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UNIT 8

Chemical Equilibrium Focussing on Acid-Base Systems

Teaching Unit 8: Chemical Equilibrium Focussing on Acid-Base Systems

Chapter 16 *Chemical Equilibrium*

Answers to Questions for Comprehension, notes on Figures and Web Links appear throughout the chapter.

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Teaching Unit 8: Chemical Equilibrium Focussing on Acid-Base Systems

(25% of the course time; approximately 30 hours)

Student Textbook pages 624–727

General Outcomes

- explain that there is a balance of opposing reactions in chemical equilibrium systems
- determine quantitative relationships in simple equilibrium systems

Contents

Chapter 16: Chemical Equilibrium
Chapter 17: Acid-Base Equilibrium Systems

Content Summary

To this point in their studies of chemistry, students have been taught to assume that reactants change completely to products in chemical reactions. That is, the majority of students think that reactions proceed in one direction only. In fact, a closed reacting system produces an equilibrium mixture of reactants and products, and this unit introduces students to these systems. Students first learn how to recognize equilibrium, and the conditions necessary for equilibrium to become established. Next, they learn how to measure an equilibrium constant and how to interpret its value. Le Châtelier's principle allows qualitative predictions to be made about shifts in the position of equilibrium when changes are made to the system, and students will learn how this principle is used by some chemical industries in Alberta. Students will broaden their understanding of acid-base chemistry from earlier studies as they apply equilibrium concepts to the ionization of weak acids and bases. They also examine indicators, acid-base titrations and buffer solutions, and how buffer systems help to maintain the pH of body fluids such as blood within a very narrow range of pH.

Chapter 16 introduces students to the concept of reversible changes and chemical reactions at equilibrium. The macroscopic and microscopic changes that occur in reversible systems are illustrated and the conditions that lead to equilibrium are outlined. In Investigation 16.A: Modelling Equilibrium (pp. 635-636), students model a reversible change to discover the criteria that apply to chemical systems at equilibrium. Systems in equilibrium have a constant ratio between the concentrations of products and reactants. In Section 16.2, this ratio is examined and is quantified through the definitions of the equilibrium expression and equilibrium constant. Le Châtelier's principle is introduced, and in Investigation 16.B: Disturbing Equilibrium students use the principle to check their predictions for the direction of equilibrium shift. In Section 16.3, students use the equilibrium expression in solving a number of different equilibrium problems involving the concentrations of products and reactants in homogeneous systems. Investigation 16.C: Using Experimental Data to Determine an Equilibrium Constant will enable students to collect data for a system in equilibrium and determine the equilibrium constant for a mixture of $Fe^{3+}(aq)$ and $SCN^{-}(aq)$ ions in solution.

The final section in Chapter 16 looks at real-life applications of equilibrium systems. Physical and chemical equilibrium are important factors involving the gases dissolved in blood, and students learn how this applies to scuba diving and carbon monoxide poisoning. Students will learn how the formation of cave structures and corals depends on the concentration of dissolved carbonate ions in water, and that the burning of fossil fuels is changing this equilibrium. Next, Alberta's important role in the chemical industry is outlined. Students learn how the Haber-Bosch process applies equilibrium principles in the manufacture of ammonia. In addition, Alberta extracts millions of tonnes of sulfur each year from natural gas and petroleum. The contact process uses equilibrium principles to manufacture sulfuric acid, a chemical that is essential for many industries. Finally, the equilibrium principles used to manufacture syngas are examined, using the example of the production of methanol at the Celanese plant in Edmonton.

Chapter 17 extends students' understanding of acid-base chemistry by applying the principles of chemical equilibrium to weak acids and weak bases. The chapter begins by examining the causes of acid deposition and the role of rocks and soil as natural buffers. In Section 17.2, acid-base reactions are described by the Brønsted-Lowry theory as proton transfer reactions. Students learn how to write Brønsted-Lowry equations, identify conjugate acid-base pairs, and how to use the strength of an acid or a base to predict the direction of an acid-base reaction. In the Section 17.3, the acid and base dissociation constants (K_a and K_b) are defined. In Investigation 17.A: Determining K_a for Acetic Acid, students perform a titration and measure pH to determine K_a for ethanoic acid. Acid-base ionization constants are then used to perform calculations to determine pH, pOH, $[H_3O^+]$, and $[OH^-]$ for aqueous solutions. The relationship between conjugate acid-base pairs ($K_a \times K_b = K_w$) allows students to calculate acid or base constants for conjugate ions. They will use this information to predict the direction of an acid-base reaction between ions in solution.

In Section 17.4, students learn how equilibrium applies to the choice of indicators for acid-base titrations. They examine titration curves for various combinations of strong or weak acids and bases, and in Thought Lab 17.1: Analyzing a Weak Acid-Strong Base Titration, students will use a spreadsheet program to plot a titration curve and determine the pH at the equivalence point. Students will prepare a buffer solution and investigate its properties. Finally, students learn how different buffer systems are important in the control of blood pH and the transfer of oxygen to muscle tissues.

Curriculum Fit

This unit builds on:

- *Science 8*, Unit 1: Solutions and Substances

- *Science 9*, Unit B: Matter and Chemical Change
- *Science 10*, Unit A: Energy and Matter in Chemical Change
- *Chemistry 20*, Unit A: Matter as Solutions, Acids, Bases
- *Chemistry 20*, Unit D: Quantitative Relationships in Chemical Changes

Core Concepts

Concept	**Outcome**	**Text Reference**
All isolated systems, and closed systems at constant temperature, eventually reach a state of equilibrium	30–D1.1k	Section 16.1, p. 637
In systems at equilibrium, forward and reverse changes are taking place at the same rate.	30–D1.1k	Section 16.1, p. 634; p. 637
Equilibrium involves dynamic change at the molecular level but no change at the macroscopic level.	30–D1.1k	Section 16.1, p. 634; p. 637
Equilibrium can be approached from either direction.	30–D1.1k	Section 16.1, pp. 637-638; Section 16.2, p. 642
In a chemical system at equilibrium, there is a constant ratio between the concentrations of the products and the concentrations of the reactants: $K_c = \frac{[C]^c[D]^d}{[A]^a[B]^b}$ (Law of Chemical Equilibrium).	30–D1.4k	Section 16.2, pp. 639-642
The concentration of a pure solid or a pure liquid is constant, and is not included in the equilibrium equation.	30–D1.4k	Section 16.2, pp. 644-645
The extent of a particular reaction at equilibrium is indicated by the equilibrium constant, K: ■ $K > 1$, products are favoured and the equilibrium lies to the right. If $K > 10^{10}$ the reaction is usually regarded as proceeding to completion. ■ $K \approx 1$, there are approximately equal concentrations of reactants and products at equilibrium. ■ $K < 1$, reactants are favoured and the equilibrium lies to the left. Reactions in which K is smaller than 10^{-10} are usually regarded as not taking place at all.	30–D1.2k	Section 16.2, p. 646
A dynamic equilibrium tends to respond so as to relieve the effect of any change in the conditions that affect the equilibrium (Le Châtelier's principle).	30–D1.3k	Section 16.2, pp. 646-649
A change in temperature is the only change that causes a change in the equilibrium constant.	30–D1.4k	Section 16.2, pp. 648-649
When the initial concentration of a reactant is much greater than the equilibrium constant, the change in the concentration of the reactant at equilibrium is not significant.	30–D2.3k	Section 16.3, p. 659 Section 17.3, p. 698
Equilibrium theories and principles apply to a variety of phenomena in nature.	30–D1.1sts	Section 16.1, p. 636 Section 16.4, pp. 666-668 Chapter 17 opener, p. 678 Section 17.1, pp. 682-683
Equilibrium principles are routinely applied in industrial processes to increase yields, but the rate of reaction is another important factor.	30–D1.3sts 30–D2.4s	Chapter 16 opener, p. 632 Section 16.4, pp. 668-674

Concept	Outcome	Text Reference
Acid deposition is the total effect of acid falling in rain, in snow, or as fine solid particles. The most important cause of acid deposition is the presence of $SO_x(g)$ and $NO_x(g)$.	30–D1.1sts 30–D2.1sts	Section 17.1, pp. 680-682
In the Brønsted-Lowry theory, an acid-base reaction involves the transfer of a proton: an acid is a substance that can donate a proton, H^+. a base is a substance that can accept a proton, H^+.	30–D1.5k	Section 17.2, pp. 684-688
An acid-base reaction always contains two conjugate acid-base pairs, which differ by one proton.	30–D1.7k	Section 17.2, pp. 685-687
Amphiprotic substances can act as a Brønsted-Lowry acid in one reaction and as a Brønsted-Lowry base in a different reaction.	30–D1.7k	Section 17.2, p. 688
A polyprotic acid has more than one proton that can ionize; a polyprotic base is a substance capable of accepting more than one proton.	30–D1.7k	Section 17.2, p. 688
An acid-base reaction usually proceeds to the right if the stronger acid and stronger base are on the left side of the equation, and products are favoured.	30–D1.6k	Section 17.2, p. 689
For the ionization in aqueous solution of an acid HA(aq): $HA(aq) + H_2O(\ell) \rightleftharpoons H_3O^+(aq) + A^-(aq)$ the acid ionization constant is $K_a = \frac{[H_3O^+][A^-]}{[HA]}$	30–D2.2k	Section 17.3, p. 691
The percent ionization of a weak acid is the fraction of molecules that ionize compared with the initial concentration of the acid, expressed as a percentage.	30–D2.2k	Section 17.3, p. 692
The equilibrium constant for the ionization of water $2H_2O(aq) \rightleftharpoons H_3O^+(aq) + OH^-(aq)$ is given the symbol K_w $K_w = [H_3O^+][OH^-]$	30–D2.2k	Section 17.3, pp. 695-696
pH + pOH = 14.0 (at 25 °C)	30–D2.1k 30–D2.2k	Section 17.3, p. 696
A weak base, B, reacts with water to form an equilibrium solution of ions: $B(aq) + H_2O(\ell) \rightleftharpoons HB^+(aq) + OH^-(aq)$	30–D2.2k	Section 17.3, p. 696
For a conjugate acid-base pair, $K_a \times K_b = K_w$.	30–D2.2k	Section 17.3, p. 701
An indicator is a weak organic acid that is one colour in acid form and a different colour in its conjugate base form.	30–D1.6k	Section 17.4, p. 705
The endpoint of a titration occurs when the indicator changes colour and you stop adding acid or base.	30–D1.9k	Section 17.4, p. 705
A titration curve is a graph of the pH of the titration mixture against the volume of added acid (or base).	30–D1.9k	Section 17.4, pp. 706-709
A buffer is a solution containing two solutes that keep the pH of the solution almost constant when moderate amounts of acid or base are added to it.	30–D1.8k	Section 17.4, pp. 710-713
Different buffer systems help to maintain the pH of body fluids such as blood within a very narrow range of pH.	30–D1.1sts	Section 17.4, pp. 714-717

Related Skills

Skills	Outcome	Text Reference
Students will describe procedures for safe handling, storage and disposal of materials used in the laboratory, with reference to WHMIS and consumer product labelling information.	30–D2.1s	Investigation 16.B, Section 16.2, pp. 652-654 Investigation 16.C, Section 16.3, pp. 662-665 Chapter 17 Launch Lab, p. 679 Investigation 17.A, section 17.3, pp. 694-695 Investigation 17.B, Section 17.4, pp. 710-711
Students will investigate a reversible change, ask questions about observed relationships, and develop hypotheses about the system.	30–D1.2sts 30–D1.1s	Chapter 16 Launch Lab, Section 16.1, p. 633
Students will model a reversible change, interpret graphical data to determine when equilibrium is established, and determine the criteria that apply to chemical systems at equilibrium.	30–D1.3s 30–D1.4s	Investigation 16.A, Section 16.1, pp. 635-636.
Students will write the equilibrium expression for a homogeneous chemical system at equilibrium.	30–D1.2k	Section 16.2, p. 641
Given the chemical equation, and data describing the amounts or concentrations of species present at equilibrium, students will calculate the equilibrium constant K_c.	30–D1.4k	Section 16.2, pp. 642-644
Students will use Le Châtelier's principle to predict shifts in equilibrium caused by changes in temperature, pressure, volume, concentration or the addition of a catalyst, and describe how these changes affect the equilibrium constant.	30–D1.3k 30–D1.2s	Section 16.2, pp. 649-651
Students will use Le Châtelier's principle to predict the effect of change on an equilibrium system, and will design an experiment to test predictions of equilibrium shifts.	30–D1.1s 30–D2.2s	Investigation 16.B, Section 16.2, pp. 652-654.
Students will apply stoichiometric principles to chemical equations to solve equilibrium problems.	30–D2.3k	Section 16.3, pp. 656-660
Students will determine the value of an equilibrium constant from experimental data.	30–D2.3s 30–D2.3k 30–D2.1s	Thought Lab 16.1, Section 16.3, p. 661. Investigation 16.C, Section 16.3, pp. 662-665.
Students will investigate an experimental model for the buffering effect of different types of rock and soil on acid rain.	30–D1.2s	Chapter 17 Launch Lab, p. 679
Students will identify the conjugate acid-base pairs in a Brønsted-Lowry reaction.	30–D1.7k	Section 17.2, p. 687
Students will write Brønsted-Lowry equations that illustrate a molecule or an ion acting as an amphiprotic substance.	30–D1.5k 30–D1.7k	Section 17.2, p. 688
Students will write Brønsted –Lowry equations for polyprotic acids and polyprotic bases.	30–D1.7k	Section 17.2, p. 688
Students will write Brønsted –Lowry equations and predict whether reactants or products are favoured for acid-base equilibrium reactions.	30–D1.6k	Section 17.2, pp. 689-690 Section 17.3, pp. 702-703
Students will write acid ionization expressions and use the value of K_a to compare the strengths of weak acids.	30–D2.2k	Section 17.3, pp. 691-692

Skills	Outcome	Text Reference
Given the concentration of a weak acid in water and the pH or pOH of the solution, students will calculate K_a and the percent ionization of a weak acid.	30–D2.1k 30–D2.2k 30–D2.3k	Section 17.3, pp. 692-693
Students will determine the concentration of an aqueous solution of a weak acid, measure the pH of the solution, and use the data to calculate K_a and the percent ionization of the weak acid.	30–D2.3k 30–D2.3s	Investigation 17.A, Section 17.3, pp. 694-695
Given the concentration of a weak base in water and the pH or pOH of the solution, students will calculate K_b.	30–D2.2k 30–D2.3k	Section 17.3, pp. 697-698
Given the concentration of solution and K_a (or K_b) for a weak acid (or base), students will calculate the following: pH, pOH, $[H_3O^+]$, $[OH^-]$.	30–D2.2k 30–D2.3k	Section 17.3, pp. 698-701
Given the acid (or base) ionization constant, students will determine the ionization constant for the conjugate base (or acid).	30–D1.7k 30–D2.2k	Section 17.3, pp. 702-703
Students will use a spreadsheet to graph the results of a weak acid-strong base titration.	30–D1.9k	Thought Lab 17.1, Section 17.4, p. 709
Students will distinguish between the equivalence point and the endpoint for a titration, and qualitatively explain the pH of an acid-base titration at equivalence.	30–D1.9k	Section 17.4, pp. 706-710
Students will design and prepare a buffer to investigate the relative abilities of a buffer and a control to resist a pH change when a small amount of strong acid or strong base is added.	30–D1.1s 30–D1.2s	Investigation 17.B, Section 17.4, pp. 710-712

Activities and Target Skills

Activity	Target Skills
Chapter 16: Chemical Equilibrium	
Launch Lab: The Chemical Blues, p. 633	■ Observing an example of a reversible change that involves a change of colour
Investigation 16.A: Modelling Equilibrium, pp. 635-636	■ Illustrating reversible chemical reactions ■ Interpreting data from a graph to determine when equilibrium is established ■ Determining the criteria that apply to chemical systems at equilibrium
Investigation 16.B: Disturbing Equilibrium, pp. 652-653	■ Predicting the effect of change on a system at equilibrium using Le Châtelier's principle ■ Designing an experiment to illustrate and test predictions of equilibrium shifts ■ Applying equilibrium theories and principles to analyze a variety of phenomena
Thought Lab 16.1: Finding an Equilibrium Law, p. 661	■ Determining equilibrium constants from experimental data
Investigation 16.C: Using Experimental Data to Determine an Equilibrium Constant, pp. 662-663	■ Using experimental data to calculate equilibrium constants

Activity	Target Skills
Chapter 17: Acid-Base Equilibrium Systems	
Launch Lab: Buffering Ground Water, p. 679	■ Modelling different types of ground water ■ Comparing the impact of acid rain on models of ground water
Investigation 17.A: Determining Ka for Ethanoic Acid, pp. 694-695	■ Using experimental data to calculate equilibrium constants
Thought Lab 17.1: Analyzing a Weak Acid-Strong Base Titration, p. 709	■ Using a spreadsheet to graph the results of a weak acid–strong base reaction and analyzing the outcome of the titration
Investigation 17.B: Preparing a Buffer and Investigating Its Properties, pp. 710-711	■ Designing a buffering system ■ Preparing a buffer to investigate the relative abilities of a buffer and a control to resist a pH change when a small amount of strong acid or strong base is added

Conceptual Challenges

Chemical equilibrium is one of the more difficult concepts in general chemistry. It involves ideas that are difficult to conceptualize, and students must discard the notion that reactions proceed in one direction only. The principles of chemical equilibrium are introduced in Chapter 16, and it is essential that the concepts introduced in this chapter be firmly established.

- A chemical system in dynamic equilibrium is difficult to conceptualize because the changes that continue to take place within the system are invisible (i.e., are taking place only at the microscopic level). At a macroscopic level, the system appears to be static, and students will not understand the dynamic processes that are taking place. Investigation 16.A (pp. 635-636) is an excellent activity to model equilibrium systems. The activity models dynamic change and constant macroscopic properties at equilibrium, and students should see a constant ratio (within experimental error) for the volume of liquid in the "product" graduated cylinder to the volume of liquid in the "reactant" graduated cylinder. The activity also models what happens when a change is introduced to the system at equilibrium and provides a concrete basis for understanding Le Châtelier's principle. One additional exercise that will help students understand the dynamic nature of equilibrium is outlined in the extension activity attached to Investigation 16.A Modelling Equilibrium. Gifted students (who may be able to solve equilibrium problems, yet have the same fundamental misconception) can be challenged to model equilibrium in a computer program or in a series of sketches.
- A few students will write the equilibrium expression and calculate the equilibrium constant with reactant terms divided by product terms. Remind students to double-check their expressions and calculations to avoid this mistake.
- Students who take the word "constant" too literally may believe that the value of the equilibrium constant does not depend on temperature. Remind students that the rate of reaction depends on temperature. The activation energy for the forward reaction will not be the same as the activation energy for the reverse reaction. Thus, changing the temperature of a system at equilibrium changes the rates of the forward and reverse reactions differently. Changing the temperature of a system at equilibrium will result in a new equilibrium to be established with a different value for the equilibrium constant.
- Sometimes students use the equilibrium expression for a system that is not, in fact, at equilibrium. For example, students may assume that when more products are added to an equilibrium system at constant temperature, K_c will increase. At the moment when a change is made to the concentration of any of the reacting species at equilibrium, the system is no longer at equilibrium. After some time (which may be short or long, depending on the reaction), a new equilibrium will be established with the same value of K_c as before.

A number of misconceptions are commonly found that relate to Le Châtelier's principle:

- Students may believe that introducing a catalyst to a reaction at constant temperature will change the equilibrium constant. This is usually because these students believe that a catalyst changes only the rate of the forward reaction. In fact, a catalyst changes both the forward and reverse reactions by the same factor, and the value of K_c remains unchanged.
- Students often believe that changing the amount of a solid or a pure liquid will affect an equilibrium mixture. For example, they may claim that increasing the amount of a solid ionic substance that is already at equilibrium with its dissolved ions will cause more dissolved ions to be produced. Inform students that the concentration of a solid is much like its density; both are properties that do not depend on the physical amount of solid present. Thus, the concentration terms for solids (and pure liquids) in a chemical reaction are themselves constants and are not included in the equilibrium expression.
- Predicting the direction of shift in a system at equilibrium when the temperature is changed causes more difficulty

than any other type of change. For example, many students believe that when the temperature is changed, the direction of an equilibrium shift can be predicted without knowing whether the reaction is endothermic or exothermic. Students may argue that increasing the temperature of an equilibrium system will increase the number of collisions, favouring formation of relatively more products than more reactants. Another argument sometimes made is that increasing the temperature of a gaseous equilibrium system at constant volume will increase the pressure of the system (true); the resulting pressure increase causes the equilibrium to shift to the side of the chemical equation with fewer moles of gas (false). Emphasize that increasing the temperature adds energy to the surroundings, so the system will shift toward the endothermic direction of reaction to relieve the increase in temperature.

- Some students believe that all acids and bases are harmful and poisonous, or that all concentrated acids will burn the skin, "eat" metals, etc. Remind students that many foods (e.g., citrus fruits) are acidic, that vinegar is an acid, and that some people enjoy tonic water, a base. Whether an acid or a base is caustic depends on two factors: strength and concentration.
- Remind students not to confuse the strength of acids and bases with their concentration. Reinforce the idea that "concentrated" (having a high concentration of particles in mol/L) is different from "strong" (dissociates 100%), and that "weak" (dissociates much less than 100%) is different from "dilute" (having a low concentration of particles in mol/L). A strong acid is always strong, but in solution it may or may not be concentrated. It is interesting to note that the percentage of acid molecules that normally ionize in a weak acid actually increases as the acid concentration is decreased through dilution with water. Thus, ethanoic acid actually becomes slightly stronger as the solution is diluted.
- Some students will think of a higher pH value as meaning more acidic, rather than more basic. Perhaps the best approach is to have students memorize a few "reference markers" on the pH scale, and use them as a memory aid. For example, stomach acid has a pH of about 3, while a glass cleaner using aqueous ammonia (a base) has a pH of about 10. All students should know that neutral water has pH of 7. Thus, a substance with a pH of 4 must be acidic because it is closer to 3 than to 10.
- A number of misconceptions exist linking the formula of a substance to whether it is acidic or basic. For example, students may believe that chemicals having formulas with hydrogen are acids and those having formulas with the hydroxyl group are bases. Use examples such as table sugar (sucrose), $C_{12}H_{22}O_{11}$, and methanol, CH_3OH, to show that this is not true.
- It is important that students understand what is meant when a Brønsted-Lowry acid is described as a proton donor. It is the single proton in a hydrogen ion, H^+, not a proton from the nucleus of some other atom that is donated. Although chemists commonly refer to acids and bases as proton donors and proton acceptors, it is important to realize that H^+ is being moved from one chemical species to another, not a proton from one nucleus to another. Furthermore, if nuclear protons *were* transferred, then elements would change their identities in acid-base reactions, since the atomic numbers of the atoms involved would be altered.
- Another misconception related to the formula of an acid is the belief that as the number of hydrogen atoms increases in the formula of an acid, its acidity becomes stronger. In fact, increasing the number of oxygen atoms in the formula of an acid increases the strength of the acid. This is because oxygen is very electronegative, and increasing the number of oxygen atoms in the molecule makes the O-H bond more polar. Water molecules are more strongly attracted to the H atom, and the acid ionizes more easily.
- Some students confuse endpoint and equivalence point or treat them as the same thing. The point in a titration where the chosen indicator changes colour is the endpoint. The equivalence point (also known as the stoichiometric point) is the point at which the amount of H^+ ions that were originally present in the flask equals the amount of OH^- ions that have been added from the burette. A good choice of indicator for a titration is one that will provide an endpoint that occurs at exactly the same point as the equivalence point, or very close to it.
- A common misconception is the idea that at the end of all neutralization reactions, there is neither H^+ nor OH^- ions in the resulting solutions. Take the opportunity to have students examine carefully the titration curve for a weak acid-strong base titration and the titration curve for a strong acid-weak base titration. At equivalence, these solutions are not neutral.
- While examining titration curves, emphasize that at equivalence the solution contains an aqueous solution of a salt. Students who might have thought that all salts form aqueous solution that are neutral will see that some salt solutions are neutral, others acidic, and some are basic.

Using the Unit 8 Preparation Feature

The unit opener and the Focussing Questions introduce the concept of equilibrium in the context of industrial processes, asking students to begin to consider what is happening in a system at equilibrium and how scientists predict shifts in the equilibrium of a system.

To be successful in this unit (and prepare for post-secondary studies in related areas), students must be proficient in solving stoichiometry problems. Chapter 17 relies on a good understanding of acids and bases from *Chemistry 20*, Unit A. The Unit 8 Preparation feature includes a brief review of Arrhenius Theory of acids and bases, the basic principles relating to strong and weak acids and bases, acid-base tritration, describing acidic and basic solutions, and significant digits and pH.

Encourage students to take the Unit Prequiz (found at **www.albertachemistry.ca,** Online Learning Centre, Student Edition) to gauge their recall of these concepts and methods, noting that if they are familiar with the background science, their experience with this unit will be much easier.

UNIT 8: COURSE MATERIALS

Chapter, Section	Item Description	Suggested Quantity (assume 40 in class)	Text Activity
Chapters 16, 17	safety goggles	40 pairs	Chapter 16 Launch Lab; Chapter 17 Launch Lab; Investigations: 16.B (all 3 parts), 16.C, 17.A, 17.B
Chapters 16, 17	nonlatex disposable gloves	40 pairs × 6 investigations	Chapter 16 Launch Lab; Chapter 17 Launch Lab; Investigations: 16.B (all 3 parts), 16.C, , 17.A, 17.B
Chapters 16, 17	aprons	40	Chapter 16 Launch Lab; Chapter 17 Launch Lab; Investigations: 16.B (all 3 parts), 16.C, 17.A, 17.B
Chapter 16, Chapter Opener	5% glucose solution (5 g/100 mL)	250 mL	Launch Lab: The Chemical Blues, p. 633
	8 mol/L NaOH(aq) [*Note: poisonous and corrosive*]	7.5 mL	
	dropper bottle of methylene blue solution	1–2 drops	
	500 mL Erlenmeyer flask with a stopper	1	
	medium test tube with a stopper (ensure that the stopper is clean because contaminated stoppers can introduce errors into the experiment)	1	
	ice bath	1	
	warm-water bath	1	
Chapter 16, Section 16.1	water coloured with food dye	30 mL per group	Investigation 16.A: Modelling Equilibrium, pp. 635–636
	25 mL graduated cylinders	2 per group	
	glass tubes of different diameters (e.g., 10 mm and 6 mm; approx. 20 cm long)	2 per group	
	labels or a grease pencil	2 labels per group	
Chapter 16, Section 16.2	Part 1		Investigation 16.B: Disturbing Equilibrium, Part 1: Changes to a Base Equilibrium System, p. 652
	0.01 mol/L $NH_3(aq)$ [*Note: corrosive and poisonous*]	approx. 10 mL per group	
	phenolphthalein solution [*Note: flammable*]	1 bottle , approx. 10 mL per group	
	$NH_4Cl(s)$ [*Note: corrosive and poisonous*]	1 bottle	
	6.0 mol/L HCl(aq) [*Note: corrosive and toxic*]	1 per group	
	25 mL beaker	1 per group	
	white paper	2 per group	
	test tubes	1 per group	
	test-tube rack	1 per group	
	scoopula		
	Part 2		Investigation 16.B: Disturbing Equilibrium, Part 2: Concentration and Temperature Changes, pp. 652–653
	$CoCl_2$ dissolved in a solution of water and ethanol [*Note: flammable, corrosive and poisonous*]	1–1.2 L	
	concentrated HCl(aq) in a dropper bottle [*Note: corrosive and poisonous*]	1 bottle	
	$AgNO_3(aq)$ 0.1 mol/L in a dropper bottle [*Note: flammable, corrosive and poisonous*]	1 bottle	
	distilled water in a dropper bottle	1 bottle	
	hot water bath	1–2	
	cold water bath	1–2	
	25 mL or 50 mL beaker	1 per group	
	small test tubes	4 per group	
	test-tube rack	1 per group	
	test-tube holder	1 per group	

Chapter, Section	Item Description	Suggested Quantity (assume 40 in class)	Text Activity
Chapter 16, Section 16.2	Part 3 small piece of copper concentrated nitric acid [*Note: corrosive, toxic and oxidizing*] boiling water ice water test tube test-tube rack one-hole stopper glass delivery tube short length of rubber tubing (approx. 10 cm) syringe with a cap or rubber stopper to seal the tip $NO_2(g)/N_2O_4(g)$ tubes [*Note: poisonous and corrosive*]	 1 1 mL 1 boiling water bath 1 ice water bath 1 1 1 1 10 cm 1 3	Investigation 16.B: Disturbing Equilibrium, Part 3: Investigating Gaseous Equilibriums, pp. 653–654
Chapter 16, Section 16.3	0.0020 mol/L KSCN (aq) 0.0020 mol/L $Fe(NO_3)_3(aq)$ (acidified) [*Note: corrosive and poisonous*] 0.200 mol/L $Fe(NO_3)_3(aq)$ (acidified) [*Note: corrosive and poisonous*] distilled water test tubes (18 mm × 150 mm) flat-bottom vials beakers (100 mL) test-tube rack labels or grease pencil pipette (20.0 mL) pipettes (5.0 mL) pipette bulb stirring rod paper towel thermometer (alcohol or digital) strip of paper diffuse light source, such as a light box medicine dropper	30 mL per group 30 mL per group 25 mL per group ~1.2 L 5 per group 5 per group 3 per group 1 per group 10 1 per group 3 per group 1 per group 1 per group 5 rolls 1 10 1 1 per group	Investigation 16.C: Using Experimental Data to Determine an Equilibrium Constant, pp. 662–665
Chapter 17, Chapter Opener	dilute vinegar solution to simulate acid rain tap water granite or quartz chips marble chips sample of local soil or rock 250 mL beakers universal indicator paper (pH paper) pH meter (optional) 100 mL graduated cylinder 50 mL burette stirring rod retort stand burette clamp a label or a grease pencil	120 mL per group 100 mL per group 1 bag/box 1 bag/box 1 bag/box 5 per group 200 pieces (if available to be shared by class) 1 per group 1 per group 1 per group 1 per group 1 per group 1 per group 1 per group	Launch Lab: Buffering Groundwater: A Delicate Balance, p. 679 [Note: Acid rain can be simulated by diluting 20 mL of vinegar to 1 L with tap water. The solution pH should be about 3.3. If marble chips are not available, dissolve a Tums™ tablet in 200 mL of tap water. If granite or quartz chips are not available, use distilled water.]

Chapter, Section	Item Description	Suggested Quantity (assume 40 in class)	Text Activity
Chapter 17, Section 17.3	unknown concentration ethanoic acid solution, $CH_3COOH(aq)$ [*Note: flammable and corrosive*] 0.75 mol/L (or appropriate for the concentration of ethanoic acid that is used) sodium hydroxide solution, NaOH(aq) [*Note: corrosive and poisonous*] dropper bottle containing phenolphthalein [*Note: flammable*] distilled water 10 mL pipette labels 100 mL beakers pH meter or pH paper 250 mL beaker for waste solutions burette and burette clamp retort stand meniscus reader funnel pipette bulb or pipette pump 150 mL Erlenmeyer flask sheet of white paper	40 mL per group 70 mL per group 1 2 L 1 per group 2 per group 2 per group 1 piece of paper per group (if meter is available to be shared by class) 1 per group 1 per group 1 per group 1 per group 1 per group 1 per group 1 per group 1 per group	Investigation 17.A: Determining Ka for Ethanoic Acid, pp. 694–695
Chapter 17, Section 17.4	distilled water 0.20 mol/L ethanoic acid, $CH_3COOH(aq)$ [*Note: flammable and corrosive*] 0.20 mol/L sodium hydroxide, NaOH (aq) [*Note: corrosive and poisonous*] 0.20 mol/L HCL(aq) [*Note: corrosive and poisonous*] 50 mL graduated cylinder 50 mL beakers 100 mL beaker universal indicator paper (pH paper) (if no pH meter is available pH meter (optional) clean straw stirring rod burette	840 mL 1.8 L 840 mL 1 L 1 per group 4 per group 1 per group minimum of 400 pieces (if meter is available to be shared by class) 2 per group 1 per group 1 per group	Investigation 17.B: Preparing a Buffer and Investigating Its Properties, pp. 710–712

CHAPTER 16 CHEMICAL EQUILIBRIUM

Curriculum Correlation

(Note: This correlation includes Chapters 16 and 17. Chapter 16 references are in bold.)

General Outcome 1: Students will explain that there is a balance of opposing reactions in chemical equilibrium systems.

	Student Textbook	Assessment Options
Outcomes for Knowledge		
30–D1.1k define equilibrium and state the criteria that apply to a chemical system in equilibrium, i.e., closed system, constancy of properties, equal rates of forward and reverse reactions	**Chemical Equilibrium, Section 16.1, p. 634** **Conditions That Apply to All Equilibrium Systems, Section 16.1, pp. 636–637**	**Questions for Comprehension: 1–3, Section 16.1, p. 636** **Section 16.1 Review: 1–11, p. 638** **Chapter 16 Review: 1–3, 6, 7, 22, pp. 676–677** **Chapter 16 Test** Unit 8 Review: 1–3, 5–11, 20, pp. 724–727
30–D1.2k identify, write and interpret chemical equations for systems at equilibrium	**The Law of Chemical Equilibrium, Section 16.2, pp. 639–641** **Equilibrium Law Expression, Section 16.2, p. 641** **Sample Problem: Writing Equilibrium Expression for Homogeneous Chemical Reactions, Section 16.2, p. 641** **Sample Problem: Calculating an Equilibrium Constant, Section 16.2, pp. 642–643** **Sample Problem: Calculating an Equilibrium Concentration, Section 16.2, p. 643**	**Practice Problems: 1–5, Section 16.2, p. 641** **Practice Problems: 6–10, Section 16.2, p. 644** **Section 16.2 Review: 1–8, p. 655** **Chapter 16 Review: 4, 5, 8–12, 14–19, 24, pp. 676–677** **Chapter 16 Test** Unit 8 Review: 2–4, 7–10, 12, 19, 26–31, 33, 37, 38, 41, 50, 52, 53, pp. 724–727
30–D1.3k predict, qualitatively, using Le Châtelier's principle, shifts in equilibrium caused by changes in temperature, pressure, volume, concentration or the addition of a catalyst, and describe how these changes affect the equilibrium constant	**Le Châtelier's Principle, Section 16.2, pp. 646–647** **Sample Problem: Temperature and the Extent of a Reaction, Section 16.2, p. 640** **Sample Problem: Using Le Châtelier's Principle, Section 16.2, p. 651**	**Practice Problems: 11–15, Section 16.2, pp. 649–650** **Practice Problems: 16–20, Section 16.2, pp. 651** **Section 16.2 Review: 3, 5–8, p. 655** **Chapter 16 Review: 10–14, 18, 19, 24, pp. 676–677** **Chapter 16 Test** Unit 8 Review: 9, 26–30, 34, pp. 724–727
30–D1.4k define K_c and write equilibrium law expressions for given chemical equations, using lowest whole–number coefficients	**The Law of Chemical Equilibrium, Section 16.2, pp. 639–641** **Equilibrium Law Expression, Section 16.2, p. 641** **Sample Problem: Writing Equilibrium Expression for Homogeneous Chemical Reactions, Section 16.2, p. 641** **Sample Problem: Using Stoichiometry to Calculate Kc, Section 16.3, pp. 657–658** **Sample Problem: Using the Approximation Method, Section 16.3, pp. 659–660**	**Practice Problems: 1–5, Section 16.2, p. 641** **Practice Problems: 6–10, Section 16.2, p. 644** **Section 16.2 Review: 1–8, p. 655** **Practice Problems: 21–24, Section 16.3, p. 658** **Practice Problems: 25–29, Section 16.3, p. 660** **Section 16.3 Review: 1–5, p. 665** **Chapter 16 Review: 3, 5–7, pp. 676–677** **Chapter 16 Test** Unit 8 Review: 4, 7, 10, 11, 26–30, pp. 724–727

	Student Textbook	Assessment Options
30–D1.5k describe Brönsted–Lowry acids as proton donors and bases as proton acceptors	Understanding Acids and Bases, Section 17.2, p. 684 Sample Problem: Conjugate Acid–Base Pairs, Section 17.2, pp. 686–687	Questions for Comprehension: 3–5, Section 17.2, p. 686 Practice Problems: 1–4, Section 17.2, p. 687 Section 17.3 Review: 1–6, p. 704 Chapter 17 Test Unit 8 Review: 16–19, 21–23, 25, 31–39, 44–47, 50, 52, 53, pp. 724–727
30–D1.6k write Brönsted–Lowry equations and predict whether reactants or products are favoured for acid–base equilibrium reactions (including indicators, polyprotic acids, and polyprotic bases)	Predicting the Direction of Reaction for an Acid–Base Reaction, Section 17.2, p. 689 Sample Problem: Predicting the Direction of an Acid–Base Reaction, Section 17.2, p. 689	Practice Problems: 8–10, Section 17.2, p. 690 Section 17.2 Review: 1–4, p. 690 Practice Problems: 11–13, Section 17.3, p. 692 Chapter 17 Review: 7, p. 720–721 Chapter 17 Test Unit 8 Review: 16–19, 21–23, 25, 31–39, 44–47, 50, 52, 53, pp. 724–727
30–D1.7k identify polyprotic acids, polyprotic bases, conjugate pairs and amphiprotic substances	Both an Acid and a Base: Amphiprotic Substances, Section 17.2, p. 688 Sample Problem: An Amphiprotic Ion, Section 17.2, p. 688	Practice Problems: 5–7, Section 17.2, p. 688 Section 17.2 Review: 3, p. 690 Chapter 17 Test Unit 8 Review: 16–19, 23, 25, 31, 34, pp. 726–729
30–D1.8k define a buffer as relatively large amounts of a weak acid and its conjugate base in equilibrium that maintain a relatively constant pH when small amounts of acid or base are added	Buffer Solutions, Section 17.4, pp. 712–713	Questions for Comprehension: 1, 2, Section 17.1, p. 682 Section 17.4 Review: 6–8, p. 718 Chapter 17 Review: 11, 26, pp. 720–721 Chapter 17 Test Unit 8 Review: 24, 25, 43, 47, pp. 724–727
30–D1.9k sketch and qualitatively interpret titration curves of monoprotic and polyprotic acids and bases identifying equivalence points and regions of buffering for weak acid–strong base, strong acid–weak base, and strong acid–strong base.	**Thought Lab 16.1: Finding and Equilibrium Law, Section 16.3, p. 661** Titration Curves and Buffers, Section 17.4, pp. 705–708	**Thought Lab 16.1: Finding and Equilibrium Law: 1–7, Section 16.3, p. 661** Questions for Comprehension: 9–12, Section 17.4, p. 710 Section 17.4 Review: 1–8, p. 718 Chapter 17 Review: 25, pp. 720–721 Chapter 17 Test Unit 8 Review: 39, 44–47, pp. 724–727
Outcomes for Science, Technology and Society (Emphasis on nature of science)		
30–D1.1sts demonstrate an understanding that the goal of science is knowledge about the natural world by ■ *applying equilibrium theories and principles to analyze a variety of phenomena, e.g.,* – *carbon dioxide escaping from an open bottle/can of carbonated beverage* – *role of the oceans in the carbon cycle* – *solubility of oxygen gas in lake water* – *acid precipitation (deposition)* – *blood gases in deep–sea diving* – *buffers in living systems*	**Chapter 16 Launch Lab: The Chemical Blues, p. 633** **Connections: The Development of Equilibrium Theories, Section 16.2, p. 645**	**Chapter 16 Launch Lab: The Chemical Blues: 1, 2, p. 633** **Connections: The Development of Equilibrium Theories, Section 16.2, p. 645** Chapter 17 Review: 2, 21, 22, 27–29, pp. 720–721 Unit 8 Review: 13–15, 39–53, pp. 724–727

	Student Textbook	Assessment Options
30–D1.2sts demonstrate an understanding that scientific knowledge and theories develop through hypotheses, the collection of evidence through experimentation and the ability to provide explanations, e.g., *research how equilibrium theories and principles developed*	**Chapter 16 Launch Lab: The Chemical Blues, p. 633** **Connections: The Development of Equilibrium Theories, Section 16.2, p. 645**	**Chapter 16 Launch Lab: The Chemical Blues: 1, 2, p. 633** **Connections: The Development of Equilibrium Theories: 1–3, Section 16.2, p. 645** Chapter 17 Review: 2, 21, 22, pp. 720–721 Unit 8 Review: 13–15, 39–53, pp. 724–727
30–D1.3sts demonstrate an understanding that the goal of technology is to provide solutions to practical problems by ■ *analyzing how equilibrium principles have been applied in industrial processes, e.g.,* – *the Haber–Bosch process for making ammonia* – *the Solvay process for making sodium carbonate* – production of methanol.	**Chapter 16 Launch Lab: The Chemical Blues, p. 633** **Connections: The Development of Equilibrium Theories, Section 16.2, p. 645** **Investigation 16.B: Disturbing Equilibrium, Section 16.2, pp. 652–654**	**Chapter 16 Launch Lab: The Chemical Blues: 1, 2, p. 633** **Connections: The Development of Equilibrium Theories: 1–3, Section 16.2, p. 645** **Investigation 16.B: Disturbing Equilibrium: 1–8, Section 16.2, pp. 652–654** **Section 16.4 Review: 1–4, p. 674** Section 17.1 Review: 1–6, p. 683 Unit 8 Review: 13–15, 39–53, pp. 724–727
Skill Outcomes (Focus on scientific inquiry)		
Initiating and Planning		
30–D1.1s ask questions about observed relationships and plan investigations of questions, ideas, problems and issues by ■ predicting variables that can cause a shift in equilibrium ■ designing an experiment to show equilibrium shifts, e.g., *colour change, temperature change, precipitation* ■ describing procedures for safe handling, storage and disposal of materials used in the laboratory, with reference to WHMIS and consumer product labelling information ■ *designing a buffering system.*	**Chapter 16 Launch Lab: The Chemical Blues, p. 633** **Investigation 16.A: Modelling Equilibrium, Section 16.1, pp. 635–636** **Investigation 16.B: Disturbing Equilibrium, Section 16.2, pp. 652–654** **Investigation 16.C: Using Experimental Data to Determine an Equilibrium Constant, Section 16.3, pp. 662–665**	**Chapter 16 Launch Lab: The Chemical Blues: 1, 2, p. 633** **Investigation 16.A: Modelling Equilibrium: 1–10, Section 16.1, pp. 635–636** **Investigation 16.B: Disturbing Equilibrium: 1–8, Section 16.2, pp. 652–654** **Investigation 16.C: Using Experimental Data to Determine an Equilibrium Constant: 1–4, Section 16.3, pp. 662–665** Unit 8 Review: 39, 40, 44, 47, pp. 724–727
Performing and Recording		
30–D1.2s conduct investigations into relationships among observable variables and use a broad range of tools and techniques to gather and record data and information by ■ performing an experiment to test, qualitatively, predictions of equilibrium shifts, e.g., *colour change, temperature change, precipitation, and gas production* ■ *preparing a buffer to investigate the relative abilities of a buffer and a control; i.e., water, to resist a pH change when a small amount of strong acid or strong base is added.*	**Investigation 16.A: Modelling Equilibrium, Section 16.1, pp. 635–636** **Investigation 16.B: Disturbing Equilibrium, Section 16.2, pp. 652–654** **Investigation 16.C: Using Experimental Data to Determine an Equilibrium Constant, Section 16.3, pp. 662–665**	**Investigation 16.A: Modelling Equilibrium: 1–10, Section 16.1, pp. 635–636** **Investigation 16.B: Disturbing Equilibrium: 1–8, Section 16.2, pp. 652–654** **Investigation 16.C: Using Experimental Data to Determine an Equilibrium Constant: 1–4, Section 16.3, pp. 662–665** Unit 8 Review: 39, 40, 44, 47, pp. 724–727
Analyzing and Interpreting		
30–D1.3s analyze data and apply mathematical and conceptual models to develop and assess possible solutions by ■ writing the equilibrium law expression for a given equation ■ analyzing, qualitatively, the changes in concentrations of reactants and products after an equilibrium shift ■ interpreting data from a graph to determine when equilibrium is established, and determining the cause of a stress on the system.	**Investigation 16.B: Disturbing Equilibrium, Section 16.2, pp. 652–654** **Thought Lab 16.1: Finding an Equilibrium Law, Section 16.3, p. 661** **Investigation 16.C: Using Experimental Data to Determine an Equilibrium Constant, Section 16.3, pp. 662–665**	**Investigation 16.B: Disturbing Equilibrium: 1–8, Section 16.2, pp. 652–654** **Thought Lab 16.1: Finding an Equilibrium Law, Section 16.3, p. 661** **Investigation 16.C: Using Experimental Data to Determine an Equilibrium Constant: 1–4, Section 16.3, pp. 662–665** Unit 8 Review: 10, 26–30, 32–34, 39, 47, 49–53, pp. 724–727

	Student Textbook	Assessment Options
Communication and Teamwork		
30–D1.4s work collaboratively in addressing problems and apply the skills and conventions of science in communicating information and ideas and in assessing results by ■ *working collaboratively with team members to develop an illustration and explanation of reversible reactions* ■ *using advanced menu features within a word processor to develop a group report on equilibrium systems.*	**Investigation 16.B: Disturbing Equilibrium, Section 16.2, pp. 652–654**	**Investigation 16.B: Disturbing Equilibrium: 1–8, Section 16.2, pp. 652–654**

General Outcome 2: Students will determine quantitative relationships in simple equilibrium systems.

	Student Textbook	Assessment Options
Outcomes for Knowledge		
30–D2.1k recall the concepts pH and hydronium ion concentration, pOH and hydroxide ion concentration in relation to acids and bases	Describing Acidic and Basic Solutions, Unit 8 Preparation, pp. 630–631	**Chapter 16 Review: 23, pp. 676–677** Chapter 17 Review: 1, 3–6, 8, 13, 15, 19, pp. 720–721 Chapter 17 Test Unit 8 Review: 31, 32, 34, 36, 39, 43, 47, 49, 50, pp. 724–727
30–D2.2k define K_w , K_a , K_b and use these to determine pH, pOH, $[H_3O^+]$, $[OH^-]$ of acidic and basic solutions	Acid–Base Equilibriums, Section 17.3, p. 691	
	Sample Problem: Acid Ionization Expressions, Section 17.3, p. 691	Practice Problems: 11–13, Section 17.3, p. 692
	Sample Problem: Determining K_a and Percent Ionization, Section 17.3, pp. 692–693	Practice Problems: 14–18, Section 17.3, p. 693
	The Ion Product Constant for Water, Section 17.3, pp. 695–696	Questions for Comprehension: 6–8, Section 17.3, p. 696
	The Base Ionization Constant K_b, Section 17.3, pp. 696–697	
	Sample Problem: Calculating K_b for a Weak Base, Section 17.3, p. 697	Practice Problems: 19–23, Section 17.3, p. 698
	Calculating the pH of a Solution of a Weak Acid or a Weak Base, Section 17.3, p. 698	
	Sample Problem: Calculating the pH of a Weak Acid, Section 17.3, p. 699	Practice Problems: 24–27, Section 17.3, p. 699 Practice Problems: 28–33, Section 17.3, p. 701
	Sample Problem: Calculating the pH of a Weak Base, Section 17.3, p. 700	Section 17.3 Review: 1, 3–6, p. 704 Chapter 17 Review: 1, 3–6, 8, 12–20, 23, 24, 28, 29, pp. 720–721 Chapter 17 Test Unit 8 Review: 20, 21, 31, 32–36, 39, 43, 47, 49, 50, pp. 724–727

	Student Textbook	Assessment Options
30–D2.3k calculate equilibrium constants and concentrations for homogeneous systems and Brönsted–Lowry acids and bases (excluding buffers) when ■ concentrations at equilibrium are known ■ initial concentrations and one equilibrium concentration are known ■ the equilibrium constant and one equilibrium concentration are known **Note:** Examples that require the application of the quadratic equation are excluded; however, students may use this method in responding to open–ended questions.	The Relationship Between Conjugate Acid–Base Pairs, Section 17.3, p. 701 Sample Problem: Equilibrium Constants for Conjugate Acid–Base Pairs, Section 17.3, p. 702 Sample Problem: Predicting the Direction of an Acid–Base Reaction, Section 17.3, pp. 702–703	Practice Problems: 34–38, Section 17.3, p. 703 Section 17.3 Review: 1, 4, p. 704 Chapter 17 Review: 23, 28, 29, pp. 720–721 Chapter 17 Test Unit 8 Review: 20, 21, 27, 29, 31–39, 47, 49, 51, pp. 724–727
Outcomes for Science, Technology and Society (Emphasis on nature of science)		
30–D2.1sts develop an understanding that technological development may involve the creation of prototypes and testing, as well as application of knowledge from related scientific and interdisciplinary fields by ■ *analyzing, on the basis of chemical principles, the application of equilibrium in, e.g.,* *– industrial processes or medical sciences* *– antacid tablets, buffering in living systems* *– acid precipitation.*	Chapter 17 Launch Lab: Buffering Ground Water: A Delicate Balance, p. 679 Connections: Buffers in the Blood, Section 17.4, p. 717	Chapter 17 Launch Lab: Buffering Ground Water: A Delicate Balance: 1–3, p. 679 Connections: Buffers in the Blood: 1–3, Section 17.4, p. 717
Skill Outcomes (Focus on scientific inquiry)		
Initiating and Planning		
30–D2.1s ask questions about observed relationships and plan investigations of questions, ideas, problems and issues by ■ designing an experiment to show qualitative equilibrium shifts in concentration under a given set of conditions ■ describing procedures for safe handling, storage and disposal of materials used in the laboratory, with reference to WHMIS and consumer product labelling information.	Investigation 17.B: Preparing a Buffer and Investigating Its Properties, Section 17.4, pp. 710–712	Investigation 17.B: Preparing a Buffer and Investigating Its Properties: 1–8, Section 17.4, pp. 710–712
Performing and Recording		
30–D2.2s conduct investigations into relationships among observable variables and use a broad range of tools and techniques to gather and record data and information by ■ performing an experiment to show equilibrium shifts in concentration	Investigation 17.A: Determining K_a for Ethanoic Acid, Section 17.2, pp. 694–695 Investigation 17.B: Preparing a Buffer and Investigating Its Properties, Section 17.4, pp. 710–712	Investigation 17.A: Determining K_a for Ethanoic Acid: 1–8, Section 17.2, pp. 694–695 Investigation 17.B: Preparing a Buffer and Investigating Its Properties: 1–8, Section 17.4, pp. 710–712

	Student Textbook	Assessment Options
Analyzing and Interpreting		
30–D2.3s analyze data and apply mathematical and conceptual models to develop and assess possible solutions by ■ using experimental data to calculate equilibrium constants.	Investigation 17.A: Determining K_a for Ethanoic Acid, Section 17.2, pp. 694–695 Thought Lab 17.1: Analyzing a Weak Acid–Strong Base Titration, Section 17.4, p. 709 Investigation 17.B: Preparing a Buffer and Investigating Its Properties, Section 17.4, pp. 710–712	Investigation 17.A: Determining K_a for Ethanoic Acid: 1–8, Section 17.2, pp. 694–695 Thought Lab 17.1: Analyzing a Weak Acid–Strong Base Titration: 1–4, Section 17.4, p. 709 Investigation 17.B: Preparing a Buffer and Investigating Its Properties: 1–8, Section 17.4, pp. 710–712
Communication and Teamwork		
30–D2.4s work collaboratively in addressing problems and apply the skills and conventions of science in communicating information and ideas and in assessing results by ■ *using advanced menu features within a word processor to develop a group report on equilibrium applications in Alberta industries.*	Thought Lab 17.1: Analyzing a Weak Acid–Strong Base Titration, Section 17.4, p. 709 Investigation 17.B: Preparing a Buffer and Investigating Its Properties, Section 17.4, pp. 710–712	Thought Lab 17.1: Analyzing a Weak Acid–Strong Base Titration: 1–4, Section 17.4, p. 709 Investigation 17.B: Preparing a Buffer and Investigating Its Properties: 6, 7, Section 17.4, pp. 710–712

Chapter 16

Chemical Equilibrium

Student Textbook pages 632–676

Chapter Concepts

Section 16.1 Chemical Equilibrium

- When a system is at equilibrium, no further change takes place in the concentration of reactants or products.

Section 16.2 Equilibrium Expressions and Le Châtelier's Principle

- The equilibrium expression for a chemical system can be written from the chemical equation.
- The equilibrium constant can be calculated from the concentration of reactants and products at equilibrium.
- A reaction at equilibrium can be disturbed by changing factors, such as concentration, pressure, volume, or temperature.

Section 16.3 Equilibrium Calculations for Homogeneous Systems

- The equilibrium constant for a reaction can be calculated when equilibrium concentrations are known or when initial concentrations and one equilibrium concentration are known.
- The concentration of each substance in a reaction at equilibrium can be calculated if the equilibrium constant and the concentration of one reactant are known.

Section 16.4 Applications of Equilibrium Systems

- Principles of equilibrium apply to many biological systems.
- Modern chemical processes are the result of extensive experiments and may involve building models that rely on equilibrium concepts.

Common Misconceptions

- A common misconception held by students is that there is a simple arithmetical relationship between the concentrations of reactants and products. For example, students may think the concentration of reactants equals the concentration of products, or the concentrations of products and reactants are related to the stoichiometric coefficients in the chemical equation. The coefficients fix the relative amount that will react, not the amounts that are present. Investigation 16.A, Modelling Equilibrium, should counter this misconception. There are a number of other modelling exercises (e.g., using coins or dice) in the literature that students can perform in small groups, or which can be demonstrated using an overhead projector.
- Students may believe that the forward reaction rate is always equal to the reverse reaction rate. Remind students that the rate of reaction depends on the concentration of reactants. In the student textbook, the rate expressions for a simple reversible reaction are shown in the Chemistry File on page 637.
- In *Chemistry 20*, students performed laboratory activities and stoichiometric calculations assuming that reactions went to completion. Some students may believe that similar concepts are true in equilibrium mixtures, and that the forward reaction is completed before the reverse reaction has begun. Address this misconception by referring again to the modelling exercise. Remind students that the double reaction arrows illustrate the fact that both reactants and products remain in a system at equilibrium.
- Students may believe the equilibrium constant, K_c, changes if more reactant is added, or some product is removed. Remind students that in step 7 of Investigation 16.A, Modelling Equilibrium, they modelled the effect of adding more reactant to an equilibrium mixture. In analysis step 7, the ratio of volume of product to volume of reactant is a model for the equilibrium constant, and students should find the ratio the same, within experimental error.
- Another misconception about K_c is that it does not depend on temperature. Remind students that the rate of a reaction depends on the number of molecules with energy greater than the activation energy for the reaction. The activation energy for a given reaction in the forward direction is not the same as the activation energy in the reverse direction. Consequently, changing the temperature has a different effect on the rate of the forward reaction compared with the rate of the reverse reaction.
- Students may use Le Châtelier's principle incorrectly in heterogeneous systems. This is likely to occur when the principle is applied by rote, without considering that the concentration of a pure liquid or a pure solid is constant. Tell students that the concentration of a pure solid or a pure liquid is a property that is similar to density. Most students will know that the density of a pure solid or a pure liquid does not depend on the amount present. Thus, the concentration terms for pure solids and pure liquids are constants and are not included in the equilibrium expression.
- Many students will incorrectly apply Le Châtelier's principle when an inert gas is added to a gaseous system at equilibrium in a container with constant volume. Remind students that the inert gas is not included in the equilibrium expression, and therefore has no effect on the value of K_c. If the argument is made that the pressure of the system changes due to the addition of the inert gas (which, of course, is true), ask if the concentrations of the reacting gases has changed. Some students may argue that the volume of the container must be effectively reduced, because the inert gas molecules occupy some volume. Remind these students that one of the assumptions of the kinetic theory of gases is that the volume occupied by gas molecules is negligible compared with the volume of the container.

Helpful Resources

Books and Journal Articles

- Bartholow, M. "Modelling Dynamic Equilibrium with Coins." *Journal of Chemical Education*. Vol. 83, No. 1, January 2006, pages 48A–48B.
- Bucat, R.B., Tyson, L., and Treagust, D.F. "The Complexity of Teaching and Learning Chemical Equilibrium." *Journal of Chemical Education*. Vol. 76, No. 4, April 1999, pages 554–558.
- Voska, K.W., Heikkinen, H.W. "Identification and Analysis of Student Conceptions Used To Solve Chemical Equilibrium Problems." *Journal of Research in Science Teaching*. Vol. 37, No. 2, February 2000, pages 160–176.
- Walter R., Vandaveer, I.V., and Mosher, M. "The Blue Bottle Revisited." *Journal of Chemical Education*. Vol. 74, No. 4, April 1997, page 402.
- Volker, B. and Thomsen, E. "Le Châtelier's Principle in the Sciences." *Journal of Chemical Education*. Vol. 77, No. 2, February 2000, pages 173–176.
- Glickstein, N. "Putting a Human Face on Equilibrium." *Journal of Chemical Education*. Vol. 82, No. 3, March 2005, pages 391–392.
- Lainchbury, A., Stevens, J. and Thompson, A. *ILPAC Physical: Equilibrium I – Principles*. Vol. 6, John Murray (Publishers).

Web Sites

Web links related to chemical equilibrium can be found at **www.albertachemistry.ca**. Go to the Online Learning Centre, and log on to the Instructor Edition. Choose Teacher Web Links.

List of BLMs

Blackline masters (BLMs) have been prepared to support the material in this chapter. The BLMs are either for assessment (AST); use as overheads (OH); use as handouts (HAND), in particular to support activities; or to supply answers (ANS) for assessment or handouts. The BLMs are in digital form, stored on the CD-ROM that accompanies this Teacher's Resource or on the web site at **www.albertachemistry.ca**, Online Learning Centre, Instructor Edition, BLMs.

Number (Type) Title
16.0.1 (HAND) Launch Lab: The Chemical Blues

16.0.1A (ANS) Launch Lab: The Chemical Blues Answer Key
16.1.1 (HAND) Investigation 16.A: Modelling Equilibrium
16.1.1A (ANS) Investigation 16.A: Modelling Equilibrium Answer Key

16.2.1 (OH) Disturbing Equilibrium
16.2.2 (HAND) Investigation 16.B: Disturbing Equilibrium
16.2.2A (ANS) Investigation 16.B: Disturbing Equilibrium Answer Key

16.3.1 (HAND) Thought Lab 16.1: Finding an Equilibrium Law
16.3.1A (ANS) Thought Lab 16.1: Finding an Equilibrium Law Answer Key
16.3.2 (OH) Using Stoichiometry to Calculate K_c
16.3.3 (HAND) Investigation 16.C: Using Experimental Data to Determine an Equilibrium Constant
16.3.3A (ANS) Investigation 16.C: Using Experimental Data to Determine an Equilibrium Constant Answer Key

16.4.1 (OH) Haber-Bosch Process

16.5.1 (AST) Chapter 16 Test
16.5.1A (ANS) Chapter 16 Test Answer Key

Using the Chapter 16 Opener

Student Textbook pages 632–633

Teaching Strategies

- Read the Chapter Opener with your students, reviewing the concept of chemical change.
- Ask students to brainstorm different changes that occur in science and nature, and to decide whether these changes are reversible or irreversible.

Launch Lab The Chemical Blues

Student Textbook page 633

Purpose

The purpose of this lab is to introduce students to a reversible chemical reaction. They will also learn that systems at equilibrium have constant properties. Encourage students to identify that the reactions take place in a closed container. The Erlenmeyer flask contains air above the solution, while the test tube does not.

Outcomes

- 30-D1.4s

Advance Preparation

When to Begin	What to Do
3 to 4 weeks before	■ Order the required supplies.
2 to 3 days before	■ Assemble the required glassware, stoppers, and timers.
1 day before	■ Prepare the solution for the reaction. Store the solution in a sealed container. ■ Photocopy **BLM 16.0.1 (HAND) Launch Lab: The Chemical Blues.**

Materials

For student experiments:
- 500 mL <5% glucose solution (5 g/100 mL)
- 500 mL < 8.0 NaOH(aq)
- dropper bottle of methylene blue indicator
- 500 mL Erlenmeyer flasks with stoppers
- medium test tubes with stoppers
- ice bath
- warm water bath

Time Required

- 40 minutes

Helpful Tips

- The solution concentrations are not very critical to the success of the experiment or demonstration. However, if the class is doing the experiments, the NaOH(aq) should be made close to the day of the lab and stored in a sealed container to prevent the solution from absorbing $CO_2(g)$ from the air. If the reaction is to be demonstrated, the solution can easily be prepared on the day it will be used.
- Make sure the stoppers are clean.
- Brief details of this reaction can be found at http://www.chem.utas.edu.au/staff/yatesb/bluebottle.html
- Use **BLM 16.0.1 (HAND) Launch Lab: The Chemical Blues** to support this activity. Modify it as necessary.
- ***Expected Results:*** If the solution has been sitting for a while, it will be clear. When the solution is shaken in the presence of some air, it will turn blue. The colour will gradually fade upon sitting. If the solution is shaken in the absence of air (stoppered with no air between the solution and the stopper) the blue colour will not appear. Oxygen is necessary for the solution to turn blue.

Safety Precautions

- Ensure that students have read "Safety in Your Chemistry Laboratory and Classroom" pages xii-xv in the student textbook. Ensure that students follow all safety precautions outlined for this investigation in the student textbook.
- Sodium hydroxide is harmful if swallowed or if the fumes are inhaled. It causes irritation to the skin and eyes. Spills of 0.5 mol/L NaOH(aq) on skin or clothing can be rinsed with plenty of cold water. Spills on the lab bench or floor can be mopped up, and the mop rinsed with plenty of cold water.
- Methylene blue will stain clothing and may also cause eye irritation.
- Remind students to wash their hands when they have completed the lab.

Answers to Analysis Questions

1. Initially, the solution in the Erlenmeyer flask is blue due to the addition of methylene blue, which is a redox indicator. Swirling the flask contents should quickly cause the glucose to reduce methylene blue to its colourless form. When the flask is shaken, the solution turns blue, but turns colourless again on standing. The colour changes of the solution indicate a reversible reaction.
2. The solution in the test tube should remain colourless because there is no oxygen present to oxidize the methylene blue; oxidized methylene blue is blue in colour. Students should realize that the air above the solution in the Erlenmeyer flask is the key difference. Since $O_2(g)$ is much more reactive than $N_2(g)$, some students should hypothesize that the methylene blue in the solution in the Erlenmeyer flask is reacting with oxygen. They could test this hypothesis in a number of ways, for example, by pouring the solution from the test tube into a larger container and shaking the solution to see if it changes colour. A better test would be to generate oxygen gas and bubble it through the solution.

Assessment Options

- Collect and assess students' answers to Analysis questions.
- Use Assessment Checklist 2, Laboratory Report from Appendix A.

16.1 Chemical Equilibrium

Student Textbook pages 634–638

Section Outcomes

Students will:
- identify physical and chemical systems at equilibrium
- describe the conditions necessary for a system to be at equilibrium
- experiment with a model of an equilibrium system
- define a system at equilibrium in terms of the rates of opposing change
- distinguish between the constant macroscopic properties of an equilibrium system and the changes taking place at the molecular level
- design an experiment to show equilibrium shifts

Key Terms

constant macroscopic properties
dynamic equilibrium
homogeneous equilibrium
heterogeneous equilibrium

Chemistry Background

- Equilibrium is a state of a system in which the rates of opposing changes are equal. At equilibrium, the macroscopic properties of a system such as colour, pH, pressure and temperature, are constant. Changes continue to take place at the microscopic level, so adjusting any of the factors affecting the equilibrium will cause a reaction to shift direction to re-establish equilibrium. The equilibrium constant can be interpreted qualitatively to describe the extent of a reaction, and it can be used to calculate the concentrations of an equilibrium mixture.

- Physical equilibrium is usually observed in saturated solutions or in changes of state (e.g., vapour pressure, saturated solutions containing excess solute). Chemical equilibrium is reached when the rates of the forward and reverse reactions are equal and the concentrations of the products are no longer changing with time.
- A reaction at equilibrium may appear to have stopped because the mixture will have constant macroscopic properties. At the molecular level, chemical reactions are still occurring in both forward and reverse directions at the same rate. The state is described as a dynamic equilibrium.
- Homogeneous equilibrium systems involve substances that are all in the same physical state. These are usually gaseous reactions. Heterogeneous equilibrium systems involve substances in different physical states.
- Equilibrium can only be reached in a closed system at constant temperature.

Teaching Strategies

- The subject matter of this chapter is new to most students, and it will form the basis of their understanding of equilibrium in acid-base systems. If possible, dedicate a full class period to Investigation 16.A: Modelling Equilibrium. This activity introduces key concepts: dynamic equilibrium and constant macroscopic properties. In addition, the activity lays the groundwork for the equilibrium constant and Le Châtelier's principle.
- Challenge students to come up with their own analogies for dynamic equilibrium, and check for comprehension and/or misconceptions.
- Explain to students that all reactions are, in principle, reversible. Most of the reactions they have seen before have an equilibrium that lies far to the right, favouring products.
- Stress the significance (and insist that students do the same) of using a double arrow for equilibrium reactions, and a single arrow for reactions that go to completion.
- Several times during the chapter, (e.g., when discussing an experiment, or taking up a problem) encourage students to review "The Four Conditions That Apply to All Equilibrium Systems" in the shaded box on page 637.
- The BLMs prepared for this section support the activities. You will find them with the Chapter 16 BLMs on the CD-ROM that accompanies this Teacher's Resource or at **www.albertachemistry.ca,** Online Learning Centre, Instructor Edition, BLMs.

SUPPORTING DIVERSE STUDENT NEEDS

- Students for whom English is a second language face the added challenge of a unit with specialized terminology and abstract concepts. Carefully enunciate "equilibrium" and emphasize that it is not "equalibrium." Encourage students to list new words, and words with difficult or abstract meanings. Students with language difficulties should be encouraged to draw a series of sketches to show their understanding of a system reaching equilibrium.

Investigation 16.A: Modelling Equilibrium

Student Textbook pages 635–636

Purpose

The purpose of this activity is to model a dynamic equilibrium by transferring water between two graduated cylinders. Students will gain insight into the initial rate of forward and reverse reactions, and will learn that equilibrium is reached when these rates are equal. Students will see that the volumes of water in the two graduated cylinders are not equal at equilibrium. They will model the addition of more reactant to a system at equilibrium, and find that more product is formed. The activity should demonstrate that the ratio, $\frac{\text{volume of product}}{\text{volume of reactant}}$, which models the equilibrium constant, remains unchanged.

Outcomes

- 30-D1.1k
- 30-D1.3s
- 30-D1.4s

Advance Preparation

When to Begin	What to Do
1 week before	■ Cut and fire polish glass tubing. ■ Prepare water and food dye.
1–2 days before	■ Assemble graduated cylinders and labels or grease pencils. ■ Photocopy **BLM 16.1.1 (HAND) Investigation 16.A: Modelling Equilibrium.**

Materials

For each student pair:
- water, coloured with food dye
- 2 graduated cylinders (25 mL)
- 2 glass tubes of different diameters (e.g., 10 mm and 6 mm)
- 2 labels or a grease pencil

Time Required

- 50 to 60 minutes

Helpful Tips

- The glass tubing should be cut to equal lengths, approximately 20 cm long if 25 mL graduated cylinders are used. Straws with different diameters can be used, but some teachers believe students may be more likely to use their mouths to draw up liquid.
- Caution students about glass tubing rolling off the bench. If available, use pipette racks to store the tubing when not being used.
- Encourage students to predict the outcome (and sketch a graph of volume vs. number of transfers) before they begin this activity.
- This activity can be done as a classroom demonstration using two large aquariums and beakers with different volumes. Give one student a small beaker (250 mL) to scoop water from an aquarium into a similar size, empty aquarium. Another student should use a larger beaker (600 mL) to simultaneously scoop water from the second (initially empty) aquarium into the first. Continue the simultaneous transfer of water until equilibrium is established.
- A modeling activity using coins is described in *Journal of Chemical Education*. Vol. 83, No. 1, January 2006, pages 48A–48B.
- Use **BLM 16.1.1 (HAND) Investigation 16.A: Modelling Equilibrium** to support this activity. Remove sections as necessary to meet the needs of students in your class.
- *Extension to help students understand the dynamic nature of equilibrium* and model the concept of the dynamic nature of equilibrium, have students replace the volumes of water in the established equilibrium—the reactant volume coloured with food colouring, the product volume with uncoloured water. Have students continue to perform transfers from both cylinders. As they will see, although the volumes remain constant, confirming that equilibrium is not disturbed, the actual molecules, represented by the coloured/uncoloured water do, in fact, "react," by moving between the two cylinders.
- As the students begin to transfer the water from the "reactants" cylinder to the "products" cylinder, the water level in the "reactants" cylinder will decrease and that in the "products" cylinder will increase. Eventually, the water levels will reach a point at which the amount of water in each of the transfer tubes will be the same and the levels will no longer change. This occurs because the water level will be higher in the small bore glass tube and lower in the large bore tube making the volumes the same.

Safety Precautions

- Remind students of your procedures for handling and disposing of broken glass.

Answers to Analysis Questions

1. The graph for the reactant volume should begin at 25.0 mL and zero transfer. It should have a negative slope that decreases in magnitude with increasing transfer number. The graph for the product volume should begin at 0 mL and zero transfer. It should have a positive slope that decreases in magnitude with increasing transfer number. At equilibrium, the graphs should be horizontal and the reactant volume should be different from the product volume.
2. The reaction rates are equal to the slopes of the graphs. The graph will show that the rate of the forward reaction decreased with number of transfers, while the rate of the reverse reaction increased over the same interval.
3. At the point where the two curves cross, the volume of water in each cylinder is the same. This is not the same as the rate of reaction. The magnitudes of the slopes of the graph are different, so the rate of the forward reaction is not equal to the rate of the reverse reaction at this point.
4. In this activity, the system is at equilibrium when the volumes of water in each cylinder no longer change. In other words, the system has constant macroscopic properties. Students might identify this as the region on the graph where the volume has constant value. They should also state that the system is at equilibrium because the rate of the forward reaction is equal to the rate of the reverse reaction.
5. The volumes of water in the two cylinders were not equal at equilibrium. This indicates that concentrations of reactant and product do not have to be equal at equilibrium. In fact, it would be very rare to find a reaction at equilibrium where concentrations of reactant and product were equal.
6. The addition of more water is analogous to increasing the concentration of a reactant in a chemical reaction. When equilibrium is re-established, there will be a greater concentration of product. The volume of water in the product cylinder should increase as a result of adding more water to the reactant graduated cylinder.
7. Within experimental error, the ratio $\frac{\text{volume of product}}{\text{volume of reactant}}$ at the end of the first equilibrium should equal the ratio at the end of the second equilibrium. This is analogous to a chemical equilibrium that is changed by the addition of more reactant at constant temperature. The equilibrium

constant, K_c, will be unchanged when equilibrium has been re-established.

8. The relative volumes of water in each cylinder were determined by the diameter of the glass tubes. Students can easily devise an experiment to test this hypothesis if glass tubing with a different diameter is substituted for one of the original tubes. A wider glass tube is analogous to a reaction with lower activation energy. Students will probably suggest temperature as a factor that affects the relative concentrations of reactants and products at equilibrium. If students also suggest pressure, accept the answer but be sure to refer back to this exercise after they have learned about Le Châtelier's principle.

9. The experiment is a model of a closed system, because during the exercise there are negligible changes in the volume and temperature of the water.

Answer to Conclusion Question

10. In their conclusion, students should mention the dynamic nature of chemical equilibrium and describe how the rates of forward and reverse reactions change, eventually becoming equal at equilibrium. They should also mention that when reactant is added to a system at equilibrium, the system re-establishes equilibrium with the same ratio of product to reactant as before.

Assessment Options

- Collect and assess students' answer to Analysis and conclusion questions.
- Use Assessment Checklist 2, Laboratory Report from Appendix A.

Answers to Questions for Comprehension

Student Textbook page 636

Q1. Macroscopic properties are observable properties such as temperature, colour, and pH. Microscopic properties are only present at the molecular level.

Q2. Equilibrium can only be established in a closed system. Equilibrium is established when forward and reverse reactions are taking place at the same rate.

Q3. During a dynamic equilibrium, changes are still occurring at the microscopic level. Individual molecules continue to react even though the system displays constant macroscopic properties.

Section 16.1 Review Answers

Student Textbook page 638

1. The physical changes most likely to be mentioned involve changes of state. Answers should indicate an open system, probably involving energy changes. Examples include boiling water in an open saucepan on a stove, or melting a block of ice. Another common answer would involve dissolving a solute in water to make an unsaturated solution. Chemical processes that are examples of reversible changes not at equilibrium include the electrolysis of water and the voltaic cell reactions students learned in the unit on electrochemistry.

2. Investigation 16.A modelled a reversible change in a closed system that had constant macroscopic properties at equilibrium. The fourth condition, that equilibrium can be approached from either direction was not modelled. Students should suggest simple modifications to the activity, such as adding water to the product cylinder, or removing water from either one of the cylinders. If time permits, they could repeat the exercise making the changes they have suggested. They should repeat the calculation for the ratio $\frac{\text{volume of product}}{\text{volume of reactant}}$ to make sure it has not changed, within experimental errors.

3. The rate of a reaction depends on the concentration of reactants and the rate constant. If the reaction is started with only reactants present, the rate of the forward reaction is relatively large while the rate of the reverse reaction is essentially zero. As reactants form products, the concentration of reactants decreases and the rate of the forward reaction also decreases. As the concentration of products increases, the rate of the reverse reaction also increases. At equilibrium, both forward and reverse reactions are taking place at the same rate. At the molecular level, particles are constantly colliding and taking part in reactions, even though there is no overall change in the reaction mixture. This is the nature of a dynamic equilibrium.

4. (a) The sealed drink bottle contains an aqueous solution of carbon dioxide in equilibrium with $CO_2(g)$. The concentration of $CO_2(g)$ above the liquid is greater than the concentration of $CO_2(aq)$. Some $CO_2(g)$ molecules will strike the aqueous solution with sufficient energy to be dissolved. At the same time, some $CO_2(aq)$ molecules will have enough energy to leave the solution and enter the space above the liquid as a gas. On average, the number of $CO_2(g)$ molecules entering the solution in a given time will equal the number of $CO_2(aq)$ molecules leaving the solution.

(b) The system includes the solution and the gas above it. The following properties will be constant: temperature; pH of the solution; concentration of dissolved carbon dioxide; appearance of the solution; pressure of $CO_2(g)$ above the liquid.

5. Ice and water form an equilibrium mixture (slush) at 0 °C and 1 atmosphere pressure. At temperatures above 0 °C, the ice will melt. At temperatures below 0 °C, the water will freeze.

6. (a) The system has reached equilibrium. The forward reaction has formed NOBr(g), and the decomposition of NOBr(g) produces NO(g) and $Br_2(g)$. Because the forward reaction rate is the same as the reverse reaction rate, a mixture of $NO(g)$, $Br_2(g)$ and $NOBr(g)$ will be present in the flask.

(b) This is a homogeneous equilibrium because all the substances taking part in the reactions are gaseous.

7. (a) $PCl_5(g) \rightleftharpoons PCl_3(g) + Cl_2(g)$ homogeneous equilibrium

(b) $CaCO_3(s) \rightleftharpoons CaO(s) + CO_2(g)$ heterogeneous equilibrium

(c) $HF(aq) + H_2O(\ell) \rightleftharpoons H_3O^+(aq) + F^-(aq)$ heterogeneous equilibrium

(d) $N_2O_3(g) \rightleftharpoons NO_2(g) + NO(g)$ homogeneous equilibrium

(e) $C(s) + CO_2(g) \rightleftharpoons 2CO(g)$ heterogeneous equilibrium

8. The double arrow, $\rightleftharpoons$, in a chemical equation indicates that an equilibrium mixture will be formed containing quantities of reactants and products.

9. False. Equilibrium occurs when rates of opposing changes are equal.

10. (a) The towel in the gym bag remains damp because it is in a relatively closed system. There is not enough moisture leaving the system to dry the towel.

(b) A running clothes dryer that is vented to the outside is an open system. The vent gives water vapour a way out.

11. False. Adding heat to a system at equilibrium will change equilibrium.

16.2 Equilibrium Expressions and Le Châtelier's Principle

Student Textbook pages 639–655

Section Outcomes

Students will:

- define the equilibrium constant and write equilibrium law expressions for given chemical equations
- determine the equilibrium constant for a homogeneous equilibrium when given the concentration of reactants and products at equilibrium
- interpret the meaning of equilibrium constants as they apply to the relative concentrations of products and reactants at equilibrium
- identify factors, such as concentration, pressure, volume, or temperature that can be adjusted to affect the relative concentrations of reactants and products at equilibrium
- design an experiment to qualitatively test predictions of shifts in equilibrium

Key Terms

law of chemical equilibrium
equilibrium constant, K_c
equilibrium law expression
Le Châtelier's principle

Chemistry Background

- A state of equilibrium can be established by the simultaneous formation of products from reactants, and reactants from products.
- Substances whose concentrations do not experience appreciable changes (such as pure solids or pure liquids) are not included when expressing the equilibrium constant.
- For a given system at equilibrium, the value of the equilibrium constant depends only on temperature.
- If K_c is very large ($>10^{10}$), a reaction is usually regarded as proceeding to completion. If K_c is very small ($<10^{-10}$), a reaction is usually regarded as not taking place at all.
- According to Le Châtelier's principle, a dynamic equilibrium tends to respond so as to relieve the effect of any change in the conditions that affect the equilibrium. Changes that affect equilibrium include change in temperature, addition or removal of products/reactants, and changes in pressure for a reaction that involve one or more gases.

Teaching Strategies

- Use the example and images on page 640 (Figure 16.5) of the student textbook to help students visualize the meaning of the law of chemical equilibrium.
- Emphasize that square brackets around the formula of a substance (e.g., $[NO_2]$) should be read as the concentration of that substance. Tell students that the majority of problems they will solve relating to the equilibrium law will require equilibrium concentrations of the reacting substances.
- Take time to discuss why a pure solid or a pure liquid is not included in the equilibrium expression, as it is often a source of confusion. Carefully distinguish between a pure liquid and an aqueous solution. Remind students to carefully check the state of each substance in a chemical equation, and insist they include states in chemical equations. If a reaction contains water, they will need to be careful to distinguish among $H_2O(\ell)$, $H_2O(g)$, and aqueous solutions.
- Ensure students understand that for a given chemical reaction, K_c depends only on the temperature. Verify students' understanding that the rate of a reaction depends on the activation energy, and that the activation energies for the forward and reverse reactions differ. Consequently, changing the temperature of a system at equilibrium affects the forward and reverse rates of reaction in different ways.
- Discuss why the units for K_c are dropped. The units of K_c depend on the coefficients in the chemical equation.
- There is a wide range between the values of equilibrium constants for different reactions. The guideline values $K > 10^{10}$ (completion) and $K < 10^{-10}$ (no reaction) are worth trying to remember. However, mention that $K \approx 1$, to

indicate approximately equal concentrations of reactants and products, needs to be interpreted carefully, because the equilibrium expression may involve power terms.

- Students should be made aware that Le Châtelier's principle is only valid for systems initially at equilibrium, and does not apply to irreversible processes.
- Tell students that Le Châtelier's principle is used in other areas of science and economics. The article in *Journal of Chemical Education*, Vol. 77, No. 2, February 2000, pages 173–176, discusses some of these applications. Students who have studied electromagnetism will probably be familiar with Lenz's Law.
- **BLM 16.2.1 (OH) Disturbing Equilibrium** has been prepared for this section. You will find it with the Chapter 16 BLMs on the CD-ROM that accompanies this Teacher's Resource or at **www.albertachemistry.ca**, Online Learning Centre, Instructor Edition, BLMs.

SUPPORTING DIVERSE STUDENT NEEDS

- A possible activity for students to see the way scientific concepts connect would be to construct a web showing how Le Châtelier's principle applies in a non-chemical system. For example, how would some defined change (such as the introduction of an exotic species, or the removal of a dead tree from a forest) affect a steady-state ecological system. Large changes (such as clear-cutting a forest) are usually irreversible, or only recover after a long time. Artistically inclined students could depict the effects of the same perturbations graphically.

Answers to Practice Problems 1–5

Student Textbook page 641

For full solutions to the practice problems, visit **www.albertachemistry.ca,** Online Learning Centre, Instructor Edition, Full Solutions.

1. $$K_c = \frac{[CH_3COOCH_2CH_3][H_2O]}{[CH_3CH_2OH][CH_3COOH]}$$

2. $$K_c = \frac{[NO]^2}{[N_2][O_2]}$$

3. $$K_c = \frac{[H_2O]^2}{[H_2]^2[O_2]}$$

4. $$K_c = \frac{[Fe^{2+}]^2[I_2]}{[Fe^{3+}]^2[I^-]^2}$$

5. $$K_c = \frac{[NO]^4[H_2O]^6}{[NH_3]^4[O_2]^5}$$

Answers to Practice Problems 6–10

Student Textbook page 644

For full solutions to the practice problems, visit **www.albertachemistry.ca,** Online Learning Centre, Instructor Edition, Full Solutions.

6. $K_c = 1.9 \times 10^{-2}$

7. $K_c = 1.2 \times 10^{2}$

8. $K_c = 0.013$

9. $[S_2(g)] = 0.046$ mol/L

10. $[CH_4(g)] = 0.11$ mol/L

Connections (Nature of Science): The Development of Equilibrium Theories

Student Textbook page 645

Teaching Strategies

- The questions here could lead to a class discussion about the nature of science. You may wish to work through these questions with your students.

Answers to Questions

1. The chemical equation should be changed by replacing the single direction arrow with the double direction, equilibrium arrow to show that the reaction takes place in both directions.
2. You would begin by identifying variables that might affect the observation you want to understand. Then you would change one of these variables, keeping the others constant, and observe what effect the change had. After gathering some observations, you would try to explain, or hypothesize, what was happening. Next you would predict the effect of other changes, or combinations of change, and perform experiments to see if your predictions were correct. If the predictions were wrong, this would indicate your understanding was incomplete and you would have to change one or more of your hypotheses.
3. Science advances most effectively when experiments are designed to confirm or disprove accepted theories. As a result of such an experiment, a new hypothesis may conflict with the accepted understanding of an event for a number of reasons. A new hypothesis may be required to explain the observations of an experiment that the previous hypothesis could not explain. New observations may result in a generalization to a more limited hypothesis, and technological advances often allow more critical tests of older hypotheses. Researchers publish the results of their experiments so that others can check their methods, duplicate the results, and see how the results fit established theories.

Chemistry File: FYI

Student Textbook page 648

Faster breathing will lower the concentration of carbonate ions in the chicken's blood: $Ca^{2+}(aq) + CO_3^{2-}(aq) \rightleftharpoons CaCO_3(s)$. This will reduce the concentration of carbonate ions available to form calcium carbonate in the egg shell. Supplying the chickens with carbonated water increases the concentration of carbonate ions. According to Le Châtelier's principle, the equilibrium will shift to the right to form calcium carbonate.

Answers to Practice Problems 11–15

Student Textbook pages 649–650

For full solutions to the practice problems, visit **www.albertachemistry.ca,** Online Learning Centre, Instructor Edition, Full Solutions.

11. The value of K_c is smaller at 1100 K. The position of equilibrium lies farther to the left and favours the formation of $H_2(g)$ and $I_2(g)$ at the higher temperature of 1100 K.

12. The order favouring the formation of products is reaction (III) > (II) > (I).

13. **(a)** proceeds to completion
  (b) essentially does not proceed to right
  (c) proceeds to completion

14. There is essentially no dissociation of HCl(g) into $H_2(g)$ and $Cl_2(g)$.

15. $[Ag(NH_3)_2]^+(aq)$ is more stable

Answers to Practice Problems 16–20

Student Textbook page 651

For full solutions to the practice problems, visit **www.albertachemistry.ca,** Online Learning Centre, Instructor Edition, Full Solutions.

16. The reaction shifts to the left.

17. **(a)** As the reaction proceeds there is no change in the number of gas molecules on each side of the equation. Therefore, increasing the volume of the container has no effect on the position of equilibrium.
  (b) There are more gas molecules on the left side of the equation. Therefore, increasing the volume of the container causes the reaction to shift to the left.
  (c) There are more gas molecules on the right side of the equation. Therefore, increasing the volume of the container causes the reaction to shift to the right.
  (d) There are more gas molecules on the right side of the equation. Therefore, increasing the volume of the container causes the reaction to shift to the right.

18. **(a)** left
  (b) right
  (c) no change
  (d) left
  (e) no change

19. Change **(d)** affects the value.

20. If the reactant, methylcyclohexane, $C_7H_{14}(g)$, is added, the equilibrium will shift to the right because this will use up this reactant and minimize the change. Therefore, increasing $[C_7H_{14}(g)]$ increases the production of toluene. If either product, toluene ($C_7H_8(g)$) or hydrogen ($H_2(g)$), is removed, the equilibrium will shift to the right because this will produce more of these products and minimize the change. Therefore, decreasing either $[C_7H_8(g)]$ or $[H_2(g)]$, or both, increases the production of toluene. Since the number of gas molecules increases as the reaction proceeds from left to right, an increase in the volume of the system in the reaction vessel (decrease in pressure) will result in a shift to the right and increase the production of toluene. Since the reaction is endothermic from left to right, an increase in temperature will cause a shift to the right to minimize this change. Therefore, an increase in temperature will increase the production of toluene.

Investigation 16.B: Disturbing Equilibrium

Student Textbook pages 652–654

Purpose

The purpose of this activity is to provide students with the opportunity to use Le Châtelier's principle to predict the effect of changing one factor that affects a system at equilibrium. Students will then design an experiment to illustrate and test predictions of equilibrium shifts.

Outcomes

- 30-D1.3k
- 30-D1.1sts
- 30-D1.2sts
- 30-D1.1s
- 30-D1.2s

Advance Preparation

When to Begin	What to Do
3–4 weeks before	■ Order materials.
1–2 days before	■ Prepare solutions and materials. ■ Photocopy **BLM 16.2.2 (HAND) Investigation 16.B: Disturbing Equilibrium.**

Materials

- 0.01 mol/L $NH_3(aq)$
- 6.0 mol/L HCl(aq)
- concentrated HCl(aq) in a dropper bottle
- $AgNO_3(aq)$ 0.1 mol/L
- $CoCl_2$ dissolved in a solution of water and ethanol
- phenolphthalein solution
- $NH_4Cl(s)$
- distilled water in a dropper bottle
- 25 mL beaker
- 50 mL beaker (optional)
- 4 test tubes
- test-tube rack
- test-tube holder
- scoopula
- white paper
- hot water bath
- cold water bath

For the optional teacher demonstrations:

- small piece of copper
- concentrated nitric acid
- boiling water
- ice water
- test tube
- test-tube rack
- one-hole stopper
- glass delivery tube
- short length of rubber tubing
- syringe with a cap or rubber stopper to seal the tip
- $NO_2(g)/N_2O_4(g)$ tubes

Time Required

- 60 minutes

Helpful Tips

- While student experiments are preferred, the first two parts could be done as a teacher demonstration. If this is the case, invite suggestions from the class about how to test Le Châtelier's principle and the change they expect to observe.
- The various parts of this investigation can also be split over more than one class as the students are learning about Le Châtelier's principle, or to provide a break from problem-solving.
- When immersed in an ice–water mixture, the colour of gases in a tube containing $NO_2(g)/N_2O_4(g)$ will lighten somewhat. The colour will be noticeably lighter if the tubes are immersed in a solution of acetone and dry ice. If the school does not have the equipment to make dry ice, it can often be purchased at moderate cost from a store that makes ice cream.
- Use **BLM 16.2.2 (HAND) Investigation 16.B: Disturbing Equilibrium** to support this activity. Remove sections as necessry to meet the needs of students in your class.
- ***Expected Results:***
  - Part I: Students methods will vary. They might choose to dilute the solution thus pushing the reaction to the right. Also, they could allow the ammonia to evapourate, drawing the equilibrium to the left.
  - Part II: Cooling the cobalt chloride solution will pull the equilibrium to the left, making the colour pink. Heating the solution will push the equilibrium to the right, causing the colour to become blue or purple.
  - Part III: Cooling the nitrogen dioxide/dinitrogen tetroxide gas mixture will pull the equilibrium to the left, increasing the amount of colourless N_2O_4. Heating the gas will push the equilibrium to the right, increasing the amount of the brown NO_2.

  When pressure is suddenly applied to the sealed syringe of the nitrogen dioxide/dinitrogen tetroxide gas mixture, it is heated causing an increase in the brown NO_2. While maintaining the pressure, the temperature gradually lowers and the amount of the colourless N_2O_4 increases again.

Safety Precautions

- Ensure that students have read "Safety in Your Chemistry Laboratory and Classroom" pages xii-xv in the student textbook. Ensure that students follow all safety precautions outlined for this investigation in the student textbook.
- Ammonia solution is harmful if swallowed or inhaled and is irritating to skin and mucous membranes. The solution is a dilute, weak base. Spills to skin or clothing can be rinsed with cool water.
- Hydrochloric acid and aqueous ammonia are corrosive to eyes and skin and harmful if swallowed or inhaled. The HCl(aq) is concentrated and spills on the bench or floor should be treated with baking soda. If the acid is spilled on the skin, rinse with copious quantities of cool water.
- Ammonium chloride is harmful if inhaled or swallowed and causes skin and eye irritation. Rinse spills with cool water.
- Phenolphthalein solution may irritate skin, eyes, and mucous membranes. Phenolphthalein is a laxative. The solution contains ethanol and is flammable.
- Cobalt chloride and silver nitrate solutions are toxic and corrosive. Keep away from eyes and skin. Wash spills on the skin with plenty of cool water. Silver nitrate will discolour the skin. While not dangerous, the discolouration is unsightly and will remain visible for a few days.
- Ethanol is flammable. Keep ethanol and solutions containing ethanol away from open flames.
- Concentrated nitric acid is highly corrosive and a strong oxidizing agent.

Answers to Procedure Questions

Le Châtelier's principle helps to predict the direction in which an equilibrium reaction will shift to re-establish equilibrium when a change in concentration, pressure, volume or temperature is made. Le Châtelier's principle states that a dynamic equilibrium tends to respond so as to relieve the effect of any change in the conditions that affect the equilibrium.

PART 1

$NH_3(aq) + H_2O(\ell) \rightleftharpoons NH_4^+(aq) + OH^-(aq)$

Initially, the solution is basic and the phenolphthalein will be pink. The following summarizes how students might change the equilibrium and their predicted observations:

- Add HCl(aq). The addition of $H^+(aq)$ removes $OH^-(aq)$ and the equilibrium shifts to the right. A reduction in $OH^-(aq)$ will cause the solution to become clear.
- Add $NH_4Cl(s)$. The solid is soluble in water, and ionizes. The addition of $NH_4^+(aq)$ will shift the equilibrium to the left. The solution will become clear.
- Add $H_2O(\ell)$. The addition of $H_2O(\ell)$ will shift the equilibrium to the right. The solution will become more pink.
- Add $NH_3(aq)$. This will shift the equilibrium to the right and the solution will become pinker.

PART 2

$$\text{Heat} + \underset{\text{pink}}{Co(H_2O)_6^{2+}(aq/et)} + 4Cl^-(aq/et) \rightleftharpoons \underset{\text{blue or purple}}{CoCl_4^{2-}(aq/et)} + 6H_2O(\ell/et)$$

Initially, the solution is blue–purple. The following summarizes how students might change the equilibrium and their predicted observations:

- Add $H_2O(\ell)$. The addition of $H_2O(\ell)$ will shift the equilibrium to the left. The solution will become more pink.
- Add HCl(aq). The addition of $Cl^-(aq)$ will shift the equilibrium to the right. The solution will become more blue or purple.
- Add $AgNO_3(aq)$. The addition of $Ag^+(aq)$ will precipitate AgCl(s), a white solid. The removal of of $Cl^-(aq)$ will shift the equilibrium to the left. The solution will become more pink.
- Place the test tube containing solution in a hot water bath. Heating the solution will shift the equilibrium to the right. The solution will become more blue/purple.
- Place the test tube containing solution in a cold water bath. Cooling the solution will shift the equilibrium to the left. The solution will become more pink.

PART 3

$$\text{Heat} + \underset{\text{colourless}}{N_2O_4(g)} \rightleftharpoons \underset{\text{brown}}{2NO_2(g)}$$

At room temperature, the gaseous mixture is brown.

- When placed into boiling water, the equilibrium shifts to the right and the mixture of gases will become darker brown.
- When placed into ice-cold water, the equilibrium shifts to the left and the mixture of gases will become lighter brown.
- When the mixture of gases is present in a syringe, pressing the plunger down sharply will concentrate the gases and the colour will darken. If the plunger is held in a fixed position, the colour will change over a few seconds as the system re-establishes equilibrium. The equilibrium shifts to the left and the mixture of gases will become lighter in colour.

Unnumbered Figure

Student Textbook page 653

The tubes contain a mixture of gases. The tube on the left is in ice water, so heat has been removed from the environment. According to Le Châtelier's principle, the equilibrium will respond so as to relieve the effect of this change. In this case, the shift will be to the left, as the system adds heat to the environment. This will increase the concentration of $N_2O_4(g)$ and reduce the concentration of $NO_2(g)$. Because $N_2O_4(g)$ is colourless, the colour of the gases inside the tube will become lighter. For the tube on the right in boiling water, heat has been added to the environment. In this case, the shift in equilibrium will be to the right, as the system removes heat from the environment. This will increase the concentration of $NO_2(g)$ and reduce the concentration of $N_2O_4(g)$. Because $NO_2(g)$ is brown, the colour of the gases inside the tube will become darker.

Unnumbered Figure

Student Textbook page 654

Decreasing the volume of the gases heats the mixture and causes the colour to become darker. This can be explained using the same reasoning that was used for the colour change to the tube placed into boiling water. A few seconds later, after the tube has cooled, the colour of the gaseous mixture in the photograph on the right is lighter than before. When the plunger in the syringe was forced down, the pressure on the reacting gases was increased. Le Châtelier's principle predicts that the system will respond to relieve this effect. The equilibrium will shift towards the side with fewer gas molecules, increasing the concentration of $N_2O_4(g)$ and decreasing the concentration of $NO_2(g)$. The gaseous mixture becomes lighter in colour.

Answers to Analysis Questions

1. The accuracy of students' predictions may be used as a means for assessing their understanding of Le Châtelier's principle both before and after the activity.
2. In Part 1, the concentration of a reactant was increased by adding $H_2O(\ell)$ and by adding $NH_3(aq)$. In Part 2, the concentration of a reactant was increased by adding HCl(aq). Increasing the concentration of a reactant caused each equilibrium to respond by forming more product and less reactant.

In Part 1, the concentration of a product was increased by adding $NH_4Cl(s)$. In Part 2, the concentration of a product was increased by adding water. In each case, the equilibrium responded by forming more reactant, and less product.

3. In Part 2, the concentration of a reactant was decreased by adding $AgNO_3(aq)$. This caused the equilibrium to respond by forming more reactant, and less product.

   In Part 1, the concentration of a product was decreased by adding HCl(aq). The equilibrium responded by forming more product, and less reactant.

4. **(a)** Each system is endothermic when read from left to right.

   (b) When heated, each system shifted to the right. In terms of the energy change, the observed shift in equilibrium was toward the endothermic side of the reaction.

   (c) The value of K_c changed when the equilibrium mixture was heated. In these reactions, because the equilibrium shifts to the right, the value of K_c is expected to increase at greater temperature.

5. **(a)** The total pressure of the mixture increased when the plunger was pushed down.

   (b) The pressure of a gas, or a mixture of gases, is directly proportional to the number of gas molecules present.

   (c) Pushing the plunger down increases the pressure of the gaseous mixture. Le Châtelier's principle predicts the system will respond by shifting to reduce the gas pressure. Because there are fewer gas molecules on the left side of the chemical equation, the equilibrium will shift to the left and form more $N_2O_4(g)$. Students will observe (or see from the photographs on page 654) the mixture of gases becoming lighter in colour.

6. Discuss any variations to the anticipated experiments, using Le Châtelier's principle to interpret the results.

Answer to Conclusion Question

7. In small groups, students should discuss and resolve any differences between their results and predictions. Encourage students to think beyond experimental error when considering any differences.

Answer to Application Question

8. Changing the volume has no effect on this equilibrium system because there are the same numbers of gas molecules on both sides of the chemical equation.

Assessment Options

- Collect and assess students' Observations and answers to Prediction, Analysis, Conclusion, and Application questions.
- Use Assessment Checklist 2 Laboratory Reports from Appendix A.

Chemistry File: Try This

Student Textbook page 655

The toy bird will bob up and down as long as its beak remains wet. Water evapourates from the wet beak and cools the top chamber of the bird. This lowers the vapour pressure in the top part of the bird and the column of liquid rises in the stem. The rising liquid raises the centre of gravity until it is above the tipping point of the bird. When the bird tips, the vapour pressure and temperature become uniform, and the bird returns to its upright position.

The equilibrium in the closed system inside the bird involves the physical change, liquid $\rightleftharpoons$ vapour. The position of this equilibrium depends on the temperature of the beak, which in turn depends on evapouration of water from the beak. Students could try a number of ways to change the position of the equilibrium inside the toy bird. For example, they could enclose the bird and the drinking reservoir inside a larger container such as a bell jar. In this case, the relative humidity in the jar will rise, less water will evapourate from the beak, and the bird will stop bobbing after a few minutes. Alternatively, they could add a little ethanol or methanol to the drinking reservoir. This will increase the rate of evapouration of liquid from the beak, and students should be able to predict the effect on the bobbing of the bird.

Section 16.2 Review Answers

Student Textbook page 655

1. **(a)** $K_c = \dfrac{[SbCl_3(g)][Cl_2(g)]}{[SbCl_5(g)]}$

   (b) $K_c = \dfrac{[N_2(g)][H_2O(g)]^2}{[H_2(g)]^2[NO(g)]^2}$

   (c) $K_c = \dfrac{[H_2(g)]^4[CS_2(g)]}{[H_2S(g)]^2[CH_4(g)]}$

2. **(a)** $K_c = \dfrac{[N_2(g)][H_2(g)]^3}{[NH_3(g)]^2}$

   (b) Equilibrium concentrations must be calculated from the information given in the problem.

$$[N_2(g)] = \frac{0.10\text{ mol}}{0.50\text{ L}} = 0.20\text{ mol/L}$$

$$[H_2(g)] = \frac{0.30\text{ mol}}{0.50\text{ L}} = 0.60\text{ mol/L}$$

$$[NH_3(g)] = \frac{0.80\text{ mol}}{0.50\text{ L}} = 1.6\text{ mol/L}$$

   Substituting into the expression

$$K_c = \frac{(0.20)(0.60)^3}{(1.6)^2} = 0.017$$

The equilibrium constant is 0.017.

3. The equilibrium expression is $K_c = \frac{[PCl_5(g)]}{[PCl_3(g)][Cl_2(g)]}$.

Equilibrium concentrations are given in the problem. Substituting into the expression

$$K_c = 49 = \frac{[PCl_5(g)]}{(0.035)(0.035)}$$

$[PCl_5] = 0.060$ mol/L

4. The equilibrium constant is small and [NO] is in the numerator of the equilibrium expression. Therefore, the equilibrium concentration of NO(g) at this temperature is very small.

5. The equilibrium constant is 14.5 and $[CH_3OH]$ is in the numerator of the equilibrium expression. Therefore, the formation of methanol is favourable at this temperature.

6. (a) right
(b) right
(c) no effect
(d) right
(e) left

7. (a) high temperature
(b) low temperature
(c) high temperature
(d) low temperature

8. (a) no effect
(b) reaction shifts to the right
(c) reaction shifts to the left

16.3 Equilibrium Calculations for Homogeneous Systems

Student Textbook pages 656–665

Section Outcomes

Students will:

- determine equilibrium constants and concentrations for homogeneous systems
- apply stoichiometric principles to chemical equations at equilibrium
- solve equilibrium problems using an approximation method
- determine an equilibrium constant by preparing equilibrium mixtures with different concentrations

Chemistry Background

- The equilibrium concentration of each reactant and product can be calculated using the stoichiometry of the reaction, if K_c and the initial concentrations are known, and the concentration of one substance at equilibrium.
- When K_c is small compared with the initial concentration, it may be possible to neglect the small change in concentration that would result at equilibrium. Divide the initial concentration by the value of K_c. If the answer is greater than 1000, the approximation will give an error less than 5% and neglecting the small change in concentration is justified.

Teaching Strategies

- *Please note:* In order to comply with an Alberta Education preference to use the "rule of 1000" *only* with acid-base equilibria, an alternative method for making approximations (to avoid the use of quadratic equations) has been used questions relating to Section 16.3. The "rule of 1000" will be used for questions relating to Section 17.3.
- Continue to insist that students write the states of substances in chemical reactions. While the student textbook is limited to homogeneous problems, students who continue their chemistry studies beyond high school will certainly encounter heterogeneous problems to solve. The habit of checking the state of substances in equilibrium reactions should be well established.
- Emphasize the use of ICE tables. Sometimes, students assign "x" to the first substance with a change in concentration in their ICE table. This is not necessary and may lead to fractional values of x for other substances in the table. Students should always avoid fractional values of x to denote a change in concentration. Although the problem could be solved using fractional values of x, mistakes occur more frequently.
- Students are expected to be able to solve three types of equilibrium problems: simple, perfect square, and approximation. Be sure to work through practice problems of each type.
- The approximation method used in the student textbook is to neglect a change in the initial concentration of a reactant when its concentration is more than 1000 times greater than K_c. This value (1000) is somewhat arbitrary, but conservative. Some texts use the same value, others use 500, and some use 100.
- **BLM 16.3.2 (OH) Using Stoichiometry to Calculate K_c** has been prepared for this section. You will find it with the Chapter 16 BLMs on the CD-ROM that accompanies this Teacher's Resource or at **www.albertachemistry.ca**, Online Learning Centre, Instructor Edition, BLMs.

Answers to Practice Problems 21–24

Student Textbook page 658

For full solutions to the practice problems, visit **www.albertachemistry.ca,** Online Learning Centre, Instructor Edition, Full Solutions.

21. 30

22. $[I_2(g)] = [Cl_2(g)] = 0.015$ mol/L
$[ICl(g)] = 0.14$ mol/L

23. 4.7

24. $[HBr(g)] = 0.045$ mol/L

$[H_2(g)] = [Br_2(g)] = 2.9 \times 10^{-6}$ mol/L

Answers to Practice Problems 25–29

Student Textbook page 660

For full solutions to the practice problems, visit **www.albertachemistry.ca,** Online Learning Centre, Instructor Edition, Full Solutions.

25. $[CO(g)] = [Cl_2(g)] = 1.8 \times 10^{-4}$ mol/L

26. $[H_2(g)] = 9.4 \times 10^{-2}$ mol/L

27. $[CO_2(g)] = 0.60$ mol/L

$[CO(g)] = 1.1 \times 10^{-2}$ mol/L

$[O_2(g)] = 5.6 \times 10^{-3}$ mol/L

28. 6.0×10^{-5} mol/L

29. $[NO(g)] = 1.2 \times 10^{-5}$ mol/L

Thought Lab 16.1: Finding an Equilibrium Law

Student Textbook page 661

Purpose

This activity is designed to illustrate that substituting equilibrium concentrations into the equilibrium expression results in a constant value for a given reaction at constant temperature.

Outcomes

- 30-D1.4s
- 30-D2.3s

Time Required

- 45 minutes

Helpful Tips

- Take time to describe how the experiments would have been performed and the data collected.
- The following reference will be useful when using a spreadsheet for this activity: http://ferl.becta.org.uk/display.cfm?resid=1788&printable=1)
- Use **BLM 16.3.1 (HAND) Thought Lab 16.1: Finding an Equilibrium Law** to support this activity. Remove sections as necessary to meet the needs of students in your class.

Answers to Analysis Questions

2. The data should be entered in a spreadsheet program and show calculated values for the initial $[CH_3COOH(aq)]$, the initial $[CH_3CH_2OH(aq)]$, and the equilibrium $[CH_3COOH(aq)]$. Calculated values should show three decimal places to reduce round-off errors later.

Experiment	Initial CH_3COOH (mol)	Initial CH_3CH_2OH (mol)	Equilibrium CH_3COOH (mol)	Total volume (mL)	Initial $[CH_3COOH]$ (mol/L)	Initial $[CH_3CH_2OH]$ (mol/L)	Equilibrium $[CH_3COOH]$ (mol/L)
1	0.22	0.114	0.125	38.1	5.774	2.992	3.281
2	0.184	0.115	0.0917	40.3	4.566	2.854	2.275
3	0.152	0.121	0.0631	39.4	3.858	3.071	1.602
4	0.214	0.132	0.11	42.6	5.023	3.099	2.582
5	0.233	0.137	0.122	41.5	5.614	3.301	2.940

4. The five ICE tables are shown below.

$$CH_3COOH(aq) + CH_3CH_2OH(aq) \rightleftharpoons CH_3COOCH_2CH_3(aq) + H_2O(aq)$$

Experiment 1

	$[CH_3COOH(aq)]$ (mol/L)	$[CH_3CH_2OH(aq)]$ (mol/L)	$[CH_3COOCH_2CH_3(aq)]$ (mol/L)	$[H_2O(aq)]$ (mol/L)
Initial	5.774	2.992	0	0
Change	–2.493	–2.493	2.493	2.493
Equilibrium	3.281	0.499	2.493	2.493

$$CH_3COOH(aq) + CH_3CH_2OH(aq)) \rightleftharpoons CH_3COOCH_2CH_3(aq) + H_2O(aq)$$

Experiment 2

	$[CH_3COOH(aq)]$ (mol/L)	$[CH_3CH_2OH(aq)]$ (mol/L)	$[CH_3COOCH_2CH_3(aq)]$ (mol/L)	$[H_2O(aq)]$ (mol/L)
Initial	4.566	2.854	0	0
Change	–2.291	–2.291	2.291	2.291
Equilibrium	2.275	0.563	2.291	2.291

$CH_3COOH(aq) + CH_3CH_2OH(aq) \rightleftharpoons CH_3COOCH_2CH_3(aq) + H_2O(aq)$

Experiment 3

	$[CH_3COOH(aq)]$ (mol/L)	$[CH_3CH_2OH(aq)]$ (mol/L)	$[CH_3COOCH_2CH_3(aq)]$ (mol/L)	$[H_2O(aq)]$ (mol/L)
Initial	3.858	3.071	0	0
Change	–2.256	–2.256	2.256	2.256
Equilibrium	1.602	0.815	2.256	2.256

$CH_3COOH(aq) + CH_3CH_2OH(aq) \rightleftharpoons CH_3COOCH_2CH_3(aq) + H_2O(aq)$

Experiment 4

	$[CH_3COOH(aq)]$ (mol/L)	$[CH_3CH_2OH(aq)]$ (mol/L)	$[CH_3COOCH_2CH_3(aq)]$ (mol/L)	$[H_2O(aq)]$ (mol/L)
Initial	5.023	3.099	0	0
Change	–2.441	–2.441	2.441	2.441
Equilibrium	2.582	0.658	2.441	2.441

$CH_3COOH(aq) + CH_3CH_2OH(aq) \rightleftharpoons CH_3COOCH_2CH_3(aq) + H_2O(aq)$

Experiment 5

	$[CH_3COOH(aq)]$ (mol/L)	$[CH_3CH_2OH(aq)]$ (mol/L)	$[CH_3COOCH_2CH_3(aq)]$ (mol/L)	$[H_2O(aq)]$ (mol/L)
Initial	5.614	3.301	0	0
Change	–2.674	–2.674	2.674	2.674
Equilibrium	2.940	0.627	2.674	2.674

5. The calculated values for K_c are shown below.

$$K_c = \frac{[CH_3COOCH_2CH_3(aq)][H_2O(aq)]}{[CH_3COOH(aq)][CH_3CH_2OH(aq)]}$$

Experiment

1 3.80 2 4.10 3 3.90 4 3.51 5 3.88

6. At this point, students should round their calculated values to three significant digits. The average value for K_c is 3.84 and the values are reasonably close. The value furthest from the average is 3.51, but without further details on the experiment students cannot be expected to know if experiment 4 had different experimental or procedural errors compared with the other experiments.

7. Other mathematical relationships among the equilibrium concentrations include the following:

$$K_c = \frac{[CH_3COOCH_2CH_3(aq)] + [H_2O(aq)]}{[CH_3COOH(aq)] + [CH_3CH_2OH(aq)]}$$

$$K_c = \frac{[CH_3COOCH_2CH_3(aq)] + [H_2O(aq)]}{[CH_3COOH(aq)][CH_3CH_2OH(aq)]}$$

$$K_c = \frac{[CH_3COOCH_2CH_3(aq)] \div [H_2O(aq)]}{[CH_3COOH(aq)] \div [CH_3CH_2OH(aq)]}$$

Other mathematical combinations are possible, but the only combination that will give a reasonably constant value for the five experiments is

$$K_c = \frac{[CH_3COOH(aq)][CH_3CH_2OH(aq)]}{[CH_3COOCH_2CH_3(aq)][H_2O(aq)]}$$

This, of course, is the inverse of K_c.

Assessment Options

- Collect and assess student tables from the spreadsheet program and answers to the Analysis questions.

Investigation 16.C: Using Experimental Data to Determine an Equilibrium Constant

Student Textbook pages 662–665

Purpose

In this activity, students will determine the value of K_c at room temperature for the reaction
$Fe^{3+}(aq) + SCN^{-}(aq) \rightleftharpoons Fe(SCN)^{2+}(aq)$

Outcomes

- 30-D2.3s
- AI-NS1
- ICT F1-4.3

Figure

Student Textbook page 663

The top of the figure shows a side view of the solutions. The solution in the graduated cylinder on the right is the least concentrated because the colour intensity of this solution is the lowest. When looking vertically through the solutions the colour intensity is the same because colour intensity is proportional to concentration × depth of solution: $C \times d =$ constant. In this figure, $C_1d_1 = C_2d_2$. The linear relationship between absorbance and concentration is known as the Beer-Lambert law (or sometimes, just Beer's law).

Advance Preparation

When to Begin	What to Do
3 to 4 weeks before	■ Order materials.
1 to 2 days before	■ Prepare solutions. ■ Photocopy **BLM 16.3.3 (HAND) Investigation 16.C: Using Experimental Data to Determine an Equilibrium Constant.**

Materials

- 30 mL 0.0020 mol/L KSCN (aq)
- 30 mL 0.0020 mol/L $Fe(NO_3)_3$(aq) (acidified)
- 25 mL 0.200 mol/L $Fe(NO_3)_3$(aq) (acidified)
- distilled water
- 5 test tubes (18 mm × 150 mm)
- 5 flat-bottom vials

Materials

- 3 beakers (100 mL)
- test-tube rack
- labels or grease pencil
- 1 pipette (20.0 mL)
- 3 pipettes (5.0 mL)
- pipette bulb
- stirring rod
- paper towel
- thermometer (alcohol or digital)
- strip of paper
- diffuse light source, such as a light box
- medicine dropper

Time Required

- 60 to 90 minutes

Helpful Tips

- Each student should predict the equilibrium expression for the reaction being studied:

$$K_c = \frac{[\mathrm{Fe(SCN)^{2+}(aq)}]}{[\mathrm{Fe^{3+}(aq)}][\mathrm{SCN^{-}(aq)}]}$$

- A Probeware version of this experiment is available on the web site if a spectrophotometer or a colourimeter is available. These instruments can be used to measure the absorbance of each equilibrium mixture. By comparing the absorbance of each solution to that of a standard, the equilibrium concentrations of [$FeSCN^{2+}$(aq)] can be more accurately measured.
- The solutions of $Fe(NO_3)_3$ should be prepared in approximately 1 mol/L HNO_3(aq) rather than in water. The presence of the acid will help prevent the formation of $Fe(OH)^{2+}$(aq) (a brown-coloured complex), which can interfere with making absorbance or colour intensity measurements.
- Flat-bottom vials eliminate error due to the curvature of a test tube. However, if these are not available, small test tubes (e.g., 16 mm × 125 mm) may be substituted.
- Remind students that the beakers used to obtain their stock solutions must be clean and dry; otherwise they will dilute the solution they obtained. It is not practical to dry the vials or test tubes used to prepare the equilibrium mixtures. Before making a measurement of an equilibrium mixture, students should rinse the vial or test tube with some of the solution to be measured, then discard the rinse.
- Emphasise that the calculation of the concentration of ions must use the total volume of solution present. Neglecting to do this is a common calculation mistake.
- Tell students that this is an "order of magnitude" experiment, and discuss what this means.
- Preparation of stock solutions:
  - To prepare 0.0020 mol/L KSCN (aq), dissolve 0.19 g KSCN(s) in distilled water and make up to 1.0 L in a volumetric flask. NaSCN(aq) may be substituted by using 0.16 g NaSCN(s).
  - To prepare 0.0020 mol/L $Fe(NO_3)_3$(aq) (acidified), dissolve 0.48 g $Fe(NO_3)_3$(s) in about 500 mL of 1.0 mol/L HNO_3(aq). Make up to 1.0 L in a volumetric flask. If $Fe(NO_3)_3{\bullet}9H_2O$(s) is used, substitute 0.81 g and follow the same procedure.
  - To prepare 0.20 mol/L $Fe(NO_3)_3$(aq) (acidified), dissolve 48 g $Fe(NO_3)_3$(s) in about 500 mL of 1.0 mol/L HNO_3(aq). Make up to 1.0 L in a volumetric flask. If $Fe(NO_3)_3{\bullet}9H_2O$(s) is used, substitute 81 g and follow the same procedure.
- This investigation can be performed on a microscale using 1 mL microtip pipettes and a 24-well plate. Solutions mixed in wells can be stirred using a toothpick. The microscale procedure has the advantage of reducing the quantities of solution used (and that needs to be disposed) and eliminating the preparation of 0.0020 mol/L $Fe(NO_3)_3$(aq).
  - Place 5 drops of KSCN(aq) into wells C3, D2, D3 and D4. Mix the reference solution in well D3 by adding 5 drops of $Fe(NO_3)_3$(aq).
  - Use wells A1 – A4 to dilute the $Fe(NO_3)_3$(aq) solution. Place 6 drops of water into wells A1 – A4. Add 4 drops $Fe(NO_3)_3$(aq) to well A1. Use a clean 1 mL pipette to draw up the solution from well A1. Dispense 4 drops of this solution into well A2, then 5 drops into well C3.
  - Mix the solution in well A2. Dispense 4 drops of the solution from A2 into well A3, then 5 drops into well D4.
  - Mix the solution in well A3. Dispense 4 drops of the solution from A3 into well A4, then 5 drops into well E3.
  - Mix the solution in well A4. Dispense 5 drops of the solution from A4 into well D2.
  - There should be five solutions in wells with the reference solution in the centre. Use a clean microtip pipette to draw up the reference solution from well D3. Place the well plate over a bright light source covered with a sheet of white to disperse the light. Add the reference solution back into its well, drop by drop. Record the number of drops when the solution matches the colour in well D2, then E3, D4, and finally C3.
  - The ratio of the drop count to the reference (10 drops) can be used to calculate the concentration of the coloured ion at equilibrium. It is equivalent to measuring the heights of solution in test tubes.
- Use **BLM 16.3.3 (HAND) Investigation 16.C: Using Experimental Data to Determine an Equilibrium Constant** to support this activity. Remove sections as necessary to meet the needs of students in your class.
- ***Expected Results:*** The expected value of the equilibrium constant is about 1.4×10^2. Since it will be difficult to determine the concentration with accuracy, the students' results will probably vary greatly from this value.

Safety Precautions

- The $Fe(NO_3)_3$ solution is prepared with nitric acid. It should be handled with care. Nitric acid is corrosive and will damage eyes, skin, and mucous membranes. Immediately wash any spills on your skin or clothing with plenty of water.
- Provide containers for chemical disposal so that students do not pour the solutions down the drain.
- Ensure that students have read "Safety in Your Chemistry Laboratory and Classroom" on pages xii-xv in the student textbook. Ensure that students follow all safety precautions outlined for this investigation in the student textbook.

Answer to Prediction Question

$$K_c = \frac{[Fe(SCN)^{2+}(aq)]}{[Fe^{3+}(aq)][SCN^-(aq)]}$$

Answers to Analysis Questions

1. As the volume of $SCN^-(aq)$ in the vials increases, the colour intensity of the $Fe(SCN)^{2+}(aq)$ should increase because the equilibrium is being shifted to the right.

$Fe^{3+}(aq) + SCN^-(aq) \rightleftharpoons Fe(SCN)^{2+}(aq)$

2. Accept values of K_c that are within an order of magnitude. Values of K_c will vary greatly from one group to another; the range is typically from 100 to 800. If a spectrophotometer or a colourimeter is used, the precision should be much improved.

3. (a) The equilibrium concentration of the product would be half the concentration of the product in the actual reaction. This is because in the supposed reaction, each mol of $SCN^-(aq)$ forms half a mole of product.

(b) The value of K_c will be larger. The $[SCN^-(aq)]$ at equilibrium will be reduced and the equilibrium expression is different:

$$K_c = \frac{[Fe(SCN)_2{}^+(aq)]}{[Fe^{3+}(aq)][SCN^-(aq)]^2}$$

Answer to Conclusion Question

4. The conclusion should state that *within experimental error*, the ratio of equilibrium concentrations

$$\frac{[Fe(SCN)^{2+}(aq)]}{[Fe^{3+}(aq)][SCN^-(aq)]}$$

is constant. There will almost certainly be large variations in the calculated value for K_c even in the same group, and students should discuss the experimental errors that are present.

Assessment Options

- Collect and assess students' tables, calculations, and answers to Prediction, Analysis and Conclusion questions.
- Use Assessment Checklist 2 Laboratory Reports from Appendix A.

Section 16.3 Review Answers

Student Textbook page 665

1. These are all homogeneous gas phase equilibrium reactions and the state (g) has been dropped in each expression for K_c, which is common practice.

(a) $K_c = \dfrac{[H_2]^4[CS_2]}{[H_2S]^2[CH_4]}$

(b) $K_c = \dfrac{[P_2O_3]^2}{[P_4][O_2]^3}$

(c) $K_c = \dfrac{[N_3S_4]^2[N_2S_2]^4}{[N_2]^7[S_8]^2}$

2. The ICE table is

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

	$[NH_3(g)]$ (mol/L)	$[N_2(g)]$ (mol/L)	$[H_2(g)]$ (mol/L)
Initial	2.0	0	0
Change	$-2x$	$+x$	$+3x$
Equilibrium	1.6	0.20	0.60

$$K_c = \frac{[N_2][H_2]^3}{[NH_3]^2}$$

$$K_c = \frac{(0.20) \times (0.60)^3}{(1.6)^2}$$

$K_c = 0.017$

3. The ICE table is

$$SO_2(g) + NO_2(g) \rightleftharpoons NO(g) + SO_3(g)$$

	$[SO_2]$ (mol/L)	$[NO_2]$ (mol/L)	$[NO]$ (mol/L)	$[SO_3]$ (mol/L)
Initial	0.36	0.36	0	0
Change	$-x$	$-x$	$+x$	$+x$
Equilibrium	$0.36 - x$	$0.36 - x$	x	x

$$K_c = \frac{[NO][SO_3]}{[SO_2][NO_2]}$$

$$K_c = \frac{x \times x}{(0.36 - x)(0.36 - x)}$$

This is a perfect square.

$$\sqrt{4.8} = \frac{x}{0.36 - x}$$

$x = 0.247$

$[SO_3] = 0.247$ mol/L

In a 5.0 L container at equilibrium at the same temperature there will be

$0.247\ \text{mol/L} \times 5.0\ \text{L}$

$= 1.2\ \text{mol}\ SO_3(g)$.

4. The ICE table is

$$PCl_3(g) + Cl_2(g) \rightleftharpoons PCl_5(g)$$

	$[PCl_3]$ (mol/L)	**$[Cl_2]$ (mol/L)**	**$[PCl_5]$ (mol/L)**
Initial	0.0938	0.0938	0
Change	$-x$	$-x$	$+x$
Equilibrium	0. 0938 – 0.059	0. 0938 – 0.059	0.059
	= 0.0348	= 0.0348	

$$K_c = \frac{[PCl_5]}{[PCl_3][Cl_2]}$$

$$K_c = \frac{0.059}{(0.0348)(0.0348)}$$

$$K_c = 49$$

5. (a) The ICE table is

$$2H_2O(g) \rightleftharpoons 2H_2(g) + O_2(g)$$

	$[H_2O]$ (mol/L)	**$[H_2]$ (mol/L)**	**$[O_2]$ (mol/L)**
Initial	0.055	0	0
Change	$-2x$	$+2x$	$+x$
Equilibrium	$0.055 - 2x$	$2x$	x
	≈ 0.055		

Students should be able to see, without performing the calculation, that

$$\frac{0.055}{7.3 \times 10^{-18}} >> 1000$$

Because x (and $2x$) is very small compared with 0.055, the approximation is valid.

$$K_c = \frac{[H_2]^2[O_2]}{[H_2O]^2}$$

$$7.3 \times 10^{-18} = \frac{(2x)^2(x)}{(0.055)^2}$$

$$x = 1.77 \times 10^{-7}\ \text{mol/L}$$

$$[H_2] = 2 \times 1.77 \times 10^{-7}\ \text{mol/L}$$

$$[H_2] = 3.5 \times 10^{-7}\ \text{mol/L}$$

(b) This is not a practical method, because the equilibrium concentration of $H_2(g)$ is very small.

16.4 Applications of Equilibrium Systems

Student Textbook pages 666–674

Section Outcomes

Students will:

- identify physical and chemical equilibrium important in blood gases
- illustrate the involvement of prototypes, experiments, and theory in the development of technology
- analyze the application of equilibrium principles in an industrial process
- evaluate various factors that are important to the construction of a chemical plant

Key Terms

acidosis
alkalosis

Chemistry Background

- In the atmosphere, $CO_2(g)$ dissolves in water to form an equilibrium mixture with carbonic acid. This equilibrium is very important in the growth of cave formations and coral reefs.
- The chemical industry is the third largest manufacturing industry, behind the automotive and agricultural sectors. Canada's chemical industries are concentrated in three provinces: Ontario, Quebec, and Alberta.
- The reaction conditions used to manufacture a chemical by industry often involve compromises between several factors, including the rate of the reaction, the position of equilibrium, temperature, and pressure.

Teaching Strategies

- Students can identify items around the home that are made using products from local industries.
- Invite a chemical engineer, technologist or tradesperson to your class to talk about the work they do.
- Local industries often have posters and booklets available, or information on a website concerning their processes and products.
- **BLM 16.4.1 (OH) Haber-Bosch Process** has been prepared for this section. You will find it with the Chapter 16 BLMs on the CD-ROM that accompanies this Teacher's Resource or at **www.albertachemistry.ca**, Online Learning Centre, Instructor Edition, BLMs.

SUPPORTING DIVERSE STUDENT NEEDS

Students with an interest in scuba diving could find more information and make a display of how Boyle's, Charles's and Henry's Laws relate to their interest. They should include more information concerning the physiological effects of increased concentrations of $N_2(aq)$ and $O_2(aq)$ in the blood, and the gas mixtures used by divers at greater depth.

Students with an interest in the environment might research the effects of increased $CO_2(g)$ in the atmosphere. They could find out the predictions made by atmospheric models on global temperatures, ice cover in polar regions, changing weather patterns and the distribution of energy by ocean currents.

Students interested in knowing about jobs in the chemical industry could interview someone working in this area. These students might prepare a brochure outlining the educational path and career prospects open to them.

Figure 16.10

Student Textbook page 666

The physical equilibrium of a gas dissolved in a liquid is summarized in Henry's Law: At any given temperature, the solubility of a gas dissolved in a liquid is directly proportional to the partial pressure of this gas above the solution. When the diver breathes air at greater pressure, larger concentrations of $O_2(aq)$ and $N_2(aq)$ appear in the blood. As the text mentions, each of these increases in concentration of dissolved gas represents a potential hazard to the diver.

Section 16.4 Review Answers

Student Textbook page 674

1. (a) The solubility of a gas dissolved in water is directly proportional to the partial pressure of this gas above the solution. Therefore, the solubility of a gas in water increases with an increase in the partial pressure of the gas above the solution. When the diver breathes air at greater pressure, larger concentrations of $O_2(aq)$ and $N_2(aq)$ appear in the blood. As the student textbook mentions, each of these increases in concentration of dissolved gas represents a potential hazard to the diver.

(b) The scuba diver must take into account the physiological effects of increased concentrations of $N_2(aq)$ (nitrogen narcosis) and $O_2(aq)$ (toxic). When ascending from depths greater than about 20 m, they must plan a stop, or stops, at prescribed depth for a certain period of time to allow the dissolved gases to leave the blood safely.

(c) Increased concentration of $N_2(aq)$ causes an effect called nitrogen narcosis. The effect is similar to consuming an alcoholic drink, and a diver might experience impaired judgment, confusion, and drowsiness. At depths of 90 m, nitrogen narcosis can lead to hallucinations and unconsciousness. Oxygen is toxic at increased concentration in the blood, and toxic levels can lead to a seizure.

2. The calcium carbonate shell will react with hot acidic water from hydrothermal vents.

3. (a) To increase the yield of hydrogen gas, the reaction should take place at low pressure.

(b) The reaction is endothermic. Therefore, to increase the yield of hydrogen gas the reaction should take place at a high temperature. This will increase the rate of the reaction.

(c) The catalyst has no effect on the position of equilibrium.

(d) $K_c = \dfrac{[H_2(g)]^4[CO_2(g)]}{[CH_4(g)][H_2O(g)]^2}$

(e) The pressure is raised to increase the yield of ammonia. The reaction is exothermic, so Le Châtelier's principle predicts that the forward reaction is favoured at low temperatures. However, this would slow the rate of reaction.

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

(g) $K_c = \dfrac{[NH_3]^2}{[N_2][H_2]^3}$

(h) Substituting the equilibrium values into the expression for K_c

$$5.0 \times 10^{-2} = \frac{[NH_3]^2}{(1.2) \times (2.0)^3}$$

$$[NH_3] = 0.69 \text{ mol/L}$$

4. As the temperature increases, the value of K_c also increases. That is, as the temperature increases, the products are favoured. The reaction must be endothermic in the forward direction.

Chapter 16 Review Answers

Student Textbook pages 676–677

Answers to Understanding Concepts Questions

1. The statement is almost certainly false. At equilibrium, the rate of change in opposite directions is equal, but the amounts are determined by the position of equilibrium.

2. Chemical equilibrium is a dynamic state because, at the molecular level, both the forward and the reverse reaction continue to take place.

3. In a chemical system at equilibrium, the concentrations of reactants and products are not in the same ratio as the coefficients in the chemical equation. The coefficients in a chemical equation describe the ratio in which substances react, not the concentrations at equilibrium.

4. (a) heterogeneous

(b) homogeneous

(c) heterogeneous

5. (a) $K_c = \dfrac{[CO_2(g)]^3[H_2O(g)]^4}{[C_3H_8(g)][O_2(g)]^5}$

(b) $K_c = \dfrac{[NO(g)]^2[H_2O(g)]^2}{[N_2H_4(g)][O_2(g)]^2}$

(c) $K_c = \dfrac{[Cu(NH_3)_4{}^{2+}(aq)]}{[Cu^{2+}(aq)][NH_3(aq)]^4}$

6. For a reaction that goes to completion, K_c is very large. The equilibrium expression has product terms in the numerator, and reactant terms in the denominator. For a reaction that goes to completion, there will be a very small concentration of reactant. Thus, K_c will be very large.

7. The factors that can affect the equilibrium of a reaction are: the concentration of reactants or products; the pressure of a gaseous system where the number of gas molecules changes during the reaction; and temperature. Of these factors, only temperature affects the value of K_c.

8. The reaction shifts to the right at higher temperature. Therefore, according to Le Châtelier's principle, the reaction is endothermic.

9. (a) Le Châtelier's principle predicts the equilibrium will shift to the right if dilute acid is added. The colour of the mixture should be more orange than before the addition of acid.

(b) If dilute base is added, it will remove $H_3O^+(aq)$ from the mixture. The equilibrium will shift to the left, and the mixture will become more yellow.

10. Changing the volume of the mixture has no effect in this case because there is the same number of gas molecules on each side of the equation. A decrease in temperature will shift the equilibrium to the right because the reaction is exothermic. Le Châtelier's principle predicts that increasing the temperature will shift the position of equilibrium towards the endothermic reaction, thus reducing the effect of the change.

11. An increase in pressure will result in a shift to the left where there are fewer gas molecules.

12. Pumping in a non-reacting gas has no effect on the equilibrium. This is because there is no change in the concentrations of the reacting gases.

13. Cousteau used a mix of helium and oxygen (Heliox) with relatively small $O_2(g)$ concentration to make sure his blood oxygen concentration did not reach toxic levels.

14. (a) Increasing the pressure of methane would shift the equilibrium to the right, increasing the concentration of hydrogen.

(b) Increasing the pressure of carbon dioxide would shift the equilibrium to the left, decreasing the concentration of hydrogen.

(c) Decreasing the pressure of water vapour would shift the equilibrium to the left, decreasing the concentration of hydrogen.

(d) The reaction is endothermic and raising the temperature would shift the equilibrium to the right, increasing the concentration of hydrogen.

(e) Adding a catalyst would have no effect on the concentration of hydrogen.

(f) Adding helium gas to the mixture would have no effect on the concentration of hydrogen.

(g) Transferring the mixture of gases to a container with greater volume would have the effect of reducing the pressure of the reacting mixture. The equilibrium would shift to the right, increasing the concentration of hydrogen.

Answers to Applying Concepts Questions

15. (a) $4HCl(g) + O_2(g) \rightleftharpoons 2H_2O(g) + 2Cl_2(g)$

(b)
$$K_c = \frac{[H_2O(g)]^2[Cl_2(g)]^2}{[HCl(g)]^4[O_2(g)]}$$
$$= \frac{(7.8 \times 10^{-3})^2(3.6 \times 10^{-3})^2}{(2.7 \times 10^{-2})^4(8.6 \times 10^{-2})}$$
$$= 1.7 \times 10^{-2}$$

16. The chemical equation is $2SO_2(g) + O_2(g) \rightleftharpoons 2SO_3(g)$

$$K_c = \frac{[SO_3(g)]^2}{[SO_2(g)]^2[O_2(g)]}$$

Substituting equilibrium concentrations from experiment 1

$$K_c = \frac{(3.50)^2}{(1.50)^2(1.25)} = 4.36$$

Substituting equilibrium concentrations from experiment 2

$$K_c = \frac{(0.260)^2}{(0.590)^2(0.0450)} = 4.32$$

These values are probably equal, within experimental error.

17. $4NH_3(g) + 5O_2(g) \rightleftharpoons 4NO(g) + 6H_2O(g)$

Since $[NO(g)] = [NH_3(g)]$, and these terms are raised to the same power, they divide out.

$$K_c = \frac{(2.0)^6}{(3.0)^5} = 0.26$$

18. $K_c = \dfrac{[CH_3OH(g)]}{[CO(g)][H_2(g)]^2}$

$$14.5 = \frac{[CH_3OH(g)]}{(0.25)(0.15)^2}$$

$[CH_3OH(g)] = 0.082$ mol/L

19. The ICE table is

$$2SO_2(g) \quad + \quad O_2(g) \quad \rightleftharpoons \quad 2SO_3(g)$$

	$[SO_2]$ (mol/L)	$[O_2]$ (mol/L)	$[SO_3]$ (mol/L)
Initial	3.8	3.8	0
Change	$-2x$	$-x$	$+2x$
Equilibrium	$3.8 - 2x$	$3.8 - x$	$2x$
	≈ 3.8	≈ 3.8	

Check the assumption that x is negligible compared with 3.8

$$\frac{3.8}{3.6 \times 10^{-3}} = 1056$$

Since this is greater than 1000, the approximation is valid.

$$K_c = \frac{[SO_3(g)]^2}{[SO_2(g)]^2[O_2(g)]}$$

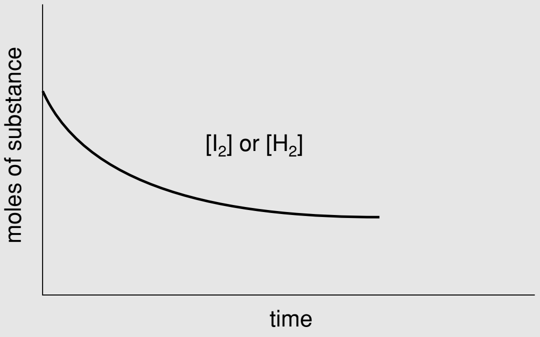
$$3.6 \times 10^{-3} = \frac{(2x)^2}{(3.8)^2(3.8)}$$

$$x = 0.222$$

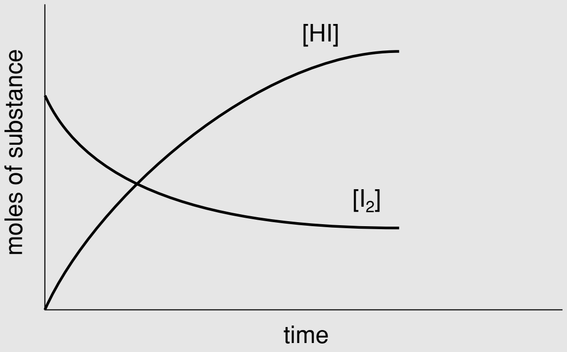
$$[SO_3(g)] = 2x = 0.44 \text{ mol/L}$$

Answers to Solving Problems Questions

20. (a)



(b). The graph of $[I_2(g)]$ and $[HI(g)]$ will be much different.



21. Carbon monoxide is dangerous because it is a highly poisonous, colourless and odourless gas. The gas forms in any situation where there is incomplete combustion caused by inadequate supply of air. The flyer should mention some of the common causes, such as operating a gasoline engine in a closed garage, or a natural gas furnace with a blocked supply of air. The flyer should list the symptoms of CO(g) poisoning (headache and drowsiness), and suggest that homeowners check their furnaces for blockages (perhaps caused by nests) before the weather turns cold. The flyer should also mention that homeowners can easily install a CO(g) detector.

Answers to Making Connections Questions

22. The value for the equilibrium constant of a reaction is not related to the rate of the reaction.

23. Sulfide ions, $S^{2-}(aq)$, will combine with hydronium ions and remove them from solution. Consequently, the urine will become more basic and the pH will increase.

24. (a) The reaction conditions that favour increasing the yield of nitrogen monoxide are low pressure and low temperature.

(b) The catalyst has no effect on the position of equilibrium. A catalyst favours the production of NO(g) because it will increase the rate of the reaction.

(c) The reaction temperature is relatively high because, although the yield decreases at higher temperature, this is more than offset by an increase in the rate of the reaction.

(d) A relatively low pressure is used because this favours the production of NO(g). The cost of increasing the pressure in the system would not be justified.

(e) The reaction is exothermic, so cooling the gases forces the equilibrium towards more product. Also, it is probable that cooling the gases is a safety measure to prevent the reaction vessel from becoming dangerously overheated. The heat extracted would be used elsewhere in the plant.

(f) The nitrogen monoxide that is formed is recycled back to react with oxygen.

(g) Nitric acid has many uses including the manufacture of fertilizers, explosives, dyes, nylon, and pharmaceuticals.

CHAPTER 17 ACID-BASE EQUILIBRIUM SYSTEMS

Curriculum Correlation

(Note: This correlation includes Chapters 16 and 17. Chapter 17 references are in bold.)

General Outcome 1: Students will explain that there is a balance of opposing reactions in chemical equilibrium systems.

	Student Textbook	Assessment Options
Outcomes for Knowledge		
30–D1.1k define equilibrium and state the criteria that apply to a chemical system in equilibrium, i.e., closed system, constancy of properties, equal rates of forward and reverse reactions	Chemical Equilibrium, Section 16.1, p. 634 Conditions That Apply to All Equilibrium Systems, Section 16.1, pp. 636–637	Questions for Comprehension: 1–3, Section 16.1, p. 636 Section 16.1 Review: 1–11, p. 638 Chapter 16 Review: 1–3, 6, 7, 22, pp. 676–677 Chapter 16 Test Unit 8 Review: 1–3, 5–11, 20, pp. 724–727
30–D1.2k identify, write and interpret chemical equations for systems at equilibrium	The Law of Chemical Equilibrium, Section 16.2, pp. 639–641 Equilibrium Law Expression, Section 16.2, p. 641 Sample Problem: Writing Equilibrium Expression for Homogeneous Chemical Reactions, Section 16.2, p. 641 Sample Problem: Calculating an Equilibrium Constant, Section 16.2, pp. 642–643 Sample Problem: Calculating an Equilibrium Concentration, Section 16.2, p. 643	Practice Problems: 1–5, Section 16.2, p. 641 Practice Problems: 6–10, Section 16.2, p. 644 Section 16.2 Review: 1–8, p. 655 Chapter 16 Review: 4, 5, 8–12, 14–19, 24, pp. 676–677 Chapter 16 Test Unit 8 Review: 2–4, 7–10, 12, 19, 26–31, 33, 37, 38, 41, 50, 52, 53, pp. 724–727
30–D1.3k predict, qualitatively, using Le Châtelier's principle, shifts in equilibrium caused by changes in temperature, pressure, volume, concentration or the addition of a catalyst, and describe how these changes affect the equilibrium constant	Le Châtelier's Principle, Section 16.2, pp. 646–647 Sample Problem: Temperature and the Extent of a Reaction, Section 16.2, p. 640 Sample Problem: Using Le Châtelier's Principle, Section 16.2, p. 651	Practice Problems: 11–15, Section 16.2, pp. 649–650 Practice Problems: 16–20, Section 16.2, pp. 651 Section 16.2 Review: 3, 5–8, p. 655 Chapter 16 Review: 10–14, 18, 19, 24, pp. 676–677 Chapter 16 Test Unit 8 Review: 9, 26–30, 34, pp. 724–727
30–D1.4k define K_c and write equilibrium law expressions for given chemical equations, using lowest whole–number coefficients	The Law of Chemical Equilibrium, Section 16.2, pp. 639–641 Equilibrium Law Expression, Section 16.2, p. 641 Sample Problem: Writing Equilibrium Expression for Homogeneous Chemical Reactions, Section 16.2, p. 641 Sample Problem: Using Stoichiometry to Calculate K_c, Section 16.3, pp. 657–658 Sample Problem: Using the Approximation Method, Section 16.3, pp. 659–660	Practice Problems: 1–5, Section 16.2, p. 641 Practice Problems: 6–10, Section 16.2, p. 644 Section 16.2 Review: 1–8, p. 655 Practice Problems: 21–24, Section 16.3, p. 658 Practice Problems: 25–29, Section 16.3, p. 660 Section 16.3 Review: 1–5, p. 665 Chapter 16 Review: 3, 5–7, pp. 676–677 Chapter 16 Test Unit 8 Review: 4, 7, 10, 11, 26–30, pp. 724–727

	Student Textbook	Assessment Options
30–D1.5k describe Brönsted–Lowry acids as proton donors and bases as proton acceptors	**Understanding Acids and Bases, Section 17.2, p. 684** **Sample Problem: Conjugate Acid–Base Pairs, Section 17.2, pp. 686–687**	**Questions for Comprehension: 3–5, Section 17.2, p. 686** **Practice Problems: 1–4, Section 17.2, p. 687** **Section 17.3 Review: 1–6, p. 704** **Chapter 17 Test** Unit 8 Review: 16–19, 21–23, 25, 31–39, 44–47, 50, 52, 53, pp. 724–727
30–D1.6k write Brönsted–Lowry equations and predict whether reactants or products are favoured for acid–base equilibrium reactions (including indicators, polyprotic acids, and polyprotic bases)	**Predicting the Direction of Reaction for an Acid–Base Reaction, Section 17.2, p. 689** **Sample Problem: Predicting the Direction of an Acid–Base Reaction, Section 17.2, p. 689**	**Practice Problems: 8–10, Section 17.2, p. 690** **Section 17.2 Review: 1–4, p. 690** **Practice Problems: 11–13, Section 17.3, p. 692** **Chapter 17 Review: 7, p. 720–721** **Chapter 17 Test** Unit 8 Review: 16–19, 21–23, 25, 31–39, 44–47, 50, 52, 53, pp. 724–727
30–D1.7k identify polyprotic acids, polyprotic bases, conjugate pairs and amphiprotic substances	**Both an Acid and a Base: Amphiprotic Substances, Section 17.2, p. 688** **Sample Problem: An Amphiprotic Ion, Section 17.2, p. 688**	**Practice Problems: 5–7, Section 17.2, p. 688** **Section 17.2 Review: 3, p. 690** **Chapter 17 Test** Unit 8 Review: 16–19, 23, 25, 31, 34, pp. 726–729
30–D1.8k define a buffer as relatively large amounts of a weak acid and its conjugate base in equilibrium that maintain a relatively constant pH when small amounts of acid or base are added	**Buffer Solutions, Section 17.4, pp. 712–713**	**Questions for Comprehension: 1, 2, Section 17.1, p. 682** **Section 17.4 Review: 6–8, p. 718** **Chapter 17 Review: 11, 26, pp. 720–721** **Chapter 17 Test** Unit 8 Review: 24, 25, 43, 47, pp. 724–727
30–D1.9k sketch and qualitatively interpret titration curves of monoprotic and polyprotic acids and bases identifying equivalence points and regions of buffering for weak acid–strong base, strong acid–weak base, and strong acid–strong base.	Thought Lab 16.1: Finding and Equilibrium Law, Section 16.3, p. 661 **Titration Curves and Buffers, Section 17.4, pp. 705–708**	Thought Lab 16.1: Finding and Equilibrium Law: 1–7, Section 16.3, p. 661 **Questions for Comprehension: 9–12, Section 17.4, p. 710** **Section 17.4 Review: 1–8, p. 718** **Chapter 17 Review: 25, pp. 720–721** **Chapter 17 Test** Unit 8 Review: 39, 44–47, pp. 724–727
Outcomes for Science, Technology and Society (Emphasis on nature of science)		
30–D1.1sts demonstrate an understanding that the goal of science is knowledge about the natural world by ■ *applying equilibrium theories and principles to analyze a variety of phenomena, e.g.,* – *carbon dioxide escaping from an open bottle/can of carbonated beverage* – *role of the oceans in the carbon cycle* – *solubility of oxygen gas in lake water* – *acid precipitation (deposition)* – *blood gases in deep–sea diving* – *buffers in living systems*	Chapter 16 Launch Lab: The Chemical Blues, p. 633 Connections: The Development of Equilibrium Theories, Section 16.2, p. 645	Chapter 16 Launch Lab: The Chemical Blues: 1, 2, p. 633 Connections: The Development of Equilibrium Theories, Section 16.2, p. 645 **Chapter 17 Review: 2, 21, 22, 27–29, pp. 720–721** Unit 8 Review: 13–15, 39–53, pp. 724–727

	Student Textbook	Assessment Options
30–D1.2sts demonstrate an understanding that scientific knowledge and theories develop through hypotheses, the collection of evidence through experimentation and the ability to provide explanations, e.g., *research how equilibrium theories and principles developed*	Chapter 16 Launch Lab: The Chemical Blues, p. 633 Connections: The Development of Equilibrium Theories, Section 16.2, p. 645	Chapter 16 Launch Lab: The Chemical Blues: 1, 2, p. 633 Connections: The Development of Equilibrium Theories: 1–3, Section 16.2, p. 645 **Chapter 17 Review: 2, 21, 22, pp. 720–721** Unit 8 Review: 13–15, 39–53, pp. 724–727
30–D1.3sts demonstrate an understanding that the goal of technology is to provide solutions to practical problems by ■ *analyzing how equilibrium principles have been applied in industrial processes, e.g.,* – *the Haber–Bosch process for making ammonia* – *the Solvay process for making sodium carbonate* – production of methanol.	Chapter 16 Launch Lab: The Chemical Blues, p. 633 Connections: The Development of Equilibrium Theories, Section 16.2, p. 645 Investigation 16.B: Disturbing Equilibrium, Section 16.2, pp. 652–654	Chapter 16 Launch Lab: The Chemical Blues: 1, 2, p. 633 Connections: The Development of Equilibrium Theories: 1–3, Section 16.2, p. 645 Investigation 16.B: Disturbing Equilibrium: 1–8, Section 16.2, pp. 652–654 Section 16.4 Review: 1–4, p. 674 **Section 17.1 Review: 1–6, p. 683** Unit 8 Review: 13–15, 39–53, pp. 724–727
Skill Outcomes (Focus on scientific inquiry)		
Initiating and Planning		
30–D1.1s ask questions about observed relationships and plan investigations of questions, ideas, problems and issues by ■ predicting variables that can cause a shift in equilibrium ■ designing an experiment to show equilibrium shifts, e.g., *colour change, temperature change, precipitation* ■ describing procedures for safe handling, storage and disposal of materials used in the laboratory, with reference to WHMIS and consumer product labelling information ■ *designing a buffering system.*	Chapter 16 Launch Lab: The Chemical Blues, p. 633 Investigation 16.A: Modelling Equilibrium, Section 16.1, pp. 635–636 Investigation 16.B: Disturbing Equilibrium, Section 16.2, pp. 652–654 Investigation 16.C: Using Experimental Data to Determine an Equilibrium Constant, Section 16.3, pp. 662–665	Chapter 16 Launch Lab: The Chemical Blues: 1, 2, p. 633 Investigation 16.A: Modelling Equilibrium: 1–10, Section 16.1, pp. 635–636 Investigation 16.B: Disturbing Equilibrium: 1–8, Section 16.2, pp. 652–654 Investigation 16.C: Using Experimental Data to Determine an Equilibrium Constant: 1–4, Section 16.3, pp. 662–665 Unit 8 Review: 39, 40, 44, 47, pp. 724–727
Performing and Recording		
30–D1.2s conduct investigations into relationships among observable variables and use a broad range of tools and techniques to gather and record data and information by ■ performing an experiment to test, qualitatively, predictions of equilibrium shifts, e.g., *colour change, temperature change, precipitation, and gas production* ■ *preparing a buffer to investigate the relative abilities of a buffer and a control; i.e., water, to resist a pH change when a small amount of strong acid or strong base is added.*	Investigation 16.A: Modelling Equilibrium, Section 16.1, pp. 635–636 Investigation 16.B: Disturbing Equilibrium, Section 16.2, pp. 652–654 Investigation 16.C: Using Experimental Data to Determine an Equilibrium Constant, Section 16.3, pp. 662–665	Investigation 16.A: Modelling Equilibrium: 1–10, Section 16.1, pp. 635–636 Investigation 16.B: Disturbing Equilibrium: 1–8, Section 16.2, pp. 652–654 Investigation 16.C: Using Experimental Data to Determine an Equilibrium Constant: 1–4, Section 16.3, pp. 662–665 Unit 8 Review: 39, 40, 44, 47, pp. 724–727
Analyzing and Interpreting		
30–D1.3s analyze data and apply mathematical and conceptual models to develop and assess possible solutions by ■ writing the equilibrium law expression for a given equation ■ analyzing, qualitatively, the changes in concentrations of reactants and products after an equilibrium shift ■ interpreting data from a graph to determine when equilibrium is established, and determining the cause of a stress on the system.	Investigation 16.B: Disturbing Equilibrium, Section 16.2, pp. 652–654 Thought Lab 16.1: Finding an Equilibrium Law, Section 16.3, p. 661 Investigation 16.C: Using Experimental Data to Determine an Equilibrium Constant, Section 16.3, pp. 662–665	Investigation 16.B: Disturbing Equilibrium: 1–8, Section 16.2, pp. 652–654 Thought Lab 16.1: Finding an Equilibrium Law, Section 16.3, p. 661 Investigation 16.C: Using Experimental Data to Determine an Equilibrium Constant: 1–4, Section 16.3, pp. 662–665 Unit 8 Review: 10, 26–30, 32–34, 39, 47, 49–53, pp. 724–727

	Student Textbook	Assessment Options
Communication and Teamwork		
30–D1.4s work collaboratively in addressing problems and apply the skills and conventions of science in communicating information and ideas and in assessing results by ■ *working collaboratively with team members to develop an illustration and explanation of reversible reactions* ■ *using advanced menu features within a word processor to develop a group report on equilibrium systems.*	Investigation 16.B: Disturbing Equilibrium, Section 16.2, pp. 652–654	Investigation 16.B: Disturbing Equilibrium: 1–8, Section 16.2, pp. 652–654

General Outcome 2: Students will determine quantitative relationships in simple equilibrium systems.

	Student Textbook	Assessment Options
Outcomes for Knowledge		
30–D2.1k recall the concepts pH and hydronium ion concentration, pOH and hydroxide ion concentration in relation to acids and bases	Describing Acidic and Basic Solutions, Unit 8 Preparation, pp. 630–631	Chapter 16 Review: 23, pp. 676–677 **Chapter 17 Review: 1, 3–6, 8, 13, 15, 19, pp. 720–721** **Chapter 17 Test** Unit 8 Review: 31, 32, 34, 36, 39, 43, 47, 49, 50, pp. 724–727
30–D2.2k define K_w , K_a , K_b and use these to determine pH, pOH, $[H_3O^+]$, $[OH^-]$ of acidic and basic solutions	**Acid–Base Equilibriums, Section 17.3, p. 691**	
	Sample Problem: Acid Ionization Expressions, Section 17.3, p. 691	**Practice Problems: 11–13, Section 17.3, p. 692**
	Sample Problem: Determining K_a and Percent Ionization, Section 17.3, pp. 692–693	**Practice Problems: 14–18, Section 17.3, p. 693**
	The Ion Product Constant for Water, Section 17.3, pp. 695–696	**Questions for Comprehension: 6–8, Section 17.3, p. 696**
	The Base Ionization Constant K_b, Section 17.3, pp. 696–697	
	Sample Problem: Calculating K_b for a Weak Base, Section 17.3, p. 697	**Practice Problems: 19–23, Section 17.3, p. 698**
	Calculating the pH of a Solution of a Weak Acid or a Weak Base, Section 17.3, p. 698	
	Sample Problem: Calculating the pH of a Weak Acid, Section 17.3, p. 699	**Practice Problems: 24–27, Section 17.3, p. 699** **Practice Problems: 28–33, Section 17.3, p. 701**
	Sample Problem: Calculating the pH of a Weak Base, Section 17.3, p. 700	**Section 17.3 Review: 1, 3–6, p. 704** **Chapter 17 Review: 1, 3–6, 8, 12–20, 23, 24, 28, 29, pp. 720–721** **Chapter 17 Test** Unit 8 Review: 20, 21, 31, 32–36, 39, 43, 47, 49, 50, pp. 724–727

	Student Textbook	Assessment Options
30–D2.3k calculate equilibrium constants and concentrations for homogeneous systems and Brönsted–Lowry acids and bases (excluding buffers) when ■ concentrations at equilibrium are known ■ initial concentrations and one equilibrium concentration are known ■ the equilibrium constant and one equilibrium concentration are known **Note:** Examples that require the application of the quadratic equation are excluded; however, students may use this method in responding to open–ended questions.	**The Relationship Between Conjugate Acid–Base Pairs, Section 17.3, p. 701** **Sample Problem: Equilibrium Constants for Conjugate Acid–Base Pairs, Section 17.3, p. 702** **Sample Problem: Predicting the Direction of an Acid–Base Reaction, Section 17.3, pp. 702–703**	**Practice Problems: 34–38, Section 17.3, p. 703** **Section 17.3 Review: 1, 4, p. 704** **Chapter 17 Review: 23, 28, 29, pp. 720–721** **Chapter 17 Test** Unit 8 Review: 20, 21, 27, 29, 31–39, 47, 49, 51, pp. 724–727
Outcomes for Science, Technology and Society (Emphasis on nature of science)		
30–D2.1sts develop an understanding that technological development may involve the creation of prototypes and testing, as well as application of knowledge from related scientific and interdisciplinary fields by ■ *analyzing, on the basis of chemical principles, the application of equilibrium in, e.g.,* – *industrial processes or medical sciences* – *antacid tablets, buffering in living systems* – *acid precipitation.*	**Chapter 17 Launch Lab: Buffering Ground Water: A Delicate Balance, p. 679** **Connections: Buffers in the Blood, Section 17.4, p. 717**	**Chapter 17 Launch Lab: Buffering Ground Water: A Delicate Balance: 1–3, p. 679** **Connections: Buffers in the Blood: 1–3, Section 17.4, p. 717**
Skill Outcomes (Focus on scientific inquiry)		
Initiating and Planning		
30–D2.1s ask questions about observed relationships and plan investigations of questions, ideas, problems and issues by ■ designing an experiment to show qualitative equilibrium shifts in concentration under a given set of conditions ■ describing procedures for safe handling, storage and disposal of materials used in the laboratory, with reference to WHMIS and consumer product labelling information.	**Investigation 17.B: Preparing a Buffer and Investigating Its Properties, Section 17.4, pp. 710–712**	**Investigation 17.B: Preparing a Buffer and Investigating Its Properties: 1–8, Section 17.4, pp. 710–712**
Performing and Recording		
30–D2.2s conduct investigations into relationships among observable variables and use a broad range of tools and techniques to gather and record data and information by ■ performing an experiment to show equilibrium shifts in concentration	**Investigation 17.A: Determining K_a for Ethanoic Acid, Section 17.2, pp. 694–695** **Investigation 17.B: Preparing a Buffer and Investigating Its Properties, Section 17.4, pp. 710–712**	**Investigation 17.A: Determining K_a for Ethanoic Acid: 1–8, Section 17.2, pp. 694–695** **Investigation 17.B: Preparing a Buffer and Investigating Its Properties: 1–8, Section 17.4, pp. 710–712**

	Student Textbook	Assessment Options
Analyzing and Interpreting		
30–D2.3s analyze data and apply mathematical and conceptual models to develop and assess possible solutions by ■ using experimental data to calculate equilibrium constants.	**Investigation 17.A: Determining K_a for Ethanoic Acid, Section 17.2, pp. 694–695** **Thought Lab 17.1: Analyzing a Weak Acid–Strong Base Titration, Section 17.4, p. 709** **Investigation 17.B: Preparing a Buffer and Investigating Its Properties, Section 17.4, pp. 710–712**	**Investigation 17.A: Determining K_a for Ethanoic Acid: 1–8, Section 17.2, pp. 694–695** **Thought Lab 17.1: Analyzing a Weak Acid–Strong Base Titration: 1–4, Section 17.4, p. 709** **Investigation 17.B: Preparing a Buffer and Investigating Its Properties: 1–8, Section 17.4, pp. 710–712**
Communication and Teamwork		
30–D2.4s work collaboratively in addressing problems and apply the skills and conventions of science in communicating information and ideas and in assessing results by ■ *using advanced menu features within a word processor to develop a group report on equilibrium applications in Alberta industries.*	**Thought Lab 17.1: Analyzing a Weak Acid–Strong Base Titration, Section 17.4, p. 709** **Investigation 17.B: Preparing a Buffer and Investigating Its Properties, Section 17.4, pp. 710–712**	**Thought Lab 17.1: Analyzing a Weak Acid–Strong Base Titration: 1–4, Section 17.4, p. 709** **Investigation 17.B: Preparing a Buffer and Investigating Its Properties: 6, 7, Section 17.4, pp. 710–712**

Chapter 17

Acid-Base Equilibrium Systems

Student Textbook pages 678–721

Chapter Concepts

Section 17.1 Acid Deposition

- Acid deposition, which includes acid rain, is caused by natural processes in the atmosphere and by industrial activity.

Section 17.2 Understanding Acids and Bases

- In the Brønsted-Lowry theory, an acid is a substance capable of donating a proton, H^+, and a base is a substance capable of accepting a proton.

Section 17.3 Acid-Base Equilibriums

- The percent ionization of a weak acid or a weak base is related to its equilibrium constant, K_a or K_b.
- The ion product constant for pure water, K_w, is the equilibrium constant for the ionization of water into its ions.
- For a conjugate acid-base pair, $K_a \times K_b = K_w$.
- The constants K_a and K_b can be used to determine pH, pOH, $[H_3O^+(aq)]$, and $[OH^-(aq)]$ of acidic and basic solutions.
- The direction of an acid-base reaction can be predicted by using the values of K_a and K_b for an acid and base reacting together.

Section 17.4 Titration Curves and Buffers

- An acid-base titration curve is a plot of the pH of a solution as a function of the volume of titrant added.
- Buffer solutions resist change in pH when a small amount of an acid or a base is added.

Common Misconceptions

- Some students may believe that all compounds that contain hydrogen atoms are acidic, and all compounds containing OH are basic. A compound such as table sugar (sucrose, $C_{12}H_{22}O_{11}$) contains both hydrogen atoms and OH groups, but dissolves in water to form a neutral solution because no dissociation into ions occurs.
- Students often think of a solution with high pH as being more acidic, rather than being more basic. This usually results when students understand that a more acidic solution contains greater concentration of hydronium ion, but then fail to think of the definition of pH as $-\log[H_3O^+(aq)]$. Have these students memorize a few reference markers on the pH scale. For example, stomach acid has a pH of about 3, and a typical ammonia solution used for glass cleaner has a pH of about 10.
- Students can easily misunderstand what is meant when an acid is described as a proton donor. Emphasize that the proton removed from an acid comes from the acid molecule, not from the nucleus of one of the atoms in the molecule.
- Some older textbooks, and even current catalogues, refer to ammonia solutions as "ammonium hydroxide," and students may believe this substance exists. The compound NH_4OH has never been isolated.
- Many students believe that a neutral solution always has a pH of 7. Emphasize that neutral in this context means an equal number of hydronium and hydroxide ions. Like all equilibrium reactions, the equilibrium constant for the self-ionization of water ($K_w = 1.0 \times 10^{-14}$ at 25 °C) is temperature dependent. The ionization of water increases at higher temperature, so the pH of a neutral solution decreases at higher temperature. For example, the pH of neutral water at body temperature (37 °C) is 6.81, because K_w at that temperature is 2.4×10^{-14}.
- An acid-base reaction is often called a neutralization reaction, and it is natural for students to think that such reactions always result in a solution that is neutral (pH = 7). Avoid the term neutralization reaction and emphasize the concept of equivalence in an acid-base titration. The result is an aqueous solution of a salt, and the pH of the solution will depend on whether any of the ions present hydrolyze.
- Students will often say that the pH range is 0 – 14. Stress that this is the practical range, but that a concentrated solution of a strong acid will have a negative pH, and a concentrated solution of a base acid will have pH greater than 14. Ask students to calculate the pH of 2.0 mol/L HCl(aq) (pH = –0.30) and the pH of 2.0 mol/L NaOH(aq) (pH = 14.30).

Helpful Resources

Books and Journal Articles

- JCE Editorial Staff. "Acid Raindrops Keep Fallin' in My Lake." *Journal of Chemical Education.* Vol. 80, No. 1, January 2003, pages 40A – 40B.
- Weston, Ralph E. Jr. "Climate Change and its Effect on Coral Reefs." *Journal of Chemical Education.* Vol. 77, No. 12, December 2000, pages 1574 – 1577.
- Halstead, Judith A. "Spring Shock!: Impact of Spring Snowmelt on Lakes and Streams." *Journal of Chemical Education.* Vol. 75, No. 4, April 1998, pages 400A – 400B.
- Halstead, Judith A. "Rain, Lakes and Streams Investigating Acidity and Buffering Capacity in the Environment." *Journal of Chemical Education.* Vol. 74, No. 12, December 1997, pages 1456A – 1456B.
- Kousathana, M., Demerouti, M., and Tsaparlis, G. "Instructional Misconceptions in Acid-Base Equilibria: An Analysis From a History and Philosophy of Science Perspective." *Science and Education.* Vol.14, No. 2, 2005, pages 173 – 193.

- Adcock, Jamie L. “Teaching Brønsted-Lowry Acid-Base Theory in a Direct Comprehensive Way.” *Journal of Chemical Education*. Vol. 78, No. 11, November 2001, pages 1495 – 1496.
- Plumsky, Roger. “A pHorseshoe.” *Journal of Chemical Education*. Vol. 76, No. 7, July 1999, pages 935 – 936.
- Silverstein, Todd P. “Weak vs. Strong Acids and Bases: The Football Analogy.” *Journal of Chemical Education*. Vol. 77, No. 7, July 2000, pages 849 – 850.
- Barnum, Dennis. “Predicting Acid-Base Titration Curves without Calculations.” *Journal of Chemical Education*. Vol. 76, No. 7, July 1999, pages 938 – 942.
- Casiday, Richard E. et al. “Blood-Chemistry Tutorials: Teaching Biological Applications of General Chemistry Material.” *Journal of Chemical Education*. Vol. 78, No. 9, September 2001, pages 1210 – 1215.

Web Sites

Web links related to acid-base equilibrium systems can be found at **www.albertachemistry.ca**. Go to the Online Learning Centre, and log on to the Instructor Edition. Choose Teacher Web Links.

List of BLMs

Blackline masters (BLMs) have been prepared to support the material in this chapter. The BLMs are either for assessment (AST); use as overheads (OH); use as handouts (HAND), in particular to support activities; or to supply answers (ANS) for assessment or handouts. The BLMs are in digital form, stored on the CD-ROM that accompanies this Teacher’s Resource or on the web site at **www.albertachemistry.ca**, Online Learning Centre, Instructor Edition, BLMs.

Number (Type) Title

17.0.1 (HAND) Launch Lab: Buffering Ground Water
17.0.1A (ANS) Launch Lab: Buffering Ground Water Answer Key

17.1.1 (OH) Sources of Acid Deposition

17.2.1 (OH) Dissociation of Strong and Weak Acids
17.2.2 (OH) Conjugate Acid-Base Pairs

17.3.1 (OH) Determining K_a and Percent Ionization
17.3.2 (HAND) Investigation 17.A: Determining K_a for Ethanoic Acid
17.3.2A (ANS) Investigation 17.A: Determining K_a for Ethanoic Acid Answer Key
17.3.3 (OH) Calculating K_b for a Weak Base
17.3.4 (OH) Calculating the pH of a Weak Acid
17.3.5 (OH) Calculating the pH of a Weak Base
17.3.6 (AST) pH Calculations
17.3.6A (ANS) pH Calculations Answer Key

17.4.1 (OH) Titration Curves
17.4.2 (AST) Determining the Concentration of an Acid
17.4.2A (ANS) Determining the Concentration of an Acid Answer Key
17.4.3 (HAND) Thought Lab 17.1: Analyzing a Weak Acid-Strong Base Titration
17.4.3A (ANS) Thought Lab 17.1: Analyzing a Weak Acid-Strong Base Titration Answer Key
17.4.4 (AST) Determining the pH of Buffer Solutions
17.4.4A (ANS) Determining the pH of Buffer Solutions Answer Key
17.4.5 (HAND) Investigation 17.B: Preparing a Buffer and Investigating its Properties
17.4.5A (ANS) Investigation 17.B: Preparing a Buffer and Investigating its Properties Answer Key

17.5.1 (AST) Chapter 17 Test
17.5.1A (ANS) Chapter 17 Test Answer Key

Using the Chapter 17 Opener

Student Textbook pages 678–679

Teaching Strategies

- Read the Chapter Opener with your students.
- Ask students to list characteristics of acids and bases they have learned from previous science and chemistry courses.

Launch Lab Buffering Ground Water: A Delicate Balance

Student Textbook page 679

Purpose

In this activity, students model the effect of acid rain on ground water in contact with different soil and rock types using samples of granite or quartz, and marble. They should find that marble neutralizes the effect of acid rain, and then understand why rocks that contain carbonates are a natural buffer against acid deposition.

Outcomes

- 30-D1.2s
- AI-NS6
- ICT C7-4.2

Advance Preparation

When to Begin	What to Do
1 to 2 weeks before	■ Obtain samples of granite or quartz chips, marble chips, and local soil or rock.
2 to 3 days before	■ Assemble the required glassware, clamps, stands and pH paper. Check the pH meter (optional).

When to Begin	What to Do
1 day before	■ Prepare the simulated acid rain solution. ■ Photocopy **BLM 17.0.1 (HAND) Launch Lab: Buffering Ground Water: A Delicate Balance.**

Materials

- 250 mL beakers (5)
- 100 mL graduated cylinder
- 50 mL burette
- retort stand
- burette clamp
- stirring rod
- label or grease pencil
- universal indicator paper (pH paper)
- pH meter (optional)
- granite or quartz chips
- marble chips
- dilute vinegar solution to simulate acid rain
- sample of local soil or rock

Time Required

- 50 minutes

Helpful Tips

- To simulate acid rain, dilute 20 mL of vinegar to 1 L with tap water. The solution pH should be about 3.3.
- If marble chips are not available, dissolve a Tums™ tablet in 200 mL of tap water.
- If granite or quartz chips are not available, use clean sand or distilled water.
- Use **BLM 17.0.1 (HAND) Launch Lab: Buffering Ground Water: A Delicate Balance** to support this activity. Modify it as necessry.
- ***Expected Results:*** Marble or any rock containing limestone will buffer the solution and reduce the extent of the pH change when "acid rain" is added. As the solution sits for a long period of time, more calcium carbonate will dissolve and thus increase the pH of the solution. Granite and quartz will not buffer the solution and thus the pH change will be greater and will not change upon standing.

Safety Precautions

- Ensure that students have read "Safety in Your Chemistry Laboratory and Classroom" on pages xii-xv in the student handbook. Ensure students follow all safety precautions outlined for this investigation in the student textbook.

Answers to Analysis Questions

1. Students should plot their data and use either a line graph or a bar chart (histogram). Line graphs are more widely used to show data in science, but students could argue that a histogram shows comparisons better.
2. The water samples least resistant to change in pH when acid rain was added should be the water that was in contact with granite or quartz. The most resistant water samples should be those in contact with the marble chips. If the ranking differs between student groups, ask the students to suggest possible reasons. The ranking of local water supplies will depend on the source of the water and how it was stored before the experiments.
3. Water left standing in contact with marble chips is likely to increase in pH, and the same will be true if the sample of local soil or rock contains carbonates. The water left standing in contact with granite or quartz is not likely to change pH much.

Assessment Options

- Collect and assess students' answers to Analysis questions.
- Use Assessment Checklist 2 Laboratory Report from Appendix A.

17.1 Acid Deposition

Student Textbook pages 680–683

Section Outcomes

Students will:

- identify several sources of pollutants that cause acid deposition
- write chemical equations describing the formation of acids in acid rain
- apply Le Châtelier's principle to the emission of a pollutant from automobile engines
- distinguish between the ability of different rocks to buffer ground water

Key Terms

acid rain
acid deposition
liming

Chemistry Background

- Unpolluted rainwater would have a pH of about 5.3 because it contains dissolved carbon dioxide, which forms carbonic acid in aqueous solution.
- Acid deposition is the total effect of acid falling in rain, in snow, or as fine solid particles. The pH of acid rain is significantly below 5.3.
- Sulfur dioxide and sulfur trioxide (sometimes collectively referred to as SO_x and called "sox") are formed when coal

and some natural gas supplies that contain compounds of sulfur are burned. In Alberta, the pollutants that cause acid rain come mainly from removing sulfur from natural gas and oil, and from burning coal that contains sulfur compounds. These gaseous sulfur oxides dissolve in water vapour in the atmosphere to form sulfurous acid and sulfuric acid.

- Nitrogen and oxygen combine at the high temperatures generated by operating cars, trucks, and furnaces, to form oxides of nitrogen. These gaseous oxides are sometimes collectively referred to as NO_x and called "nox." When dissolved in moisture in the atmosphere, these oxides form nitrous acid and nitric acid.
- Chalk, limestone, and marble are different forms of calcium carbonate, $CaCO_3(s)$. Calcium carbonate reacts with the acids in acid rain, which reduces the concentration of hydronium ion. Much of Alberta is protected in this way from the effects of acid deposition.
- Quartz and granite do not contain carbonates to neutralize the effects of acid deposition. The Canadian Shield, which covers much of Canada, is made of granite rock and many areas have been affected by acid deposition.
- In the spring, lakes and streams can experience acid shock when acidic snow melts. The surge of acid dissolves aluminium in the soil, which blocks the gills of fish. Waterways that may be affected can be treated with limestone.

Teaching Strategies

- There are many newspaper, magazine, and Internet articles dealing with acid deposition. Groups of students could be assigned to each of these types of sources to collect and organize the information they find.
- You could begin this chapter by asking some apparently simple questions. For example, should coal and natural gas burning power plants be restricted to using low sulfur sources of fuel? Explain that burning fossil fuels that contain sulfur is a major source of the sulfur dioxide that dissolves in atmospheric water vapour to form sulfuric acid. The answer to the question posed is very complex. Does Canada have good supplies of low-sulfur fuels? If so, where are they located? How do transportation costs compare with the sources presently used? Do low sulfur fossil fuels have the same heat content? If not, more fuel will need to be burned, contributing to greater emissions of carbon dioxide. Do low sulfur fuels cost more? Could more be done with the present sources of fossil fuels to remove the sulfur before the fuel is burned, or remove the $SO_2(g)$ before it is released to the atmosphere? What are the costs involved in retrofitting plants to remove sulfur dioxide?
- **BLM 17.1.1 (OH) Sources of Acid Deposition** has been prepared for this section. You will find it with the Chapter 17 BLMs on the CD-ROM that accompanies this Teacher's Resource or at **www.albertachemistry.ca**, Online Learning Centre, Instructor Edition, BLMs.

Chemistry File: Web Link

Student Textbook page 680

Assess students' research, organization, and clarity of the argument in their letters. Ensure that students have cited sources properly.

Answers to Questions for Comprehension

Student Textbook page 682

Q1. Rocks that contain carbonates, such as limestone, are the most effective at neutralizing the effects of acid deposition.

Q2. The only way to completely eliminate the effects of acid rain is to stop putting SO_x and NO_x into the atmosphere. We would have to stop burning oil and gas in factories, power plants and home furnaces. Automobiles and trucks would be at a standstill, and metal refineries would close. We cannot go back to a time when these activities were unknown. The aid should instead point out responsible ways to reduce the emissions that cause acid deposition: using low sulfur content fuels, saving energy at home and in factories, promoting public transport and increased fuel efficiency for cars and trucks. There might be tax incentives to stimulate change, especially for industries that reduce their emissions of harmful gases. The technology already exists to reduce the emissions that cause acid deposition. The real question is, are we willing to pay for this technology? Also, since the problem is global, are we willing to help developing countries with this technology?

Section 17.1 Review Answers

Student Textbook page 683

1. Major sources of emissions that cause acid deposition are fossil-fuel combustion (releasing mainly $SO_2(g)$, but also $SO_3(g)$), from coal-fired power plants, but also from the combustion of oil or gas in furnaces. Refining metal ores also releases large quantities of $SO_2(g)$ and $SO_3(g)$. In Alberta, the major sources of sulfur dioxide emissions are the oil sands projects and power plants. Gas plant flares, oil refineries, pulp and paper mills and fertilizer plants also contribute to the problem. Transportation and power plants are the major sources of $NO(g)$ and $NO_2(g)$. Natural sources of SO_x include releases from the oceans, biological decay, forest fires, and volcanoes. The Kilauea volcano in Hawaii is one of the most active in the world, and acid rain from the volcano has resulted in a large desert area near the crater.
2. Students should identify parts of Canada with large capacities for electric power generators using fossil fuels, and areas with mining and smelting activities. These are principally located in Ontario and Quebec. Emissions that impact western Canada are located in Alberta, southern Saskatchewan and northern Manitoba.

3. A lake or stream may have a lower pH in the spring than at any other time of the year if there is a relatively sudden melting of acidic snow that feeds into the waterway.
4. Carbonic acid is a **much** weaker acid than the other acids of concern.
5. Acid rain affects stone buildings and bridges, marble building facades, concrete and mortar, because these materials contain carbonates that react with acid rain.
6. These strategies should focus on ways to consume less energy from fossil fuel sources. Strategies could include walking, cycling, or using public transport rather than an automobile; turning the thermostat down at night and consuming less energy; insulating doors, windows and attics.

17.2 Understanding Acids and Bases

Student Textbook pages 684–690

Section Outcomes

Students will:

- define acid-base reactions as proton-transfer reactions
- write Brønsted-Lowry equations for acid-base reactions and identify conjugate acid-base pairs
- identify amphiprotic compounds and write Brønsted-Lowry equations to explain their acid-base reactions
- recall the names and formulas of some polyprotic acids and polyprotic bases
- distinguish between strong acids and bases, and weak acids and bases

Key Terms

Brønsted-Lowry theory of acids and bases
conjugate acid-base pair
conjugate base
conjugate acid
amphiprotic
polyprotic acid
polyprotic base
strong acid
weak acid
strong base
weak base

Chemistry Background

- The Brønsted-Lowry theory recognizes acid-base reactions as a chemical equilibrium having both a forward and a reverse reaction that involve the transfer of a proton.
- In the Brønsted-Lowry theory, an acid is a substance from which a proton can be removed. A base is a substance that can accept a proton.
- Two molecules or ions that differ because of the transfer of a proton are called a conjugate acid-base pair.
- An amphiprotic substance is one that can act as a Brønsted-Lowry acid in one reaction and as a Brønsted-Lowry base in a different reaction.
- A polyprotic acid has more than one proton that may ionize. A polyprotic base is a substance capable of accepting more than one proton.
- An acid that ionizes 100% in water is termed a strong acid. A weak acid ionizes only slightly in water. Similarly, a strong base ionizes completely in water while a weak base ionizes only slightly in water.
- An acid-base reaction usually proceeds to the right if the stronger acid and stronger base are on the left side of the equation and products are favoured.

Teaching Strategies

- The development of acid-base theories is a very good illustration of the way scientific theories develop through hypotheses, the collection of evidence through experimentation, and the ability to provide explanations. This section could begin with a brief review of earlier acid-base theories (Lavoisier and Arrhenius) and a discussion of the limitations of the modified Arrhenius theory. Although the Brønsted-Lowry theory is more successful than the modified Arrhenius theory, it too has been replaced. Students may like to know that if they continue their studies of chemistry, they will learn an even more general theory about acids and bases, the Lewis theory.
- It may be helpful to review nomenclature and net ionic equations before starting this section.
- Some texts use the term “dissociation” and “dissociation constant” when referring to acids and bases. Because the focus in this chapter is on the formation of hydronium ions and hydroxide ions, this the student textbook uses the terms “ionization” and “ionization constant.” It could be confusing to students if the terms “dissociation” and “ionization” were used interchangeably.
- Use models of an acid and a water molecule to show how a proton is transferred from the acid to form a hydronium ion and the conjugate base.
- Ensure that students are absolutely clear on the Brønsted-Lowry theory, and especially that they understand the concept of conjugate acid-base pairs. These are key ideas that are applied later in the chapter.
- **BLM 17.2.1 (OH) Dissociation of Strong and Weak Acids** and **BLM 17.2.2 (OH) Conjugate Acid-Base Pairs** have been prepared for this section. You will find them with the Chapter 17 BLMs on the CD-ROM that accompanies this Teacher’s Resource or at **www.albertachemistry.ca**, Online Learning Centre, Instructor Edition, BLMs.

SUPPORTING DIVERSE STUDENT NEEDS

- Students with a special interest in chemistry can research the Lewis theory of acids and bases and answer the following questions: What are the Lewis definitions of acids and bases? How is the Lewis theory an improvement over the Brønsted-Lowry theory? Students should give examples of Lewis acids and bases that would not be considered Brønsted-Lowry acids and bases. Finally, they should be able to write a chemical equation illustrating a Lewis acid-base reaction.

- Students could prepare a graphic organizer to help develop their own definitions of terms related to acid/base chemistry and to see the relationships between concepts.

Answers to Questions for Comprehension

Student Textbook page 686

Q3. The Arrhenius theory is restricted because it applies only to acid-base reactions in aqueous solution. Also, the Arrhenius theory recognizes only one kind of base, the hydroxide ion. The Brønsted-Lowry theory is more general because it removes each of these restrictions.

Q4. Look for answers that clearly show a substance capable of accepting a proton, but which do not form the hydroxide ion in solution. Aqueous ammonia, $NH_3(aq)$, is one frequent answer, but any amine might be chosen. Other good answers include anions such as $HCO_3^-(aq)$ and $HSO_4^-(aq)$.

Q5. A Brønsted-Lowry acid must contain H^+, because it is defined as a proton donor. However, a Brønsted-Lowry base does not have to contain OH^-, because it is defined as a proton acceptor.

Answers to Practice Problems 1–4

Student Textbook page 687

For full solutions to the practice problems, visit **www.albertachemistry.ca,** Online Learning Centre, Instructor Edition, Full Solutions.

1. **(a)** fluoride ion, $F^-(aq)$
  (b) carbonate ion, $CO_3^{2-}(aq)$
  (c) hydrogen sulfate ion, $HSO_4^-(aq)$
  (d) hydrazine, $N_2H_4(aq)$

2. **(a)** nitric acid, $HNO_3(aq)$
  (b) water, $H_2O(\ell)$
  (c) hydronium ion, $H_3O^+(aq)$
  (d) carbonic acid, $H_2CO_3(aq)$

3. $HClO_4(aq)/ClO_4^-(aq)$ and $H_2O(\ell)/H_3O^+(aq)$.

4. **(a)** $H_2O(\ell)/OH^-(aq)$ and $HS^-(aq)/H_2S(aq)$.
  (b) $H_2O(\ell)/OH^-(aq)$ and $O^{2-}(aq)/OH^-(aq)$.
  (c) $H_2S(aq)/HS^-(aq)$ and $NH_3(aq)/NH_4^+(aq)$
  (d) $H_2SO_4(aq)/HSO_4^-(aq)$ and $H_2O(\ell)/H_3O^+(aq)$

Chemistry File: FYI

Student Textbook page 688

Students will likely think of "amphibious," meaning able to live both on land and in water. An interesting chemistry word that applies here is "amphichroic," meaning a substance that is one colour with an acid and another colour with a base. Chemical indicators such as litmus are amphichroic.

Answers to Practice Problems 5–7

Student Textbook page 688

For full solutions to the practice problems, visit **www.albertachemistry.ca,** Online Learning Centre, Instructor Edition, Full Solutions.

5. **(a)** $H_2S(aq) + H_2O(\ell) \rightleftharpoons HS^-(aq) + H_3O^+(aq)$
  $HS^-(aq) + H_2O(\ell) \rightleftharpoons S^{2-}(aq) + H_3O^+(aq)$
  (b) $HS^-(aq) + H_2O(\ell) \rightleftharpoons H_2S(aq) + OH^-(aq)$

6. **(a)** $HCO_3^-(aq) + OH^-(aq) \rightleftharpoons CO_3^{2-}(aq) + H_2O(\ell)$
  (b) $HCO_3^-(aq) + HF(aq) \rightleftharpoons H_2CO_3(aq) + F^-(aq)$

7. Acid: $H_2PO_4^-(aq) + H_2O(\ell) \rightleftharpoons HPO_4^{2-}(aq) + H_3O^+(aq)$
  Base: $H_2PO_4^-(aq) + H_2O(\ell) \rightleftharpoons H_3PO_4\ (aq) + OH^-(aq)$

Answers to Practice Problems 8–10

Student Textbook page 690

For full solutions to the practice problems, visit **www.albertachemistry.ca,** Online Learning Centre, Instructor Edition, Full Solutions.

8. **(a)** left, reactants are favoured
  (b) left, and reactants are favoured
  (c) left, reactants are favoured

9. **(a)** left, reactants are favoured
  (b) right, products are favoured
  (c) right, products are favoured

10. **(a)** product favoured
  $H_2CO_3(aq) + NH_3(aq) \rightleftharpoons NH_4^+(aq) + HCO_3^-(aq)$
  (b) reactant favoured
  $Na^+(aq) + HSO_3^-(aq) + H_2O(\ell) \rightleftharpoons Na^+(aq) + SO_3^{2-}(aq) + H_3O^+(aq)$
  (c) reactants favoured
  $K^+(aq) + NO_3^-(aq) + HF(aq) \rightleftharpoons K^+(aq) + HNO_3(aq) + F^-(aq)$
  (d) reactants favoured
  $H_3PO_4(aq) + H_2O(\ell) \rightleftharpoons H_2PO_4^-(aq) + H_3O^+(aq)$
  $H_2PO_4^-(aq) + H_2O(\ell) \rightleftharpoons HPO_4^{2-}(aq) + H_3O^+(aq)$
  $HPO4^{2-}(aq) + H_2O(\ell) \rightleftharpoons PO_4^{3-}(aq) + H_3O^+(aq)$

Section 17.2 Review Answers

Student Textbook page 690

1. The conjugate acid-base pairs are $CH_3COOH(aq)/CH_3COO^-(aq)$ and $H_2O(\ell)/OH^-(aq)$
2. The acids in this reaction are $H_2PO_4^-(aq)$ and $NH_4^+(aq)$. Looking at Table 17.1 on page 689, the stronger acid is $H_2PO_4^-(aq)$. Therefore, the reaction proceeds to the right as written, and products are favoured.
3. $H_3BO_3(aq) + H_2O(\ell) \rightleftharpoons H_3O^+(aq) + H_2BO_3^-(aq)$
   $H_2BO_3^-(aq) + H_2O(\ell) \rightleftharpoons H_3O^+(aq) + HBO_3^{2-}(aq)$
   $HBO_3^{2-}(aq) + H_2O(\ell) \rightleftharpoons H_3O^+(aq) + BO_3^{3-}(aq)$
4. The conjugate acid-base pairs are:
   (a) $H_2PO_4^-(aq)/HPO_4^{2-}(aq)$ and $HCO_3^-(aq)/CO_3^{2-}(aq)$
   (b) $HCOOH(aq)/HCOO^-(aq)$ and $HCN(aq)/CN^-(aq)$
   (c) $H_2PO_4^-(aq)/HPO_4^{2-}(aq)$ and $H_2O(\ell)/OH^-(aq)$

17.3 Acid-Base Equilibriums

Student Textbook pages 691–704

Section Outcomes

Students will:

- calculate acid ionization constants, K_a, and base ionization constants, K_b, given the concentration of a weak acid or base and the pH of the solution
- solve acid-base equilibrium problems using an approximation method
- measure an acid ionization constant by preparing a solution of a weak acid and measuring the pH of the solution
- calculate the ionization constant for a conjugate base or a conjugate acid, given K_a or K_b, respectively
- predict whether products or reactants are favoured in an acid-base reaction

Key Terms

acid ionization constant, K_a
percent ionization
base ionization constant, K_b
ion product constant for water, K_w

Chemistry Background

- The general equilibrium reaction of a weak monoprotic acid in aqueous solution can be written as:
  $HA(aq) + H_2O(\ell) \rightleftharpoons H_3O^+(aq) + A^-(aq)$
  For this reaction, the acid ionization constant, K_a, is written as

$$K_a = \frac{[H_3O^+(aq)][A^-(aq)]}{[HA(aq)]}$$

- Equilibrium constants are correctly written in terms of activities. The activity of water is assumed to be 1, and the activities of $H_3O^+(aq)$, $A^-(aq)$, and $HA(aq)$ are approximated by their molar concentrations. These assumptions are valid in dilute solutions, leading to the expression for K_a.
- The percent ionization of a weak acid is the fraction of molecules that ionize compared with the initial concentration of the acid, expressed as a percentage.
- The general equilibrium of a weak base in aqueous solution can be written as:
  $B(aq) + H_2O(\ell) \rightleftharpoons HB^+(aq) + OH^-(aq)$
  The equilibrium constant for this general reaction is given the symbol K_b.

$$K_b = \frac{[HB^+(aq)][OH^-(aq)]}{[B(aq)]}$$

- The ionization of water can be written as
  $2H_2O(\ell) \rightleftharpoons H_3O+ (aq) + OH^- (aq)$
  The equilibrium constant is given the symbol K_w.
  $K_w = [H_3O^+ (aq)][OH^- (aq)]$.
  K_w is called the ion product constant for water. At 25 °C, $K_w = 1.0 \times 10^{-14}$.
- pH + pOH = 14.0 (at 25 °C).
- For a conjugate acid-base pair, $K_w = K_a \times K_b$. Consequently, there is an inverse relationship between the strength of an acid and the strength of its conjugate base.

Teaching Strategies

- Since calculators differ, spend time working through the steps necessary to use logs and exponentials, and check that students use the correct functions.
- As mentioned in the student textbook, there is really nothing new involved in solving the problems in this section. Emphasize that K_a and K_b are expressions with the same format as for the equilibrium constant K_c. Also, the approximation used when K_a or K_b is small when compared with the initial concentration of acid or base is the same as they learned in the previous chapter.
- A number of overhead masters and a quiz have been prepared for this section. You will find them with the Chapter 17 BLMs on the CD-ROM that accompanies this Teacher's Resource or at **www.albertachemistry.ca**, Online Learning Centre, Instructor Edition, BLMs.

Number (Type) Title
17.3.1 (OH) Determining K_a and Percent Ionization
17.3.3 (OH) Calculating K_b for a Weak Base
17.3.4 (OH) Calculating the pH of a Weak Acid
17.3.5 (OH) Calculating the pH of a Weak Base
17.3.6 (AST) pH Calculations
17.3.6A (ANS) pH Calculations Answer Key

SUPPORTING DIVERSE STUDENT NEEDS

- Advanced students who used the quadratic equation for solving problems where the approximation is invalid may wish to continue to use these skills with acid-base problems.
- As an enrichment exercise, students could research and complete a project detailing how the pH of soaps, shampoos and skin products affects hair and skin care.

Answers to Practice Problems 11–13

Student Textbook page 692

For full solutions to the practice problems, visit **www.albertachemistry.ca,** Online Learning Centre, Instructor Edition, Full Solutions.

11.

(a) $K_a = \frac{[H_3O^+][NO_2^-]}{[HNO_2]} = 5.6 \times 10^{-4}$

(b) $K_a = \frac{[H_3O^+][C_6H_5COO^-]}{[C_6H_5COOH]} = 6.3 \times 10^{-5}$

(c) $K_a = \frac{[H_3O^+][H_2C_6H_5O_7^-]}{[H_3C_6H_5O_7]} = 7.4 \times 10^{-4}$

(d) $K_a = \frac{[H_3O^+][C_6H_5O_7^{3-}]}{[HC_6H_5O_7^{2-}]} = 4.0 \times 10^{-7}$

12. $K_{a2} = \frac{[HCO_3^-][H_3O^+]}{[H_2CO_3]}$ $K_{a1} = \frac{[CO_3^{2-}][H_3O^+]}{[HCO_3^-]}$

13. $C_6H_6O(aq) + H_2O(\ell) \rightleftharpoons H_3O^+(aq) + C_6H_5O^-(aq)$

Answers to Practice Problems 14–18

Student textbook page 693

For full solutions to the practice problems, visit **www.albertachemistry.ca,** Online Learning Centre, Instructor Edition, Full Solutions.

14. $K_a = 1.0 \times 10^{-4}$

percent ionization = 3.2%

15. $K_a = 1.5 \times 10^{-5}$, 3.9%

16. $K_a = 6.1 \times 10^{-10}$

17. $K_a = 1.2 \times 10^{-5}$

18. 13%

Investigation 17.A: Determining K_a for Ethanoic Acid

Student Textbook pages 694–695

Purpose

The purpose of this activity is to determine the acid ionization constant for ethanoic acid. To obtain the data for their calculation, students will have to measure the pH, and determine the concentration of a solution of ethanoic acid that they will be given.

Outcomes

- 30-D2.3s

Advance Preparation

When to Begin	What to Do
3 to 4 weeks before	■ Order materials.
1 to 2 days before	■ Prepare Solutions of $CH_3COOH(aq)$ and NaOH(aq).
1 day before	■ See Appendix H in the student text Review titration procedures. ■ Photocopy **BLM 17.3.2 (HAND) Investigation 17.A: Determining K_a for Ethanoic Acid.**

Materials

- unknown concentration $CH_3COOH(aq)$
- 0.75 mol/L NaOH(aq)
- dropper bottle containing phenolphthalein
- distilled water
- 10 mL pipette
- pipette bulb or pipette pump
- burette and burette clamp
- retort stand
- meniscus reader
- funnel
- pH meter or pH paper
- 100 mL beakers (2)
- 250 mL beaker for waste solutions
- 150 mL Erlenmeyer flask
- sheet of white paper
- labels

Time Required

- 60 to 90 minutes

Helpful Tips

- Demonstrate a titration to the class before they begin their investigation. Refer students to Appendix H for a review of the processes.
- Prepare a stock solution of 1.0 mol/L ethanoic acid, then dilute a portion of this to obtain “unknown” solutions with concentrations of 1.0 mol/L, 0.80 mol/L, 0.60 mol/L, and 0.50 mol/L.
- Take a moment to calculate the total volume of each solution you will need for your class(es).
- Sodium hydroxide is not a primary standard, but this investigation does not justify making up solutions with precise concentration. Make up 0.75 mol/L NaOH(aq) the day before it will be used, and seal the container to prevent reaction with $CO_2(g)$ from the air.
- Preparation of stock solutions:
    - **Prepare ALL ethanoic acid solutions in a fume hood. Wear gloves, gown, and safety goggles.**
    - 1.0 mol/L $CH_3COOH(aq)$. Add about 500 mL of distilled water to a 1-L volumetric flask. Measure 57.2 mL of glacial ethanoic acid (17.5 mol/L) and add slowly to the volumetric flask. Make up to the mark using distilled water.
    - 0.75 mol/L NaOH(aq). Dissolve 30 g NaOH(s) pellets in about 300 mL of distilled water. Transfer the solution to a 1-L volumetric containing about 200 mL of distilled water, then make up to the mark using distilled water. Transfer to a container that can be sealed until it is used.
    - Preparation of diluted ethanoic acid solutions. Each of the required solutions can be prepared by diluting the 1.0 mol/L solution of $CH_3COOH(aq)$. Add the quantity from the table below to about 100 mL of distilled water in a volumetric flask, then make up to the mark using distilled water.

To prepare 1.0 L of this concentration	Measure this volume of 1.0 mol/L $CH_3COOH(aq)$.
0.80 mol/L	800 mL
0.60 mol/L	600 mL
0.50 mol/L	500 mL

- Use **BLM 17.3.2 (HAND) Investigation 17.A: Determing K_a for Ethanoic Acid** to support this activity. Remove sections as necessary to meet the needs of students in your class.
- ***Expected Results:*** The K_a for ethanoic acid is 1.8×10^{-5}. The students should be able to carry out the experiment with accuracy and their results should be very close to this expected value.

Safety Precautions

- Provide containers for chemical disposal so that students do not pour the solutions down the drain.
- Ensure students have read “Safety in Your Chemistry Laboratory and Classroom” on pages xii-xv in the student handbook. Ensure students follow all safety precautions outlined for this investigation in the student textbook.
- Sodium hydroxide is toxic and is harmful if swallowed or inhaled. Both ethanoic acid and sodium hydroxide are corrosive. Immediately wash any spills on skin or clothing with plenty of cool water.
- Spills on the floor or lab bench should be handled using the appropriate chemical spill kit.
- Phenolphthalein solution may irritate skin, eyes, and mucous membranes. This solution is flammable. Keep away from open flames.
- Remind students to wash their hands when the investigation has been completed.

Answer to Prediction Question

Students should know that ethanoic acid is a weak acid. Accept any prediction that indicates this, and a small percent ionization.

Answers to Analysis Questions

1. $CH_3COOH(aq) + NaOH(aq) \rightleftharpoons CH_3COONa(aq) + H_2O(\ell)$

2. The molar concentration of the ethanoic acid is given by

$$\frac{\text{Volume NaOH(aq)} \times \text{Concentration NaOH(aq) mol/L}}{\text{Volume } CH_3COOH(aq)}$$

3. $[H_3O^+(aq)] = 10^{-pH}$. Measurements of the pH should be in the range 2.4–2.5. Therefore, the $[H_3O^+(aq)]$ should be between 4.0×10^{-3} and 3.0×10^{-3} mol/L.

4. $K_a = \dfrac{[H_3O^+(aq)][CH_3COO^-(aq)]}{[CH_3COOH(aq)]}$

5. Using a pH reading of 2.42, a typical ICE table for this experiment is

$$CH_3COOH(aq) + H_2O(\ell) \rightleftharpoons CH_3COO^-(aq) + H_3O^+(aq)$$

	$[CH_3COOH(aq)]$ (mol/L)	$[H_2O(\ell)]$ (mol/L)	$[CH_3COO^-(aq)]$ (mol/L)	$[H_3O^+(aq)]$ (mol/L)
Initial	0.80		0	~0
Change	-3.8×10^{-3}		$+3.8 \times 10^{-3}$	$+3.8 \times 10^{-3}$
Equilibrium	~0.80		3.8×10^{-3}	3.8×10^{-3}

$$K_a = \frac{[H_3O^+(aq)][CH_3COO^-(aq)]}{[CH_3COOH(aq)]} = \frac{(3.8 \times 10^{-3}) \times (3.8 \times 10^{-3})}{0.80}$$

$$K_a = 1.8 \times 10^{-5}$$

$$\text{Percent ionization} = \frac{3.8 \times 10^{-3}}{0.80} \times 100 = 4.8\%$$

Answers to Conclusion Questions

6. $\text{Percent difference} = \frac{(\text{Measured } K_a - 1.8 \times 10^{-5})}{1.8 \times 10^{-5}} \times 100$

Sources of error include error in the initial concentration of solution; measurement of the pH; measurement of the volume of base added to reach equivalence. Temperature might be mentioned, but the effect on the pH of the solution is negligible. Students sometimes mention "errors" in calculation but these, if present, are mistakes that can be avoided. Insist on reasonable measurement errors.

7. If there is an order of magnitude difference, students should look for a calculation error. Otherwise, they should identify whether the experimental value is larger, or smaller, than the accepted value. Then they should identify specific experimental errors that could account for the difference.

Answer to Application Question

8. The concentration of $H_3O^+(aq)$ should be much smaller than the concentration of $CH_3COOH(aq)$. Therefore, students should conclude that ethanoic acid is a weak acid.

Assessment Options

- Collect and assess students' tables, calculations and answers to Prediction, Analysis, Conclusion, and Application questions.
- Use Assessment Checklist 2 Laboratory Reports from Appendix A.

Answers to Questions for Comprehension

Student Textbook page 696

Q6. The ion product, K_w, is an equilibrium constant. The equilibrium constant is affected only by changes in temperature. The ions in water are in equilibrium. When an acid is added, the concentration of hydronium ion, $H_3O^+(aq)$, is increased. This will cause the equilibrium to shift, decreasing the concentration of the ions, and maintaining K_w.

Q7. The ionization of water at 25 °C results in $[H_3O^+(aq)] = [OH^-(aq)] = 1.0 \times 10^{-7}$ mol/L. These concentrations are not significant unless a very dilute solution of an acid or base is present. The concentrations of hydronium ion and hydroxide ion due to the ionization of water are not important in any problem in this text.

Q8. Diluting the acid will result in the position of equilibrium being shifted to the right. Thus, as the acid is diluted, it becomes a stronger acid.

Answers to Practice Problems 19–23

Student Textbook page 698

For full solutions to the practice problems, visit **www.albertachemistry.ca,** Online Learning Centre, Instructor Edition, Full Solutions.

19. (a) $NH_3(aq) + H_2O(\ell) \rightleftharpoons NH_4^+(aq) + OH^-(aq)$

(b) $(CH_3)_3N(aq) + H_2O(\ell) \rightleftharpoons (CH_3)_3NH^+(aq) + OH^-(aq)$

(c) $HSO_3^-(aq) + H_2O(\ell) \rightleftharpoons H_2SO_3(aq) + OH^-(aq)$

(d) $CO_3^{2-}(aq) + H_2O(\ell) \rightleftharpoons HCO_3^-(aq) + OH^-(aq)$

20. (a) $K_b = \frac{[NH_4^+][OH^-]}{[NH_3]}$

(b) $K_b = \frac{[(CH_3)_3NH^+][OH^-]}{[(CH_3)_3N]}$

(c) $K_b = \frac{[H_2SO_3][OH^-]}{[HSO_3^-]}$

(d) $K_b = \frac{[HCO_3^-][OH^-]}{[CO_3^{2-}]}$

21. Percent ionization 0.044 %

22. $K_b = 1.7 \times 10^{-6}$

23. $K_b = 1.4 \times 10^{-9}$

Answers to Practice Problems 24–27

Student Textbook page 699

For full solutions to the practice problems, visit **www.albertachemistry.ca,** Online Learning Centre, Instructor Edition, Full Solutions.

24. 0.47%, pH = 2.41

25. 2.5%, pH = 1.60

26. $K_a = 2.9 \times 10^{-8}$

27. 1.01×10^{-2} g

Answers to Practice Problems 28–33

Student Textbook page 701

For full solutions to the practice problems, visit **www.albertachemistry.ca,** Online Learning Centre, Instructor Edition, Full Solutions.

28. pH 11.14

29. pH 10.04

30. $K_b = 1.6 \times 10^{-6}$

31. $[OH^-(aq)] = 2.6 \times 10^{-2}$ mol/L; pOH = 1.58

32. $[OH^-(aq)] = 3.7 \times 10^{-3}$ mol/L, 1.7%

33. $[NH_3(aq)] = 2.8 \times 10^{-2}$ mol/L

Answers to Practice Problems 34–38

Student Textbook page 703

For full solutions to the practice problems, visit **www.albertachemistry.ca,** Online Learning Centre, Instructor Edition, Full Solutions.

34. (a) $CN^-(aq)$, $K_b = 1.6 \times 10^{-5}$

(b) $NO_2^-(aq)$, $K_b = 1.8 \times 10^{-11}$

35. The conjugate acid of CH_3NH_2 is $CH_3NH_3^+$ (aq) for which $K_a = 2.3 \times 10^{-11}$.

The conjugate acid of $C_5H_5N(aq)$ is $C_5H_5NH^+(aq)$ for which $K_a = 7.1 \times 10^{-6}$

36. $CN^-(aq)$ is the stronger base.

37. Use $K_{a2} \bullet K_b$ for $S^{2-}(aq) = 1.0 \times 10^5$

38. $NH_4^+(aq)$ is the weaker acid. The reaction proceeds to the right and products are favoured.

Section 17.3 Review Answers

Student Textbook page 704

1. (a) as an acid: $HPO_4^{2-}(aq) + H_2O(\ell) \rightleftharpoons H_3O^+(aq) + PO_4^{3-}(aq)$

as a base: $HPO_4^{2-}(aq) + H_2O(\ell) \rightleftharpoons OH^-(aq) + H_2PO_4^-(aq)$

(b) $$K_a = \frac{[H_3O^+(aq)][PO_4^{3-}(aq)]}{[HPO_4^{2-}(aq)]}$$

(c) $$K_b = \frac{[OH^-(aq)][H_2PO_4^-(aq)]}{[HPO_4^{2-}(aq)]}$$

(d) The solution is basic because K_b is greater than K_a.

2. (a) The stronger Brønsted-Lowry acid is $H_3O^+(aq)$. Polar water molecules will be more strongly attracted to positive hydronium ions than to neutral carbonic acid molecules.

(b) The stronger Brønsted-Lowry acid is $HSO_4^-(aq)$. The four oxygen atoms pull electrons away from the hydrogen atom that will ionize, making the bond polar. Sulfur is a fairly electronegative atom, and this increases the partial charge on the hydrogen atom in $HSO_4^-(aq)$.

(c) $H_2AsO_4^-(aq)$ is the stronger Brønsted-Lowry acid. It will require more energy to remove a proton from $HAsO_4^{2-}(aq)$ because it has a double negative charge.

3. (a) The stronger Brønsted-Lowry base is $SO_3^{2-}(aq)$. Polar water molecules will be more strongly attracted to negative sulfite ions.

(b) The stronger Brønsted-Lowry base is $F^-(aq)$. Although the ions have the same charge, the fluoride ion has a smaller radius. The charge density on $F^-(aq)$ is greater, and this ion is able to remove a proton from water more easily than $Cl^-(aq)$.

(c) The stronger Brønsted-Lowry base is $O_2^{2-}(aq)$. The oxide ion is smaller and has a larger negative charge than the hydroxide ion.

4. (a) The conjugate base is $C_6H_7O_6^-(aq)$.

$$K_b = \frac{K_w}{K_a} = \frac{1.0 \times 10^{-14}}{9.1 \times 10^{-5}} = 1.1 \times 10^{-10}$$

(b) The conjugate base is $CH_3COO^-(aq)$.

$$K_b = \frac{K_w}{K_a} = \frac{1.0 \times 10^{-14}}{1.8 \times 10^{-5}} = 5.6 \times 10^{-10}$$

(c) The conjugate base is $HCOO^-(aq)$.

$$K_b = \frac{K_w}{K_a} = \frac{1.0 \times 10^{-14}}{1.8 \times 10^{-4}} = 5.6 \times 10^{-11}$$

5. $CH_3NH_3^+(aq)$ is the stronger acid because $CH_3NH_2(aq)$ is the weaker base.

6. (a) Reactants are favoured because $H_3O^+(aq)$ is a stronger acid than $NH_4^+(aq)$.

(b) Reactants are favoured because $H_2SO_4(aq)$ is a stronger acid than $H_3O^+(aq)$.

(c) Reactants are favoured because $H_2S(aq)$ is a stronger acid than $HSO_3^-(aq)$.

(d) Products are favoured because HF(aq) is a stronger acid than $HCO_3^-(aq)$.

17.4 Titration Curves and Buffers

Student Textbook pages 705–718

Section Outcomes

Students will:

- interpret various titration curves and distinguish between the endpoint and the equivalence point in the titration
- identify buffers in living systems and describe them using the Brønsted-Lowry theory
- investigate a buffer solution

Key Terms

titration curve
buffer solution
buffer capacity

Chemistry Background

- In a titration, the equivalence point occurs when the stoichiometric amount of acid is equal to the stoichiometric amount of base. The endpoint of the titration occurs when the indicator changes colour.
- The reaction of a chemical compound with water is called hydrolysis. When an acid or base dissolves in water, both ionization *and* hydrolysis take place.
- The conjugate ion formed by a strong acid or by a strong base is very weak. Consequently, these conjugate ions do not hydrolyze and have no effect on the pH of water.
- The conjugate base of a weak acid hydrolyzes to make an aqueous solution basic. The conjugate acid of a weak base hydrolyzes to make an aqueous solution acidic.
- When a strong acid is titrated with a strong base, the pH at equivalence is 7. When titrating a strong acid with a weak base, the pH at equivalence is less than 7. When a weak acid is titrated with a strong base, the pH at equivalence is greater than 7.
- A graph of the pH of the titration mixture against the volume of added acid (or base) is called a titration curve. The equivalence point is taken to be halfway along the portion of the titration curve where the pH changes rapidly.
- An indicator changes colour over a range of about 2 pH units. In a titration experiment, the range over which an indicator changes colour must include the equivalence point of the titration.
- The titration curve obtained by titrating a weak diprotic acid with a strong base has two equivalence points, one for each proton that ionizes from the acid.
- A solution that contains a weak acid–conjugate base mixture, or a weak base–conjugate acid mixture, is called a buffer solution.
- The buffer capacity is generally taken to be ± 1 pH unit from the buffer pH.
- There are three types of buffers present in blood: the carbonic acid system, phosphate buffers, and proteins. These buffers must keep blood pH at 7.4 ± 0.4.
- A decrease in blood pH is called acidosis. An increase in blood pH is called alkalosis.

Teaching Strategies

- The original meaning of "buffer" is a cushion that reduces the shock from the contact between two items. Students will be familiar with the term "buffer" in other contexts. Discuss some of these, such as a green-belt buffer zone, or in a computer system referring to the area where data is stored. Use these more familiar ideas to develop an analogy with the action of chemical buffers.
- Many students will understand the concept of buffers using Le Châtelier's principle and the shifts in equilibrium resulting from the addition of acid or base to a buffer solution.
- The regulation of blood pH is covered in detail at a number of web sites. The following offer a good overview:
  http://members.aol.com/Bio50/LecNotes/lecnot39.html
  http://scifun.chem.wisc.edu/chemweek/BioBuff/BioBuffers.html
- The next site deals with pH regulation during exercise. It is longer than the previous two sites but very good.
  http://www.chemistry.wustl.edu/~edudev/LabTutorials/Buffer/Buffer.html
- The next site also deals with blood pH and buffer systems but has a short quiz for which students can easily obtain answers.
  http://www.chemistry.wustl.edu/~edudev/LabTutorials/Buffer/Buffer.html
- A more advanced site that introduces the Henderson-Hasselbalch equation used to calculate the pH of a buffer may be found at:
  http://www.lsbu.ac.uk/biology/biolchem/acids.html
- An overhead master and two quizzes have been prepared for this section. You will find them with the Chapter 17 BLMs on the CD-ROM that accompanies this Teacher's Resource or at **www.albertachemistry.ca**, Online Learning Centre, Instructor Edition, BLMs.

  Number (Type) Title
  17.4.1 (OH) Titration Curves
  17.4.2 (AST) Determining the Concentration of an Acid
  17.4.2A (ANS) Determining the Concentration of an Acid Answer Key
  17.4.4 (AST) Determining the pH of Buffer Solutions
  17.4.4A (ANS) Determining the pH of Buffer Solutions Answer Key

SUPPORTING DIVERSE STUDENT NEEDS

Encourage students who want to learn more about buffers to investigate antacids and the pH balance within the stomach. Students can research how the stomach can maintain a low pH without damaging surrounding cells. They can further their investigation by researching the ingredients of popular antacids.

Thought Lab 17.1: Analyzing a Weak Acid-Strong Base Titration

Student textbook page 709

Purpose

This activity is intended to give students practice in using a graphing program and interpreting the results of a weak acid-strong base titration. Students generally find performing this experiment tedious and time-consuming. Providing data allows greater emphasis on analyzing and interpreting the results.

Outcomes

- 30-D1.9k
- 30-D1.4s

Time Required

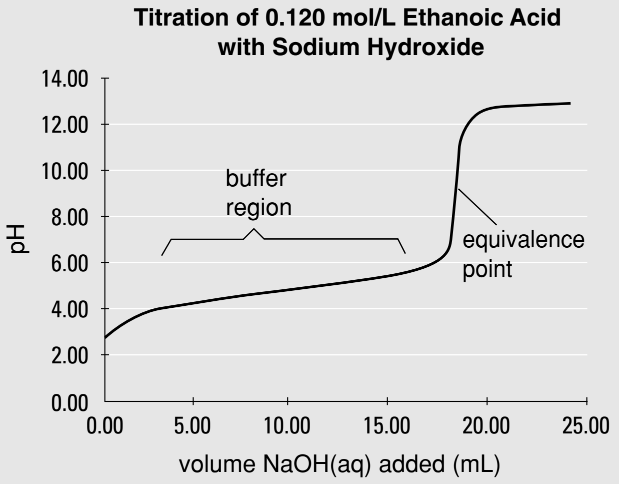
- 45 minutes

Helpful Tips

- The following reference will be useful when using a spreadsheet for this activity: http://ferl.becta.org.uk/display.cfm?resid=1788&printable=1)
- Use **BLM 17.4.3 (HAND) Thought Lab 17.1: Analyzing a Weak Acid-Strong Base Titration** to support this activity. Remove sections as necessary to meet the needs of students in your class.

Answers to Procedure Steps

1. The figure shows a typical graph drawn by a spreadsheet program for the data.



2. The pH range of a suitable indicator for this titration is 7.5 – 10.5. Suitable indicators are phenolphthalein and thymolphthalein.

3. Refer to the above figure in response to Question 1.

Answers to Analysis Questions

1. The pH and volume of base added at equivalence must be interpolated from the graph. For this reason, student answers may be slightly different, depending on their interpolation skills. The pH at equivalence is 9.25, and the volume of base added is 18.18 mL.

2. $NaOH(aq) + CH_3COOH(aq) \rightleftharpoons CH_3COONa(aq) + H_2O(\ell)$

The ratio of acid:base is 1:1

$C_{acid}V_{acid} = C_{base}V_{base}$

$0.120 \text{ mol/L} \times 25.00 \text{ mL} = C_{base} \times 18.18 \text{ mL}$

$C_{base} = [NaOH] = 0.165 \text{ mol/L}$

3. The solution is basic at equivalence because the acetate ion hydrolyzes

$CH_3COO^-(aq) + H_2O(\ell) \rightleftharpoons CH_3COOH(aq) + OH^-(aq)$

4. At 10 mL of base added, the predominant ions present are $H_3O^+(aq)$, $CH_3COO^-(aq)$, and $Na^+(aq)$. There will be a small concentration of $OH^-(aq)$. Well after the equivalence point, the most predominant ions will be $Na^+(aq)$ and $OH^-(aq)$. There will be small concentrations of $CH_3COO^-(aq)$ and $H_3O^+(aq)$.

Assessment Options

- Collect and assess students' reports.
- Use Assessment Checklist 3 Performance Task Self-Assessment and Assessment Checklist 4 Performance Task Group Assessment in Appendix A.

Answers to Questions for Comprehension

Student Textbook page 710

Q9. (a) The equivalence point in a titration is the point at which the number of moles of an unknown solution is equal to the number of moles of a standard solution. The endpoint in a titration is the point at which an indicator turns colour.

(b) When choosing an indicator, it does not need to coincide exactly. The endpoint range of the indicator should include the pH of the equivalence point.

Q10. At the equivalence point where neutralization is complete, the solution may contain ions that react with water (hydrolyze). The conjugate base of a weak acid will react with water to form a basic solution. The conjugate acid of a weak base will react with water to form an acid solution.

Q11. Checking the region of the pH scale over which the colour change for the indicators overlap, we can say that bromocresol green is blue and methyl red is orange from pH = 4.5 to pH 5.5.

Q12. Sodium hydroxide is a strong base and nitric acid is a strong acid. The pH at the equivalence point will be 7, since neither the cation nor the anion of the salt that forms from the neutralization will hydrolyze. Phenolphthalein or bromothymol blue would include pH = 7 in the range over which they change colour.

Investigation 17.B: Preparing a Buffer and Investigating Its Properties

Student Textbook pages 710–712

Purpose

The purpose of this activity is to emphasize that mid-way through the titration of a strong base/weak acid titration the flask contains a buffer solution. Students then investigate the properties of the buffer solution they will have prepared.

Outcomes

- 30-D1.1s
- 30-D1.2s

Advance Preparation

When to Begin	What to Do
2 to 3 days before	■ Gather glassware and check the pH meter (optional).
1 to 2 days before	■ Prepare solutions. ■ Photocopy **BLM 17.4.5 (HAND) Investigation 17.B: Preparing a Buffer and Investigating Its Properties.**

Materials

- 0.20 mol/L ethanoic acid, $CH_3COOH(aq)$
- 0.20 mol/L sodium hydroxide, $NaOH(aq)$
- 0.20 mol/L $HCl(aq)$
- distilled water
- 50 mL graduated cylinder
- 50 mL beakers (4)
- 100 mL beaker
- universal indicator paper (pH paper)
- pH meter (optional)
- clean straw
- stirring rod
- burettes

Time Required

- 60 to 90 minutes

Helpful Tips

- Students should do all the calculations in the Predictions steps as part of their preparation before doing the investigation.
- You can demonstrate buffer solutions: Dissolve one antacid tablet into each of three beakers each containing 100 mL of distilled water. Add a couple of drops of universal indicator to each beaker. Set one beaker aside as a control. To a second beaker, add 1.0 mol/L $HCl(aq)$ drop-wise and have students observe the change in pH by observing the indicator colour. Use the third beaker and repeat the process, adding 1.0 mol/L $NaOH(aq)$. Then repeat the entire activity without using the antacid tablets. Ask students to explain how the use of the antacid tablets buffered the change in pH as the acid and base was added.
- Preparation of stock solutions:
  - 0.20 mol/L $CH_3COOH(aq)$. Measure 11.4 mL of glacial ethanoic acid (17.5 mol/L). Add to about 500 mL of distilled water in a 1 L volumetric, then make up to the mark using distilled water.
  - 0.20 mol/L NaOH (aq). Dissolve 8.0 g $NaOH(s)$ pellets in about 500 mL of distilled water in a 1 L volumetric, then make up to the mark using distilled water.
  - 0.20 mol/L HCL(aq). Measure 16.7 mL of stock hydrochloric acid (12.0 mol/L). Add to about 500 mL of distilled water in a 1 L volumetric, then make up to the mark using distilled water.
- Use **BLM 17.4.5 (HAND) Investigation 17.B: Preparing a Buffer and Investigating Its Properties** to support this activity. Modify it as necessary.
- ***Expected Results:***
  - Part I: As the students blow into the water, the pH will decrease. It is unlikely but possible that the pH will become lower than about pH 5.
  - Part III: When students add the first mL of NaOH to the water, the pH will rise dramatically to about pH 11 or 12. Similarly, when they add HCl, the pH will drop to about pH 2 or 3.
  - Part IV: The pH of the buffer should be around pH 4.7. When students add 1.0 mL of base, the pH should increase by only a small fraction of a pH unit.
  - Part V: When students add 1.0 mL of acid, the pH should decrease by only a small fraction of a pH unit.
  - Part VI: Blowing into the buffer solution should have very little effect on the pH of the solution.

Safety Precautions

- Sodium hydroxide and hydrochloric acid are corrosive to eyes and skin and harmful if swallowed or inhaled. Wash any spills on the skin or clothing with plenty of cool water.
- Provide containers for chemical disposal so that students do not pour the solutions down the drain.
- Remind students to wash their hands after working in the laboratory.
- Ensure that students have read “Safety in Your Chemistry Laboratory and Classroom” on pages xii-xv in the student handbook. Ensure that students follow all safety precautions outlined for this investigation in the student textbook.

Answers to Predictions

(a) 40.0 mL of 0.20 mol/L $NaOH(aq)$ is required to neutralize 40.0 mL of 0.20 mol/L ethanoic acid.

(b) 20.0 mL of 0.20 mol/L $NaOH(aq)$ is required to prepare a buffer solution with 40.0 mL of 0.20 mol/L ethanoic acid.

(c) Accept any reasonable estimate that indicates the buffer is acidic.

(d) $$[\text{NaOH(aq)}] = \frac{1 \times 10^{-3}\ \text{L} \times 0.20\ \text{mol/L}}{0.021\ \text{L}} = 9.5 \times 10^{-3}\ \text{mol/L}$$

Therefore, $[\text{OH}^-(\text{aq})] = 9.5 \times 10^{-3}$ mol/L

pOH = 2.02; pH = 11.98

(e) Accept any reasonable estimate that the pH of the buffer solution changes very little. The pH may rise slightly, but the solution should still be acidic.

(f) The calculation is very similar.

$[HCl(aq)] = 9.5 \times 10^{-3}$ mol/L

Therefore, $[H_3O^+(aq)] = 9.5 \times 10^{-3}$ mol/L

pH = 2.02.

The pH of the buffer solution may fall slightly, but it should change very little.

Answers to Analysis Questions

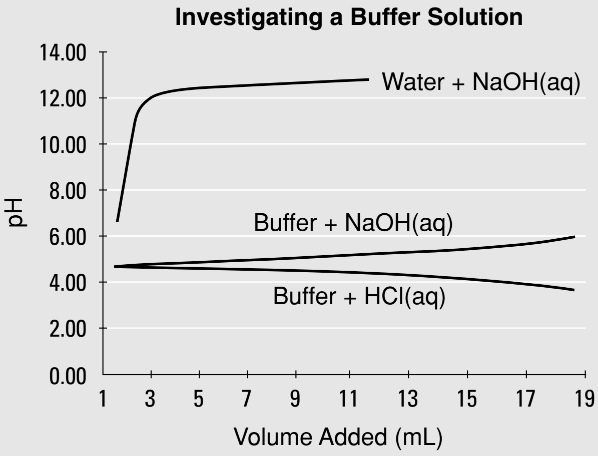
1. The distilled water may be acidic due to dissolved carbon dioxide. This should not be a problem in this activity because the amount of carbonic acid present will be very small.
2. The pH of the water will become lower. Blowing air into the water will increase the amount of dissolved carbon dioxide, and therefore the concentration of carbonic acid.

   $CO_2(aq) + H_2O(\ell) \rightleftharpoons H_2CO_3(aq)$

   $H_2CO_3(aq) + H_2O(\ell) \rightleftharpoons H_3O^+(aq) + HCO_3^-(aq)$
3. The pH of the buffer will change very little as a result of blowing air through the solution.
4. The most likely reason for any difference will be an inaccurate measurement of the volume of base added.
5. Initially, the pH of the buffer should be equal to the p*K*a of ethanoic acid, 4.7. The measured pH of the buffer may differ from the estimated pH for a number of reasons. If the difference is large, check to make sure the volumes used to make the buffer were calculated correctly. Smaller differences are likely due to errors in the measurement of the volume of acetic acid, or the volume of sodium hydroxide, or both.
6. The following is typical student data.

Volume NaOH(aq) added (mL)	pH of water + added NaOH	pH of buffer + NaOH(aq)	pH of buffer + HCl(aq)
0.0	7.00	4.74	4.74
1.0	11.98	4.78	4.72
2.00	12.26	4.83	4.69
3.00	12.42	4.87	4.67
4.00	12.52	4.92	4.64
5.00	12.60	4.96	4.62
6.00	12.66	5.01	4.59
7.00	12.71	5.06	4.55
8.00	12.76	5.11	4.52
9.00	12.79	5.16	4.48
10.00	12.82	5.22	4.44
11.00		5.28	4.39
12.00		5.34	4.34
13.00		5.41	4.28
14.00		5.49	4.22
15.00		5.59	4.14
16.00		5.69	4.04
17.00		5.83	3.92
18.00		6.02	3.74

This data is shown graphically in the figure. The buffer resists change in pH when acid or base is added. By comparison, the effect of adding base to pure water is a sudden and large increase in pH.



Answer to Conclusion Question

7. Adding a small amount of an acid or base to a buffer changes the pH of the solution by only a very small amount.

Answer to Extension Question

8. A basic pH buffer solution could be prepared by mixing aqueous ammonia with a solution of an ammonium salt, such as ammonium chloride.

Assessment Options

- Collect and assess students' reports.

- Use Assessment Checklist 3 Performance Task Self-Assessment and Checklist 4 Performance Task Group Assessment in Appendix A.

Connections (Nature of Science): Buffers in the Blood

Student Textbook page 717

Teaching Strategies

- The chemical processes that take place in the blood are directly relevant to every student. You could invite a nurse or a doctor to discuss the role of blood and how the pH is buffered by various systems.
- The following article is an excellent introduction to blood pH and the systems that buffer the blood: Richard E. Casiday, et al, "Blood-Chemistry Tutorials: Teaching Biological Applications of General Chemistry Material." *Journal of Chemical Education*, Vol. 78, No. 9, September 2001, pages 1210–1215.

Answers to Connections Questions

1. The blood transports dissolved gases between the lungs and the muscles. $CO_2(aq)$ and $H_3O^+(aq)$ produced in the muscles and dissolved in the blood are exchanged for $O_2(g)$ in the lungs. When the blood becomes too acidic, for example as the result of strenuous exercise, rapid breathing increases the removal of $CO_2(aq)$. This shifts the equilibrium to the right, removing $H_3O^+(aq)$ and raising the blood pH.

$$H_3O^+(aq) + HCO_3^-(aq) \rightleftharpoons 2H_2O(\ell) + CO_2(aq)$$

If the blood pH is too high, the kidneys remove $HCO_3^-(aq)$, shifting the equilibrium to the left, increasing the $[H_3O^+(aq)]$ and lowering the pH.

2. A change in blood pH from the optimum value for a particular enzyme will reduce the effectiveness of the enzyme. The rate of reaction catalyzed by the enzyme will be reduced.

3. Ketoacidosis results when the body shifts from using carbohydrates for fuel and instead begins using fat. The fats are converted into ketones and organic acids accumulate in the blood, raising the pH. The most common cause of ketoacidosis is Type 1 diabetes, where the pancreas makes little or no insulin. Alcoholic ketoacidosis occurs in people who abuse alcohol and as a result fail to eat enough food to supply the carbohydrates they need.

The symptoms of ketoacidosis include increased thirst and urination, and a feeling of being tired and unwell. The breath may have the fruity odour of acetone. Although someone with diabetic ketoacidosis may have increased appetite, the body is not able to use carbohydrates and the patient will lose weight. Patients with Type 1 diabetes control their symptoms by using insulin.

Section 17.4 Review Answers

Student Textbook page 718

1. (a) $KOH(aq) + HNO_3(aq) \rightarrow KNO_3(aq) + H_2O(\ell)$

(b) $H_3PO_4(aq) + 3NaOH(aq) \rightarrow Na_3PO_4(aq) + 3H_2O(\ell)$

(c) $2HBr(aq) + Ca(OH)_2(aq) \rightarrow CaBr_2(aq) + 2H_2O(\ell)$

(d) $NH_3(aq) + HClO_4(aq) \rightarrow NH_4ClO_4(aq)$

2. (a) At equivalence, none of the ions present in solution hydrolyze and the solution is neutral.

(b) At equivalence, the solution is basic because the phosphate ion hydrolyzes:

$$PO_4^{3-}(aq) + H_2O(\ell) \rightleftharpoons HPO_4^{2-}(aq) + OH^-(aq)$$

(c) At equivalence, none of the ions present in solution hydrolyze and the solution is neutral.

(d) At equivalence, the solution is acidic because the ammonium ion hydrolyzes.

$$NH_4^+(aq) + H_2O(\ell) \rightleftharpoons NH_3(aq) + H_3O^+(aq)$$

3. Congo red is not a suitable indicator for this titration because the pH at equivalence will be 7.0. The endpoint will occur after the equivalence point.

4. Titrating a strong base with a weak acid will result in a solution at equivalence that is basic. Phenolphthalein changes colour over the pH range 8–10, and thus most likely to be a suitable indicator.

5. A diprotic acid was placed in a flask and titrated against $NaOH(aq)$. The first equivalence point occurred after the addition of about 50 mL when the pH was approximately 4.0. The second equivalence point occurred after the addition of about 100 mL when the pH was approximately 9.0.

6. A buffer solution contains two components, one acidic and the other basic. A buffer solution may be made using a weak acid and a salt that contains its conjugate base, or by using a weak base and a salt that contains its conjugate acid.

7. An aqueous mixture of $NaCl(aq)$ and $HCl(aq)$ does not act as a buffer because $HCl(aq)$ is a strong acid and its conjugate base is so weak that it does not hydrolyze. An aqueous mixture of $NH_3(aq)$ and $NH_4Cl(aq)$ does act as a buffer because $NH_3(aq)$ is a weak base, and the ammonium ion hydrolyzes to act as an acid.

8. Blood buffers maintain the pH of blood within a very narrow range, pH 7.4 ± 0.4. Outside this range, very serious health consequences can result.

Chapter 17 Review Answers

Student Textbook pages 720–721

Answers to Understanding Concepts Questions

1. In the unpolluted rainwater, $[H_3O^+(aq)] = 2.51 \times 10^{-6}$. In the acid rain sample, $[H_3O^+(aq)] = 3.02 \times 10^{-5}$. Thus, the acid rain sample has more hydronium ion by a factor of

$$\frac{3.02 \times 10^{-5}}{2.51 \times 10^{-6}} = 12$$

2. Major contributions to NO_x are from automobiles and trucks. Developing countries have far fewer vehicles per person compared with developed countries such as Canada. As developing countries obtain more vehicles, the proportion of NO_x emitted will increase.

3. **(a)** The shampoo is basic.

(b)

$$K_w = [H_3O^+(aq)] \times [OH^-(aq)]$$

$$[H_3O^+(aq)] = \frac{K_w}{[OH^-(aq)]} = \frac{1.0 \times 10^{-14}}{6.8 \times 10^{-5}}$$

$$[H_3O^+(aq)] = 1.5 \times 10^{-10} \text{ mol/L}$$

(c) $pH = -\log(1.5 \times 10^{-10}) = 9.83$;
$pOH = -\log(6.8 \times 10^{-5}) = 4.17$

4. $pOH = 14.0 - pH = 14.0 - 5.53 = 8.47$
$[H_3O^+(aq)] = 10^{-pH} = 10^{-5.53} = 3.0 \times 10^{-6}$ mol/L
$[OH^-(aq)] = 10^{-pOH} = 10^{-8.47} = 3.4 \times 10^{-9}$ mol/L

5. Phenol is acidic.
$[H_3O^+(aq)] = 10^{-pH} = 10^{-4.72} = 1.9 \times 10^{-5}$ mol/L
$pOH = 14.0 - pH = 14.0 - 4.72 = 9.28$
$[OH^-(aq)] = 10^{-pOH} = 10^{-9.28} = 5.2 \times 10^{-10}$ mol/L

6. The patient's blood has $pOH = -\log(2.34 \times 10^{-7}) = 6.625$.

Therefore, the blood pH is determined to be $pH = 14.0 - 6.625 = 7.375$. This is at the very high end of the allowable range for blood pH, indicating the patient has alkalosis.

7. The conjugate base of PABA is $C_6H_4NH_2COO^-$.

8. **(a)** The conjugate acid-base pairs are $H_3BO_3(aq)$/$H_2BO_3^-(aq)$ and $H_3O^+(aq)$/$H_2O(\ell)$

(b) Boric acid is not one of the strong acids, therefore it is weak.

9. **(a)** Butyric acid is a weak acid.

(b) Hydroiodic acid is a strong acid.

(c) Potassium hydroxide is a strong base.

(d) Iron(III) oxide is a weak base.

10. **(a)** $SO_4^{2-}(aq)$

(b) $S^{2-}(aq)$

(c) $HPO_4^{2-}(aq)$

(d) $CO_3^{2-}(aq)$

11. When a small amount of acid is added to this buffer, the added hydronium ion will be removed as it combines with the methanoate ion, $HCOO^-(aq)$. The addition of a small amount of base would be neutralized by reaction with the methanoic acid.

Answers to Applying Concepts Questions

12. The ICE table for this problem is:

$$CH_3CHOHCOOH(aq) + H_2O(\ell) \rightleftharpoons CH_3CHOHCOO^-(aq) + H_3O^+(aq)$$

	$[CH_3CHOHCOOH(aq)]$ (mol/L)	$[H_2O(\ell)]$ (mol/L)	$[CH_3CHOHCOO^-(aq)]$ (mol/L)	$[H_3O^+(aq)]$ (mol/L)
Initial	0.12		0	~0
Change	–x		+x	+x
Equilibrium	0.12 –x		x	x

$$\frac{[CH_3CHOHCOOH(aq)]}{K_a} = \frac{0.12}{1.38 \times 10^{-4}} = 870$$

This is close to 1000, and the approximation may be valid. At equilibrium,
$[CH_3CHOHCOOH(aq)] \approx 0.12$ mol/L

$$K_a = \frac{[CH_3CHOHCOO^-(aq)][H_3O^+(aq)]}{[CH_3CHOHCOOH(aq)]}$$

$$= 1.38 \times 10^{-4}$$

$$\frac{x^2}{0.12} = 1.38 \times 10^{-4}$$

$$x = 4.07 \times 10^{-3}$$

$0.12 - 4.07 \times 10^{-3} = 0.116$; the approximation is valid.
$[H_3O^+(aq)] = 4.07 \times 10^{-3}$ mol/L
$pH = 2.39$

13. The molar mass of salicylic acid, $C_6H_4(OH)COOH$, is 138.1 g/mol.

$$\text{Amount } C_6H_4(OH)COOH = \frac{0.22 \text{ g}}{138.1 \text{ g/mol}} = 1.59 \times 10^{-3} \text{ mol}$$

$$[CH_3CHOHCOOH(aq)] = \frac{1.59 \times 10^{-3} \text{ mol}}{0.100 \text{ L}} = 1.59 \times 10^{-2} \text{ mol/L}$$

$[H_3O^+(aq)] = 10^{-2.43} = 3.72 \times 10^{-3}$ mol/L

$$C_6H_4(OH)COOH(aq) + H_2O(\ell) \rightleftharpoons C_6H_4(OH)COO^-(aq) + H_3O^+(aq)$$

	$[C_6H_4(OH)COOH(aq)]$ (mol/L)	$[H_2O(\ell)]$ (mol/L)	$[C_6H_4(OH)COO^-(aq)]$ (mol/L)	$[H_3O^+(aq)]$ (mol/L)
Initial	1.59×10^{-2}		0	~0
Change	-3.72×10^{-3}		$+3.72 \times 10^{-3}$	$+3.72 \times 10^{-3}$
Equilibrium	1.22×10^{-2}		3.72×10^{-3}	3.72×10^{-3}

$$K_a = \frac{[C_6H_4(OH)COO^-(aq)][H_3O^+(aq)]}{[C_6H_4(OH)COOH(aq)]}$$

$$K_a = \frac{(3.72 \times 10^{-3})^2}{1.22 \times 10^{-2}} = 1.13 \times 10^{-3}$$

$$K_a = 1.1 \times 10^{-3}$$

14.

Acid	K_a	Conjugate Base	$K_b = \frac{K_w}{K_a}$
$HCOOH(aq)$	1.8×10^{-5}	$HCOO^-(aq)$	5.6×10^{-10}
$HF(aq)$	6.3×10^{-4}	$F^-(aq)$	1.6×10^{-11}
$C_6H_5COOH(aq)$	6.3×10^{-5}	$C_6H_5COO^-(aq)$	1.6×10^{-10}
$HCN(aq)$	6.2×10^{-10}	$CN^-(aq)$	1.6×10^{-5}

The order of increasing base strength is
$F^-(aq) < C_6H_5COO^-(aq) < HCOO^-(aq) < CN^-(aq)$

15. $K_w = [H_3O^+(aq)][OH^-(aq)]$

In pure water, $[H_3O^+(aq)] = [OH^-(aq)]$

$$[H_3O^+(aq)] = [OH^-(aq)] = \sqrt{K_w} = \sqrt{2.5 \times 10^{-14}}$$
$$= 1.6 \times 10^{-7}$$

$[H_3O^+(aq)] = [OH^-(aq)] = 1.6 \times 10^{-7}$ mol/L

pH = pOH = 6.80

Pure water at this (or any other) temperature is neutral.

16. The conjugate base is $OBr^-(aq)$.

$$K_b = \frac{K_w}{K_a} = \frac{1.0 \times 10^{-14}}{2.1 \times 10^{-9}} = 4.8 \times 10^{-6}$$

17. The weaker base is ammonia. Therefore, the stronger acid will be $NH_4^+(aq)$.

18. $NH_4NO_3(aq) \rightarrow NH_4^+(aq) + NO_3^-(aq)$

Only the ammonium ion hydrolyzes.

K_a for $NH_4^+(aq) = 5.6 \times 10^{-10}$

$$\frac{[NH_4^+(aq)]}{K_a} = \frac{0.1}{5.6 \times 10^{-10}} >> 1000$$

$$NH_4^+(aq) + H_2O(\ell) \rightleftharpoons NH_3(aq) + H_3O^+(aq)$$

	$[NH_4^+(aq)]$ (mol/L)	$[H_2O(\ell)]$ (mol/L)	$[NH_3(aq)]$ (mol/L)	$[H_3O^+(aq)]$ (mol/L)
Initial	0.10		0	~0
Change	$-x$		$+x$	$+x$
Equilibrium	$0.10 - x$ ≈ 0.10		x	x

$$K_a = \frac{[NH_3(aq)][H_3O^+(aq)]}{[NH_4^+(aq)]} = 5.6 \times 10^{-10}$$

$$5.6 \times 10^{-10} = \frac{x^2}{0.10}$$

$x = [H_3O^+(aq)] = 7.48 \times 10^{-6}$

pH = 5.1

19. $NaOCl(aq) \rightarrow Na^+(aq) + OCl^-(aq)$

K_a for $HOCl(aq) = 4.0 \times 10^{-8}$.

$$K_b \text{ for } OCl^-(aq) = \frac{K_w}{K_a} = \frac{1.0 \times 10^{-14}}{4.0 \times 10^{-8}} = 2.5 \times 10^{-7}$$

$M_{NaOCl} = 74.4$ g/mol

$$n_{NaOCl} = \frac{5.0 \text{ g}}{74.4 \text{ g/mol}} = 6.72 \times 10^{-2} \text{ mol}$$

$$[OCl^-(aq)] = \frac{6.72 \times 10^{-2} \text{ mol}}{0.100 \text{ L}} = 6.72 \times 10^{-1} \text{ mol/L}$$

$$\frac{[OCl^-(aq)]}{K_b} = \frac{6.72 \times 10^{-1}}{2.5 \times 10^{-7}} >> 1000$$

$$OCl^-(aq) + H_2O(\ell) \rightleftharpoons HOCl(aq) + OH^-(aq)$$

	$[OCl^-(aq)]$ (mol/L)	$[H_2O(\ell)]$ (mol/L)	[HOCl (aq)] (mol/L)	$[OH^-(aq)]$ (mol/L)
Initial	6.72×10^{-1}		0	~0
Change	$-x$		$+x$	$+x$
Equilibrium	$6.72 \times 10^{-1} - x$ $\approx 6.72 \times 10^{-1}$		x	x

$$K_b = \frac{[HOCl(aq)][OH^-(aq)]}{[OCl^-(aq)]} = 2.5 \times 10^{-7}$$

$$2.5 \times 10^{-7} = \frac{x^2}{6.72 \times 10^{-1}}$$

$x = [OH^-(aq)] = 4.10 \times 10^{-4}$ mol/L

pOH = 3.39

pH = 10.61.

20. **(a)** $K_a = \frac{[PO_4^{3-}(aq)][H_3O^+(aq)]}{[HPO_4^{2-}(aq)]}$

(b) $K_b = \frac{[H_2PO_4^-(aq)][OH^-(aq)]}{[HPO_4^{2-}(aq)]}$

(c) An aqueous solution of sodium hydrogen phosphate will be basic because $K_b > K_a$.

Answers to Solving Problems Questions

21. There are four major types of coal: lignite, sub-bituminous, bituminous and anthracite. The coal on Alberta's plains is mainly sub-bituminous, which does not require processing for use as a fuel in electricity generating plants. Bituminous coal is mined in the Foothills and mountainous regions, and this must be upgraded or processed before it can be used. Sulfur dioxide emissions from coal fuels may be reduced during combustion or from the flue gases after combustion. Lime or limestone can be added to the coal to absorb $SO_2(g)$ as it is released during combustion. Sulfur dioxide can be removed from flue gases by passing the gas through a spray tower where calcium carbonate reacts with $SO_2(g)$ to form calcium sulfite and calcium sulfate. The reaction products are treated to form gypsum, which has a number of uses including the manufacture of wallboard for the building industry. In the ammonia scrubbing process, aqueous ammonia is used in a spray tower, and the $SO_2(g)$ reacts to form ammonium sulfate which can be recovered and sold as fertilizer.

22. Sodium hypochlorite reacts with water to produce the hypochlorous acid needed to kill germs and algae:

$$NaOCl(aq) + H_2O(\ell) \rightarrow HOCl(aq) + Na^+(aq) + OH^-(aq)$$

The relative amounts of HOCl and OCl^- change with pH, and the ideal range is 7.2 – 7.8. This is why the pH of a pool must be tested regularly. If the pH is too low, the water is too acidic and the acid may make swimmers' eyes sore. Also, the acid would attack concrete pools or the mortar between tiles, as these materials contain calcium carbonate.

23. $H_2S(aq)$: $K_a = 8.9 \times 10^{-8}$

$H_2SO_3(aq)$: $K_a = 1.4 \times 10^{-2}$

(a) HS^-: $K_b = \dfrac{[H_2S(aq)][OH^-(aq)]}{[HS^-(aq)]}$

HSO_3^-: $K_b = \dfrac{[H_2SO_3(aq)][OH^-(aq)]}{[HSO_3^-(aq)]}$

(b) $K_b = \dfrac{K_w}{K_a}$

$HS^-(aq)$: $K_b = \dfrac{1.1 \times 10^{-14}}{8.9 \times 10^{-8}} = 1.2 \times 10^{-7}$

$HSO_3^-(aq)$: $K_b = \dfrac{1.0 \times 10^{-14}}{1.4 \times 10^{-2}} = 7.1 \times 10^{-13}$

(c) $HS^-(aq)$ is the stronger base because it has a larger value for K_b.

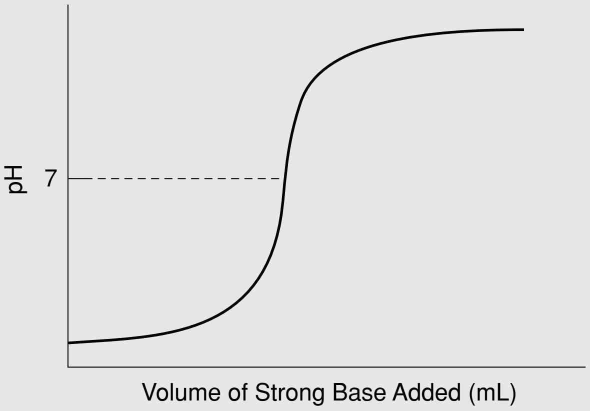
24. Each equilibrium lies to the right, therefore, the stronger acid is on the left of the equation. Thus, each equation gives the following relative strengths:

$HBrO_2(aq) > CH_3COOH(aq)$; $H_2S(aq) > H_2O(\ell)$; and $CH_3COOH(aq) > H_2S(aq)$.

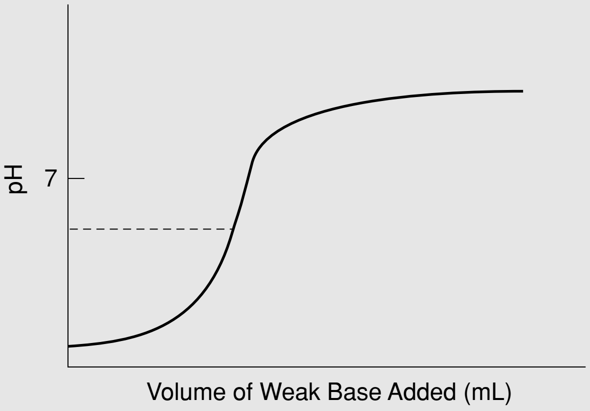
Therefore, the order is $HBrO_2(aq) > CH_3COOH(aq) > H_2S(aq) > H_2O(\ell)$.

25. (a)

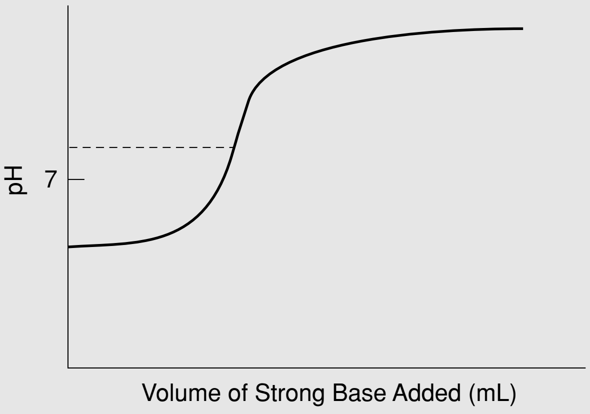
Titration Curve for a Strong Acid with a Strong Base



Titration Curve for a Weak Acid with a Weak Base



Titration Curve for a Weak Acid with a Strong Base



(b) Congo red would be a good indicator for the titration of a strong acid with a weak base.

26. Boric acid is very weak ($K_a = 5.9 \times 10^{-10}$). Therefore the borate ion is a relatively strong base. Because K_b for the borate ion is greater than K_a for boric acid, the buffer is basic.

Answers to Making Connections Questions

27. **(a)** In the student textbook, the environmental issue of acid rain is extended beyond students' previous knowledge by including chemical equilibrium and Le Châtelier's principle. Students are likely to mention facts that relate to either or both of these concepts.

(b) Some possible questions include:

Does the pH of acid deposition change with the seasons?

Design an experiment to investigate the effect of acid rain on leaves.

How do controls fitted to vehicles reduce $SO_x(g)$ and $NO_x(g)$?

Draw a map of North America showing the location of coal burning power plants, prevailing winds, and regions with low pH acid deposition.

The rubric used to assess the research project must be specific to the actual project, and should contain qualitative descriptions of performance criteria.

28. For the most dilute solution,

$$\frac{[C_6H_5COOH(aq)]}{K_a} = \frac{0.01}{6.3 \times 10^{-5}} = 159$$

Thus, the approximation may not be valid for this solution. However, the approximation will be valid for the other concentrations. Since the question is looking for a trend, assume the approximation is valid for each dilute solution.

(a) $C_6H_5COOH(aq) + H_2O(\ell) \rightleftharpoons C_6H_5COO^-(aq) + H_3O^+(aq)$

	$[C_6H_5COOH(aq)]$ (mol/L)	$[H_2O(\ell)]$ (mol/L)	$[C_6H_5COO^-(aq)]$ (mol/L)	$[H_3O^+(aq)]$ (mol/L)
Initial	1.0		0	~0
Change	–x		+x	+x
Equilibrium	1.0 – x ≈ 1.0		x	x

$$K_a = \frac{[C_6H_5COO^-(aq)][H_3O^+(aq)]}{[C_6H_5COOH(aq)]}$$

$$6.3 \times 10^{-5} = \frac{x^2}{1.0}$$

$x = [H_3O^+(aq)] = 7.94 \times 10^{-3}$ mol/L

pH = 2.10

$$\text{percent ionization} = \frac{7.94 \times 10^{-3}}{1.0} \times 100 = 0.79\%$$

(b) By similar calculations, the 0.10 mol/L solution benzoic acid has pH = 2.60 and is 2.5% ionized.

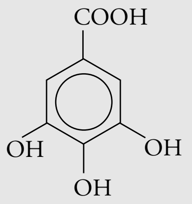
(c) The 0.01 mol/L solution has pH = 3.10 and is 7.9% ionized.

As the solution becomes more dilute, more benzoic acid molecules ionize.

$C_6H_5COOH(aq) + H_2O(\ell) \rightleftharpoons C_6H_5COO^-(aq) + H_3O^+(aq)$

Le Châtelier's principle predicts that as more water is added, the equilibrium will shift to the right causing greater ionization and increased $[H_3O^+(aq)]$. This matches the results of the calculations.

29. **(a)**



(b) The conjugate base is $C_6H_5(OH)_3COO^-$.

$$K_b = \frac{K_w}{K_a} = \frac{1.0 \times 10^{-14}}{3.9 \times 10^{-5}} = 2.6 \times 10^{-10}$$

Career Focus: Ask a Dairy Farmer

Student Textbook pages 722–723

Teaching Strategies

- If possible, invite a dairy farmer or a veterinarian to talk to the class about animal nutrition.
- Remind students that acidosis is a term meaning elevated concentration of $H_3O^+(aq)$, and just as their stomach has an optimum pH, so does the rumen of a cow. Ask how enzyme function might depend on pH, and how this would affect digestion.
- Alberta Dairy Management has information on rumen acidosis at http://www.westerndairyscience.com/html/ADM%20articles/html/Acidosis.html
- A fact sheet on rumen acidosis in dairy cows is available at http://www.omafra.gov.on.ca/english/livestock/dairy/facts/03-031.htm
- General information on occupations in health services in Alberta is available at http://nextgen.org/Portals/nextgen/media/programs/Health%20Services%20Careers%20Brochure%202005.pdf
- Alberta occupation profiles may be found at http://www.alis.gov.ab.ca/occinfo/Content/RequestAction.asp?format=html&aspAction=GetHomePage&Page=Home
- The following sites may be useful for students
    - Canadian Association of Nephrology Nurses and Technologists http://www.cannt.ca/
    - The Canadian Society of Clinical Perfusion http://www.cscp.ca/
    - A summary for those interested in a career as a Respiratory Therapist, Clinical Perfusionist or Cardio-Pulmonary Technologist. http://www1.on.hrdc-drhc.gc.ca/ojf/ojf.jsp?lang=e&section=Profile&noc=3214

Answers to Go Further… Questions

1. $HPO_4^{2-}(aq) + H_3O^+(aq) \rightleftharpoons H_2PO_4^-(aq) + H_2O(\ell)$
2. Blood pH will fall if excess carbon dioxide builds up in the blood. Le Châtelier's principle predicts the equilibrium will shift to the right with excessive build up of carbon dioxide, causing an increase in $H^+(aq)$ in the blood.

   $$CO_2(aq) + H_2O(\ell) \rightleftharpoons H_2CO_3(aq) \rightleftharpoons H^+(aq) + HCO_3^-(aq)$$
3. Humans sometimes take sodium bicarbonate to relieve the symptoms of excess stomach acidity. The same treatment should work for rumen acidosis.

   $$HCO_3^-(aq) + H^+(aq) \rightleftharpoons H_2CO_3(aq) \rightleftharpoons CO_2(g) + H_2O(\ell)$$

Unit 8 Review Answers

Student Textbook pages 724–727

Answers to Understanding Concepts Questions

Understanding Concepts

1. A system like this is not at equilibrium because it is not a closed system.
2. **(a)** heterogeneous
   (b) homogeneous
   (c) homogeneous
   (d) heterogeneous
   (e) heterogeneous
3. **(a)** heterogeneous
   (b) heterogeneous
   (c) heterogeneous
   (d) heterogeneous
4. **(a)** $K_c = \dfrac{[H_2(g)]^3[CO(g)]}{[CH_4(g)][H_2O(g)]}$

   (b) $K_c = \dfrac{[NO(g)]^4}{[N_2O(g)]^2[O_2(g)]}$

   (c) $K_c = \dfrac{[POCl_3(g)]^2}{[PCl_3(g)]^2[O_2(g)]}$

   (d) $K_c = \dfrac{[N_2(g)][H_2O(g)]^2}{[NO(g)]^2[H_2(g)]^2}$

   (e) $K_c = \dfrac{[C_7H_8(g)][H_2(g)]^3}{[C_7H_{14}(g)]}$
5. Reactions (c) and (d) will shift to the right to re-establish equilibrium if the volume of the container is decreased.
6. Reaction (a) will shift to the left to re-establish equilibrium if the temperature of the system is decreased.
7. **(a)** Reactants are favoured when the temperature is increased.
   (b) No change occurs when helium gas is added at constant volume.
   (c) Reactants are favoured when helium gas is added at constant pressure.
8. Helium does not react with any gases in the equation, and adding He(g) at constant volume leaves the concentration of reacting gases unchanged. Therefore, the equilibrium is not affected.
9. **(a)** Removing $O_2(g)$ will cause the reaction to shift to the left to re-establish equilibrium.
   (b) Adding NO(g) will cause the reaction to shift to the left.
   (c) Adding $H_2O(g)$ will cause the reaction to shift to the left.
   (d) Adding a catalyst will have no effect on the position of equilibrium.
   (e) Decreasing the volume of the container will cause the reaction to shift to the left.
   (f) Cooling the container will cause the reaction to shift to the right.
10. The reaction shifts to the right at the higher temperature. Based on this information only, Le Châtelier's principle predicts the reaction is endothermic.
11. **(a)** The statement is false. The equilibrium constant does depend on temperature.
    (b) The statement is false. The rate of a reaction does not depend on the magnitude of the equilibrium constant.
    (c) The statement is false. At equilibrium, the concentration of reactants and the concentration of products are determined by the stoichiometry and the equilibrium constant for the reaction.
    (d) The statement is false. The equilibrium constant is a ratio of product concentrations to reactant concentrations.
12. The position of equilibrium lies to the right at lower temperature. Although this increases the yield of product, the rate is slower at lower temperatures. A compromise is made between a good rate and reasonable yield.
13. Syngas is a mixture of carbon monoxide and hydrogen gas. Syngas can be produced by reacting methane with steam:

    $$CH_4(g) + H_2O(g) + \text{heat} \rightleftharpoons CO(g) + 3H_2(g)$$

    By adjusting the temperature and pressure conditions, different proportions of the reacting gases will be present at equilibrium.
14. The most important gases responsible for acid deposition are $SO_2(g)$, $SO_3(g)$, NO(g) and $NO_2(g)$.
15. In Alberta, the pollutants that cause acid rain come mainly from removing sulfur from natural gas and oil,

and from power plants that burn coal containing sulfur compounds.

16. The conjugate acids are:
- **(a)** $H_2SO_4(aq)$, sulfuric acid
- **(b)** $NH_4^+(aq)$, ammonium ion
- **(c)** $H_3O^+(aq)$, hydronium ion
- **(d)** $H_2PO_4^-(aq)$, dihydrogen phophate ion

17. The conjugate bases are:
- **(a)** $SO_4^{2-}(aq)$, sulfate ion
- **(b)** $NO_2^-(aq)$, nitrite ion
- **(c)** $HSO_3^-(aq)$, hydrogen sulfite ion.
- **(d)** $O^{2-}(aq)$, oxide ion

18. The conjugate acid-base pairs are: $HClO_2(aq)/ClO_2^-(aq)$ and $H_2O(\ell)/H_3O^+(aq)$.

19. The conjugate acid-base pairs are: $HBO_3^{2-}(aq)/H_2BO_3^-(aq)$ and $HSIO_3^-(aq)/SIO_3^{2-}(aq)$

20. **(a)** $$K_a = \frac{[H_3O^+(aq)][CN^-(aq)]}{[HCN(aq)]}$$

(b) $$K_a = \frac{[H_3O^+(aq)][HPO_4^{2-}(aq)]}{[H_2PO_4^-(aq)]}$$

(c) $$K_a = \frac{[H_3O^+(aq)][HCOO^-(aq)]}{[HCOOH(aq)]}$$

(d) $$K_a = \frac{[H_3O^+(aq)][OOCCOO^{2-}(aq)]}{[HOOCCOO^-(aq)]}$$

21. **(a)** $$K_b = \frac{[HF(aq)][OH^-(aq)]}{[F^-(aq)]}$$

(b) $$K_b = \frac{[H_2C_6H_6O_6(aq)][OH^-(aq)]}{[HC_6H_6O_6^-(aq)]}$$

(c) $$K_b = \frac{[N_2H_5^+(aq)][OH^-(aq)]}{[N_2H_4(aq)]}$$

(d) $$K_b = \frac{[PH_4^+(aq)][OH^-(aq)]}{[PH_3(aq)]}$$

22. $CH_3COOH(aq) < H_3PO_4(aq) < HF(aq) < HI(aq)$.

23. $I^-(aq) < F^-(aq) < H_2PO_4^-(aq) < CH_3COO^-(aq)$

24. (c) is a buffer, because the solution will contain $NH_3(aq)$ and $NH_4^+(aq)$. The other solutions are not buffers.

25. **(a)** Possible answers include $NH_4^+(aq)/NH_3(aq)$ and $CH_3COOH(aq)/CH_3COO^-(aq)$

(b) Possible answers include $HCl(aq)/Cl^-(aq)$ and $HNO_3(aq)/NO_3^-(aq)$

Answers to Applying Concepts Questions

26. The percentage of sulfur dioxide that reacts at this temperature is relatively large. Twice as much sulfur dioxide reacts compared with the percentage of $O_2(g)$ that reacts.

27. The ICE table is:

concentration (mol/L)

$$I_2(aq) \quad + \quad I^-(aq) \quad \rightleftharpoons \quad I_3^-(aq)$$

Initial	0.002	0.002	0
Change	$-x$	$-x$	$+x$
Equilibrium	5.0×10^{-4}	$0.002 - x$	x

$0.002 - x = 5.0 \times 10^{-4}$

$x = 0.0015$

At equilibrium, $[I_2(aq)] = 5.0 \times 10^{-4}$ mol/L; $[I^-(aq)] = 5.0 \times 10^{-4}$ mol/L; $[I_3^-(aq)] = 0.0015$ mol/L

$$K_c = \frac{[I_3^-(aq)]}{[I_2(aq)][I^-(aq)]} = \frac{0.0015}{(5.0 \times 10^{-4})(5.0 \times 10^{-4})}$$

$K_c = 6 \times 10^3$

28. Initial concentrations:

$$[H_2(g)] = \frac{3.60}{2.00} = 1.80 \text{ mol/L}$$

$$[CO_2(g)] = \frac{5.20}{2.00} = 2.60 \text{ mol/L}$$

The ICE table is:

concentration (mol/L)

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

Initial	1.80	2.60	0.00	0.00
Change	$-x$	$-x$	$+x$	$+x$
Equilibrium	$1.80 - x$	$2.60 - x$	x	x

75% of $H_2(g)$ reacts $= 0.75 \times 1.80 = 1.35$ mol/L $= x$

Equilibrium concentrations:

$[H_2(g)] = (1.80 - 1.35) = 0.45$ mol/L

$[CO_2(g)] = (2.60 - 1.35) = 1.25$ mol/L

$[H_2O(g)] = 1.35$ mol/L

$[CO(g)] = 1.35$ mol/L

$$K_c = \frac{[H_2O(g)][CO(g)]}{[H_2(g)][CO_2(g)]} = \frac{(1.35)(1.35)}{(0.45)(1.25)}$$

$K_c = 3.2$

29. Let the initial concentration of reactant gases be 1.00 mol/L. The change in concentration of $N_2(g)$ at equilibrium is $0.053 \times 1.00 = 0.053$ mol/L.

ICE table:

concentration (mol/L)

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

Initial	1.00	1.00	0
Change	– 0.053	– 0.053	+ 0.106
Equilibrium	0.947	0.947	0.106

$$K_c = \frac{[NO(g)]^2}{[N_2(g)][O_2(g)]} = \frac{(0.106)^2}{(0.947)(0.947)}$$

$K_c = 1.3 \times 10^{-2}$

30. Initial concentration:

$$[NO_2(g)] = \frac{0.100}{4.00} = 0.0250 \text{ mol/L}$$

ICE table:

concentration (mol/L) $\quad 2NO_2(g) \quad \rightleftharpoons \quad N_2O_4(g)$

Initial	0.0250	0
Change	$-2x$	$+ x$
Equilibrium	$0.0250 - 2x$	x

74% of $NO_2(g)$ reacts = 0.74 × 0.0250 = 0.0185 mol/L $= 2x$

Equilibrium concentrations:

$[NO_2(g)] = (0.0250 - 0.0185) = 0.0065$ mol/L

$[N_2O_4(g)] = 0.0092$ mol/L

$$K_c = \frac{[N_2O_4(g)]}{[NO_2(g)]^2} = \frac{0.0092}{(0.0065)^2}$$

$K_c = 2.2 \times 10^2$

31. (a) ICE table:

concentration (mol/L)

$HOCl(aq) + H_2O(\ell) \rightleftharpoons H_3O^+(aq) + OCl^-(aq)$

Initial	0.065		≈ 0	0
Change	$- x$		$+ x$	$+ x$
Equilibrium	$0.065 - x$		x	x

$$\frac{[HOCl(aq)]}{K_a} = \frac{0.065}{4.0 \times 10^{-8}} >> 1000$$

Therefore, at equilibrium, $[HOCl(aq)] \approx 0.065$ mol/L

$$K_a = \frac{[H_3O^+(aq)][OCl^-(aq)]}{[HOCl(aq)]}$$

$$4.0 \times 10^{-8} = \frac{x^2}{0.065}$$

$x = [H_3O^+(aq)] = 5.10 \times 10^{-5}$ mol/L

$pH = -\log[H_3O^+(aq)] = 4.29$

$$\text{percent ionization} = \frac{5.0 \times 10^{-5}}{0.065} \times 100 = 0.077\%$$

(b) The conjugate base of $HOCl(aq)$ is $OCl^-(aq)$.

$$K_b = \frac{K_w}{K_a} = \frac{1.0 \times 10^{-14}}{4.0 \times 10^{-8}}$$

$K_b = 2.5 \times 10^{-7}$

32. The amount that ionizes is
0.25 × 0.030 = 0.0075 mol/L.

ICE table:

concentration (mol/L)

$HA(aq) \quad + \quad H_2O(\ell) \rightleftharpoons H_3O^+(aq) \quad + A^-(aq)$

Initial	0.25		≈ 0	0
Change	–0.0075		+ 0.0075	+ 0.0075
Equilibrium	0.25 – 0.0075 = 0.242		0.0075	0.0075

$[H_3O^+(aq)] = 0.0075$ mol/L

$pH = -\log[H_3O^+(aq)] = -\log(0.0075)$

$pH = 2.12$

$pOH = 14.00 - pH = 11.88$

$[OH^-(aq)] = 10^{-pOH} = 10^{-11.88}$

$[OH^-(aq)] = 1.3 \times 10^{-12}$ mol/L

$$K_a = \frac{[H_3O^+][A^-(aq)]}{[HA(aq)]} = \frac{(0.0075)(0.0075)}{0.242}$$

$K_a = 2.3 \times 10^{-4}$

33. The conjugate acid of $HSeO_3{}^-(aq)$ is $H_2SeO_3(aq)$.

$$K_b = \frac{K_w}{K_a} = \frac{1.0 \times 10^{-14}}{3.5 \times 10^{-3}}$$

K_b for $HSeO_3{}^-(aq) = 2.9 \times 10^{-12}$

34. (a) $K_b = \dfrac{K_w}{K_a} = \dfrac{1.0 \times 10^{-14}}{6.2 \times 10^{-10}}$

K_b for $CN^-(aq) = 1.6 \times 10^{-5}$

(b) ICE table:

concentration (mol/L)

$HCN(aq) \; + \; H_2O(\ell) \rightleftharpoons H_3O^+(aq) + CN^-(aq)$

Initial	0.12		≈ 0	0
Change	$- x$		$+ x$	$+ x$
Equilibrium	$0.12 - x$		x	x

$$\frac{[HCN(aq)]}{K_a} = \frac{0.12}{6.2 \times 10^{-10}} >> 1000$$

Therefore, at equilibrium, $[HCN(aq)] \approx 0.12$ mol/L

$$K_a = \frac{[H_3O^+(aq)][CN^-(aq)]}{[HCN(aq)]}$$

$$6.2 \times 10^{-10} = \frac{x^2}{0.12}$$

$$x = 8.63 \times 10^{-6} \text{ mol/L} = [H_3O^+(aq)]$$

$$pH = -\log(8.63 \times 10^{-6}) = 5.06$$

35.

$$C_6H_5COOH(aq) + H_2O(\ell) \rightleftharpoons C_6H_5COO^-(aq) + H_3O^+(aq)$$

	$[C_6H_5COOH(aq)]$ (mol/L)	$[H2O(\ell)]$ (mol/L)	$[C_6H_5COO^-(aq)]$ (mol/L)	$[H_3O^+(aq)]$ (mol/L)
Initial	0.25		0	0
Change	–x		+x	+x
Equilibrium	0.25–x		+x	+x

$$\frac{[C_6H_5COOH]}{K_a} = \frac{0.25}{6.3 \times 10^{-4}}$$

$$= 4.0 \times 10^2$$

Since this value is less than 1000, the amount that ionizes is significant compared with the initial concentration of acid. Therefore, the approximation cannot be used. The value for x must be determined using either the quadratic equation or a graphing calculator.

$$K_a = \frac{[C_6H_5COO^-][H_3O^+]}{[C_6H_5COOH]}$$

$$6.3 \times 10^{-4} = \frac{(x)(x)}{(0.25 - x)}$$

$$6.3 \times 10^{-4}(0.25 - x) = x^2$$

$$0 = x^2 + 6.3 \times 10^{-4}x - 1.6 \times 10^{-4}$$

$$x = 0.012 \text{ or } 0.013$$

Only the positive root is possible, therefore $[H_3O^+(aq)] =$ 0.012 mol/L

$$\text{Percent Ionization} = \frac{0.012 \text{ mol/L}}{0.25 \text{ mol/L}} \times 100\%$$

$$= 4.8\%$$

36. ICE table:

concentration (mol/L)

$$BH(aq) + H_2O(\ell) \rightleftharpoons OH^-(aq) + BH_2^+(aq)$$

Initial	0.16		≈ 0	0
Change	– *x*		+ *x*	+ *x*
Equilibrium	0.16 – *x*		*x*	*x*

$$\frac{[BH(aq)]}{K_b} = \frac{0.16}{1.55 \times 10^{-4}} > 1000$$

Therefore, at equilibrium, $[BH(aq)] \approx 0.16$ mol/L

$$K_b = \frac{[BH_2^+(aq)][OH^-(aq)]}{[BH(aq)]}$$

$$1.55 \times 10^{-4} = \frac{x^2}{0.16}$$

$$x = 4.98 \times 10^{-3} \text{ mol/L} = [OH^-(aq)]$$

$$pOH = -\log[OH^-(aq)] = -\log(4.98 \times 10^{-3}) = 2.303$$

$$pH = 14.00 - pOH = 14.00 - 2.303$$

$$pH = 11.70$$

37. $H_2S(aq)$ is a stronger acid than HCN(aq), because an acid-base reaction always proceeds from the stronger to the weaker acid.

38. **(a)** $CH_3COOH(aq)$ is a stronger acid than $H_2S(aq)$. Therefore, the equilibrium mixture favours reactants.

(b) $NH_4^+(aq)$ is a stronger acid than $HCO_3^-(aq)$. Therefore, the equilibrium mixture favours products.

(c) $HSO_4^-(aq)$ is a stronger acid than HCN(aq). Therefore, the equilibrium mixture favours reactants.

(d) $H_2CO_3(aq)$ is a stronger acid than HCN(aq). Therefore, the equilibrium mixture favours products.

39. **(a)** 1, 2, 6, and 8 would definitely allow the student to solve the problem.

- Method 1 – a weak acid will have a higher pH than the expected pH of 1.0 if the acid were strong.
- Method 2 – the equivalence point will be at pH 7 if the acid is strong, above pH 7 if the acid is weak.
- Method 3 – although a strong acid will react more vigorously than a weak acid, without a comparison, the student will not be able to tell if the reaction is "more" or "less" vigorous.
- Method 4 – since the number of moles of hydrogen atoms is the same in a weak or strong acid with the same concentration, the same volume of gas will be generated regardless of whether the unknown is weak or strong.
- Method 5 – since the number of moles of hydrogen atoms is the same in a weak or strong acid with the same concentration, the same amount of energy will be released during the neutralization reaction.
- Method 6 – the final pH of the mixture will be 7 if the acid is a strong acid (the pH at neutralization for a strong acid/strong base), the pH will be greater than 7 if the acid is a weak acid (the pH at neutralization for a weak acid/strong base)
- Method 7 – although the voltage will be slightly lower for a weak acid than a strong acid (since the concentration of the hydrogen ions is lower), without a strong acid cell for comparison, the student will not be able to identify the acid.

- Method 8 – the acid will react at the same vigor as the HCl(aq) if it is a strong acid. If it is a weak acid, it will react more slowly.

(b) Method 1

- sample of acid
- beaker
- pH paper or pH meter

Method 2

- sample of acid
- 0.100 mol/L NaOH(aq) solution
- pH paper or pH meter
- distilled water
- pipette
- burette and burette clamp
- retort stand
- funnel
- pipette bulb
- Erlenmeyer flask

Method 6

- sample of acid
- graduated cylinder(s)
- 0.10 mol/L NaOH(aq)
- 100 mL beaker
- stirring rod
- pH paper or pH meter

Method 8

- sample of acid
- balance
- $NaHCO_3(s)$
- 0.100 mol/L HCl(aq)
- stopwatch
- 100 mL beaker(s)
- stirring rod

(c) The titration of 10.0 mL of 0.100 mol/L solution of an unknown acid with 0.100 mol/L NaOH(aq)

Volume of NaOH(aq) added (mL)	pH
0.00	
1.00	
2.00	
3.00	
4.00	
5.00	
6.00	
7.00	
8.00	
9.00	
10.00	
11.00	
12.00	
13.00	
14.00	
15.00	

(d) The important controlled (or fixed) variables are: the mass of $NaHCO_3(s)$, the volume of the acid used, the concentration of the acid used.

(e) If the unknown acid is a strong acid, its pH will be 1.000. ($pH=-\log[H_3O^+(aq)] = -\log(0.100) = 1.000$)

(f) $Zn(s) + 2H_3O^+(aq) \rightarrow H_2(g) + Zn^{2+}(aq) + 2H_2O(\ell)$ or $Zn(s) + 2H^+(aq) \rightarrow H_2(g) + Zn^{2+}(aq)$

(g) If the acid solution had a concentration of 1.00 mol/L, the predicted voltage would be 0.45 V. Since the concentration is only 0.100 mol/L, the voltage would likely be less.

Answers to Solving Problems Questions

40. Accept reasonable estimates, with supporting calculations. The sources used should be correctly referenced. Information about coal may be found on an Alberta government site at:
http://www.energy.gov.ab.ca/436.asp

According to information on this site, the sulfur content in the coal from Alberta mines is typically less than 0.5%.

Natural Resources Canada has information about coal based electrical generation at:
http://www.nrcan.gc.ca/es/etb/cetc/combustion/cctrm/htmldocs/map_electrical_utility_e.html

From this site, students can find further information. For example, about the Keephills generating plants in the Wabamun Lake area of Alberta:
http://www.transalta.com/website2001/tawebsite.nsf/AllDoc/1C13A445807845A4872569AE00737AF8?OpenDocument
The plants at Keephills consume 3.2 million tonnes of coal annually. The mass of sulfur burned is therefore approximately $0.005 \times 3.2 \times 10^6$ t $= 1.6 \times 10^4$ t. This produces a mass of $SO_2(g)$:

$1.6 \times 10^4 \times \frac{64}{32} = 3.2 \times 10^4$ t.

Sulfur dioxide gas can be removed in wet or dry scrubber towers to form sulfates that can be produced into by-products such as gypsum. Spraying the emission gases with ammonia removes $NO_x(g)$ as ammonium nitrate, and $SO_2(g)$ as ammonium sulfate.

41. **(a)** Acidic deposition is linked to sulfurous acid, $H_2SO_3(aq)$, sulfuric acid, $H_2SO_4(aq)$, nitrous acid, $HNO_2(aq)$, nitric acid, $HNO_3(aq)$.

(b) $SO_2(g) + H_2O(\ell) \rightleftharpoons H_2SO_3(aq)$

SO3(g) + H2O(ℓ) $\rightleftharpoons$ $H_2SO_4(aq)$

$2NO_2(g) + H_2O(\ell) \rightleftharpoons HNO_2(aq) + HNO_3(aq)$

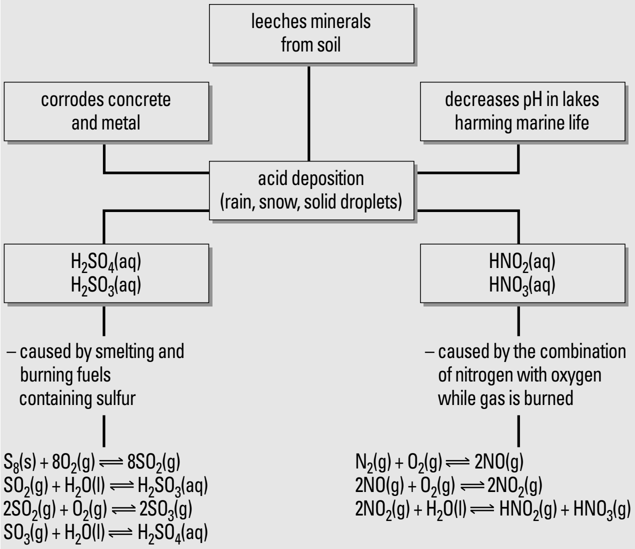
(c) $SO_x(g)$: the burning of fuels containing sulfur (especially coal), the smelting of ores (purifying metals), the removal of sulfur from natural gas and oil

$NO_x(g)$: as a byproduct in the burning of gasoline

(d) The effects of acidic deposition include:

- destruction of trees and other forms of land life
- leeching of minerals from the soil which can be toxic to marine life
- decreased pH in bodies of water leading to the death of marine life
- erosion of concrete structures
- increased rate of corrosion of metal

(e)



42. An amphiprotic substance can act either as a proton donor or a proton acceptor. The hydrogen carbonate ion is a good example. As a proton donor:

$HCO_3^-(aq) + OH^-(aq) \rightleftharpoons CO_3^{2-}(aq) + H_2O(\ell)$

As a proton acceptor:

$HCO_3^-(aq) + H_3O^+(aq) \rightleftharpoons H_2CO_3(aq) + H_2O(\ell)$

43. $HNO_2(aq) + OH^-(aq) \rightleftharpoons NO_2^-(aq) + H_2O(\ell)$

$NO_2^-(aq) + H_3O^+(aq) \rightleftharpoons HNO_2(aq) + H_2O(\ell)$

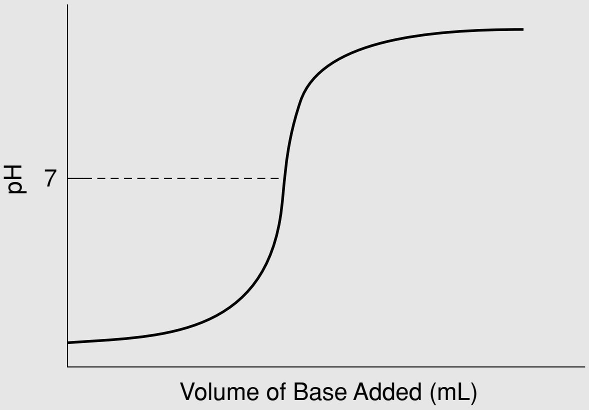
44. The endpoint in a titration is the point at which an indicator changes colour. The equivalence point in an acid-base titration is the point at which the number of moles of acid is equal to the number of moles of base. The equivalence point may occur before, after, or coincide with the endpoint, depending on the choice of indicator for a particular titration.

(a) The titration of a strong acid with a weak base will form an aqueous solution with pH < 7 at equivalence. An indicator that changes colour in the acidic range should be chosen.

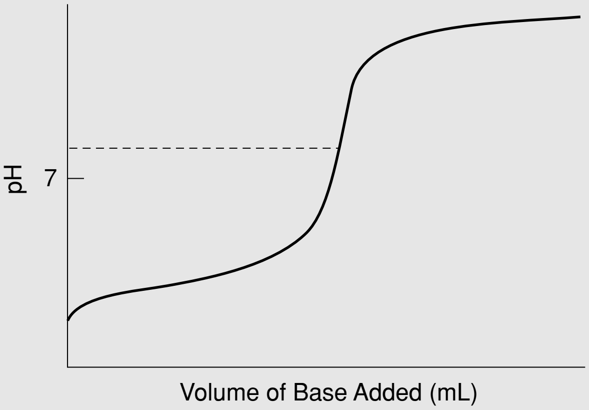
(b) A different indicator should be chosen in this case, because the solution will be basic at equivalence.

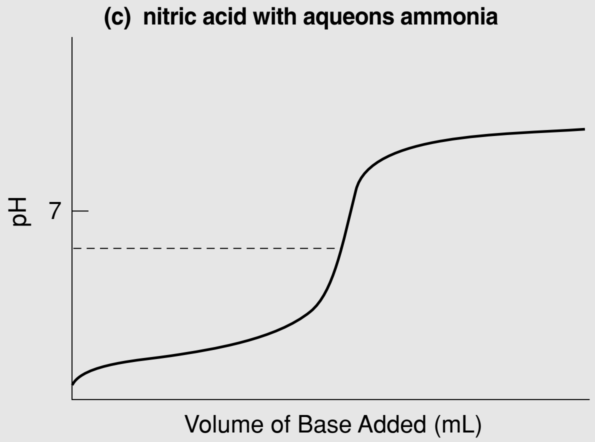
45.

(a) nitric acid with potassium hydroxide



(b) ethanoic acid with potassium hydroxide





(c) nitric acid with aqueons ammonia

46. A polyprotic acid will have two or more steep changes in pH corresponding to the number of protons the acid donates.

47. **(a)** The solution contained a weak acid based on the shape of the titration curve.

(b) The pH at equivalence was approximately 9.

(c) Yes, the solution acted as a buffer from pH 4 to pH 6.

(d) Phenolphthalein would be an appropriate indicator, since its endpoint is between pH 8.2 and pH 10.0.

(e) The solution is most likely ii, $CH_3COOH(aq)$, since this is the only weak acid in the list.

48. **OMITTED FROM STUDENT TEXT**

Answers to Making Connections Questions

49. The acidic rainwater might be caused by a number of factors, including a nearby coal-burning electric generating plant, a gas-processing facility, or even being downwind from a source of acidic emissions from far away. Students should outline the data to be collected, such as the pH of rain or snow, the date of the precipitation, and the wind direction and strength.

50. **(a)** Hydroxyapatite, is a weak base due to its hydroxy groups.

(b) $Ca_{10}(PO_4)_6(OH)_2(s) + 8H_3O^+(aq) \rightleftharpoons 6CaHPO_4(s) + 10H_2O(\ell) + 4Ca^{2+}(aq)$

(c) Bacteria in the mouth feed off the sugars in candy. The bacteria excrete acids and, as the equation shows, acids react with $CaHPO_4$ which causes teeth to decay.

51. ICE Table

concentration (mol/L)

$NH_2(CH_2)_4NH_2(aq) + H_2O(\ell) \rightleftharpoons OH^-(aq) + NH_2(CH_2)_4NH_3^+(aq)$

Initial	0.10		≈ 0	0
Change	-2.1×10^{-3}		$+ 2.1 \times 10^{-3}$	$+ 2.1 \times 10^{-3}$
Equilibrium	$0.10 - 2.1 \times 10^{-3}$ $= 0.098$		2.1×10^{-3}	2.1×10^{-3}

$$K_b = \frac{[NH_2(CH_2)_4NH_3^+(aq)][OH^-(aq)]}{[NH_2(CH_2)_4NH_2(aq)]}$$

$$= \frac{(2.1 \times 10^{-3})(2.1 \times 10^{-3})}{0.098}$$

$$K_b = 4.5 \times 10^{-5}$$

52. **(a)** $H_3VO_4(aq) + H_2O(\ell) \rightleftharpoons H_2VO_4^-(aq) + H_3O^+(aq)$

$H_2VO_4^-(aq) + H_2O(\ell) \rightleftharpoons HVO_4^{2-}(aq) + H_3O^+(aq)$

$HVO_4^{2-}(aq) + H_2O(\ell) \rightleftharpoons VO_4^{3-}(aq) + H_3O^+(aq)$

(b) As an acid:

$H_2VO_4^-(aq) + H_2O(\ell) \rightleftharpoons HVO_4^{2-}(aq) + H_3O^+(aq)$

As a base:

$H_2VO_4^-(aq) + H_2O(\ell) \rightleftharpoons H_3VO_4(aq) + OH^-(aq)$

(c) The stronger acid is $H_3VO_4(aq)$. In a polyprotic acid, acid strength always decreases as each proton is removed, because the next proton must be removed from an increasingly negatively charged ion.

53. The ammonium ion acts as an acid. From the Appendix, K_a for $NH_4^+(aq)$ is 5.6×10^{-10}. The hydrogen carbonate ion acts as a base. From the Appendix, K_a for $HCO_3^-(aq)$ is 4.7×10^{-11}.

$$K_b = \frac{K_w}{K_a} = \frac{1.0 \times 10^{-14}}{4.7 \times 10^{-11}} = 2.1 \times 10^{-4}$$

Because K_b for $HCO_3^-(aq)$ is greater than K_a for $NH_4^+(aq)$, a solution of $NH_4HCO_3(aq)$ will be slightly basic.