

Chapter Summary and Review

Key Learning Outcomes

Chapter Objectives	Assessment
Calculate the pH of a Buffer Solution (17.2)	<ul style="list-style-type: none"> Example 17.1 For Practice 17.1 For More Practice 17.1 Exercises 27, 28, 31, 32
Use the Henderson–Hasselbalch Equation to Calculate the pH of a Buffer Solution (17.2)	<ul style="list-style-type: none"> Example 17.2 For Practice 17.2 Exercises 35, 36, 37, 38, 39, 40
Calculate the pH Change in a Buffer Solution after the Addition of a Small Amount of Strong Acid or Base (17.2)	<ul style="list-style-type: none"> Example 17.3 For Practice 17.3 For More Practice 17.3 Exercises 45, 46, 47, 48
Use the Henderson–Hasselbalch Equation to Calculate the pH of a Buffer Solution Composed of a Weak Base and Its Conjugate Acid (17.2)	<ul style="list-style-type: none"> Example 17.4 For Practice 17.4 For More Practice 17.4 Exercises 35, 36, 37, 38, 39, 40
Determine Buffer Range (17.3)	<ul style="list-style-type: none"> Example 17.5 For Practice 17.5 Exercises 55, 56
Determine pH Values for a Strong Acid–Strong Base Titration (17.4)	<ul style="list-style-type: none"> Example 17.6 For Practice 17.6 Exercises 65, 66, 67, 68
Determine pH Values for a Weak Acid–Strong Base Titration (17.4)	<ul style="list-style-type: none"> Example 17.7 For Practice 17.7 Exercises 63, 64, 69, 70, 73, 75, 76, 77, 78
Calculate Molar Solubility from K_{sp} (17.5)	<ul style="list-style-type: none"> Example 17.8 For Practice 17.8 Exercises 85, 86
Calculate K_{sp} from Molar Solubility (17.5)	<ul style="list-style-type: none"> Example 17.9 For Practice 17.9 Exercises 87, 88, 89, 90, 91, 92
Calculate Molar Solubility in the Presence of a Common Ion (17.5)	<ul style="list-style-type: none"> Example 17.10 For Practice 17.10 Exercises 93, 94

Determine the Effect of pH on Solubility (17.5)	<ul style="list-style-type: none">• Example 17.11 For Practice 17.11 Exercises 95, 96, 97, 98
Predict Precipitation Reactions by Comparing Q and K_{sp} (17.6)	<ul style="list-style-type: none">• Example 17.12 For Practice 17.12 Exercises 99, 100, 101, 102
Find the Minimum Required Reagent Concentration for Selective Precipitation (17.6)	<ul style="list-style-type: none">• Example 17.13 For Practice 17.13 Exercises 103, 104
Find the Concentrations of Ions Left in Solution after Selective Precipitation (17.6)	<ul style="list-style-type: none">• Example 17.14 For Practice 17.14 Exercises 105, 106
Work with Complex Ion Equilibria (17.7)	<ul style="list-style-type: none">• Example 17.15 For Practice 17.15 Exercises 107, 108, 109, 110

Key Terms

Section 17.2

[buffer](#)
[common ion effect](#)
[Henderson–Hasselbalch equation](#)

Section 17.3

[buffer capacity](#)

Section 17.4

[acid–base titration](#)
[indicator](#)
[equivalence point](#)
[endpoint](#)

Section 17.5

[solubility-product constant \(\$K_{sp}\$ \)](#)
[molar solubility](#)

Section 17.6

[selective precipitation](#)

Section 17.7

[complex ion](#)
[ligand](#)
[formation constant \(\$K_f\$ \)](#)

Key Concepts

The Dangers of Antifreeze (17.1)

- Although buffers closely regulate the pH of mammalian blood, the capacity of these buffers to neutralize can be overwhelmed.
- Ethylene glycol, the main component of antifreeze, is metabolized by the liver into glycolic acid. The resulting acidity can exceed the buffering capacity of blood and cause acidosis, a serious condition that results in oxygen deprivation.

Buffers: Solutions That Resist pH Change (17.2)

- A buffer contains significant amounts of both a weak acid and its conjugate base (or a weak base and its conjugate acid), enabling the buffer to neutralize added acid or added base.
- Adding a small amount of acid to a buffer converts a stoichiometric amount of base to the conjugate acid. Adding a small amount of base to a buffer converts a stoichiometric amount of the acid to the conjugate base.
- We can determine the pH of a buffer solution by solving an equilibrium problem, focusing on the common ion effect, or using the Henderson–Hasselbalch equation.

Buffer Range and Buffer Capacity (17.3)

- A buffer works best when the amounts of acid and conjugate base it contains are large and approximately equal.
- If the relative amounts of acid and base in a buffer differ by more than a factor of 10, the ability of the buffer to neutralize added acid and added base diminishes. The maximum pH range at which a buffer is effective is one pH unit on either side of the acid's pK_a .

Titration Curves (17.4)

- A titration curve is a graph of the change in pH versus the volume of acid or base that is added during a titration.
- This chapter examines three types of titration curves, representing three types of acid–base reactions: a strong acid with a strong base, a weak acid with a strong base (or vice versa), and a polyprotic acid with a base.
- The equivalence point of a titration can be made visible by an indicator, a compound that changes color over a specific pH range.

Solubility Equilibria and the Solubility-Product Constant (17.5)

- The solubility-product constant (K_{sp}) is an equilibrium constant for the dissolution of an ionic compound in water.
- We can determine the molar solubility of an ionic compound from K_{sp} and vice versa. Although the value of K_{sp} is constant at a given temperature, the solubility of an ionic substance can depend on other factors such as the presence of common ions and the pH of the solution.

Precipitation (17.6)

- We can compare the magnitude of K_{sp} to the reaction quotient, Q , in order to determine the relative saturation of a solution.
- Substances with cations that have sufficiently different values of K_{sp} can be separated by selective precipitation, in which an added reagent forms a precipitate with one of the dissolved cations but not others.

Complex Ion Equilibria (17.7)

- A complex ion contains a central metal ion bound to two or more ligands.
- The equilibrium constant for the formation of a complex ion is called a formation constant and is usually quite large.

- The solubility of an ionic compound containing a metal cation that forms complex ions increases in the presence of Lewis bases that form a complex ion with the cation because the formation of the complex ion drives the dissolution reaction to the right.
- All metal hydroxides become more soluble in the presence of acids, but amphoteric metal hydroxides also become more soluble in the presence of bases.

Key Equations and Relationships

The Henderson–Hasselbalch Equation (17.2)

$$\text{pH} = \text{p}K_a + \log \frac{[\text{base}]}{[\text{acid}]}$$

Effective Buffer Range (17.3)

$$\text{pH range} = \text{p}K_a \pm 1$$

The Relation between Q and K_{sp} (17.6)

If $Q < K_{\text{sp}}$, the solution is unsaturated. More of the solid ionic compound can dissolve in the solution.

If $Q = K_{\text{sp}}$, the solution is saturated. The solution is holding the equilibrium amount of the dissolved ions, and additional solid will not dissolve in the solution.

If $Q > K_{\text{sp}}$, the solution is supersaturated. Under most circumstances, the solid will precipitate out of a supersaturated solution.

Not for Distribution

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