

- The reaction quotient, *Q*, is the ratio of the concentrations (or partial pressures) of products raised to their stoichiometric coefficients to the concentrations of reactants raised to their stoichiometric coefficients *at any point in the reaction*.
- Like *K*, *Q* can be expressed in terms of concentrations (Q_c) or partial pressures (Q_p).
- At equilibrium, *Q* is equal to *K*; therefore, we can determine the direction in which a reaction proceeds by comparing *Q* to *K*. If $Q < K$, the reaction moves in the direction of the products; if $Q > K$, the reaction moves in the reverse direction.

Finding Equilibrium Concentrations (15.8)

- There are two general types of problems in which *K* is given and one (or more) equilibrium concentrations can be found:
  1. *K* and all but one equilibrium concentration are given.
  2. *K* and *only* initial concentrations are given.
- We solve the first type by rearranging the law of mass action and substituting the given values.
- We solve the second type by creating an ICE table and using a variable *x* to represent the change in concentration.

Le Châtelier's Principle (15.9)

- When a system at equilibrium is disturbed—by a change in the amount of a reactant or product, a change in volume, or a change in temperature—the system shifts in the direction that minimizes the disturbance.

Key Equations and Relationships

Expression for the Equilibrium Constant, K_c (15.3)

$$a\mathrm{A} + b\mathrm{B} \rightleftharpoons c\mathrm{C} + d\mathrm{D}$$

$$K = \frac{[\mathrm{C}]^c[\mathrm{D}]^d}{[\mathrm{A}]^a[\mathrm{B}]^b}\left(\text{equilibrium concentrations only}\right)$$

Relationship between the Equilibrium Constant and the Chemical Equation (15.3)

1. If you reverse the equation, invert the equilibrium constant.
2. If you multiply the coefficients in the equation by a factor, raise the equilibrium constant to the same factor.
3. If you add two or more individual chemical equations to obtain an overall equation, multiply the corresponding equilibrium constants by each other to obtain the overall equilibrium constant.

Expression for the Equilibrium Constant, K_p (15.4)

$$aA + bB \rightleftharpoons cC + dD$$

$$K_p = \frac{P_C^c P_D^d}{P_A^a P_B^b} \text{(equilibrium partial pressures only)}$$

Relationship between the Equilibrium Constants, K_c and K_p (15.4)

$$K_P = K_c(RT)^{\Delta n}$$

The Reaction Quotient, Q_C (15.7)

$$aA + bB \rightleftharpoons cC + dD$$

$$Q_C = \frac{[C]^c[D]^d}{[A]^a[B]^b} \text{(concentration at any point in the reaction)}$$

The Reaction Quotient, Q_P (15.7)

$$aA + bB \rightleftharpoons cC + dD$$

$$Q_P = \frac{P\overset{c}{C}P\overset{d}{D}}{P\overset{a}{A}P\overset{b}{B}} \text{(partial pressures at any point int the reaction)}$$

Relationship of Q to the Direction of the Reaction (15.7)

$Q < K$ Reaction goes to the right.

$Q > K$ Reaction goes to the left.

$Q = K$ Reaction is at equilibrium.