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UNIT 6

Electrochemical Changes

Teaching Unit 6: Electrochemical Changes

Chapter 12 *Oxidation-Reduction Reactions*

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

Chapter 13 *Cells and Batteries*

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

Teaching Unit 6: Electrochemical Changes

(20% of the course time; approximately 25 hours)

Student Textbook pages 428–529

General Outcomes

- explain the nature of oxidation-reduction reactions
- apply the principles of oxidation-reduction to electrochemical cells

Contents

Chapter 12 Oxidation-Reduction Reactions

Chapter 13 Cells and Batteries

Content Summary

Our ability to generate and use electrical current has had a profound effect on the way we live.

Students may be curious about the construction and properties of the many different types of cells and batteries they use. They should be aware that electrical energy is used to refine some metals and to deposit a metal coating such as chromium to improve the appearance of others. To understand such applications, students must extend their knowledge of chemical reactions.

Chemical reactions fall into two categories: those reactions in which there is no exchange of electrons among the atoms, ions, or molecules involved; and those reactions in which reactants *do* transfer electrons when they react. Unit 6 deals with the second category, redox reactions. Students begin by analyzing redox reactions and balancing redox equations. They learn the electrochemistry behind modern-day voltaic cells and batteries, fuel cells, and electrolytic cells used in industrial processes such as extracting and refining metals.

Chapter 12 introduces students to the modern view of oxidation and reduction. They learn to define and recognize redox reactions and to write equations for oxidation and reduction half-reactions. The activity series of metals greatly influenced the pace of civilization, and students will explore the connection between this series and the relative oxidizing and reducing strengths of metal atoms and ions in Investigation 12.A. It is often very difficult to balance equations for redox reactions by inspection. In Section 12.2, students learn the half-reaction method for balancing redox equations, including those that take place in acidic or basic solutions. Section 12.3 extends redox reactions to those that involve molecular compounds. Students learn how to assign oxidation numbers using Lewis structures and electronegativities, then how to use oxidation numbers to define redox reactions and balance their equations. In Section 12.4, students apply their knowledge of stoichiometry to redox reactions and perform a redox titration in Investigation 12.C.

In Chapter 13, students explore the relationship between redox reactions and their applications in technological processes. Students will learn to distinguish between voltaic cells and electrolytic cells and how each type relates to battery technology. In Section 13.1, they learn how a redox reaction can generate an electric current, and the components and terminology of the voltaic cell. Students will learn how to use standard reduction potentials to determine cell potentials. In Investigation 13.A, they check their predictions for the cell potential of voltaic cells they build. In Section 13.2, students learn about dry cells, and the differences between a primary battery and a secondary battery. Next, they learn about fuel cells and the potential benefits and challenges of this technology. Corrosion is shown to be the result of a voltaic cell, and students learn a number of ways to prevent corrosion, especially the rusting of iron. Section 13.3 begins by identifying the similarities and differences between voltaic cells and electrolytic cells. Students learn that standard reduction potentials cannot be used for predictions or calculations relating to the electrolysis of aqueous solutions. In Investigation 13.B, students build a cell to electrolyze an aqueous solution of a salt. In this experiment an overpotential is applied, and students use non-standard reduction potentials for water to help predict the products of the electrolysis. This section ends with applications of electrolytic cells to industry and in rechargeable batteries. In Section 13.4, students use their knowledge of stoichiometry and electricity to explore the relationship between time, current, and the amount of product produced or reactant consumed in an electrolytic process. Students will learn about Faraday's Law and explore its application in electroplating and in refining metals.

Curriculum Fit

This unit builds on:

- *Science 9,* Unit D: Electrical Principles and Technologies
- *Science 10,* Unit A: Energy and Matter in Chemical Change
- *Chemistry 20,* Unit B: Chemical Bonding
- *Chemistry 20,* Unit D: Quantitative Relationships in Chemical Changes

Core Concepts

Concept	Outcome	Text Reference
Many chemical reactions involve electrochemical changes with important applications in medicine, industry and the environment.	30–B1.4k 30–B1.2sts	Unit 6 opener, p. 430; Chapter 12 opener, p. 434; Chapter 12 Launch Lab, p. 435; Section 12.1, p. 436; Section 12.2, p. 443; Section 12.2, pp. 454-456; Section 12.3, p. 457; Section 12.3, p. 467; Section 12.4, p. 469. Chapter 13 opener, p. 478; Section 13.2, pp. 493-497; Section 13.2, pp. 499-502; Section 13.3, p. 508; Section 13.3, pp. 512-514; Section 13.4, p. 521; Career Focus, pp. 526-527.
Oxidation is the loss of electrons; reduction is the gain of electrons.	30–B1.1k	Section 12.1, pp. 436- 437
A reactant that oxidizes another reactant is called an oxidizing agent; a reactant that reduces another reactant is called a reducing agent.	30–B1.2k	Section 12.1, pp. 438- 439
The metal activity series is a listing of the metals from most reactive to least reactive. This is the order of their ability to be oxidized during reactions in aqueous solution.	30–B1.5k	Section 12.1, pp. 439- 440
The metals first used by people to make ornaments and tools are those lowest on the activity series.	30–B1.1sts	Section 12.2, pp. 454-456
The oxidation half reaction shows the part of the redox reaction in which there is a loss of electrons by a species; the reduction half reaction shows the part of the redox reaction in which there is a gain of electrons by a species.	30–B1.3k	Section 12.1, pp. 443- 445
The chemical equation for a redox reaction is the sum of the two half reactions when written so that the electrons involved cancel out.	30–B1.7k	Section 12.1, pp. 443- 445
In a disproportionation reaction, some atoms of an element are oxidized and other atoms of the same element are reduced.	30–B1.2k	Section 12.2, pp. 452- 454
The oxidation number (or state) is the actual charge on an atom if it exists as a monatomic ion, or the charge on an atom in a molecule or ion if all the electrons in its bonds were completely held by the atom with the greatest electronegativity.	30–B1.2k	Section 12.3, pp. 457- 458
During a redox reaction, the oxidation number of at least one atom must increase, and the oxidation number of at least one atom must decrease.	30–B1.1k	Section 12.3, pp. 461- 462
Oxidation is an increase in oxidation number; reduction is a decrease in oxidation number.	30–B1.1k 30–B1.2k	Section 12.3, p. 462
When a redox reaction takes place, the algebraic sum of all the changes in oxidation numbers of the atoms in the reaction is zero.	30–B1.7k	Section 12.3, p. 464
In a voltaic cell, spontaneous redox reactions transform chemical potential energy into electrical energy.	30–B2.1k	Section 13.1, pp. 480-481
A half-cell is the portion of a voltaic cell in which a half-reaction (either oxidation at the anode or reduction at the cathode) takes place.	30–B2.1k	Section 13.1, p. 481

Concept	Outcome	Text Reference
The electric potential difference between two points in a circuit is the quantity of work done to move unit charge between the two points. The cell potential for any given cell is the electrical potential difference between the electrodes.	30–B2.6k	Section 13.1, p. 485
The half-cell potential of the hydrogen half-cell is defined as zero. The half-cell potential of all other cells is defined as the potential difference between the hydrogen half-cell and the chosen cell.	30–B2.5k	Section 13.1, p. 485; Thought Lab 13.1, section 13.1, p. 489
The standard reduction potential, E°, is the reduction potential of a half-reaction at 25 °C and one atmosphere, when all ion concentrations are 1 mol/L.	30–B2.3k	Section 13.1, pp. 486-487
The standard cell potential, E°cell, can be calculated using a table of standard reduction potentials: E°cell = E°cathode - E°anode.	30–B2.6k	Section 13.1, pp. 487-489
A battery consists of a set of voltaic cells connected in series.	30–B2.1sts	Section 13.2, p. 493
The electrical potential difference across a battery is the sum of the cell potentials of the individual cells in series.	30–B2.1sts	Section 13.2, p. 493
A primary cell (or battery) must be disposed once it has been discharged; a secondary (or storage) cell is designed to be recharged.	30–B2.1sts	Section 13.2, pp. 493-495; Section 13.3, pp. 513-514
A fuel cell is a battery that operates with a continuous supply of reactants.	30–B2.3sts	Section 13.2, pp. 495-498
Corrosion is a spontaneous redox reaction of a metal with substances in the environment.	30–B1.4k 30–B2.2sts	Section 13.2, pp. 499-501
A metal will not corrode if it is in contact with a more reactive metal.	30–B2.2sts	Section 13.2, pp. 501-502
An electrolytic cell uses an external source of electrical energy to drive a non-spontaneous redox reaction.	30–B2.1k 30–B2.2k	Section 13.3, pp. 504-508; Section 13.3, pp. 513-514
The calculated value for the cell potential to drive an electrolytic cell is always lower than the actual potential.	30–B2.3s 30–B2.4k	Section 13.3, pp. 504-508
The SI unit of electric current is the ampere (A), which is equivalent to one coulomb of charge passing each point in a conductor every second (A = C/s).	30–B2.8k	Section 13.4, p. 516
The amount of a substance produced or consumed in an electrolysis reaction is directly proportional to the amount of charge that flows through the circuit. (Faraday's law)	30–B2.8k	Section 13.4, pp. 516-519

Related Skills

Skills	Outcome	Text Reference
Students must be able to describe procedures for safe handling, storage and disposal of materials used in the laboratory, with reference to WHMIS and consumer product labelling information.	30–B1.1s	Chapter 12 Launch Lab, p. 435; Investigation 12.A, Section 12.1, pp. 440-441; Investigation 12.B, Section 12.2, pp. 450-451 Investigation 12.C, Section 12.4, pp. 472-473; Investigation 13.A, Section 13.1, pp. 490-491; Investigation 13.B, Section 13.3, pp. 510-511 Investigation 13.C, Section 13.4, pp. 519-520

Skills	Outcome	Text Reference
Students will perform single replacement reactions to identify the oxidizing agent and the reducing agent in each reaction, and will use the data to derive a simple redox table.	30–B1.5k 30–B1.3s	Investigation 12A, Section 12.1, pp. 440-441
Given the metal activity series, students will predict whether a reaction between atoms of one element and ions of another will undergo a spontaneous redox reaction.	30–B1.6k 30–B2.7k	Section 12.1, pp. 441-442
Students will predict whether certain metals can be oxidized by aqueous hydrogen ions in an acidic solution, and experimentally test their predictions.	30–B1.5k 30–B1.7k	Investigation 12B, Section 12.2, pp. 450-451
Students will write and balance equations for redox reactions using the half reaction method for reactions in neutral, acidic, or basic solution.	30–B1.7k	Section 12.1, pp. 443-450; Section 12.2, pp. 453- 454
Students will determine the oxidation numbers of atoms in molecules or ions.	30–B1.7k	Section 12.3, pp. 458- 460; Thought Lab 12.1, Section 12.3, pp. 460-461
Students will assign oxidation numbers to the atoms in a reaction to determine whether it is a redox reaction	30–B1.3k	Section 12.3, p. 463
Students will balance chemical equations using changes in the oxidation states of the atoms in the reaction.	30–B1.7k	Section 12.3, pp. 464-466
Students will solve stoichiometry problems that involve redox titrations	30–B1.7k 30–B1.8k	Section 12.4, pp. 469-471
Students will perform a redox titration experiment, first standardizing one solution, then determining the concentration of the second solution.	30–B1.2s 30–B1.4s	Investigation 12.C, Section 12.4, pp. 472-473.
Students will construct a simple voltaic cell and investigate the relationship between the voltage of cells connected in series, and cells connected in parallel.		Chapter 13 Launch Lab, p. 479.
Students will be able to define and correctly label each of the following parts of a voltaic cell: anode, cathode, anion, cation, salt bridge/porous cup, electrolyte, internal circuit, direction of ion flow, external circuit, direction of electron flow	30–B2.1k	Section 13.1, pp. 481-482
Students will use the shorthand method (cell diagram) to write the components of voltaic cells.	30–B2.1k	Section 13.1, pp. 483-484
Given a table of standard reduction potentials, students will calculate the standard cell potential for a voltaic cell.	30–B2.3k 30–B2.6k	Section 13.1, p. 489; Investigation 13.A, Section 13.1, pp. 490-491.
Students will construct some voltaic cells, measure their cell potentials, and use the results to derive a simple reduction table.	30–B2.1s 30–B2.2s	Investigation 13.A, Section 13.1, pp. 490-491.
Students will predict the products of the electrolysis of: ■ an aqueous solution of a salt ■ a molten salt	30–B1.6k 30–B2.3k	Section 13.3, pp. 509-512.
Students will construct electrolytic cells and qualitatively and quantitatively check their predictions for the products of the electrolysis.	30–B2.1s 30–B2.3s 30–B2.4s	Investigation 13.B, Section 13.3, pp. 510-511. Investigation 13.C, Section 13.4, pp. 519-520
Students will calculate mass, amounts, current and time in single voltaic and electrolytic cells by applying Faraday's law and stoichiometry.	30–B2.8k	Section 13.4, p. 518; Investigation 13.C, Section 13.4, pp. 519-520

Activities and Target Skills

Activity	Target Skills
Chapter 12: Oxidation-Reduction Reactions	
Launch Lab: Penny Chemistry, p. 433	■ Using common items to clean tarnished pennies ■ Infer the chemical reactions involved in the change in the pennies
Investigation 12.A: Testing Relative Oxidizing and Reducing Strengths of Metal Atoms and Ions, p. 438	■ Determining the reactivity of various metals ■ Analyzing data to derive a simple reduction table ■ Using appropriate numeric and symbolic modes of representation to communicate equations
Investigation 12.B: Redox Reactions and Balanced Equations, pp. 448-450	■ Describing procedures for the safe handling and disposal of acids and products of hydrocarbon combustion ■ Analyzing data from an experiment on redox reactions ■ Using appropriate numeric and symbolic representations to communicate equations for redox reactions
Thought Lab 12.1: Oxidation Numbers and Lewis Structures, pp. 458-459	■ Analyze Lewis structures and use electronegativities to determine oxidation numbers for atoms in compounds
Investigation 12.C: Measuring the Concentration of Vitamin C in Orange Juice, pp. 470-471	■ Standardizing a solution of a oxidizing agent ■ Performing a titration to determine the concentration of a reducing agent
Chapter 13: Cells and Batteries	
Launch Lab: What Determines Voltage?, p. 477	■ Testing the characteristics of lemon batteries ■ Comparing cells and batteries
Thought Lab 13.1: Assigning Reference Values, p. 487	■ Explaining that values of standard reduction potentials are all relative
Investigation 13.A: Measuring Cell Potentials of Voltaic Cells, pp. 488-489	■ Evaluating data from an experiment to derive a simple reduction table ■ Interpreting patterns and trends in data derived from redox reactions ■ Constructing and observing electrochemical cells
Investigation 13.B: Electrolysis of Aqueous Potassium Iodide, pp. 508-509	■ Identifying the products of an electrolysis cell ■ Evaluating experimental designs for an electrolytic cell and suggesting alternatives ■ Using appropriate SI notation and fundamental and derived units to communicate answers to problems related to functioning electrolytic cells
Investigation 13.C: Electroplating, pp. 517-518	■ Describing procedures for safe handling and disposal of materials used in the laboratory ■ Identifying limitations of data collected ■ Comparing predictions with observations of electrochemical cells

Conceptual Challenges

Chapter 12 Oxidation-Reduction Reactions

- Remind students that reductions and oxidations do not occur in isolation. Although we write and manipulate half-reactions to help balance chemical equations, half-reactions do not occur on their own. A reduction must be accompanied by an oxidation. Point out that a half-reaction equation represents just that: half a reaction. For example, use the term "oxidation half-reaction," rather than "oxidation reaction," and insist on similar language from the students. Similarly, students may believe that half-cell potentials can be used to predict the spontaneity of half-cell reactions. Half-cell reactions do not occur by themselves.
- Ensure students are clear about the difference between identifying reducing agents or oxidizing agents and identifying the atom in a compound that is reduced or oxidized. Emphasize that an agent causes change so that, for example, a reducing agent will cause another species to be reduced. It is customary to identify the oxidizing or reducing agent as whatever chemical species contains the relevant element. If equations are written in molecular form, the chemical species is a compound. If equations are written in ionic form, the chemical species may be a compound or an ion.

Chapter 13 Cells and Batteries

- The chemistry definitions of anode and cathode differ from those learned in physics in describing the photoelectric effect. Ensure students have a solid understanding that the anode is the site of oxidation and the cathode is the site of reduction.
- Understanding why the predicted reaction is not always the reaction that occurs (e.g., the electrolysis of brine) is difficult for students. When teaching the topic of spontaneity and predicting reactions, ensure students understand that all electrochemical reactions MAY take place and that the table of standard electrode potentials helps to predict the likely reaction. Extend this concept of likelihood to non-spontaneous reactions when discussing the chlor-alkali process.

Using the Unit 6 Preparation Feature

The opener for Unit 6 uses the example of the heart pacemaker to begin the discussion about the uses of electrical energy. The Focussing Questions: "What is an electrochemical change?" And "How have scientific knowledge and technological innovation been integrated in the field of electrochemistry?" are designed to get students thinking about the connection between chemical reactions and electrical energy and about the chemical reactions that take place inside batteries.

The Unit Preparation feature has been included in order to ensure that students are familiar with the science from previous courses that relates specifically to the material they are about to study. Encourage students to take the Unit Prequiz (found at **www.albertachemistry.ca**, Online Learning Centre, Student Edition) to gauge their recall, noting that if they are familiar with the background science, their experience with this unit will be much easier.

The Unit 6 Preparation feature is brief, but it is important that students recall their knowledge of types of chemical reactions and how to determine subscripts for ionic compounds, and review types of equations and the concept of electronegativity. Students who have difficulty with the Unit Prequiz should use this section to ensure they have the necessary chemistry background for the unit. **BLM 12.0.1 (AST) Reviewing the Periodic Table** has been provided to help students with their recall. It can be found on the CD-ROM that accompanies this Teacher's Resource or at **www.albertachemistry.ca**, Online Learning Centre, Instructor Edition (password required).

UNIT 6: COURSE MATERIALS

Chapter, Section	Item Description	Suggested Quantity (assume 40 in class)	Text Activity
Chapters 12, 13	safety goggles	40 pairs × 7 investigations	Chapter 12 Launch Lab; Investigations: 12.A, 12.B, 12.C, 13.A, 13.B, 13.C
Chapters 12, 13	nonlatex disposable gloves	40 pairs × 5 investigations	Investigations: 12.A, 12.B, 12.C, 13.A, 13.C
Chapters 12, 13	aprons	40 × 6 investigations	Chapter 12 Launch Lab; Investigations: 12.A, 12.B, 12.C, 13.A, 13.C
Chapter 12, Chapter Opener	white vinegar	65 mL per group	Launch Lab: Penny Chemistry, p. 433
	table salt	1 tsp per group	
	measuring cup (to measure 1/4 cup)	1 per group	
	measuring spoons (to measure 1 tsp)	1 per group	
	stirring rod	1 per group	
	small, clear glass or plastic bowl	1 per group	
	dull pennies	6 per group	
	paper towel	5 rolls	
	marker	1 per group	
	clean steel screw or nail	1 per group	
Chapter 12, Section 12.1	small pieces of each of the following metals:	4 pieces per group of each metal	Investigation 12.A: Testing Relative Oxidizing and Reducing Strengths of Metal Atoms and Ions, p. 438
	- aluminium foil		
	- thin copper wire or tiny copper beads		
	- iron filings		
	- magnesium		
	- zinc		
	dropper bottles containing dilute solutions of	5 dropper bottles of each	
	- aluminium sulfate		
	- copper(II) sulfate		
	- iron(II) sulfate		
	- magnesium sulfate	1 per group	
	- zinc nitrate	1 per group	
	well plate		
	white paper		
	steel wool (optional)		
Chapter 12, Section 12.2	small pieces of each of the following metals:	3–4 pieces of each per group	Investigation 12.B: Redox Reactions and Balanced Equations, pp. 448–449
	- magnesium		
	- zinc		
	- copper		
	- aluminium		
	hydrochloric acid (1 mol/L) [Note: corrosive and poisonous]	10 mL per group	
	sulfuric acid (1 mol/L) [Note: corrosive and poisonous]	10 mL per group	
	well plate	1 per group	
	small test tubes (optional)	8 per group (if repeat test)	
	test tube rack (optional)		
	Bunsen burner	1 per group	
	candle	1 per group	
	steel wool (optional)	1 per group	
	micro-pipette or eyedropper (optional)	2 per group	
	grease pencil (optional)	1 per group	

Chapter, Section	Item Description	Suggested Quantity (assume 40 in class)	Text Activity
Chapter 12, Section 12.4	iodine solution [Note: 2.59 g of iodine and 26 g of potassium iodine in 50 mL of water; when dissolved, dilute to 6 L] [Note: skin irritant] 5.68 mmol/L ascorbic acid standard solution dropper bottle containing starch indicator [0.2% W/V starch in water] orange juice, pulp removed [Note: Orange juice with pulp can be used but must be measured with a graduated cylinder as the pulp clogs the pipette and is difficult to remove. Sunny Delight works well and also has the concentration listed on the label, whereas most Canadian brands do not.] deionized water burette burette clamp ring stand meniscus reader volumetric pipette (10.00 mL) suction bulb 125 mL Erlenmeyer flasks 250 mL beakers labels sheet of white paper funnel sodium thiosulfate	6 L 500 mL 1 per group or 3–5 to share among class 1.5 L as required 1 per group 1 per group 1 per group 1 per group 1 per group 1 per group 3 per group 3 per group 3 per group 1 per group 1 per group should be available to remove any stains that occur	Investigation 12.C: Measuring the Concentration of Vitamin C in Orange Juice, pp. 470–471
Chapter 13, Chapter Opener	zinc strips (1 cm × 5 cm) copper strips (1 cm × 5 cm) lemons fine sandpaper electrical leads with alligator clips voltmeter (set to scale of 0 V to 5 V) small flashlight bulb	4 per group 4 per group 4 per group 1 sheet per group 8 per group 1 per group 1 per group	Launch Lab: What Determines Voltage?, p. 477

Chapter, Section	Item Description	Suggested Quantity (assume 40 in class)	Text Activity
Chapter 13, Section 13.1	Styrofoam™ or clear plastic egg carton with 12 wells (or well plate) 0.1 mol/L solutions of each of the following: – $Mg(NO_3)_2$(aq) – $Cu(NO_3)_2$(aq) – $Al(NO_3)_3$(aq) – $Ni(NO_3)_2$(aq) – $Zn(NO_3)_2$(aq) – $SnSO_4$(aq) – $Fe(NO_3)_3$(aq) – $AgNO_3$(aq) – HNO_3(aq) [Note: corrosive and oxidizing] saturated NaCl(aq) solution 5 cm strip of Mg ribbon 1 cm × 5 cm strips of each of the following: -Cu(s), Al(s), Ni(s), Zn(s), Sn(s), Fe(s), and Ag(s) 1.0 mol/L KNO_3 thick graphite pencil lead or graphite rod (or graphite electrodes) sandpaper clear aquarium rubber tubing (Tygon®; internal diameter: 4–6 mm) cotton batting disposable pipette black and red electrical leads with alligator clips voltmeter set to a scale of 0 V to 20 V paper towel marker	1 per student 5 mL per group of each solution 5 mL per group 1 per group 1 per group of each 15 mL per group 5 cm length per group 1 sheet per group 25 cm length per group 1 roll 1 per group 1 per group 40 5 rolls 5	Investigation 13.A: Measuring Cell Potentials of Voltaic Cells, pp. 488–489
Chapter 13, Section 13.3	1 mol/L KI 20 gauge (or 14 gauge) platinum wire graphite pencil lead, 2 cm long (or carbon electrode) 1% starch solution 1% phenolphthalein [Note: flammable] white paper beaker (600 mL or 400 mL) elastic band clear aquarium rubber tubing (Tygon®; internal diameter: 4–6 mm) disposable pipettes toothpicks wire leads (black and red) with alligator clips 9-V battery or variable power source set to 9 V	10 mL per group 2 cm per group 1 per group 1 drop per group 1 drop per group 1 sheet per group 1 per group 1 per group 25 cm per group 3 per group 3 per group 2 per group 1 per group	Investigation 13.B: Electrolysis of Aqueous Potassium Iodide, pp. 508–509

Chapter, Section	Item Description	Suggested Quantity (assume 40 in class)	Text Activity
Chapter 13, Section 13.4	3 cm × 12 cm × 1 mm Cu strip 1.0 mol/L HNO_3 in a 250 mL beaker [Note: corrosive and oxidizing] deionized water in a wash bottle 16-gauge bare solid copper wire acidified 0.50 mol/L $CuSO_4$ solution with: -6 mol/L H_2SO_4 [Note: corrosive and poisonous] -0.1 mol/L HCl [Note: corrosive and poisonous] [Note: to each 120 mL of 0.50 mol/L $CuSO_4$ solution add 5 mL of 6 mol/L H_2SO_4 and 3 mL of 0.1 mol/L HCl] fine sandpaper 250 mL beaker electrical leads with alligator clips adjustable D.C. power supply with ammeter drying oven electronic balance	1 per group 150 mL per group 20 50 cm per group 120 mL per group 5 mL per 120 mL $CuSO_4$ solution 3 mL per 120 mL $CuSO_4$ solution 1 sheet per group 1 per group 2 per group 1 per group 1 4	Investigation 13.C: Electroplating, pp. 517–518

CHAPTER 12 OXIDATION-REDUCTION REACTIONS

Curriculum Correlation

(Note: This correlation includes Chapters 12 and 13. Chapter 12 references are in bold.)

General Outcome General Outcome 1: Students will explain the nature of oxidation–reduction reactions.

	Student Textbook	Assessment Options
Outcomes for Knowledge		
30–B1.1k define oxidation and reduction operationally and theoretically	**Oxidation, Section 12.1, pp. 434–435** **Reduction, Section 12.1, p. 435** **Redox Reactions, Section 12.1, pp. 436–437**	**Questions for Comprehension: 1–4, Section 12.1, p. 437** **Section 12.1 Review: 1, 2, p. 440** **Chapter 12 Review: 4, p. 474** **Chapter 12 Test** Unit 6 Review: 1, 4, 6, 33, pp. 526–529
30–B1.2k define the following terms: oxidizing agent, reducing agent, oxidation number, half–reaction, disproportionation	**Reducing Iron Ore, Section 12.2, pp. 452–454** **Sample Problem: Assigning Oxidation Numbers, Section 12.3, pp. 457–458** **Balancing Equations Using the Oxidation Number Method, Section 12.3, p. 462**	**Questions for Comprehension: 1–4, Section 12.1, p. 437** **Section 12.1 Review: 1, 2, p. 440** **Chapter 12 Test** Unit 6 Review: 1, 4, 6, 33, pp. 526–529
30–B1.3k differentiate between oxidation–reduction (redox) reactions and other reactions using half–reactions and oxidation numbers	**Balancing Equations Using Half–Reactions, Section 12.1, pp. 443–445** **Balancing a Redox Equation in Basic Solution, Section 12.3, p. 463**	**Questions for Comprehension: 9–11, Section 12.2, p. 442** **Chapter 12 Test** Unit 6 Review: 1, 4–7, 11, 26, 29, 33–36, 44, | pp. 526–529
30–B1.4k identify electron transfer, oxidizing agents and reducing agents in redox reactions that occur in everyday life in both living and non–living systems, e.g., corrosion, *cellular respiration, photosynthesis*	**Redox Reactions, Section 12.1, p. 436** **Balancing Equations Using Half–Reactions, Section 12.2, p. 443** **Redox Reactions Involving Molecular Compounds, Section 12.2, pp. 454–456** **Sample Problem: Assigning Oxidation Numbers, Section 12.3, p. 457** **Quantitative Analysis of Redox Reactions, Section 12.4, p. 467–469** Fuel Cell History, Section 13.2, pp. 493–497 Corrosion Prevention, Section 13.2, pp. 499–502	**Questions for Comprehension: 3, Section 12.1, p. 437** **Section 12.1 Review: 7, p. 440** **Practice Problems: 1, 2, Section 12.2, p. 448** **Practice Problems: 12, 13, Section 12.3, p. 461** **Section 12.3 Review: 6, p. 466** **Chapter 12 Review: 1, 2, 5, 10, 21, pp. 474–475** **Chapter 12 Test** Questions for Comprehension: 4, 6, Section 13.1, p. 482 Section 13.2 Review: 1, p. 501 Chapter 13 Review: 7, 8, 16, pp. 522–523 Unit 6 Review: 2, 27, 36, 37, 44, 46, pp. 526–529
30–B1.5k compare the relative strengths of oxidizing and reducing agents from empirical data	**Investigation 12.A: Testing Relative Oxidizing and Reducing Strengths of Metal Atoms and Ions, Section 12.1 p. 438** **Predicting the Spontaneity of Redox Reactions, Section 12.1, pp. 439–440**	**Investigation 12.A: Testing Relative Oxidizing and Reducing Strengths of Metal Atoms and Ions: 1–4, Section 12.1 p. 438** **Questions for Comprehension: 5–8, Section 12.1, p. 440** **Chapter 12 Test**

	Student Textbook	Assessment Options
30–B1.6k predict the spontaneity of a redox reaction based on standard reduction potentials, and compare predictions to experimental results	**Spontaneity of Redox Reactions, Section 12.1, p. 437**	
	Predicting the Spontaneity of Redox Reactions, Section 12.1, p. 439	**Questions for Comprehension: 5–8, Section 12.1, p. 440** **Section 12.1 Review: 5, 6, p. 440** **Chapter 12 Test** Chapter 13.3 Review: 3, p. 513 Chapter 13 Review: 11, pp. 522–523 Unit 6 Review: 3, 25, 38, pp. 526–529
30–B1.7k write and balance equations for redox reactions in acidic, basic and neutral solutions, including disproportionation reactions, by ■ using half–reaction equations obtained from a standard reduction potential table	**Writing Balanced Half–Reactions, Section 12.2, pp. 441–442** **Balancing Equations Using Half–Reactions, Section 12.2, pp. 442–443** **Balancing Equations for Reactions That Occur in Acidic or Basic Solutions, Section 12.2, pp. 444–445**	**Section 12.1 Review: 5, p. 440** **Questions for Comprehension: 9–11, Section 12.2, p. 442**
	Sample Problem: Balancing an Equation for a Reaction That Occurs in an Acidic Solution, Section 12.2, p. 446 **Sample Problem: Balancing an Equation for a Reaction That Occurs in an Basic Solution, Section 12.2, pp. 447–448** **Disproportionation Reactions, Section 12.2, p. 450**	**Practice Problems: 1, 2, Section 12.2, p. 448**
	Sample Problem: Balancing a Disproportionation Reaction, Section 12.2, p. 451	**Practice Problems: 3–6, Section 12.2, p. 452** **Section 12.2 Review: 1–5, p. 454** **Section 12.3: 6, p. 466** **Practice Problems: 11–14, Section 12.3, p. 461** **Practice Problems: 22, Section 12.3, p. 469** **Chapter 12 Review: 1, 9, 10, 14–22, pp. 474–475** **Chapter 12 Test** Section 13.2 Review: 16, p. 501 Questions for Comprehension: 17, Section 13.3, p. 510 Section 13.4 Review: 6, p. 520 Chapter 13 Review: 6, p. 522 Unit 6 Review: 1, 8, 30, 34, 35, 37, 39, 46, p. 526–529
■ equations from information provided about redox changes **Note:** Students are expected to add water molecules, hydrogen ions and hydroxide ions to skeleton equations, as appropriate.	**Writing Balanced Half–Reactions, Section 12.2, pp. 441–442** **Balancing Equations Using Half–Reactions, Section 12.2, pp. 442–443** **Balancing Equations for Reactions That Occur in Acidic or Basic Solutions, Section 12.2, pp. 444–445** **Sample Problem: Balancing an Equation for a Reaction That Occurs in an Acidic Solution, Section 12.2, p. 446** **Sample Problem: Balancing an Equation for a Reaction That Occurs in an Basic Solution, Section 12.2, pp. 447–448**	**Questions for Comprehension: 9–11, Section 12.2, p. 442** **Practice Problems: 1, 2, Section 12.2, p. 448** **Section 12.2 Review: 1–5, p. 454** **Chapter 12 Review: 1, 14, pp. 474–475** **Chapter 12 Test** Questions for Comprehension: 5, 6, Section 13.1, p. 482 Practice Problems: 1–4, Section 13.1, p. 487 Questions for Comprehension: 13, Section 13.2, p. 500 Section 13.2 Review: 3, 8, 16, p. 501 Practice Problems: 5, Section 13.3, p. 508 Questions for Comprehension: 17, 18, Section 13.3, p. 510 Section 13.3 Review: 4, p. 513 Chapter 13 Review: 3, 6, 11, 13, 29, pp. 522–523 Unit 6 Review: 7, 11, 34, 36, 44, p. 526–529

	Student Textbook	Assessment Options
30–B1.7k (continued) ■ assigning oxidation numbers, where appropriate, to the species undergoing chemical change	**Assigning Oxidation Numbers, Section 12.3, pp. 455–456** **Sample Problem: Assigning Oxidation Numbers, Section 12.3, p. 457** **Applying Oxidation Numbers to Redox Reactions, Section 12.3, pp. 459–460**	**Practice Problems: 7–10, Section 12.3, p. 457** **Practice Problems: 15–18, Section 12.3, p. 464** **Section 12.3: 5–10, p. 466** **Chapter 12 Review: 3, 6, 7, 15, 20, pp. 474–475** **Chapter 12 Test** Unit 6 Review: 5–7, 35, pp. 526–529
30–B1.8k perform calculations to determine quantities of substances involved in redox titrations.	**Stoichiometry and Redox Titrations, Section 12.4, pp. 467–468** **Sample Problem: Redox Titrations, Section 12.4, pp. 468–469** **Investigation 12.C: Measuring the Concentration of Vitamin C in Orange Juice, Section 12.4, pp. 470–471**	**Practice Problems: 19–22, Section 12.4, p. 469** **Investigation 12.C: Measuring the Concentration of Vitamin C in Orange Juice: 1–8, Section 12.4, pp. 470–471** **Section 12.4 Review: 1–5, p. 472** **Chapter 12 Test** Unit 6 Review: 39, 40, pp. 526–529
Outcomes for Science, Technology and Society (Emphasis on science and technology)		
30–B1.1sts explain how the goal of technology is to provide solutions to practical problems by ■ *describing how the process of trial and error was used by early peoples to extract metals from their ores*	**Chapter 12 Launch Lab: Penny Chemistry, p. 433** **Investigation 12.A: Testing Relative Oxidizing and Reducing Strengths of Metal Atoms and Ions, Section 12.1, p. 438** **Connections: How Green Is White Paper? Section 12.3, p. 465**	**Chapter 12 Launch Lab: Penny Chemistry: 1–6, p. 433** **Investigation 12.A: Testing Relative Oxidizing and Reducing Strengths of Metal Atoms and Ions: 4, Section 12.1, p. 438** **Connections: How Green Is White Paper? 1–3, Section 12.3, p. 465** Unit 6 Review: 45–48, pp. 526–529
30–B1.2sts explain that technological problems often lend themselves to multiple solutions that involve different designs, materials and processes and have intended and unintended consequences by ■ *analyzing redox reactions used in industry and commerce, e.g., pulp and paper, textiles, water treatment, food processing.*	**Investigation 12.B: Redox Reactions and Balanced Equations, Section 12.2, pp. 448–450** **Connections: How Green Is White Paper? Section 12.3, p. 465**	**Investigation 12.B: Redox Reactions and Balanced Equations: 1–19, Section 12.2, pp. 448–450** **Connections: How Green Is White Paper? 1–3, Section 12.3, p. 465** Unit 6 Review: 45–48, pp. 526–529
Skill Outcomes (Focus on problem solving)		
Initiating and Planning		
30–B1.1s ask questions about observed relationships and plan investigations of questions, ideas, problems and issues by ■ designing an experiment to determine the reactivity of various metals ■ describing procedures for safe handling, storage and disposal of materials used in the laboratory, with reference to WHMIS and consumer product labelling information	**Investigation 12.B: Redox Reactions and Balanced Equations, Section 12.2, pp. 448–450**	**Investigation 12.B: Redox Reactions and Balanced Equations, Section 12.2: 1–19, pp. 448–450**

	Student Textbook	Assessment Options
Performing and Recording		
30–B1.2s conduct investigations into relationships among observable variables and use a broad range of tools and techniques to gather and record data and information by ■ selecting and correctly using the appropriate equipment to perform a redox titration experiment ■ using a standard reduction potential table as a tool in predicting the spontaneity of redox reactions and their products ■ *creating charts, tables or spreadsheets which present the results of redox experiments.*	**Investigation 12.C: Measuring the Concentration of Vitamin C in Orange Juice, Section 12.4, pp. 470–471** **Investigation 12.A: Testing Relative Oxidizing and Reducing Strengths of Metal Atoms and Ions, Section 12.1, p. 438**	**Investigation 12.C: Measuring the Concentration of Vitamin C in Orange Juice: 1–8, Section 12.4, pp. 470–471** **Investigation 12.A: Testing Relative Oxidizing and Reducing Strengths of Metal Atoms and Ions: 1–4, Section 12.1, p. 438** **Investigation 12.A: Testing Relative Oxidizing and Reducing Strengths of Metal Atoms and Ions: 3, Section 12.1, p. 438**
Analyzing and Interpreting		
30–B1.3s analyze data and apply mathematical and conceptual models to develop and assess possible solutions by ■ evaluating data from an experiment to derive a simple reduction table ■ *interpreting patterns and trends in data derived from redox reactions* ■ *identifying the limitations of data collected.*	**Investigation 12.A: Testing Relative Oxidizing and Reducing Strengths of Metal Atoms and Ions, Section 12.1, p. 438** **Investigation 12.B: Redox Reactions and Balanced Equations, Section 12.2, pp. 448–450** **Thought Lab 12.1: Oxidation Numbers and Lewis Structures, Section 12.3, pp. 458–459**	**Investigation 12.A: Testing Relative Oxidizing and Reducing Strengths of Metal Atoms and Ions: 3, Section 12.1, p. 438** **Investigation 12.B: Redox Reactions and Balanced Equations: 1–19, Section 12.2, pp. 448–450** **Thought Lab 12.1: Oxidation Numbers and Lewis Structures, Section 12.3, pp. 458–459**
Communication and Teamwork		
30–B1.4s work collaboratively in addressing problems and apply the skills and conventions of science in communicating information and ideas and in assessing results by ■ selecting and using appropriate numeric, symbolic, graphic and linguistic modes of representation to communicate equations for redox reactions and answers to problems related to redox titrations	**Investigation 12.B: Redox Reactions and Balanced Equations, Section 12.2, pp. 448–450** **Thought Lab 12.1: Oxidation Numbers and Lewis Structures, Section 12.3, pp. 458–459** **Investigation 12.C: Measuring the Concentration of Vitamin C in Orange Juice, Section 12.4, pp. 470–471**	**Investigation 12.B: Redox Reactions and Balanced Equations, Section 12.2: 1–19, pp. 448–450** **Thought Lab 12.1: Oxidation Numbers and Lewis Structures, Section 12.3, pp. 458–459** **Investigation 12.C: Measuring the Concentration of Vitamin C in Orange Juice: 1–8, Section 12.4, pp. 470–471**

General Outcome 2: Students will apply the principles of oxidation–reduction to electrochemical cells.

	Student Textbook	Assessment Options
Outcomes for Knowledge		
30–B2.1k define anode, cathode, anion, cation, salt bridge/porous cup, electrolyte, external circuit, power supply, voltaic cell and electrolytic cell	The Voltaic Cell, Section 13.1, p. 478 Electrolytic Cells, Section 13.3, p. 502	Questions for Comprehension: 2, 3, Section 13.1, p. 480 Questions for Comprehension: 9, Section 13.2, p. 493 Chapter 13 Review: 1, pp. 522–523 Chapter 13 Test

	Student Textbook	**Assessment Options**
30–B2.2k identify the similarities and differences between the operation of a voltaic cell and that of an electrolytic cell	The Voltaic Cell, Section 13.1, p. 478 Electrolytic Cells, Section 13.3, p. 502	Questions for Comprehension: 2, 3, Section 13.1, p. 480 Questions for Comprehension: 8, 9, Section 13.2, p. 493 Section 13.2 Review: 9, p. 501 Chapter 13 Review: 1, 13, pp. 522–523 Chapter 13 Test Unit 6 Review: 13, 23, 24, 28, 32, pp. 526–529
30–B2.3k predict and write the half–reaction equation that occurs at each electrode in an electrochemical cell	Electrolytic Cells, Section 13.3, p. 502	Practice Problems: 5–8, Section 13.3, p. 508 Questions for Comprehension: 17–20, Section 13.3, p. 510 Chapter 13 Review: 13, 19, 20, pp. 522–523 Chapter 13 Test Unit 6 Review: 11, 31, 36, 44, pp. 526–529
30–B2.4k recognize that predicted reactions do not always occur, e.g., the production of chlorine gas from the electrolysis of brine	Predicting the Products of Electrolysis of Aqueous Solutions, Section 13.3, p. 507 Sample Problem: Electrolysis of an Aqueous Solution, Section 13.3, p. 507	Practice Problems 5–8, Section 13.3, p. 508 Questions for Comprehension: 17, 18, Section 13.3, p. 510 Chapter 13 Review: 5, 14, p. 513 Chapter 13 Test Unit 6 Review: 9, 17, 20, pp. 526–529
30–B2.5k explain that the values of standard reduction potential are all relative to zero volts set for the hydrogen electrode at standard conditions	Cell Potentials, Section 13.1, pp. 482–483	Thought Lab 13.1: Assigning Reference Values: 1–5, Section 13.1, p. 487 Section 13.1 Review: 5–7, p. 490 Questions for Comprehension: 13–16, Section 13.2, p. 500 Chapter 13 Review: 15, pp. 522–523 Chapter 13 Test Unit 6 Review: 9, 13, 23, 25, 31, pp. 526–529
30–B2.6k calculate the standard cell potential for electrochemical cells	Calculating Standard Cell Potentials, Section 13.1, pp. 485–486 Thought Lab 13.1: Assigning Reference Values: 1–5, Section 13.1, p. 487 Sample Problem: Calculating a Standard Cell Potential, Given a Net Ionic Equations, Section 13.1, p. 486 Sample Problem: Calculating a Standard Cell Potential, Given a Chemical Reaction, Section 13.1, pp. 486–487	Thought Lab 13.1: Assigning Reference Values: 1–5, Section 13.1, p. 487 Practice Problems: 1–4, Section 13.1, p. 487 Section 13.1 Review: 3–7, p. 490 Questions for Comprehension: 13, Section 13.2, p. 500 Chapter 13 Review: 6, 11, 12, 18, 21, pp. 522–523 Chapter 13 Test Unit 6 Review: 23, 28, 36, 38, pp. 526–529

	Student Textbook	Assessment Options
30–B2.7k predict the spontaneity or non–spontaneity of redox reactions based on standard cell potential and the relative positions of half–reaction equations on a standard reduction potential table	Predicting the Products of Electrolysis of Aqueous Solutions, Section 13.3, p. 507 Sample Problem: Electrolysis of an Aqueous Solution, Section 13.3, p. 507	**Questions for Comprehension: 5–8, Section 12.1, p. 440** **Section 12.1 Review; 3, 5, 6, p. 440** Practice Problems: 5-8, Section 13.3, p. 508 Section 13.3 Review: 3, p. 513 Chapter 13 Review: 11, pp. 522–523 Chapter 13 Test Unit 6 Review: 3, 38, pp. 526–529
30–B2.8k calculate mass, amounts, current and time in single voltaic and electrolytic cells by applying Faraday's law and stoichiometry.	Stoichiometry and Faraday's Law, Section 13.4, p. 514 Sample Problem: Calculating the Mass of an Electrolysis Product, Section 13.4, pp. 515–516 Faraday's law, Section 13.4, pp. 516–517	Practice Problems: 9–12, Section 13.4, p. 516 Section 13.4 Review: 1–6, p. 520 Chapter 13 Review: 16, 21–26, pp. 522–523 Chapter 13 Test Unit 6 Review: 41–43,, pp. 526–529
Outcomes for Science, Technology and Society (Emphasis on science and technology)		
30–B2.1sts describe the ways in which scientific knowledge may lead to the development of new technologies and new technologies may lead to scientific discoveries by ■ *analyzing the relationship of scientific knowledge and technological development in the applications of voltaic and electrolytic cells for, e.g., batteries, electroplating, refining metals from ores, electrowinning, sanitizing swimming pools with chlorine compounds*	**Connections: How Green Is White Paper? Section 12.3, p. 465** Chapter 13 Launch Lab: What Determines Voltage? p. 477	**Connections: How Green Is White Paper? 1–3, Section 12.3, p. 465** Chapter 13 Launch Lab: What Determines Voltage? 1–5, p. 477 Unit 6 Review: 45–48, pp. 526–529
30–B2.2sts describe applications of science and technology that have developed in response to human and environmental needs by ■ *investigating the use of technology to solve practical problems related to corrosion; e.g., galvanizing, metallurgy, magnesium coupling, painting*	**Chapter 12 Launch Lab: Penny Chemistry, p. 433** **Connections: How Green Is White Paper? Section 12.3, p. 465** Chapter 13 Launch Lab: What Determines Voltage? p. 477	**Chapter 12 Launch Lab: Penny Chemistry: 1–6, p. 433** **Connections: How Green Is White Paper? 1–3, Section 12.3, p. 465** Chapter 13 Launch Lab: What Determines Voltage? 1–5, p. 477 Unit 6 Review: 45–48, pp. 526–529
30–B2.3sts illustrate how science and technology are influenced and supported by society and have influenced and been influenced by historical development and societal needs by ■ *assessing the economic importance of electrochemical cells, particularly fuel cells, to modern society, predicting their future importance in transportation, the recycling of metals and in reducing emissions from smokestacks.*	**Connections: How Green Is White Paper? Section 12.3, p. 465**	**Connections: How Green Is White Paper? 1–3, Section 12.3, p. 465** Unit 6 Review: 45–48, pp. 526–529

	Student Textbook	Assessment Options
Skill Outcomes (Focus on problem solving)		
Initiating and Planning		
30–B2.1s ask questions about observed relationships and plan investigations of questions, ideas, problems and issues by ■ designing an experiment, including a labelled diagram, to test predictions regarding spontaneity, products and the standard cell potential for reactions occurring in electrochemical cells ■ describing procedures for safe handling, storage and disposal of materials used in the laboratory, with reference to WHMIS and consumer product labelling information ■ *developing a plan to build a battery, seeking feedback, testing and reviewing the plan, and making revisions to the plan.*	Investigation 13.C: Electroplating, Section 13.4, pp. 517–518	Investigation 13.C: Electroplating: 1–8, Section 13.4, pp. 517–518 Unit 6 Review: 22, 44, pp. 526–529
Performing and Recording		
30–B2.2s conduct investigations into relationships between and among observable variables and use a broad range of tools and techniques to gather and record data and information by ■ constructing and observing electrochemical cells ■ *investigating the issue of the disposal of used batteries and proposing alternative solutions to this problem* ■ *compiling and displaying evidence and information about voltaic and electrochemical cells, by hand or using technology, in a variety of formats, including diagrams, flow charts, tables, graphs and scatterplots.*	Chapter 13 Launch Lab: What Determines Voltage? p. 477 Investigation 13.A: Measuring Cell Potentials of Voltaic Cells, Section 13.1, pp. 488–489	Chapter 13 Launch Lab: What Determines Voltage? 1–5, p. 477 Investigation 13.A: Measuring Cell Potentials of Voltaic Cells: 1–13, Section 13.1, pp. 488–489 Unit 6 Review: 22, 44, pp. 526–529
Analyzing and Interpreting		
30–B2.3s analyze data and apply mathematical and conceptual models to develop and assess possible solutions by ■ identifying the products of electrochemical cells ■ comparing predictions with observations of electrochemical cells ■ identifying the limitations of data collected on a electrochemical cell ■ explaining the discrepancies between theoretical and the actual cell potential ■ *assessing the efficiencies and practicalities of various electrochemical configurations as batteries* ■ *evaluating experimental designs for voltaic and electrolytic cells and suggesting improvements and alternatives*	Chapter 13 Launch Lab: What Determines Voltage? p. 477 Thought Lab 13.1: Assigning Reference Values, Section 13.1, p. 487 Investigation 13.A: Measuring Cell Potentials of Voltaic Cells, Section 13.1, pp. 488–489 Investigation 13.B: Electrolysis of Aqueous Potassium Iodide, Section 13.3, pp. 508–509	Chapter 13 Launch Lab: What Determines Voltage? 1–5, p. 477 Thought Lab 13.1: Assigning Reference Values: 1–5, Section 13.1, p. 487 Investigation 13.A: Measuring Cell Potentials of Voltaic Cells: 1–13, Section 13.1, pp. 488–489 Investigation 13.B: Electrolysis of Aqueous Potassium Iodide: 1–9, Section 13.3, pp. 508–509 Unit 6 Review: 10, 39, 40–43, pp. 526–529

	Student Textbook	Assessment Options
Communication and Teamwork		
30–B2.4s work collaboratively in addressing problems and apply the skills and conventions of science in communicating information and ideas and in assessing results by ■ using appropriate SI notation, fundamental and derived units to communicate answers to problems related to functioning electrolytic cells. ■ creating multiple–linked documents, selecting and integrating information from various print and electronic sources or from several parts of the same source, to prepare a presentation on the use of hydrogen fuel cells for transportation and heating.	Investigation 13.B: Electrolysis of Aqueous Potassium Iodide, Section 13.3, pp. 508–509	Investigation 13.B: Electrolysis of Aqueous Potassium Iodide: 1–9, Section 13.3, pp. 508–509

Chapter 12

Oxidation-Reduction Reactions

Student Textbook pages 432–475

Chapter Concepts

Section 12.1 Characterizing Oxidation and Reduction

- Oxidation is a loss of electrons. Reduction is a gain of electrons.
- When one atom is oxidized, another must be reduced.
- You can experimentally determine the relative strength of atoms or ions as reducing or oxidizing agents.

Section 12.2 Redox Reactions Involving Ionic Compounds

- An equation for a half-reaction includes only those compounds that are reduced or only those that are oxidized.
- A set of rules guides you through the process of balancing half-reactions ad total reactions.

Section 12.3 Redox Reactions Involving Molecular Compounds

- Oxidation numbers describe the oxidation state of atoms in both ionic and molecular compounds.
- A set of rules allows you to assign oxidation numbers to all atoms in a compound and balance equations.

Section 12.4 Quantitative Analysis of Redox Reactions

- Because some elements change colour with a change in oxidation state, you can titrate redox reactions.
- Quantities measured in redox titrations allow you to perform quantitative calculations on redox reactions.

Common Misconceptions

- Students may believe that oxidation numbers represent an ionic charge in a covalent molecule or polyatomic ion. However, this is not the case. Assigning oxidation numbers to elements in a covalent molecule or polyatomic ion is done on the bases of pretending, for the purposes of electron tracking, that bonds in these species are ionic. To help address this misconception, ensure that ionic charge is always written with the charge following the number and oxidation numbers are always written with the charge in front of the number. For example, an ion with a –2 charge, such as S^{2-}, has an oxidation number of –2.
- There is a common misconception that technological innovations always follow scientific understanding. Early applications of metals are a good example of the opposite. Humans extracted and used metals long before the underlying science was understood. In fact, oxidation-reduction technologies have often given names to periods, such as the Bronze Age, Copper Age, or Iron Age. Have students brainstorm or research technologies often associated with these ages.

Helpful Resources

Books and Journal Articles

- Huddle, P.A. et al, "Using a Teaching Model To Correct Known Misconceptions In Electrochemistry," *Journal of Chemical Education.* Vol 77, 2000, p. 104.
- Birss, V.I. and Truax, D.R. "An Effective Approach to Teaching Electrochemistry," *Journal of Chemical Education.* Vol 67, 1990, p. 403.

Web Sites

Web links related to oxidation-reduction reactions 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
12.0.1 (AST) Reviewing the Periodic Table
12.0.1A (ANS) Reviewing the Periodic Table Answer Key
12.0.2 (HAND) Launch Lab Penny Chemistry
12.0.2A (ANS) Launch Lab Penny Chemistry Answer Key
12.1.1 (OH) Oxidizing and Reducing in a Redox Reaction
12.1.2 (HAND) Investigation 12.A Testing Relative Oxidizing and Reducing Strengths of Metal Atoms and Ions
12.1.2A (ANS) Investigation 12.A Testing Relative Oxidizing and Reducing Strengths of Metal Atoms and Ions Answer Key
12.2.1 (OH) Balancing and Equation for a Reaction that Occurs in an Acidic Solution
12.2.2 (AST) Half-Reaction Method of Balancing Equations
12.2.2A (ANS) Half-Reaction Method of Balancing Equations Answer Key
12.2.3 (AST) Oxidation Number Method of Balancing Equations
12.2.3A (ANS) Oxidation Number Method of Balancing Equations Answer Key
12.2.4 (HAND) Investigation 12.B Redox Reactions and Balanced Equations
12.2.4A (ANS) Investigation 12.B Redox Reactions and Balanced Equations Answer Key
12.2.5 (OH) Balancing a Disproportionation Reaction
12.2.6 (OH) The Smelting Process
12.2.7 (AST) Writing Balanced Ionic, Net Ionic, and Half-Reaction Equations

12.2.7A (ANS) Writing Balanced Ionic, Net Ionic, and Half-Reaction Equations Answer Key
12.3.1 (OH) Assigning Oxidation Numbers
12.3.2 (HAND) Thought Lab 12.1 Oxidation Numbers and Lewis Structures
12.3.2A (ANS) Thought Lab 12.1 Oxidation Numbers and Lewis Structures Answer Key
12.3.3 (OH) Balancing a Redox Equation in Basic Solution
12.4.1 (OH) Redox Titrations
12.4.2 (HAND) Investigation 12.C Measuring the Concentration of Vitamin C in Orange Juice
12.4.2A (ANS) Investigation 12.C Measuring the Concentration of Vitamin C in Orange Juice Answer Key

Using the Chapter 12 Opener

Student Textbook pages 432–433

Teaching Strategies

- The principles of redox rely on electronegativies. Have students create a bingo-style game as a means to review this. Instead of columns with the letters B, I, N, G, O, they can pick common electronegativity values, such as 0.8, 0.9. 2.0, 2.1, and 2.2, and list the elements beneath them.
- Have students draw or create models based on Lewis structures.
- Practice stoichiometry and unit-analysis.

Launch Lab Penny Chemistry

Student Textbook page 433

Purpose

The purpose of this lab is for students to gain an understanding of the fact that oxidation-reduction reactions are responsible for many commonplace occurrences, such as rusting and the tarnishing of coins.

Outcomes

- 30-B1.1s

Advance Preparation

When to Begin	What to Do
1–2 weeks before	■ Order supplies.
1 day before	■ Gather materials. ■ Photocopy **BLM 12.0.2: Launch Lab**

Materials

- white vinegar
- table salt
- $\frac{1}{4}$ cup measure or 50 mL beakers
- measuring teaspoon
- stirring rod
- glass bowl or large beaker
- 6 dull pennies per group
- paper towel
- marker
- clean steel screw or nail

Time Required

- 60 minutes

Helpful Tips

- Have students bring in items like pennies, measuring spoons, nails, and bowls.
- Use **BLM 12.0.2 (HAND) Launch Lab: Penny Chemistry** to support this activity. Remove sections as appropriate to meet the needs of students in your class.
- ***Expected Results:*** The acid in the vinegar will dissolve the copper oxide that has made the pennies look dull and the pennies will become shiny. The pennies that were rinsed and placed on the paper towel will remain shiny. They will eventually tarnish but not during the time period of the observations. The pennies that were not rinsed will begin to tarnish because the salt and vinegar will increase the rate of the reaction with oxygen in the air.

  The copper ions removed from the pennies remain in the vinegar/salt solution. When a steel nail or screw is placed in the solution, a redox reaction will take place between the copper ions and the iron atoms in the nail or screw. The copper will replace the iron and the copper will be coated on the nail or screw. Hydrogen ions from the acid might also react and form bubbles on the surface of the nail or screw.

Safety Precautions

- A vinegar spill on skin or clothing should be washed immediately using cool water.
- Students should thoroughly wash their hands once the lab is completed.
- Ensure that students have read "Safety in Your Chemistry Laboratory and Classroom" on pages xii-xv in the student textbook.

Answers to Analysis Questions

1. The formation of copper oxide causes the pennies to become dull.

2. The vinegar-and-salt mixture acts as a catalyst for the reaction between copper and oxygen.
3. The chemical reaction is $2Cu(s) + O_2(g) \rightarrow 2CuO(s)$.
4. The precipitate on the screw is the formation of solid copper.

Assessment Options

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

12.1 Characterizing Oxidation Reduction

Student Textbook pages 434–440

Section Outcomes

Students will:
- define oxidation and reduction operationally and theoretically
- define oxidizing agent, reducing agent, and oxidation number
- conduct an investigation to determine the relative reactivity of various metals
- predict the spontaneity of oxidation and reduction reactions

Key Terms

oxidation
reduction
oxidation-reduction reactions
redox reactions
ionic equation
spectator ions
net ionic equation
oxidizing agent
reducing agent

Chemistry Background

- Oxidation is said to occur when an atom or ion "loses" one or more electrons. The oxidized atom or ion is referred to as a reducing agent because it donates electrons. "Losing" an electron can mean effectively transferring an electron, as in the case of ionic compounds. In a molecular compound, an atom "loses" an electron when it shares the electron with an atom or a more electronegative element.
- Reduction is said to occur when the atom or ion invovlved "gains" some or more electrons. The reduced atom or ion is referred to as an oxidizing agent because it receives electrons. "Gaining" an electron can mean accepting an electron, as in the case of ionic compounds. In a molecular compound, an atom "gains" an electron when it shares the electron with an atom of a less electronegative element.
- Redox reactions are always coupled. If one atom or ion is oxidized, then another must be reduced in order to balance the number of electrons gained and lost. Electrons are not destroyed or created in chemical reactions.

Teaching Strategies

- Figure 12.3 on page 436 of the student textbook shows a quick demonstration that can catch students' interest and lead into a discussion of oxidation and reduction. If you decide to carry out the demonstration, sand the zinc strip to remove any oxide build-up. The copper (II) sulfate solution should be around 0.5 mol/L to ensure a quick reaction time. Have students make qualitative observations about the colour of the solution and the colour of the deposit on the zinc.
- To contrast with the demonstration in Figure 12.3, place a metallic copper bar in a solution of zinc sulfate. Students will observe that there is no deposit of material on the copper and no change in the appearance of the clear, colourless zinc sulfate solution. This will illustrate that copper, below zinc in the activity series, does not displace the zinc ions from solution. In other words, a redox reaction does not take place. Students will examine the link between activity series and redox reactions in Investigation 12.A. Note: Zinc sulfate is banned by some school boards. Therefore, do not allow students to work with zinc sulfate.
- Many students will need an organized review of previous chemistry topics to understand redox reactions. Review topics such as electron structure, Lewis structures, electronegativity, ions and net ionic equations, bond types, chemical equations, polyatomic molecules, and complex ions before you begin this section.
- **BLM 12.1.1 (OH) Oxidizing and Reducing in a Redox Reaction** has been provided for this section. You will find it with the Chapter 12 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

Some students may need help learning how to organize data and concepts before attempting to answer questions or perform labs related to redox reactions. This organization is key to reducing student confusion associated with redox reactions. Encourage students to refer to concept organizers, and to create their own.

Answers to Questions for Comprehension

Student Textbook page 437

Q1. During oxidation, a given substance loses electrons. Traditionally this loss was attributed to the presence of oxygen, or compounds containing oxygen. However, any substance that has the ability to gain electrons can cause oxidation to occur.

Q2. The original definition of reduction came from the reduction in mass of a metallic ore as it was transformed into a pure metal. The theoretical definition of reduction involves a substance gaining electrons. The key difference is the focus on a theoretical change in electron distribution as opposed to an observed change in mass. However, since the pure metal has a lower molar mass than the metallic oxide, the breakdown of the metal oxide is described by both definitions.

Q3. The oxidizing agent is the $Fe^{2+}(aq)$ ion, while the reducing agent is the Al(s).

Q4. A comparison of ionic charges among the reactants and products reveals that there has been no change in the distribution of electrons, meaning no redox reaction has taken place. Typically, in double replacement reactions, where all substances remain in the same ionic form, a redox reaction does not occur.

Investigation 12.A: Testing Relative Oxidizing and Reducing Strengths of Metal Atoms and Ions

Student Textbook page 438

Purpose

The purpose of this investigation is for students to detemine the oxidizing and reducing strengths of various metal atoms and ions by observing whether reactions occur between solid metals and metal ions in solution.

Outcomes

- 30-B1.2s

Advance Preparation

When to Begin	What to Do
2 – 3 weeks before	■ Check availability of material.
1– 2 days before	■ Cut metals into strips. ■ Prepare solutions. ■ Photocopy **BLM 12.1.2 Investigation 12.A.**

Materials

- 4 small pieces of each of the following metals:
  - – aluminium foil
  - – thin copper wire or tiny copper beads
  - – iron filings
  - – magnesium
  - – zinc

Materials

- dropper bottles containing dilute solutions of
  - – aluminium sulfate
  - – copper(II) sulfate
  - – iron(II) sulfate
  - – magnesium sulfate
  - – zinc nitrate
- well plate
- white paper
- test tube rack
- 4 small test tubes
- steel wool

Time Required

- 50 minutes

Helpful Tips

- To save time, you can brush the metal strips with steel wool as you prepare them. If you do this, wear gloves to prevent splinters from the steel wool.
- Have students prepare a grid on paper, similar to that shown on page 438 of the student textbook. They can then place the well plate on the paper to indicate which solutions or metals are in each well. Alternatively, students could use a grease pencil to mark the well plates.
- Use **BLM 12.1.2 (HAND) Investigation 12.A: Testing Relative Oxidizing and Reducing Strengths of Metal Atoms and Ions** to support this activity. Remove sections as appropriate to meet the needs of students in your class.
- ***Expected Results:*** Reactions should be made visible by changes in colour of solutions or changes in appearance of the metals. The table is filled in below:

Compound / Metal	$Al_2(SO_4)_3(aq)$ $[CH_3COOH(aq)]$ (mol/L)	$CuSO_4$ (aq)	$FeSO_4$ (aq)	$MgSO_4$ (aq)	$Zn(NO_3)_2$ (aq)
Al(s)		y	y	n	y
Cu(s)	n		n	n	n
Fe(s)	n	y		n	n
Mg(s)	y	y	y		y
Zn(s)	n	y	y	n	

Safety Precautions

- General safety rules for correct lab practices must be followed and students must dispose of waste liquids in containers provided by the teacher. You may wish to review these procedures if they are specific to your school.
- Ensure that students have read “Safety in Your Chemistry Laboratory and Classroom” on pages xii-xv in the student textbook.

Answers to Analysis Questions

1. (a) Students should write balanced complete equations for instances in which they observed reactions taking place. Equations below correspond to the reactions students should see.

(1) $2Al(s) + 3CuSO_4(aq) \rightarrow Al_2(SO_4)_3(aq) + 3Cu(s)$

(2) $2Al(s) + 3FeSO_4(aq) \rightarrow Al_2(SO_4)_3(aq) + 3Fe(s)$

(3) $2Al(s) + 3Zn(NO_3)_2(aq) \rightarrow 2Al(NO_3)_3(aq) + 3Zn(s)$

(4) $Fe(s) + CuSO_4(aq) \rightarrow FeSO_4(aq) + Cu(s)$

(5) $3Mg(s) + Al_2(SO_4)_3(aq) \rightarrow 3MgSO_4(aq) + 2Al(s)$

(6) $Mg(s) + CuSO_4(aq) \rightarrow MgSO_4(aq) + Cu(s)$

(7) $Mg(s) + FeSO_4(aq) \rightarrow MgSO_4(aq) + Fe(s)$

(8) $Mg(s) + Zn(NO_3)_2(aq) \rightarrow Mg(NO_3)_2(aq) + Zn(s)$

(9) $Zn(s) + CuSO_4(aq) \rightarrow ZnSO_4(aq) + Cu(s)$

(10) $Zn(s) + FeSO_4(aq) \rightarrow ZnSO_4(aq) + Fe(s)$

(b) Students should write ionic equations for all reactions they observe.

(1) $2Al(s) ++ 3Cu^{2+}(aq) + 3SO_4^{2-}(aq) \rightarrow 2Al^{3+}+(aq) + 3SO_4^{2-}(aq) + 3Cu(s)$

(2) $2Al(s) + 3Fe^{2+}(aq) + 3SO_4^{2-}(aq) \rightarrow 2Al^{3+}+(aq) + 3SO_4^{2-}(aq) + 3Fe(s)$

(3) $2Al(s) + 3Zn^{2+}(aq) + 6NO_3^{-}(aq) \rightarrow 2Al^{3+}+(aq) + 6NO_3^{-}(aq) + 3Zn(s)$

(4) $Fe(s) + Cu^{2+}(aq) + SO_4^{2-}(aq) \rightarrow Fe^{2+}(aq) + SO_4^{2-}(aq) + Cu(s)$

(5) $3Mg(s) + 2Al^{3+}+(aq) + 3SO_4^{2-}(aq) \rightarrow 3Mg^{2+}(aq) + 3SO_4^{2-}(aq) + 2Al(s)$

(6) $Mg(s) + Cu^{2+}(aq) + SO_4^{2-}(aq) \rightarrow Mg^{2+}(aq) + SO_4^{2-}(aq) + Cu(s)$

(7) $Mg(s) + Fe^{2+}(aq) + SO_4^{2-}(aq) \rightarrow Mg^{2+}(aq) + SO_4^{2-}(aq) + Fe(s)$

(8) $Mg(s) + Zn^{2+}(aq) + 2NO_3^{-}(aq) \rightarrow Mg^{2+}(aq) + 2NO_3^{-}(aq) + Zn(s)$

(9) $Zn(s) + Cu^{2+}(aq) + SO_4^{2-}(aq) \rightarrow Zn^{2+}(aq) + SO_4^{2-}(aq) + Cu(s)$

(10) $Zn(s) + Fe^{2+}(aq) + SO_4^{2-}(aq) \rightarrow Zn^{2+}(aq) + SO_4^{2-}(aq) + Fe(s)$

(c) Students should write net ionic equations for all reactions they observe.

(1) $2Al(s) + 3Cu^{2+}(aq) \rightarrow 2Al^{3+}+(aq) + 3Cu(s)$

(2) $2Al(s) + 3Fe^{2+}(aq) \rightarrow 2Al^{3+}+(aq) + 3Fe(s)$

(3) $2Al(s) + 3Zn^{2+}(aq) \rightarrow 2Al^{3+}+(aq) + 3Zn(s)$

(4) $Fe(s) + Cu^{2+}(aq) \rightarrow Fe^{2+}(aq) + Cu(s)$

(5) $3Mg(s) + 2Al^{3+}+(aq) \rightarrow 3Mg^{2+}(aq) + 2Al(s)$

(6) $Mg(s) + Cu^{2+}(aq) \rightarrow Mg^{2+}(aq) + Cu(s)$

(7) $Mg(s) + \; + Fe^{2+}(aq) \rightarrow Mg^{2+}(aq) + Fe(s)$

(8) $Mg(s) + \; + Zn^{2+}(aq) \rightarrow Mg^{2+}(aq) + Zn(s)$

(9) $Zn(s) + Cu^{2+}(aq) \rightarrow Zn^{2+}(aq) + Cu(s)$

(10) $Zn(s) + Fe^{2+}(aq) \rightarrow Zn^{2+}(aq) + Fe(s)$

2. (1) Reducing agent: Al(s); Oxidizing agent: $Cu^{2+}(aq)$

(2) Reducing agent: Al(s); Oxidizing agent: $Fe^{2+}(aq)$

(3) Reducing agent: Al(s); Oxidizing agent: $Zn^{2+}(aq)$

(4) Reducing agent: Fe(s); Oxidizing agent: $Cu^{2+}(aq)$

(5) Reducing agent: Mg(s); Oxidizing agent: $Al^{3+}+(aq)$

(6) Reducing agent Mg(s); Oxidizing agent: $Cu^{2+}(aq)$

(7) Reducing agent: Mg(s); Oxidizing agent: $Fe^{2+}(aq)$

(8) Reducing agent: Mg(s); Oxidizing agent: $Zn^{2+}(aq)$

(9) Reducing agent: Zn(s); Oxidizing agent: $Cu^{2+}(aq)$

(10) Reducing agent: Zn(s); Oxidizing agent: $Fe^{2+}(aq)$

3.

Metal ion	Metal atom
Mg^{2+}	Mg
Al^{3+}	Al
Zn^{2+}	Zn
Fe^{2+}	Fe
Cu^{2+}	Cu

Answer to Conclusion Question

4. Students should note that metal atoms that are good reducing agents make poor oxidizing agents as ions, and vice versa. Metals that are highly reactive are good reducing agents.

Assessment Options

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

Answers to Questions for Comprehension

Student Textbook page 440

Q5. (a) nonspontaneous

(b) spontaneous

(c) nonspontaneous

(d) spontaneous

Q6. Use **(b)** because the reaction will proceed spontaneously.

Q7. Any of the following metals will work: Fe, Zn, Cr, Al, Mg, Na, Ca, Ba, Li.

The reactions for Fe, Zn, Mg, Ca, and Ba, are similar:

$Fe(s) + Cd^{2+}(aq) + 2Cl^{-}(aq) \rightarrow Fe^{2+}(aq) + 2Cl^{-}(aq) + Cd(s)$

The reactions for Al and Cr are similar:

$2Al(s) + 3Cd^{2+}(aq) + 2Cl^{-}(aq) \rightarrow 2Al^{3+}(aq) + 3Cl^{-}(aq) + 3Cd(s)$

The reactions for Na and Li are also similar:

$2Na(s) + Cd^{2+}(aq) + 2Cl^{-}(aq) \rightarrow 2Na^{+}(aq) + Cl^{-}(aq) + Cd(s)$

Q8. The soluble salts could include chlorides of any metal ion listed below tin in the reactivity series. Below are three examples:

$2Ag^{+}(aq) + Cl^{-}(aq) + Sn(s) \rightarrow 2Ag(s) + Sn^{2+}(aq) + 2Cl^{-}(aq)$

$Cu^{2+}(aq) + 2Cl^{-}(aq) + Sn(s) \rightarrow Cu(s) + Sn^{2+}(aq) + 2Cl^{-}(aq)$

$Pb^{2+}(aq) + 2Cl^{-}(aq) + Sn(s) \rightarrow Pb(s) + Sn^{2+}(aq) + 2Cl^{-}(aq)$

Section 12.1 Review Answers

Student Textbook page 440

1. In an oxidation half-reaction, electrons appear on the product side (the right side, by convention). In an oxidation half-reaction, the reducing agent is the reactant, and the reducing agent gives up electrons. In a reduction half-reaction, the electrons appear on the reactant side (the left side, by convention). In a reduction half-reaction, the oxidizing agent is the reactant. Since the oxidizing agent gains electrons, electrons must also appear on the reactant side.
2. In a redox reaction, a molecule, atom, or ion gains electrons, while another molecule, atom, or ion loses electrons. Oxidation means losing electrons. For one chemical species to be oxidized, the other chemical species must accept electrons. Therefore, the chemical species that accepts electrons, facilitating oxidation, is the oxidizing agent.
3. Calcium ions will spontaneously react with Ba(s), K(s), Li(s).

   $Ca^{2+}(aq) + Ba(s) \rightarrow Ba^{2+}(aq) + Ca(s)$

   $Ca^{2+}(aq) + K(s) \rightarrow K^{+}(aq) + Ca(s)$

   $Ca^{2+}(aq) + Li(s) \rightarrow Li^{+}(aq) + Ca(s)$
4. A combination reaction is any reaction in which two reactants combine to form a single product. Lithium is the most reactive of metals. Thus, it has the greatest tendency of all metals to give up its electrons in a chemical reaction. Therefore, if lithium reacts with another element or compound, it will be oxidized, and thus act as a reducing agent.
5. **(a)** The reaction will occur.

   The complete balanced equation is:

   $2AgNO_3(aq) + Cd(s) \rightarrow 2Ag(s) + Cd(NO_3)_2(aq)$

   The ionic equation is:

   $2Ag + (aq) + Cd(s) \rightarrow 2Ag(s) + Cd^{2+} + (aq)$

   The net ionic equation is:

   $2Ag^{+}(aq) + 2NO_3^{-}(aq) + Cd(s) \rightarrow 2Ag(s) + Cd^{2+}(aq) + 2NO_3^{-}(aq)$

   (b) A reaction will not occur.

   (c) The reaction will occur**.** The complete balanced reaction is:

   $2Al(s) + 3HgCl_2(aq) \rightarrow 2AlCl_3(aq) + 3Hg(s)$

   The net ionic equation is:

   $2Al(s) + 3Hg^{2+}(aq) \rightarrow 2Al^{3+}(aq) + 3Hg(s)$

   The net ionic equation is:

   $2Al(s) + 3Hg^{2+}(aq) + 6Cl^{-}(aq) \rightarrow 2Al^{3+}(aq) + 3Hg(s) + 6Cl^{-}(aq)$
6. Possible answers for each part are shown below. In general, equations must be balanced, oxidizing agents must gain electrons, and reducing agents must lose electrons. Oxidizing agents react with metals higher in the activity series; reducing agents react with metals lower in the activity series.

   (a) $2Al(s) + 3Fe^{2+}(aq) \rightarrow 2Al^{3+}(aq) + 3Fe(s)$

   (b) $2Al(s) + 3Fe^{2+}(aq) \rightarrow 2Al^{3+}(aq) + 3Fe(s)$

   (c) $Al(s) + Au^{3+} + (aq) \rightarrow Al^{3+}(aq) + Au(s)$

   (d) $3Cu(s) + 2Au^{3+} + (aq) \rightarrow 3Cu^{2+}(aq) + 2Au(s)$
7. **(a)** $Na(\ell) + K^{+}(\ell) \rightarrow K(\ell) + Na^{+}(\ell)$

   (b) The potassium ions are the oxidizing agents. Sodium is the reducing agent.

12.2 Redox Reactions Involving Ionic Compounds

Student Textbook pages 441–454

Section Outcomes

Students will:

- define half-reaction and disproportionation
- write and balance equations for redox reactions using the half-reaction method
- develop equations for half-reactions from information about redox changes
- perform investigations on redox reactions

Key Terms

half-reactions
disproportionation reaction
smelting
refining

Chemistry Background

- Oxidation numbers may be used as tools for balancing redox equations. They provide a method of tracing the gain and loss of electrons that, in many cases, is simpler to use than half-reactions.
- In disproportionation reactions, a single element undergoes both oxidation and reduction. These reactions involve elements with more than one oxidation number.
- Elements or compounds can undergo two or more different reductions or oxidations simultaneously. For example, when iron forms magnetite, a combination of iron(II) and iron(III) oxides result, which is represented as Fe_3O_4. When this occurs, the oxidation numbers that are found using the rules from Table 12.3 on page 456 of the student textbook are not always whole numbers.

Teaching Strategies

- A number of overhead masters and handouts have been prepared for this section. You will find them with the Chapter 12 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
  12.2.1 (OH) Balancing and Equation for a Reaction that Occurs in an Acidic Solution
  12.2.2 (AST) Half-Reaction Method of Balancing Equations
  12.2.2A (ANS) Half-Reaction Method of Balancing Equations Answer Key
  12.2.3 (AST) Oxidation Number Method of Balancing Equations
  12.2.3A (ANS) Oxidation Number Method of Balancing Equations Answer Key
  12.2.5 (OH) Balancing a Disproportionation Reaction
  12.2.6 (OH) The Smelting Process
  12.2.7 (AST) Writing Balanced Ionic, Net Ionic, and Half-Reaction Equations
  12.2.7A (ANS) Writing Balanced Ionic, Net Ionic, and Half-Reaction Equations Answer Key
- Provide students with copies of **BLM 12.2.2 (HAND) Half-Reaction Method of Balancing Equations** for practice.
- Hold group problem-solving competitions to help break up the monotony of working with half-reactions. Ensure stronger students are paired with students who need extra help.
- Invite groups of students to make a redox chemistry board game based on identifying oxidizing and reducing agents and their relative strengths, as well as balancing reactions.
- Ensure students realize that although we write and manipulate half-reactions to help balance chemical equations, half-reactions do not occur on their own. A reduction must be accompanied by an oxidation.
- By the end of Chapter 12, students should be able to summarize the relationships between redox reactions and other reaction types. Students have been given opportunities to find these relationships for themselves (e.g., in Investigation 12.B and in Section 12.2 Review question 5). However, you may want to summarize these relationships. They are:
  - Single displacement reactions are redox reactions
  - Double displacement reactions, including neutralization reactions, are not redox reactions
  - Combustion reactions are redox reactions
  - Synthesis reactions and decomposition reactions may or may not be redox reactions

SUPPORTING DIVERSE STUDENT NEEDS

Some students may need help learning how to organize data and concepts before attempting to answer questions or perform investigations related to redox reactions. This organization is a key to reducing student confusion associated with redox reactions. Use **BLM 12.2.2 The Half-Reaction Method of Balancing Equations**, and **BLM 12.2.3 The Oxidation Number Method of Balancing Equations**, to provide practice opportunities. Encourage students to refer to concept organizers and to create their own.

Chemistry File: Try This

Student Textbook page 442

A half-reaction cannot occur on its own, because there are electrons present as either reactants or products. Since electrons do not exist independently under normal circumstances, both sides of the reaction need to be shown to account for their presence.

Answers to Questions for Comprehension

Student Textbook page 442

Q9. The balanced half-reactions are as follows:
Oxidation: $Al(s) \rightarrow Al^{3+}(aq) + 3e^-$
Reduction: $Fe^{3+}(aq) + 3e^- \rightarrow Fe(s)$

Q10. The balanced half-reactions are as follows:
(a) Oxidation: $Fe(s) \rightarrow Fe^{2+}(aq) + 2e^-$
Reduction: $Cu^{2+}(aq) + 2e^- \rightarrow Cu(s)$
(b) Oxidation: $Cd(s) \rightarrow Cd^{2+}(aq) + 2e^-$
Reduction: $Ag^+(aq) + e^- \rightarrow Ag(s)$

Q11. The balanced half-reactions are as follows:
(a) Oxidation: $Sn(s) \rightarrow Sn^{2+}(aq) + 2e^-$
Reduction: $Pb^{2+}(aq) + 2e^- \rightarrow Pb(s)$
(b) Oxidation: $Ag(s) \rightarrow Ag^+(aq) + e^-$
Reduction: $Au^{3+}(aq) + 3e^- \rightarrow Au(s)$
(c) Oxidation: $Zn(s) \rightarrow Zn^{2+} + 2e^-$
Reduction: $Fe^{3+}(aq) + 3e^- \rightarrow Fe(s)$

Answers to Practice Problems 1–2

Student Textbook page 448

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

1. **(a)** $5Ag(s) + MnO^{4-}(aq) + 8H^{+}(aq) \rightarrow$
   $5Ag^{+}(aq) + Mn^{2+}(aq) + 4H_2O(\ell)$

   (b) $Hg(\ell) + 4Cl^{-}(aq) + 2NO_3^{-}(aq) + 4H^{+}(aq) \rightarrow$
   $HgCl_4^{2-}N(s) + 2NO_2(g) + 2H_2O(\ell)$

   (c) $AsH_3(g) + 4H_20(\ell) + 4Zn^{2+}(aq) \rightarrow$
   $H_3AsO_4(aq) + 8H^{+}(aq) + 4Zn(s)$

   (d) $I_2(s) + H_2O(\ell) + 5OCl^{-}(aq) \rightarrow$
   $2IO_3^{-}(aq) + 2H^{+}(aq) + 5Cl^{-}(aq)$

2. **(a)** $I^{-}(aq) + 6OH^{-}(aq) + 6MnO_{4-}(aq) \rightarrow$
   $6MnO_4^{2-}(aq) + IO^{3-}(aq) + 3H_2O(\ell)$

   (b) $3H_2O_2(aq) + 2OH^{-}(aq) + 2ClO_2(aq) \rightarrow$
   $2ClO^{-} + 3O_2(g) + 4H_2O(\ell)$

   (c) $2CrO_2^{-}(aq) + 6ClO^{-}(aq) + 2H_2O(\ell) \rightarrow$
   $3Cl_2(g) + 4OH^{-}(aq) + 2CrO_4^{2-}\ (aq)$

   (d) $4Al(s) + OH^{-}(aq) + 3NO^{-}(Aq) + 4H_2O(\ell) \rightarrow$
   $3NH_3(g) + 4AlO_2^{-}(aq)$

Investigation 12.B: Redox Reactions and Balanced Equations

Student Textbook pages 448–450

Purpose

The purpose of this lab is for students to use half-reactions to determine whether a redox reaction has occurred.

Outcomes

- 30-B1.7k
- 30-B1.1s
- 30-B1.2s
- 30-B1.4s

Advance Preparation

When to Begin	What to Do
2 – 3 weeks before	■ Check availability of material.
1 – 2 days before	■ Cut metals into strips. ■ Prepare solutions. ■ Photocopy **BLM 12.2.4 Investigation 12.B.**

Materials

- small pieces of each of the following metals:
  – magnesium
  – zinc
  – copper
  – aluminium
- 10 mL hydrochloric acid (1 mol/L)
- 10 mL sulfuric acid (1 mol/L)
- 1 well plate
- 4 small test tubes
- 1 test tube rack
- 1 Bunsen burner
- 1 candle
- steel wool
- 1 micro-pipette or eyedropper

Time Required

- 60 minutes

Helpful Tips

- Use **BLM 12.2.4 (HAND) Investigation 12.B: Redox Reactions and Balanced Equations** to support this activity. Remove sections as appropriate to meet the needs of students in your class.
- To save time, you can brush the metal strips with steel wool as you prepare them. If you do this, wear gloves to prevent splinters from the steel wool.
- Review complete and incomplete combustion with your students before beginning the investigation.
- Have your students prepare a grid on paper, similar to the grid shown on page 438 of the student textbook, to keep track of the metal-acid reactions. Students can then place the well plate on the paper to indicate which solutions and metals are in each well. Alternatively, students could use a grease pencil to mark the well plates.
- If your classroom does not have multiple gas outlets, setup one Bunsen burner and a candle as a demonstration station.
- To accompany Part 1, you may want to carry out the following demonstration of the reaction of copper with nitric acid. **Caution:** The demonstration uses concentrated nitric acid, which is extremely corrosive. The reaction produces the poisonous gas nitrogen dioxide. Students must not carry out this procedure. This procedure must be carried out in a properly ventilated fumehood. Wear thick gloves, safety glasses, and an apron.

Demonstration Procedure:

1. In a fumehood, add 2 or 3 tiny copper balls to a 25 mL beaker. Place the beaker on white paper for visibility.
2. Slowly add about 10 mL of concentrated nitric acid. Bubbling of $NO_2(g)$ and the green colour of copper ions should appear instantly. The yellow-brown gas ($NO_2(g)$) appears in 5–10 s.

3. **Gently** swirl the reaction vessel after 5 minutes. Allow the reaction to proceed for 10 minutes further. Then dispose of waste in a clearly marked container.

- To accompany Part 2, you may want to carry out the following demonstration: Hold a beaker of cold water over a Bunsen burner. If you have students helping you, ensure they are wearing gloves, safety glasses, and aprons. When the flame is blue, students should observe water condensing on the outside of the beaker (a product of complete combustion). When the flame is yellow, students should observe soot (a product of incomplete combustion), as well as water on the outside of the beaker.
- ***Expected Results:*** Acids react with many metals to form hydrogen gas. The acids should react with magnesium, zinc, and aluminium but not with copper.

  When the Bunsen burner is adjusted to allow maximum oxygen to combine with the gas, the flame will be blue. When the windows in the Bunsen burner are closed to reduce the amount of oxygen to mix with the gas, the flame will be yellow/orange. The colour is due to the presence of incompletely combusted carbon compounds. The candle flame is orange, similar to the flame in the Bunsen burner with minimal oxygen available. Again, there is soot in the flame of the candle.

Safety Precautions

- Ensure that students have read “Safety in Your Chemistry Laboratory and Classroom” on pages xii-xv in the student textbook.
- The acid solutions are corrosive. Students should handle them with care.
- If a student accidentally spills a solution on their skin, they should wash the area immediately with copious amounts of cool water. If they get any acid in their eyes, they should use the eye wash station immediately. You should assist them in doing this.
- Before lighting a Bunsen burner or candle, students should make sure that flammable liquids are not present in the laboratory. They should also tie back long hair and confine any loose clothing.
- Students should thoroughly wash their hands when they have completed the investigation.

Answers to Prediction Questions

- Based on students’ experience with the activity series, they will likely predict that all metals mentioned, except copper, are oxidized by aqueous hydrogen ions.
- Metals that cannot be oxidized by hydrogen ions would not dissolve in acid. To dissolve, metals need to form ions—they must be oxidized.
- Hydrocarbon combustion is a redox reaction. In complete combustion, oxygen is reduced from 0 in the oxygen molecule to –2 in the carbon dioxide and water molecules.

  Carbon is oxidized from a negative oxidation number in the hydrocarbon to a positive oxidation number in the carbon dioxide.

Answers to Analysis Questions

PART 1

1. $Mg(s) + 2HCl(aq) \rightarrow MgCl_2(aq) + H_2(g)$

   $Mg(s) + H_2SO_4(aq) \rightarrow MgSO_4(aq) + H_2(g)$

   $Zn(s) + 2HCl(aq) \rightarrow ZnCl_2(aq) + H_2(g)$

   $Zn(s) + H_2SO_4(aq) \rightarrow ZnSO_4(aq) + H_2(g)$

   $Cu(s) + HCl(aq) \rightarrow NR$

   $Cu(s) + H_2SO_4(aq) \rightarrow NR$

   $2Al(s) + 6HCl(aq) \rightarrow 2AlCl_3(aq) + 3H_2(g)$

   $2Al(s) + 3H_2SO_4(aq) \rightarrow Al_2(SO_4)_3(aq) + 3H_2(g)$
2. Note: for each metal, the net ionic equations are the same for both acids. Therefore, only one equation is shown for each metal.

   $Mg(s) + 2H^+(aq) \rightarrow Mg^{2+}(aq) + H_2(g)$

   $Zn(s) + 2H^+(aq) \rightarrow Zn^{2+}(aq) + H_2(g)$

   $Cu(s) + H^+(aq) \rightarrow NR$

   $2Al(s) + 6H^+(aq) \rightarrow 2Al^{3+}(aq) + 3H_2(g)$
3. The reactions of aluminium, zinc, and magnesium with acid are redox reactions.
4. Note that for each metal, the half-reactions are the same for both acids. Therefore, only one reduction and oxidation is shown for each metal. The half-reactions are as follows:

Oxidation	**Reduction**
$Al(s) \rightarrow Al^{3+}(aq) + 3e^-$	$2H^+ + 2e^- \rightarrow H_2(g)$
$Mg(s) \rightarrow Mg^{2+}(aq) + 2e^-$	$2H^+ + 2e^- \rightarrow H_2(g)$
$Zn(s) \rightarrow Zn^{2+}(aq) + 2e^-$	$2H^+ + 2e^- \rightarrow H_2(g)$

5. Students may notice that copper did not react with the hydrogen ions. They may recall that copper ions are fairly good oxidizing agents and so could suggest that copper ions are better oxidizing agents than hydrogen ions.
6. The hydrogen ions are reduced. Therefore, they act as the oxidizing agent.
7. A neutralization reaction is a double displacement reaction. Therefore, it is not a redox reaction. The hydrogen ions have an oxidation number of +1 on both the reactant and product sides of the equation. In other words, they do not act as oxidizing agents.
8. Refer to the Helpful Tips section of this investigation for the demonstration of concentrated nitric acid reacting with copper. Do not allow students to carry out the procedure. Carry out the demonstration only in a properly-ventilated fumehood. If you cannot carry out the demonstration, inform students that when nitric acid is added to copper the reaction mixture turns green (copper ions) and a yellow-brown gas (NO_2) is produced.

The balanced chemical equation is:

$$Cu(s) + 4HNO_3(aq) \rightarrow Cu(NO_3)_2(aq) + 2NO_2(g) + 2H_2O(\ell)$$

The balanced net ionic equation is:

$$Cu(s) + 2NO_3^-(aq) + 4H^+(aq) \rightarrow Cu^{2+}(aq) + 2NO_2(g) + 2H_2O(\ell)$$

In this reaction, copper metal is the reducing agent. Its oxidation number increases from 0 to +2. Nitrate ion is the oxidizing agent. Its oxidation number of nitrogen decreases from +5 to +4.

9. The reaction of nitrate ions and copper atoms generates a positive cell potential, while the reaction of copper atoms with hydrogen ions does not.

PART 2

10. The products for the complete combustion of a hydrocarbon are water and carbon dioxide.

$$CH_4(g) + 2O_2(g) \rightarrow CO_2(g) + 2H_2O(g)$$

11. Assuming that the incomplete combustion does not involve the production of carbon monoxide, the balanced chemical equation is:

$$CH_4(g) + O_2(g) \rightarrow C(s) + 2H_2O(g)$$

12. Assuming that the incomplete combustion does not involve the production of carbon, the balanced chemical equation is:

$$2CH_4(g) + 3O_2(g) \rightarrow 2CO(g) + 4H_2O(g)$$

13. The balanced chemical equation for the complete combustion of paraffin wax, $C_{25}H_{52}(g)$ is:

$$C_{25}H_{52}(g) + 38O_2(g) \rightarrow 25CO_2(g) + 26H_2O(g)$$

14. Two balanced equations for the incomplete combustion of paraffin wax are:

$$C_{25}H_{52}(g) + 13O_2(g) \rightarrow 25C(s) + 26H_2O(g)$$

$$2C_{25}H_{52}(g) + 51O_2(g) \rightarrow 50CO(g) + 52H_2O(g)$$

15. It is usually possible to observe the production of soot (carbon) when a candle burns.

16. All combustion reactions involve the reduction of oxygen. In any combustion reaction, oxygen with an oxidation number of 0 forms compounds with hydrogen and possibly carbon. Consequently, its oxidation number decreases. Therefore, since every reduction must be accompanied by a corresponding oxidation, and since both complete and incomplete combustion involve the reduction of oxygen, all combustion reactions must be redox reactions.

Answer to Conclusion Question

17. If it is clear that a neutral element (such as a metal or hydrogen) is produced, the reaction must be a redox reaction. The oxidation number of an atom of an element is zero, while the oxidation number of an atom of an element in a compound is not. Electron transfer cannot be observed directly.

Answers to Applications Questions

18. $Au(s) + 3NO_3^-(aq) + 6H^+(aq) + 4Cl^-(aq) \rightarrow AuCl_4^-(aq) + 3NO_2(g) + 3H_2O(\ell)$

The oxidizing agent is the nitrate ion. The reducing agent is gold.

19. Complete combustion maximizes the energy release from the fuel used, prevents the formation of deadly carbon monoxide gas, and prevents the production of soot. To ensure complete combustion, there must be adequate, unrestricted air flow to the furnace.

Assessment Options

- Collect and assess students' Predictions and answers to Analysis, Conclusion, and Extension questions.
- Use Assessment Checklist 2: Laboratory Reports from Appendix A.

SUPPORTING DIVERSE STUDENT NEEDS

Gifted Students: The electronegativities of the metals in the activity series often follow the trend of greatest electronegativity for the least reactive metals to smallest electronegativity for the most reactive metals. However, this is not always the case. Have strong students research an explanation.

Answers to Practice Problems 3–6

Student Textbook page 452

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

3. $2PbSO_4(aq) + 2H_2O(\ell) \rightarrow Pb(s) + PbO_2(aq) + 2H_2SO_4(aq)$

4. $3NO_2(g) + H_2O(\ell) \rightarrow NO(g) + 2HNO_3(aq)$

5. $Cl_2(g) + 2OH^-(aq) \rightarrow Cl^-(aq) + ClO^-(aq) + H_2O(\ell)$

6. $3I_3^-(aq) + 3H_2O(\ell) \rightarrow IO_3^-(aq) + 6H^+(aq) + 8I^-(aq)$

Section 12.2 Review Answers

Student Textbook page 454

1. **(a)** $Cr_2O_7^{2-} + 14H^+(aq) + 6e^- \rightarrow 2Cr^{3+}(aq) + 7H_2O(\ell)$; oxidation half-reaction

(b) $2S_2O_3^{2-}(aq) \rightarrow S_4O_6^{2-}(aq) + 2e^-$; oxidation half-reaction

(c) Arsenic is reduced, so this is a reduction half-reaction. First balance the atoms other than oxygen and hydrogen.

$$4AsO_4(aq) \rightarrow As_4O_6(aq)$$

Add water to balance the oxygen atoms.

$$4AsO_4^{3-}(aq) \rightarrow As_4O_6(aq) + 10H_2O(\ell)$$

Add hydrogen ions (acidic solution) to balance the hydrogen atoms.

$4AsO_4^{3-}(aq) + 20H^+(aq) \rightarrow As_4O_6(aq) + 10H_2O(\ell)$

Add electrons to balance the charges on both sides.

$AsO_4^{3-}(aq) + 20H^+(aq) + 8e^- \rightarrow As_4O_6(aq) + 10H_2O(\ell)$

(d) Bromine is oxidized, so this is an oxidation half-reaction. First balance the atoms other than oxygen and hydrogen.

$Br_2(\ell) \rightarrow 2BrO_3^-(aq)$

Add water to balance the oxygen atoms.

$Br_2(\ell) + 6H_2O(\ell) \rightarrow 2BrO_3^-(aq)$

Add hydrogen ions (acidic solution) to balance the hydrogen atoms.

$Br_2(\ell) + 6H_2O(\ell) \rightarrow 2BrO_3^-(aq) + 12H^+(aq)$

Adjust for basic conditions by adding hydroxide ions to both sides.

$Br_2(\ell) + 6H_2O(\ell) + 12OH^-(aq) \rightarrow 2BrO_3^-(aq) + 12H^+(aq) + 12OH^-(aq)$

Combine hydrogen and hydroxide ions and cancel any water found on both sides of the equation.

$Br_2(\ell) + 12OH^-(aq) \rightarrow 2BrO_3^-(aq) + 6H_2O(\ell)$

Add electrons to balance the charges on both sides.

$Br_2(\ell) + 12OH^-(aq) \rightarrow 2BrO_3^-(aq) + 6H_2O(\ell) + 10e^-$

2. (a) Balance this equation by balancing charge:

$3Co^{2+}(aq) + 2\ Au(s) \rightarrow 3Co(s) + 2Au^{3+}(aq)$

(b) Under acidic conditions, the balanced half-reactions are:

$Cu(s) \rightarrow Cu^{2+}(aq) + 2e^-$

$NO_3^-(aq) + 3e^- + 4H^+(aq) \rightarrow NO(g) + 2H_2O(\ell)$

Write half-reactions with the LCM.

$3Cu(s) \rightarrow 3Cu^{2+}(aq) + 6e^-$

$2NO_3^-(aq) + 6e^- + 8H^+(aq) \rightarrow 2NO(g) + 4H_2O(\ell)$

Add half-reactions, remove electrons and chemical species present on both sides:

$3Cu(s) + 2NO_3^-(aq) + 8H^+(aq) \rightarrow 3Cu^{2+}(aq) + 2NO(g) + 4H_2O(\ell)$

(c) Under basic conditions, the balanced half-reactions are:

$Al(s) + 4OH^-(aq) \rightarrow AlO_2^-(aq) + 2H_2O(\ell) + 3e^-$

$NO_3^-(aq) + 6H_2O(\ell) + 8e^- \rightarrow NH_3(g) + 9OH^-(aq)$

Write half-reactions with the LCM.

$8Al(s) + 32OH^-(aq) \rightarrow 8AlO_2^-(aq) + 24e^- + 16H_2O(\ell)$

$3NO_3^-(aq) + 24e^- + 18H_2O(\ell) \rightarrow 3NH_3(g) + 27OH^-(aq)$

Add half-reactions, remove electrons and chemical species present on both sides:

$8Al(s) + 5OH^-(aq) + 3NO_3^-(aq) + 2H_2O(\ell) \rightarrow 8AlO_2^-(aq) + 3NH_3(g)$

3. $Hg^+(aq) \rightarrow Hg^{2+}(aq) + e^-$

$Hg^+(aq) + e^- \rightarrow Hg(\ell)$

Combined, the net ionic equation would be:

$2Hg^+(aq) \rightarrow Hg^{2+}(aq) + Hg(\ell)$

4. (a) The balanced chemical reaction is:

$3Fe_2O_3(s) + CO(g) \rightarrow 2Fe_3O_4(s) + CO_2(g)$

(b) The balanced chemical reaction is:

$CaCO_3(s) \rightarrow CaO(s) + CO_2(g)$

(c) The balanced chemical reaction is:

$Fe_3O_4(s) + CO(g) \rightarrow 3FeO(s) + CO_2(g)$

(d) The balanced chemical reaction is:

$FeO(s) + CO(g) \rightarrow Fe(\ell) + CO_2(g)$

5. Since electrons do not exist on their own, a gain or loss of electrons by one substance needs to correspond to a loss or gain of electrons by another substance. Thus, if one substance undergoes oxidation, another must simultaneously undergo reduction.

12.3 Redox Reactions Involving Molecular Compounds

Student Textbook pages 455–466

Section Outcomes

Students will:

- differentiate between redox reactions and other reactions using oxidation numbers
- write and balance redox reactions by assigning oxidation numbers

Key Terms

oxidation numbers

Chemistry Background

- It is not always easy to determine whether or not a redox reaction has occurred for reactions involving molecular compounds. Using Lewis structures in association with the concepts of electronegativity and valence electrons is one way of making this determination. Identifying whether an atom gains or loses electrons is based on the relative electronegativities of elements involved.
- Redox reactions do not always take place in neutral environments. When an acid or base is present in a reaction solution, the hydroxyl and hydrogen ions must be

considered when determining the balanced chemical equation.

Teaching Strategies

- Students must be familiar with ionic, covalent, and polar covalent bonding. They will benefit from a brief review of Lewis structures and counting valence electrons, the concept of electronegativity, and chemical bonding that they learned in Chapter 1 of *Chemistry 20*.
- Before providing students with rules for assigning oxidation numbers, use the Thought Lab 12.1 on page 458 of the student textbook to let students develop their own rules.
- Note that the table of oxidation number rules in Table 12.3 on page 458 of the student textbook does not include the oxidation number values for oxygen in peroxides or in oxygen difluoride. They are –1 and +2 respectively. You may want to ask students to recreate the table of rules in their notebooks, adding the values for oxygen in peroxides and oxygen on their own.
- In this section, students are sometimes asked to identify reducing agents and oxidizing agents. Ensure students are clear about the difference between identifying reducing agents or oxidizing agents and identifying the atom in a compound that is reduced or oxidized. For example, in part (a) of the Sample Problem on page 461 of the student textbook, the element carbon is oxidized, but methane, which contains the carbon, is described as the reducing agent. It is customary to identify the oxidizing or reducing agent as whatever chemical species contains the relevant element. If equations are written in molecular form, the chemical species is a compound. If equations are written in ionic form, the chemical species may be a compound or an ion.
- Use the Haber Process to connect the concepts of redox reactions and equilibrium reactions. Connecting material from different units emphasizes the ways in which chemical processes can be analyzed from different viewpoints.
- Overhead masters and a quiz have been prepared for this section. You will find them with the Chapter 12 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
  12.3.1 (OH) Assigning Oxidation Numbers
  12.3.3 (OH) Balancing a Redox Equation in Basic Solution

SUPPORTING DIVERSE STUDENT NEEDS

For students having difficulty with Lewis structures, molecular modelling kits could be used to build the substances listed above. Students could still use electronegativities by disassembling the models to determine which substance keeps the electron pair.

Answers to Practice Problems 7–10

Student Textbook page 457

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

7. **(a)** +3
(b) 0
(c) +6
(d) +5
(e) 0
(f) +2
8. **(a)** H = +1, S = +4, O = –2
(b) H = +1, O = –2
(c) H = +1, P = +5, O = –2
9. **(a)** O = +2
(b) O = –1
10. **(a)** A1 = +3, H = +1, C = +4, O = –2
(b) N = –3, H = +1, P = +5, O = –2
(c) K = +1, H = +1, I = +7, O = –2

Thought Lab 12.1: Oxidation Numbers and Lewis Structures

Student Textbook pages 458–459

Purpose

The purpose of this lab is for students to gain experience using Lewis structures to assign oxidation numbers to the atoms in molecular compounds.

Outcomes

- 30-B1.7k

Time Required

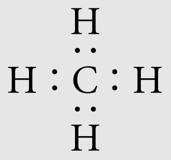
- 30 minutes

Helpful Tips

- Before having students begin the investigation, use the chalkboard to work through the examples at the beginning of the ThoughtLab.
- Students may need some help drawing Lewis structures for more complex molecules such as PCl_5. Remind students that Chapter 5 provides practice for drawing Lewis structures.
- Use **BLM 12.3.2 (HAND) Thought Lab 12.1: Oxidation Numbers and Lewis Structures** to support this activity. Modify it as necessary.

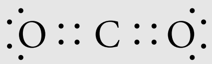
Answers to Procedure Questions

2. (a)



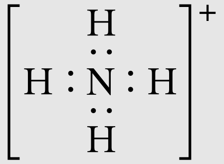
carbon = –4,
hydrogen = +1

(b)



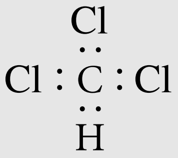
carbon = +4,
oxygen = –2

(c)



nitrogen = –3,
hydrogen = +4

(d)



carbon = +2,
hydrogen = +1,
chlorine = –1

3. carbon = 0, hydrogen = +1, oxygen = –2

Answer to Analysis Question

1. One of the things that students may notice is that through the use of Lewis structures it becomes possible to predict products that occur when one of the reactants breaks down. For interested students, this could serve as an introduction to the concept of Lewis acids and bases.

Assessment Options

- Collect and assess students' answers to Procedure and Analysis questions. You may also have students volunteer to draw their structures on the chalkboard.

Answers to Practice Problems 11–14

Student Textbook page 461

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

11. (a) This is a redox reaction.

(b) This is not a redox reaction.

(c) This is a redox reaction.

(d) This is a redox reaction and a disproportionation reaction.

12. (a) Oxidizing agent: $H_2O_2(aq)$. Reducing agent: $Fe(OH)_2(s)$

(b) There is no oxidizing or reducing agent

(c) Oxidizing agent: $O_2(g)$. Reducing agent: $C_2H_6(g)$.

(d) $NO_2(g)$ is both the oxidizing and reducing agent.

13. The oxidation number of Br decreases from 0 in $Br_2(\ell)$ to –1 in $2Br^-$ (aq). This is reduction. $Br_2(\ell)$ is an oxidizing agent. The oxidation number of Cl increases from +3 in $ClO_2^-(aq)$ to +4 in $ClO_2(aq)$. This is oxidation. $ClO_2^-(aq)$ is a reducing agent.

14. The oxidation number of nickel decreases from +2 in the sulfide to 0 in its elemental form. This is a reduction. A reducing agent must be used to achieve this change. This is a redox reaction. The same analysis can be used with CuS and Cu.

Answers to Practice Problems 15-18

Student Textbook page 464

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

15. $CS_2(g) + 3O_2(g) \rightarrow CO_2(g) + 2SO_2(g)$

16. (a) $B_2O_3(aq) + 6Mg(s) \rightarrow 3MgO(s) + Mg_3B_2(aq)$

(b) $8H_2S(g) + 8H_2O(aq) \rightarrow S_8(s) + 16H_2O(\ell)$

17. (a) $Cr_2O_7^{2-}(aq) + 6Fe^-(aq) + 14H(aq) \rightarrow 2Cr^{3+}(aq) + 6Fe^{3+}(aq) + 7H_2O(\ell)$

(b) $I_2(g) + 10NO_3^-(aq) + 8H^+(aq) \rightarrow 2IO_3^-(aq) + 10NO_2(g) + 4H_2O(\ell)$

(c) $2PbSO_4(aq) + 2H_2O(\ell) \rightarrow Pb(s) + PbO_2(aq) + 2SO_4^{2-}(aq) + 4H^+(aq)$

18. (a) $3Cl^-(aq) + 2CrO_4^{2-}(aq) + H_2O(\ell) \rightarrow 3ClO^-(aq) + 2CrO_2^-(aq) + 2OH^-(aq)$

(b) $3Ni(s) + 2MnO_4^-(aq) + H_2O(\ell) \rightarrow 3NiO(s) + 2MnO_2(s) + 2OH^-(aq)$

(c) $I^-(aq) + 6Ce^{4+}(aq) + 6OH^-(aq) \rightarrow IO^{3-}(aq) + 6Ce^{3+}(aq) + 3H_2O(\ell)$

Connections (Science and Technology): How Green is White Paper?

Student Textbook page 465

Teaching Strategies

- The pulp and paper industry is a major component of the Albertan (and Canadian) economy. Have students once again consider the implications when environmental and economic needs seem to be working at cross purposes and how industry is beginning to find ways to respond.

Answers to Connections Questions

1. The Fourdrinier machine is the general name given to the machines that complete the final stages of paper processing. When students describe images of paper coming off an assembly line in large rolls they are thinking of the final stage in the Foudrinier machine where the pulp-paper product is pressed into firm sheets, and the excess moisture is removed.

2. There are two main advantages to use the TAML™ hydrogen peroxide bleaching system over chlorine systems. First, it does not appear to be as harmful to the environment as the chlorine-containing waste products, such as dioxins. Second, it does not cause the same degree of waste discolouration as the chlorine systems.

3. Most ECF processes produce ClO_2, and students could inquire about factors that a pulp and paper mill would consider before implementing such measures. Encourage students to brainstorm technological, economic, political, and environmental factors.

Section 12.3 Review Answers

Student Textbook page 466

1. (a) The reaction is a redox reaction because the oxidation numbers of the hydrogen and iodine atoms change.

$H_2(g) + I_2(g) \rightarrow 2HI(aq)$

0 0 +1 –1

(b) The reaction is not a redox reaction because the oxidation numbers of the atoms do not change.

$2NaHCO_3(aq) \rightarrow Na_2CO_3(aq) + H_2O(\ell) + CO_2(g)$

+1 +1 +4 –2 + +1 +4 –2 +1 –2 + +4 –2

(c) The reaction is not a redox reaction because the oxidation numbers of the atoms do not change.

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

+1 –1 +2 –2 +1 +2 –1 +1–2

(d) The reaction is a redox reaction because the oxidation numbers of the phosphorus atom and two of the chlorine atoms change.

$PCl_5 \rightarrow PCl_3 + Cl_2(g)$

+5 – 1 +3 –1 0

2. Students may suggest definitions similar to the following:

- A redox reaction is a chemical reaction in which there is a change in the oxidation number of some or all of the atoms of the reactants.
- A redox reaction is a chemical reaction in which atoms of one element are oxidized and atoms of another element are reduced.
- A redox reaction is a chemical reaction in which electrons are transferred among reactants. The number of electrons gained by the oxidizing agent must equal the number of electrons lost by the reducing agent.

3. Fluorine is the most electronegative element in the periodic table. Therefore, it does not "lose" electrons in chemical bonds. It has seven valence electrons, and thus forms bonds in which it gains one valence electron, satisfying the octet rule. Because it is in the second row of the periodic table, its outer occupied energy level cannot accommodate more than eight electrons (as bromine, iodine, or chlorine can). Therefore, it does not form bonds in which it gains more than one electron, and its oxidation number is always –1.

4. When one element combines with another element in a chemical reaction, the reaction is always a redox reaction. The atoms in uncombined elements have an oxidation number of 0. However, different elements have different electronegativities.. Therefore, when they combine in a compound, atoms of the less electronegative element are oxidized and atoms of the more electronegative element are reduced.

5. (a) Since the polyatomic ion contains oxygen, rule 4 applies. Therefore, students should assign an oxidation number of –2 to the oxygen atoms. Rule 7 also applies. Letting the oxidation number of sulfur be x, students can write:

$2x + 3(-2) = -2$

$x = +2$

Therefore, the oxidation number of sulfur in the thiosulfate ion is +2.

(b) Using a Lewis structure, students may arrive at two different oxidation numbers for sulfur using the following reasoning. The central sulfur atom usually has 6 electrons in its valence shell. The three oxygen atoms, however, have a greater electronegativity than sulfur. Therefore, they can be considered to have taken these electrons, leaving the central sulfur with a +6 oxidation number. The sulfur atom bonded to the central sulfur atom may be considered to share its electrons with the central sulfur. Therefore, it will have an oxidation number of –1. Since the central sulfur (+6) can now be considered to have one shared electron it will have a total oxidation number of +6 + (–1) = +5.

(c) Using the first method, the oxidation number for both sulfur atoms is +2. Using the second method, the oxidation number for one sulfur atom is –1, and for the other is +5. (Notice the average of the second result gives the first result).

(d) Using Lewis structures allows for the fact that atoms of the same element in a molecule or polyatomic ion may have different oxidation numbers, depending on the bonding arrangement. The disadvantage to this method is that it is more difficult and time consuming than using the rules.

(e) The advantage to using the oxidation number rules is that they apply to most cases and are quick and easy to use. The disadvantage to using these rules is that they do not allow for the possibility of different oxidation numbers for the same atoms in a molecule or polyatomic ion.

6. (a) $N_2 + 3H_2(g) \rightarrow 2NH_3(g)$

0 0 –3 +1

Nitrogen is the oxidizing agent, and hydrogen is the reducing agent.

(b) $NH_3(aq) + HNO_3 \rightarrow NH_4NO_3(aq)$

–3+1 +1 +5 –2 –3 +1+5 –2

The oxidation numbers of the atoms do not change. Therefore, the reaction is not a redox reaction.

7. (a) Because the reaction takes place in neutral conditions and there are no charges to worry about, the equation may be balanced by inspection. (Students should already be accustomed to balancing combustion equations by inspection.) Students should balance carbon first, then hydrogen, and finally oxygen. Students may also balance the equation using the oxidation number method.

$CH_3COOH(aq) + 2O_2(g) \rightarrow 2CO_2(g) + 2H_2O(\ell)$

(b) Students may use either the half-reaction method or the oxidation number method to balance the equation. The oxidation number method is shown here.

$O_2(g) + H_2SO_3(aq) \rightarrow HSO_4^-(aq)$

0 + 1 +4 –2 +1 +6 –2

Sulfur's oxidation number increases by 2. Oxygen's oxidation number decreases by 2. Therefore, the ratio of sulfur atoms being oxidized to oxygen atoms being reduced must be 1:1.

$O_2(g) + 2H_2SO_3(aq) \rightarrow 2HSO_4^-(aq)$

Add hydrogen ions to balance atoms and charges. The resulting equation is already simplified.

$O_2(g) + 2H_2SO_3(aq)+ \rightarrow 2HSO_4^-(aq) + 2H^+(aq)+$

8. (a) $NH_3(g) + Cl_2(g) \rightarrow NH_4Cl(aq) + N_2(g)$

–3 +1 0 –3 +1 –1 0

Nitrogen's oxidation number increases by 3. Chlorine's oxidation number decreases by 1.

Therefore, the ratio of nitrogen atoms being oxidized to chlorine atoms being reduced must be 1:3. Since 2 nitrogen atoms are oxidized, 6 chlorine atoms must be reduced.

$NH_3(g) + 3Cl_2(g) \rightarrow 6NH_4Cl(aq) + N_2(g)$

Balance nitrogen atoms by inspection.

$8NH_3(g) + 3Cl_2(g) \rightarrow 6NH_4Cl(aq) + N_2(g)$

(b) $Mn_3O_4(aq) + Al(s) \rightarrow Al_2O_3 + Mn(s)$

+8/3 –2 0 +3 –2 0

The oxidation number of manganese decreases by 8/3 and aluminium's oxidation number increases by 3. (The fraction occurs because the manganese oxide is considered to be composed of two different oxides, MnO and Mn_2O_3.) Aluminium is oxidized and each aluminium atom, on reduction, provides three electrons. The ratio of aluminium atoms oxidized to manganese atoms reduced must therefore be 8/3:3, or 8:9.

$3Mn_3O_4(aq) + 8Al(s) \rightarrow Al_2O_3(aq) + Mn(s)$

Balance the atoms by inspection.

$3Mn_3O_4(aq) + 8Al(s) \rightarrow 4Al_2O_3(aq) + 9Mn(s)$

9. The total increase in oxidation numbers corresponds to the number of electrons lost by oxidation. The total decrease in oxidation numbers corresponds to the number of electrons gained by reduction. In a reaction, net gain of electrons must be equal to net loss of electrons. Thus, the total increase in oxidation numbers must equal the total decrease in oxidation numbers.

10. The individual half-reactions are:

$2H_2O(\ell) + NH_3(g) \rightarrow NO_2(g) + 7e^- + 7H^+(aq)$

$4H^+(aq) + 2O_2(g) + 4e^- \rightarrow 2H_2O(\ell)$

The balanced half-reactions are:

$8H_2O(\ell) + 4NH_3(g) \rightarrow 4NO_2(g) + 28e^- + 28H^+(aq)$

$28H^+(aq) + 14O_2(g) + 28e^- \rightarrow 14H_2O(\ell)$

Giving:

$4NH_3(g) + 14O_2(g) \rightarrow 4NO_2(g) + 6H_2O(\ell)$

Which reduces to:

$2NH_3(g) + 7O_2(g) \rightarrow 2NO_2(g) + 3H_2O(\ell)$

The key assumption that students have to make is that an acidic environment is created in the ammonia half-reaction, whereas the oxygen reaction starts in an acidic environment. Since the hydrogen ions and extra waters cancel out, these assumptions do not affect the overall reaction.

12.4 Quantitative Analysis of Redox Reactions

Student Textbook pages 467–472

Section Outcomes

Students will:

- describe calculations to determine quantities of substance involved in redox titrations
- perform a titration experiment
- analyze data from a titration experiment
- calculate the concentration of a reducing agent from titration data

Chemistry Background

- Many metal ions form coloured complexes. When these ions are reduced they undergo colour changes, which is useful for signaling the endpoints of titration reactions.

Teaching Strategies

- Students may need a review of the colours of various transition metal ions. This can be done by having students practise making standard solutions.
- Students will need a review of proper titration technique.

- Students may also need a review of stoichiometry. Remind students of the importance of balancing reactions, then identifying mole ratios based on the transfer of electrons.
- **BLM 12.4.1 (OH) Redox Titrations** has been prepared for this section. You will find it with the Chapter 12 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

In laboratory activities involving colour changes, ensure students with colour-blindness are properly accommodated, either through suitable partnering or by avoiding red-green and yellow-blue colour changes.

Answers to Practice Problems 19–22

Student Textbook page 469

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

19. **(a)** 0.067 74 g
(b) 5.276%

20. 0.138 7%

21. 0.011 0 mol/L

22. 0.174 3 mol/L

Investigation 12.C: Measuring the Concentration of Vitamin C in Orange Juice

Student Textbook pages 470–471

Purpose

The purpose of this investigation is for students to determine the concentration of vitamin C in orange juice, based on the principles of oxidation-reduction reactions. Students will gain experience with the technique of titration, which is a very common and invaluable tool for chemical analysis.

Outcomes

- 30-B1.2s
- 30-B1.3s

Advance Preparation

When to Begin	What to Do
2–3 weeks before	■ Check availability of material.
1–2 days before	■ Prepare iodine and starch solutions. (See **BLM 5.4.4 Preparing Solutions** for tips) ■ Photocopy **BLM 12.4.2 Investigation 12.C.**
1 night before	■ Prepare ascorbic acid solution.

Materials

- iodine solution: 2.59 g of iodine, 26 g of potassium iodide in 50 mL of water. When dissolved, dilute to 6 L.
- 5.68 mmol/L ascorbic acid standard solution
- dropper bottle containing starch indicator: 0.2% w/v starch in water
- orange juice, pulp removed
- de-ionized water
- burette
- burette clamp
- ring stand
- meniscus reader
- volumetric pipette (10.00 mL)
- suction bulb
- 3 Erlenmeyer flasks (125 mL)
- 3 beakers (250 mL)
- labels
- sheet of white paper
- funnel

Time Required

- 80 minutes

Helpful Tips

- Orange juice with pulp can be used but must be measured with a graduated cylinder as the pulp clogs the pipette and is difficult to remove. Sunny Delight® works well and also has the concentration listed on the label; most Canadian brands do not.
- Review proper titration techniques with students. Students commonly make mistakes reading measurements off the burette, or continuing to titrate past the 50.0 mL mark.
- Remind students that they are looking for consistent trials when the green colour first appears after swirling. Furthermore, it is common to ignore a trial where they have gone past the endpoint in their concentration calculations.
- ***Expected Results:*** Endpoints are often difficult to recognize in activities such as this one. Nevertheless, the students should obtain results that are relatively accurate.

Safety Precautions

- The iodine solution can cause stains on clothing and skin. Students should wash any spills with plenty of cool water.
- Students should dispose of the iodine solution used to rinse the burette and any waste iodine solution into the container provided.
- Sodium thiosulfate should be made available to remove any stains that occur.
- Students should thoroughly wash their hands when they have completed the investigation.
- Ensure that students have read "Safety in Your Chemistry Laboratory and Classroom" on pages xii-xv in the student textbook.

Answer to Prediction Question

- This will depend upon the choice of product and label information.

Answers to Analysis Questions

1. The balanced reaction is:

$$C_6H_8O_6(aq) \rightarrow C_6H_6O_6(aq) + 2H^+(aq) + 2e^-$$

$$I_2(aq) + 2e^- \rightarrow 2I^-(aq)$$

$$C_6H_8O_6(aq) + I_2(aq) \rightarrow C_6H_6O_6(aq) + 2I^-(aq) + 2H^+(aq)$$

2. Students will be able to solve for the concentration in the following way:

$$10.00\text{ mL } C_6H_8O_6(aq) \times \frac{\cancel{1\text{ L}}}{\cancel{1000\text{ mL}}} \times \frac{5.68\ \cancel{\text{mmol}}}{\cancel{1\text{ L}}}\ C_6H_8O_6(aq) \times \frac{1\text{ mol}}{1000\ \cancel{\text{mmol}}} \times \frac{1\text{ mol } I_2(aq)}{1\text{ mol } C_6H_8O_6(aq)} \times \frac{1}{\text{average } I_2(aq)\text{ mL}} \times \frac{\cancel{1000\text{ mL}}}{1\text{ L}}$$

3. Students will be able to solve for the concentration in the following way:

$$\text{average } I_2(aq)\ \cancel{\text{mL}} \times \frac{\cancel{1\text{ L}}}{1000\ \cancel{\text{mL}}} \times \cancel{\text{concentration } I_2(aq)\text{ calculated in Question 2 in}}\ \frac{\cancel{\text{mol}}}{\text{L}} \times \frac{1\text{ mol } C_6H_8O_6(aq)}{\cancel{1\text{ mol } I_2(aq)}} \times \frac{1}{10.00\ \cancel{\text{mL}}\text{ orange juice}} \times \frac{1000\ \cancel{\text{mL}}}{\cancel{1\text{ L}}}$$

Answers to Conclusion Questions

4. Typically student answers will depend on their skill in titration.

5. It is unlikely that student answers will match predicted values.

6. Exposure to heat, light, air, and preservatives denatures vitamin C, reducing its concentration. Student accuracy in titrating may also be a considerable source for errors.

7. One reason for this addition is to ensure that the concentration of vitamin C will remain high, even after some of it has been denatured.

Answer to Extension Question

8. Some other common antioxidants are vitamin B-12 in beets and cranberries, vitamin E in green leafy vegetables and whole grains, and beta-carotene in broccoli and carrots. As their name implies, antioxidants counteract the oxidation process. They are beneficial to humans because oxidation in tissues can cause the production of harmful free radicals.

Assessment Options

- Collect and assess students' Predictions, Data Tables and answers to Analysis, Conclusion and Extension questions.
- Use Assessment Checklist 2: Laboratory Report from Appendix A.

Section 12.4 Review Answers

Student Textbook page 472

1. The permanganate ion is a good choice of oxidizing agent because it is very reactive, and it undergoes a colour change. The $MnO_4^-(aq)$ ion is dark purple and turns colourless as it is reduced to Mn^{2+} ions.

2. The starch is added to the mixture to act as a visual indicator.

3. If orange juice (or any antioxidant) has been exposed to the air, it becomes oxidized by oxygen. The concentration of ascorbic acid decreases as the compound reacts to form dehydroascorbic acid.

4. The chlorine oxidized the iodide ions back to triiodide ions, which regenerated some molecular iodine.

5. (a) Students may suggest that iodine could escape as a gas or react with oxygen or organic matter. Therefore, the concentration of the solution is unstable.

(b) Crush a tablet of vitamin C and dissolve it in water. Titrate with iodine. Record the endpoint. You now know the volume of iodine solution required to titrate the amount of vitamin C contained in the tablet, according to the package label.

(c) Over time, the vitamin C in the tablets is oxidized by oxygen in the air.

Chapter 12 Review Answers

Student Textbook pages 474–475

Answers to Understanding Concepts Questions

1. (a) **(i)** $Zn(s) + 2AgNO_3(aq) \rightarrow Zn(NO_3)_2(aq) + 2Ag(s)$

(ii) $3CoBr_2(aq) + 2Al(s) \rightarrow 3Co(s) + 2AlBr_3(aq)$

(iii) $Cd(s) + SnCl_2(aq) \rightarrow CdCl_2(aq) + Sn(s)$

(b) **(i)** Total ionic equation:

$Zn(s) + 2Ag^+(aq) + 2NO_3^-(aq) \rightarrow Zn^{2+}(aq) + 2NO_3^-(aq) + 2Ag(s)$

Net ionic equation:

$Zn(s) + 2Ag^+(aq) \rightarrow Zn^{2+}(aq) + 2Ag(s)$

(ii) Total ionic equation:

$3Co^{2+}(aq) + 6Br^-(aq) + 2Al(s) \rightarrow 3Co(s) + 2Al^{3+}+(aq) + 6Br^-(aq)$

Net ionic equation:

$3Co^{2+}(aq) + 2A1(s) \rightarrow 3Co(s) + 2Al^{3+}(aq)$

(iii) Total ionic equation:

$Cd(s) + Sn^{2+}(aq) + 2Cl^-(aq) \rightarrow Cd^{2+}(aq) + 2Cl^-(aq) + Sn(s)$

Net ionic equation:

$Cd(s) + Sn^{2+}(aq) \rightarrow Cd^{2+}(aq) + Sn(s)$

(c) **(i)** oxidizing agent: $Ag^+(aq)$; reducing agent: $Zn(s)$

(ii) oxidizing agent: $Co^{2+}(aq)$; reducing agent: $Al(s)$

(iii) oxidizing agent: $Sn^{2+}(aq)$; reducing agent: $Cd(s)$

(d) **(i)** oxidation half-reaction: $Zn(s) \rightarrow Zn^{2+}(aq) + 2e^-$

reduction half-reaction: $Ag^+(aq) + e^- \rightarrow Ag(s)$

(ii) oxidation half-reaction: $Al(s) \rightarrow Al^{3+}(aq) + 3e^-$

reduction half-reaction: $Co^{2+}(aq) + 2e^- \rightarrow Co(s)$

(iii) oxidation half-reaction: $Cd(s) \rightarrow Cd^{2+}(aq) + 2e^-$

reduction half-reaction: $Sn^{2+}(aq) + 2e^- \rightarrow Sn(s)$

2. (a) metallic element

(b) nonmetallic element

(c) nonmetallic element

(d) metallic element

3. (a) Ba +2; Cl –1

(b) Al +3; C –4

(c) K +1; C +2; N –3

(d) Li +1; N +3; O –2

(e) N –3; H +1; C +3; O –2

(f) S 0

(g) As +3; O –2

(h) V +5; O –2

(i) Xe +6; O –2; F –1

(j) S +2.5; O –2

4. In its ore, a metal is found in the oxidized form. In other words, it has a positive oxidation number. Pure metal, on the other hand, has an oxidation number of 0. Thus, when pure metal is extracted from ore, the oxidation number of the metal decreases—the metal is reduced.

5. (a) $O_2(g)$ is the oxidizing agent, and $C_6H_6(\ell)$ is the reducing agent.

(b) The reaction is not a redox reaction because the oxidation numbers do not change.

(c) $I_2(s)$ is the oxidizing agent, and $H_2(g)$ is the reducing agent.

(d) $KMnO_4(aq)$ is the oxidizing agent, and $CuCl(s)$ is the reducing agent.

(e) $Cu(s)$ undergoes oxidation, and $Ag^+(aq)+$ undergoes reduction.

(f) The reaction is not a redox reaction because the oxidation numbers do not change.

(g) $Mn^{2+}(aq)$ undergoes oxidation, and BiO_3^- undergoes reduction.

6. (a) oxidation numbers of vanadium are +5 in V_2O_5, VO_2^+, VO_3^- , VO_4^{3-}, $V_3O_9^{3-}$; and +4 in VO_2, VO^{2+}

(b) The reaction is not a redox reaction because the oxidation numbers do not change.

Answers to Applying Concepts Questions

7. The chlorite ion, ClO_2^-, is an example of an ion in which chlorine has an oxidation number of +3.

8. (a) Any synthesis reaction involving changes of oxidation numbers is a redox reaction. The easiest ones to come up with are ones that include elements in the reactants.

$2H_2(g) + O_2(g) \rightarrow 2H_2O(\ell)$

(b) Any synthesis reaction that does not involve electron transfer is not a redox reaction.

$CaO(s) + H_2O(\ell) \rightarrow Ca(OH)_2(aq)$

(c) Any decomposition that involves electron transfer is a redox reaction.

$2C_7H_5(NO_2)_3(s) \rightarrow 12CO(g) + 2C(s) + 5H_2(g) + 3N_2(g)$

(d) Any decomposition that does not involve electron transfer is not a redox reaction.

$CaCO_3(s) \rightarrow CaO(s) + CO_2(g)$

(e) Double displacement reactions are never redox reactions.

(f) Any double displacement reaction will work.

$FeS(s) + 2HCl(aq) \rightarrow H_2S(g) + FeCl_2(aq)$

9. (a) $2P_4(s) + 12H_2O(\ell) \rightarrow 5PH_3(s) + 3H_3PO_4(aq)$

(b) Phosphorus is both reduced and oxidized. Phosphorus atoms, with oxidation number 0, are reduced in

$PH_3(s)$, with oxidation number –3. They are oxidized in $H_3PO_4(aq)$, with oxidation number +5.

10. **(a)** $2Al(s) + Fe_2O_3(s) \rightarrow 2Fe(s) + Al_2O_3(s)$

(b) The oxidizing agent is $Fe_2O_3(S)$; the reducing agent is Al(s).

11. Students should describe a procedure similar to Investigation 12.A, in which they use single displacement reactions to decide whether tin or nickel is the better reducing agent. They will need samples of each metal, and solutions containing ions of the metals. They can then compare observations of tin metal in a solution of nickel ions, and nickel metal in a solution of tin ions. The metal that is the better reducing agent will react in a solution containing ions of the other metal.

12. Oxygen is consumed and carbon dioxide is produced. Therefore, some of the oxygen taken in by the body is reduced, indicating a redox reaction must have taken place.

13. In order for sulfide ions to act as oxidizing agents, they need to gain electrons. Since the sulfide ion has an octet of electrons, it is unlikely that the ion will gain more electrons to act as an oxidizing agent.

Answers to Solving Problems Questions

14. **(a)** $MnO_2(s) + Cl^-(aq) \rightarrow Mn^{2+}(aq) + Cl_2(g)$ (acidic conditions)

$2Cl^-(aq) \rightarrow Cl_2(g)$

$MnO_2(s) \rightarrow Mn^{2+}(aq)$

Balance the oxygen by adding water molecules.

$2Cl^-(aq) \rightarrow Cl_2(g)$

$MnO_2(s) \rightarrow Mn^{2+}(aq) + 2H_2O(\ell)$

Balance the hydrogen atoms by adding hydrogen ions.

$2Cl^-(aq) \rightarrow Cl_2(g)$

$MnO_2(s) + 4H^+(aq) \rightarrow Mn^{2+}(aq) + 2H_2O(\ell)$

Balance charge by adding electrons.

$2Cl^-(aq) \rightarrow Cl_2(g) + 2e^-$

$MnO_2(s) + 4H^+(aq) + 2e^- \rightarrow Mn^{2+}(aq) + 2H_2O(\ell)$

LCM is in place. Therefore, all that remains is to add the half-reaction equations.

$2Cl^-(aq) + MnO_2(s) + 4H^+(aq) \rightarrow Cl_2(g) + Mn^{2+}(aq) + 2H_2O(\ell)$

(b) $NO(g) + Sn(s) \rightarrow NH_2OH(aq) + Sn^{2+}(aq)$ (acidic conditions)

$Sn(s) \rightarrow Sn^{2+}(aq) + NO(g) \rightarrow NH_2OH(aq)$

Balance the hydrogen atoms by adding hydrogen ions.

$Sn(s) \rightarrow Sn^{2+}(aq)$

$NO(g) + 3H^+(aq) \rightarrow NH_2OH(aq)$

Balance charge by adding electrons.

$Sn(s) \rightarrow Sn^{2+}(aq) + 2e^-$

$NO(g) + 3H^+(aq) + 3e^- \rightarrow NH_2OH(aq)$

Determine the LCM for the electrons and rewrite the half–reactions.

$3Sn(s) \rightarrow 3Sn^{2+}(aq) + 6e^-$

$2NO(g) + 6e^- + 6H^+(aq) + \rightarrow 2NH_2OH(aq)$

Add the half-reaction equations.

$2NO(g) + 3Sn(s) + 6H^+(aq) \rightarrow 2NH_2OH(aq) + 3Sn^{2+}(aq)$

(c) $Cd^{2+}(aq) + V^{2+}(aq) \rightarrow Cd(s) + VO_3^-(aq)$ (acidic conditions)

$V^{2+}(aq) \rightarrow VO_3^-(aq)$

$Cd^{2+}(aq) \rightarrow Cd(s)$

Balance the oxygen atoms by adding water molecules.

$V^{2+}(aq) + 3H_2O(\ell) \rightarrow VO_3^-(aq)$

$Cd^{2+}(aq) + \rightarrow Cd(s)$

Balance the hydrogen atoms by adding hydrogen ions.

$V^{2+}(aq) + 3H_2O(\ell) \rightarrow VO_3^- + (aq) + 6H^+(aq)$

$Cd^{2+}(aq) \rightarrow Cd(s)$

Balance charge by adding electrons.

$V^{2+}(aq) + 3H_2O(\ell) \rightarrow VO_3^-(aq) + 6H^+(aq) + 3e^-$

$Cd^{2+}(aq) + 2e^- \rightarrow Cd(s)$

Determine the LCM for the electrons and rewrite the half-reactions.

$2V^{2+}(aq) + 6H_2O(\ell) \rightarrow 2VO_3^-(aq) + 6e^- + 12H^+(aq)$

$3Cd^{2+}(aq) + 6e^- \rightarrow 3Cd(s)$

Add the half-reaction equations.

$3Cd^{2+}(aq) + 2V^{2+}(aq) + + 6H_2O(\ell) \rightarrow 3Cd(s) + 2VO_3^-(aq) + 12H^+(aq)$

(d) $Cr(s) \rightarrow Cr(OH)_4^-(aq) + H_2(g)$ (basic conditions)

$Cr(s) \rightarrow Cr(OH)_4^-(aq)$

$2H^+(aq) \rightarrow H_2(g)$

Balance the oxygen atoms by adding water molecules.

$Cr(s) + 4H_2O(\ell) \rightarrow Cr(OH)_4^-(aq)$

$2H^+(aq) \rightarrow H_2(g)$

Balance the hydrogen atoms by adding hydrogen ions.

$Cr(s) + 4H_2O(\ell) \rightarrow Cr(OH)_4^-(aq) + 4H^+(aq)$

$2H^+(aq) + \rightarrow H_2(g)$

Add hydroxide ions to neutralize hydrogen ions.

$Cr(s) + 4H_2O(\ell) + 4OH^-(aq) \rightarrow Cr(OH)_4^-(aq) + 4H^+(aq) + 4OH^-(aq)$

$2H^+(aq) \rightarrow H_2(g)$

Combine the hydrogen and hydroxide ions and remove any identical molecules that are common to both sides of the equation.

$Cr(s) + 4OH^-(aq) \rightarrow Cr(OH)_4^-(aq)$

$2H^+(aq) \rightarrow H_2(g)$

Balance charge by adding electrons.

$Cr(s) + 4OH^-(aq) \rightarrow Cr(OH)_4^-(aq) + 3e^-$

$2H^+(aq) + 2e^- \rightarrow H_2(g)$

Determine the LCM for the electrons and rewrite the half-reactions.

$2Cr(s) + 8OH^-(aq) \rightarrow 2Cr(OH)_4^-(aq) + 6e^-$

$6H^+(aq) + 6e^- \rightarrow 3H_2(g)$

Add the half-reaction equations. (Combine the hydrogen ions with the hydroxide ions to form water on the reactant side).

$2Cr(s) + 6H_2O(\ell) + 2OH^-(aq) \rightarrow 2Cr(OH)_4^-(aq) + 3H_2(g)$

(e) $S_2O_3^{2-}(aq) + NiO_2(s) \rightarrow Ni(OH)_2(s) + SO_3^{2-}(aq)$ (basic conditions)

Nickel is reduced (+4 to +2) and sulfur is oxidized (+2 to +4).

$S_2O_3^{2-}(aq) \rightarrow SO_3^{2-}(aq)$

$NiO_2(s) \rightarrow Ni(OH)_2(s)$

Balance the number of atoms other than hydrogen and oxygen.

$S_2O_3^{2-}(aq) \rightarrow 2SO_3^{2-}(aq)$

$NiO_2(s) \rightarrow Ni(OH)_2(s)$

Balance the number of oxygen atoms by adding water.

$3H_2O(\ell) + S_2O_3^{2-}(aq) \rightarrow 2SO_3^{2-}(aq)$

$NiO_2(s) \rightarrow Ni(OH)_2(s)$

Balance the number of hydrogen atoms by adding hydrogen ions.

$3H_2O(\ell) + S_2O_3^{2-}(aq) \rightarrow 2SO_3^{2-}(aq) + 6H^+(aq)$

$2H^+(aq) + NiO_2(s) \rightarrow Ni(OH)_2(s)$

Add hydroxide ions to neutralize hydrogen ions.

$3H_2O(\ell) + S_2O_3^{2-}(aq) + 6OH^-(aq) \rightarrow 2SO_3^{2-}(aq) + 6H^+(aq) + 6OH^-(aq)$

$2OH^-(aq) + 2H^+(aq) + NiO_2(s) \rightarrow Ni(OH)_2(s) + 2OH^-(aq)$

Combine the hydrogen and hydroxide ions and remove water molecules common to both sides of the equation.

$S_2O_3^{2-}(aq) + 6OH^-(aq) \rightarrow 2SO_3^{2-}(aq) + 3H_2O(\ell)$

$2H_2O(\ell) + NiO_2(s) \rightarrow Ni(OH)_2(s) + 2OH^-(aq)$

Balance charge by adding electrons.

$S_2O_3^{2-}(aq) + 6OH^-(aq) \rightarrow 2SO_3^{2-}(aq) + 3H_2O(\ell) + 4e^-$

$2H_2O(\ell) + NiO_2(s) + 2e^- \rightarrow Ni(OH)_2(s) + 2OH^-(aq)$

Determine the LCM for the electrons and rewrite the half-reactions.

$S_2O_3^{2-}(aq) + 6OH^-(aq) \rightarrow 2SO_3^{2-}(aq) + 3H_2O(\ell) + 4e^-$

$4H_2O(\ell) + 2NiO_2(s) + 4e^- \rightarrow 2Ni(OH)_2(s) + 4OH^-(aq)$

Add the half-reaction equations and remove water molecules and hydroxide ions common to both sides of the equation.

$S_2O_3^{2-}(aq) + H_2O(\ell) + 2OH^-(aq) + 2NiO_2(s) \rightarrow 2SO_3^{2-} + 2Ni(OH)_2(s)$

(f) $Sn^{2+}(aq) + O_2(g) \rightarrow Sn^{4+}(aq)$ (basic conditions)

$Sn^{2+}(aq) \rightarrow Sn^{4+}(aq)$

$O_2(g) + 2H_2O(\ell) \rightarrow 4OH^-(aq)$

The number of atoms, including hydrogen and oxygen, are balanced. Balance charge and determine the LCM for the electrons and rewrite the half-reactions.

$2Sn^{2+}(aq) \rightarrow 2Sn^{4+}(aq) + 4e^-$

$O_2(g) + 2H_2O(\ell) + 4e^- \rightarrow 4OH^-(aq)$

Add the half-reaction equations.

$2Sn^{2+}(aq) + O_2(g) + 2H_2O(\ell) \rightarrow 2Sn^{4+}(aq) + 4OH^-(aq)$

15. (a) $SiCl_4(aq) + Al(s) \rightarrow Si(s) + AlCl_3(aq)$

+4 –1   0   0   +3–1

Silicon gains four electrons and aluminium loses three electrons. The ratio of silicon tetrachloride to aluminium is 3:4.

$3SiCl_4(aq) + 4Al(s) \rightarrow 3Si(s) + 4AlCl_3(aq)$

(b) $PH_3(aq) + O_2(g) \rightarrow P_4O_{10}(s) + H_2O(\ell)$

–3+1   0   +5 –2   +1 –2

The oxidation number of phosphorus increases by 8. The oxidation number of oxygen decreases by 2. Therefore, the ratio of oxygen atoms to phosphorus atoms must be 8:2, or 4:1. Since there are already 4 phosphorus atoms, the ratio is 16:4.

$4PH_3(s) + 8O_2(g) \rightarrow P_4O_{10}(s) + H_2O(\ell)$

To complete the equation, balance the oxygen atoms on both sides.

$4PH_3(s) + 8O_2(g) \rightarrow P_4O_{10}(s) + 6H_2O(\ell)$

(c) $I_2O_5(s) + CO(g) \rightarrow I_2(s) + CO_2(g)$

+5 –2   +2 –2   0   +4–2

The oxidation number of iodine decreases by 5. The oxidation number of carbon increases by 2.

Therefore, the ratio of carbon to iodine is 5:2.

$I_2O_5(s) + 5CO(g) \rightarrow I_2(s) + 5CO_2(g)$

(d) $SO_3^{2-}(aq) + O_2(g) \rightarrow SO_4^{2-}(aq)$

$+4\ -2 \qquad 0 \qquad +6\ -2$

The oxidation number of sulfur increases by 2. The oxidation number of oxygen decreases by 2. Therefore, the ratio of oxygen reduced to sulfur oxidized is 1:1.

$SO_3^{2-}(aq) + O_2(g) \rightarrow SO_4^{2-}(aq)$

Balance oxygen atoms.

$SO_3^{2-}(aq) + \frac{1}{2}O_2(g) \rightarrow SO_4^{2-}(aq)$, or

$2SO_3^{2-\ +}\ (aq) + O_2(g) \rightarrow 2SO_4^{2-}(aq)$

16. (a) Nitrogen is reduced (+4 to +3) and nitrogen is oxidized (+4 to +5) under acidic conditions.

$NO_2(g) \rightarrow NO_3^-(aq)$; reduction reaction

$NO_2(g) \rightarrow NO_2^-(aq)$; oxidation reaction

Students should balance the equation via the half-reaction method, resulting in the following net ionic equation:

$2NO_2(g) + H_2O(\ell) \rightarrow NO_2^-(aq) + NO_3^-(aq) + 2H^+(aq)$

(b) Chlorine is oxidized (0 to +1) and it is also reduced (0 to –1) under basic conditions.

$Cl_2 \rightarrow 2Cl^-(aq)$; the reduction reaction

$Cl_2(g) \rightarrow ClO^-(aq)$; the oxidation reaction.

Students should balance the equation via the half-reaction method, resulting in the following net ionic equation:

$Cl_2(g) + 2OH^-(aq) \rightarrow ClO^- + (aq) + H_2O(\ell) + Cl^-(aq)$

17. (a) First, students should balance the equation by inspection.

$2Co^{3+}(aq) + Cd(s) \rightarrow 2Co^{2+}(aq) + Cd^{2+}(aq)$

In this case, students can just add the spectator ions, with the following result:

$2Co(NO_3)_3(aq) + Cd(s) \rightarrow 2Co(NO_3)_2(aq) + Cd(NO_3)_2(aq)$

(b) In this equation, silver is reduced and sulfur is oxidized under acidic conditions. Students should follow the half-reaction method to obtain the following equation:

$2Ag^+(aq) + SO_2(g) + 2H_2O(\ell) \rightarrow 2Ag(s) + SO_4^{2-} + (aq) + 4H^+(aq)$

Students can now include the spectator ions and states.

$2AgNO_3(aq) + SO_2(g) + 2H_2O(\ell) \rightarrow 2Ag(s) + H_2SO_4(aq) + 2HNO_3(aq)$

(c) Aluminium is oxidized (0 to +3) and chromium is reduced (+6 to +3) in a basic solution. Students should follow the half-reaction method to obtain the following equation:

$Al(s) + CrO_4^{2-}(aq) + 4H_2O(\ell) \rightarrow A1(OH)_3(s) + Cr(OH)_3(s) + 2OH^-(aq)$

Students can now add the spectator ions and states.

$Al(s) + Na_2CrO_4(aq) + 4H_2O(\ell) \rightarrow Al(OH)_3(s) + Cr(OH)_3(s) + 2NaOH(aq)$

18. (a) $I_2(s) + 10HNO_3(aq) \rightarrow 2HIO_3(aq) + 10NO_2(g) + 4H_2O(\ell)$

(b) $$\text{mass iodine} = \frac{28.0\text{ L }NO_2(g)}{22\ 4\text{ L/mol }NO_2(g)} \times \frac{1\text{ mol }I_2(s)}{10\text{ mol }NO_2(g)} \times \frac{253.8\text{ g }I_2(s)}{1\text{ mol }I_2(s)}$$

$$= 3.17\text{ g }I_2(s)$$

19. (a) Students should use the half-reaction or oxidation number method to balance the reaction, remembering to use basic conditions. The solution is:

$P_4(s) + 3H_2O(\ell) + 3OH^-(aq) \rightarrow 3H_2PO^{2-}(aq) + PH_3(g)$

(b) Phosphorus is both oxidized (0 to +1) and reduced (0 to –3).

(c) $$10.0\text{ kg }P_4(s) \times \frac{1000\text{ g}}{1\text{ kg}} \times \frac{\text{mol }P_4(s)}{123.88\text{ g}} \times \frac{1\text{ mol }PH_3(g)}{1\text{ mol }P_4(s)} \times \frac{34\text{ g}}{1\text{ mol }PH_3(g)}$$

$$= 2744\text{ g }PH_3(g)\text{ or }2.74 \times 10^3\text{ g }PH_3(g)$$

Answers to Making Connections Questions

20. (a) Na +1; Al +3; O –2; H +1; C +4

(b) $NaAl(OH)_2CO_3(s) + 4HCl(aq) \rightarrow NaCl(aq) + AlCl_3(aq) + 3H_2O(\ell) + CO_2(g)$

(c) The oxidation numbers were not particularly useful, since the reaction is not a redox reaction.

(d) The reaction is a neutralization reaction, not a redox reaction.

(e) Student answers will be based on the product they choose. Most antacids contain sodium, calcium, aluminium, or magnesium salts.

21. (a) $2KClO_3(s) \rightarrow 2KCl(s) + 3O_2(g)$

$S(s) + O_2(g) \rightarrow SO_2(g)$

(b) In the first reaction the oxidizing agent and the reducing agent is $KClO_3(s)$. In the second reaction oxygen is the oxidizing agent and sulfur is the reducing agent.

(c) No element in potassium chlorate undergoes disproportionation. Each element appears only once in the formulas of the reactants. Therefore, each element undergoes either reduction or oxidation, not both.

(d) The safety match was invented in 1844 by Swedish professor Gustaf Erik Pasch. Until then, matches had been made using poisonous yellow phosphorus. Their toxicity was not their only drawback. Because the matches could light when struck on any surface, they had a tendency to burst into flames unexpectedly. For his safety match, Pasch used red phosphorus instead of yellow phosphorus. Instead of adding the phosphorus to the match, Pasch impregnated a sandpaper strip on the side of the matchbox with the phosphorus. The match, which had the other reactants on its head, would only light when struck on the special sandpaper strip.

22. (a) The nitrogen is oxidized (–3 to +5) and the oxygen is reduced (0 to –2). Students should assume an acid environment, and balance the equation by the method of their choice.

$$NH_4^+(aq) + 2O_2(g) \rightarrow NO_3^- + (aq) + H_2O(\ell) + 2H^+(aq)$$

(b) Student answers should include references to relevant chemistry, as well as economic and social considerations.

CHAPTER 13 CELLS AND BATTERIES

Curriculum Correlation

(Note: This correlation includes Chapters 12 and 13. Chapter 13 references are in bold.)

General Outcome General Outcome 1: Students will explain the nature of oxidation–reduction reactions.

	Student Textbook	Assessment Options
Outcomes for Knowledge		
30–B1.1k define oxidation and reduction operationally and theoretically	Oxidation, Section 12.1, pp. 434–435 Reduction, Section 12.1, p. 435 Redox Reactions, Section 12.1, pp. 436–437	Questions for Comprehension: 1–4, Section 12.1, p. 437 Section 12.1 Review: 1, 2, p. 440 Chapter 12 Review: 4, p. 474 Chapter 12 Test Unit 6 Review: 1, 4, 6, 33, pp. 526–529
30–B1.2k define the following terms: oxidizing agent, reducing agent, oxidation number, half–reaction, disproportionation	Reducing Iron Ore, Section 12.2, pp. 452–454 Sample Problem: Assigning Oxidation Numbers, Section 12.3, pp. 457–458 Balancing Equations Using the Oxidation Number Method, Section 12.3, p. 462	Questions for Comprehension: 1–4, Section 12.1, p. 437 Section 12.1 Review: 1, 2, p. 440 Chapter 12 Test Unit 6 Review: 1, 4, 6, 33, pp. 526–529
30–B1.3k differentiate between oxidation–reduction (redox) reactions and other reactions using half–reactions and oxidation numbers	Balancing Equations Using Half–Reactions, Section 12.1, pp. 443–445 Balancing a Redox Equation in Basic Solution, Section 12.3, p. 463	Questions for Comprehension: 9–11, Section 12.2, p. 442 Chapter 12 Test Unit 6 Review: 1, 4–7, 11, 26, 29, 33–36, 44, | pp. 526–529
30–B1.4k identify electron transfer, oxidizing agents and reducing agents in redox reactions that occur in everyday life in both living and non–living systems, e.g., corrosion, *cellular respiration, photosynthesis*	Redox Reactions, Section 12.1, p. 436 Balancing Equations Using Half–Reactions, Section 12.2, p. 443 Redox Reactions Involving Molecular Compounds, Section 12.2, pp. 454–456 Sample Problem: Assigning Oxidation Numbers, Section 12.3, p. 457 Quantitative Analysis of Redox Reactions, Section 12.4, p. 467–469 **Fuel Cell History, Section 13.2, pp. 493–497** **Corrosion Prevention, Section 13.2, pp. 499–502**	Questions for Comprehension: 3, Section 12.1, p. 437 Section 12.1 Review: 7, p. 440 Practice Problems: 1, 2, Section 12.2, p. 448 Practice Problems: 12, 13, Section 12.3, p. 461 Section 12.3 Review: 6, p. 466 Chapter 12 Review: 1, 2, 5, 10, 21, pp. 474–475 Chapter 12 Test **Questions for Comprehension: 4, 6, Section 13.1, p. 482** **Section 13.2 Review: 1, p. 501** **Chapter 13 Review: 7, 8, 16, pp. 522–523** Unit 6 Review: 2, 27, 36, 37, 44, 46, pp. 526–529
30–B1.5k compare the relative strengths of oxidizing and reducing agents from empirical data	Investigation 12.A: Testing Relative Oxidizing and Reducing Strengths of Metal Atoms and Ions, Section 12.1 p. 438 Predicting the Spontaneity of Redox Reactions, Section 12.1, pp. 439–440	Investigation 12.A: Testing Relative Oxidizing and Reducing Strengths of Metal Atoms and Ions: 1–4, Section 12.1 p. 438 Questions for Comprehension: 5–8, Section 12.1, p. 440 Chapter 12 Test

	Student Textbook	Assessment Options
30–B1.6k predict the spontaneity of a redox reaction based on standard reduction potentials, and compare predictions to experimental results	Spontaneity of Redox Reactions, Section 12.1, p. 437	
	Predicting the Spontaneity of Redox Reactions, Section 12.1, p. 439	Questions for Comprehension: 5–8, Section 12.1, p. 440 Section 12.1 Review: 5, 6, p. 440 Chapter 12 Test **Chapter 13.3 Review: 3, p. 513** **Chapter 13 Review: 11, pp. 522–523** Unit 6 Review: 3, 25, 38, pp. 526–529
30–B1.7k write and balance equations for redox reactions in acidic, basic and neutral solutions, including disproportionation reactions, by ■ using half–reaction equations obtained from a standard reduction potential table	Writing Balanced Half–Reactions, Section 12.2, pp. 441–442 Balancing Equations Using Half–Reactions, Section 12.2, pp. 442–443 Balancing Equations for Reactions That Occur in Acidic or Basic Solutions, Section 12.2, pp. 444–445 Sample Problem: Balancing an Equation for a Reaction That Occurs in an Acidic Solution, Section 12.2, p. 446 Sample Problem: Balancing an Equation for a Reaction That Occurs in an Basic Solution, Section 12.2, pp. 447–448 Disproportionation Reactions, Section 12.2, p. 450 Sample Problem: Balancing a Disproportionation Reaction, Section 12.2, p. 451	Section 12.1 Review: 5, p. 440 Questions for Comprehension: 9–11, Section 12.2, p. 442 Practice Problems: 1, 2, Section 12.2, p. 448 Practice Problems: 3–6, Section 12.2, p. 452 Section 12.2 Review: 1–5, p. 454 Section 12.3: 6, p. 466 Practice Problems: 11–14, Section 12.3, p. 461 Practice Problems: 22, Section 12.3, p. 469 Chapter 12 Review: 1, 9, 10, 14–22, pp. 474–475 Chapter 12 Test **Section 13.2 Review: 16, p. 501** **Questions for Comprehension: 17, Section 13.3, p. 510** **Section 13.4 Review: 6, p. 520** **Chapter 13 Review: 6, p. 522** Unit 6 Review: 1, 8, 30, 34, 35, 37, 39, 46, p. 526–529
■ equations from information provided about redox changes **Note:** Students are expected to add water molecules, hydrogen ions and hydroxide ions to skeleton equations, as appropriate.	Writing Balanced Half–Reactions, Section 12.2, pp. 441–442 Balancing Equations Using Half–Reactions, Section 12.2, pp. 442–443 Balancing Equations for Reactions That Occur in Acidic or Basic Solutions, Section 12.2, pp. 444–445 Sample Problem: Balancing an Equation for a Reaction That Occurs in an Acidic Solution, Section 12.2, p. 446 Sample Problem: Balancing an Equation for a Reaction That Occurs in an Basic Solution, Section 12.2, pp. 447–448	Questions for Comprehension: 9–11, Section 12.2, p. 442 Practice Problems: 1, 2, Section 12.2, p. 448 Section 12.2 Review: 1–5, p. 454 Chapter 12 Review: 1, 14, pp. 474–475 Chapter 12 Test **Questions for Comprehension: 5, 6, Section 13.1, p. 482** **Practice Problems: 1–4, Section 13.1, p. 487** **Questions for Comprehension: 13, Section 13.2, p. 500** **Section 13.2 Review: 3, 8, 16, p. 501** **Practice Problems: 5, Section 13.3, p. 508** **Questions for Comprehension: 17, 18, Section 13.3, p. 510** **Section 13.3 Review: 4, p. 513** **Chapter 13 Review: 3, 6, 11, 13, 29, pp. 522–523** Unit 6 Review: 7, 11, 34, 36, 44, p. 526–529

	Student Textbook	Assessment Options
30–B1.7k (continued) ■ assigning oxidation numbers, where appropriate, to the species undergoing chemical change	Assigning Oxidation Numbers, Section 12.3, pp. 455–456 Sample Problem: Assigning Oxidation Numbers, Section 12.3, p. 457 Applying Oxidation Numbers to Redox Reactions, Section 12.3, pp. 459–460	Practice Problems: 7–10, Section 12.3, p. 457 Practice Problems: 15–18, Section 12.3, p. 464 Section 12.3: 5–10, p. 466 Chapter 12 Review: 3, 6, 7, 15, 20, pp. 474–475 Chapter 12 Test Unit 6 Review: 5–7, 35, pp. 526–529
30–B1.8k perform calculations to determine quantities of substances involved in redox titrations.	Stoichiometry and Redox Titrations, Section 12.4, pp. 467–468 Sample Problem: Redox Titrations, Section 12.4, pp. 468–469 Investigation 12.C: Measuring the Concentration of Vitamin C in Orange Juice, Section 12.4, pp. 470–471	Practice Problems: 19–22, Section 12.4, p. 469 Investigation 12.C: Measuring the Concentration of Vitamin C in Orange Juice: 1–8, Section 12.4, pp. 470–471 Section 12.4 Review: 1–5, p. 472 Chapter 12 Test Unit 6 Review: 39, 40, pp. 526–529
Outcomes for Science, Technology and Society (Emphasis on science and technology)		
30–B1.1sts explain how the goal of technology is to provide solutions to practical problems by ■ *describing how the process of trial and error was used by early peoples to extract metals from their ores*	Chapter 12 Launch Lab: Penny Chemistry, p. 433 Investigation 12.A: Testing Relative Oxidizing and Reducing Strengths of Metal Atoms and Ions, Section 12.1, p. 438 Connections: How Green Is White Paper? Section 12.3, p. 465	Chapter 12 Launch Lab: Penny Chemistry: 1–6, p. 433 Investigation 12.A: Testing Relative Oxidizing and Reducing Strengths of Metal Atoms and Ions: 4, Section 12.1, p. 438 Connections: How Green Is White Paper? 1–3, Section 12.3, p. 465 Unit 6 Review: 45–48, pp. 526–529
30–B1.2sts explain that technological problems often lend themselves to multiple solutions that involve different designs, materials and processes and have intended and unintended consequences by ■ *analyzing redox reactions used in industry and commerce, e.g., pulp and paper, textiles, water treatment, food processing.*	Investigation 12.B: Redox Reactions and Balanced Equations, Section 12.2, pp. 448–450 Connections: How Green Is White Paper? Section 12.3, p. 465	Investigation 12.B: Redox Reactions and Balanced Equations: 1–19, Section 12.2, pp. 448–450 Connections: How Green Is White Paper? 1–3, Section 12.3, p. 465 Unit 6 Review: 45–48, pp. 526–529
Skill Outcomes (Focus on problem solving)		
Initiating and Planning		
30–B1.1s ask questions about observed relationships and plan investigations of questions, ideas, problems and issues by ■ designing an experiment to determine the reactivity of various metals ■ describing procedures for safe handling, storage and disposal of materials used in the laboratory, with reference to WHMIS and consumer product labelling information	Investigation 12.B: Redox Reactions and Balanced Equations, Section 12.2, pp. 448–450	Investigation 12.B: Redox Reactions and Balanced Equations, Section 12.2: 1–19, pp. 448–450

	Student Textbook	Assessment Options
Performing and Recording		
30–B1.2s conduct investigations into relationships among observable variables and use a broad range of tools and techniques to gather and record data and information by ■ selecting and correctly using the appropriate equipment to perform a redox titration experiment ■ using a standard reduction potential table as a tool in predicting the spontaneity of redox reactions and their products ■ *creating charts, tables or spreadsheets which present the results of redox experiments.*	Investigation 12.C: Measuring the Concentration of Vitamin C in Orange Juice, Section 12.4, pp. 470–471 Investigation 12.A: Testing Relative Oxidizing and Reducing Strengths of Metal Atoms and Ions, Section 12.1, p. 438	Investigation 12.C: Measuring the Concentration of Vitamin C in Orange Juice: 1–8, Section 12.4, pp. 470–471 Investigation 12.A: Testing Relative Oxidizing and Reducing Strengths of Metal Atoms and Ions: 1–4, Section 12.1, p. 438 Investigation 12.A: Testing Relative Oxidizing and Reducing Strengths of Metal Atoms and Ions: 3, Section 12.1, p. 438
Analyzing and Interpreting		
30–B1.3s analyze data and apply mathematical and conceptual models to develop and assess possible solutions by ■ evaluating data from an experiment to derive a simple reduction table ■ *interpreting patterns and trends in data derived from redox reactions* ■ *identifying the limitations of data collected.*	Investigation 12.A: Testing Relative Oxidizing and Reducing Strengths of Metal Atoms and Ions, Section 12.1, p. 438 Investigation 12.B: Redox Reactions and Balanced Equations, Section 12.2, pp. 448–450 Thought Lab 12.1: Oxidation Numbers and Lewis Structures, Section 12.3, pp. 458–459	Investigation 12.A: Testing Relative Oxidizing and Reducing Strengths of Metal Atoms and Ions: 3, Section 12.1, p. 438 Investigation 12.B: Redox Reactions and Balanced Equations: 1–19, Section 12.2, pp. 448–450 Thought Lab 12.1: Oxidation Numbers and Lewis Structures, Section 12.3, pp. 458–459
Communication and Teamwork		
30–B1.4s work collaboratively in addressing problems and apply the skills and conventions of science in communicating information and ideas and in assessing results by ■ selecting and using appropriate numeric, symbolic, graphic and linguistic modes of representation to communicate equations for redox reactions and answers to problems related to redox titrations	Investigation 12.B: Redox Reactions and Balanced Equations, Section 12.2, pp. 448–450 Thought Lab 12.1: Oxidation Numbers and Lewis Structures, Section 12.3, pp. 458–459 Investigation 12.C: Measuring the Concentration of Vitamin C in Orange Juice, Section 12.4, pp. 470–471	Investigation 12.B: Redox Reactions and Balanced Equations, Section 12.2: 1–19, pp. 448–450 Thought Lab 12.1: Oxidation Numbers and Lewis Structures, Section 12.3, pp. 458–459 Investigation 12.C: Measuring the Concentration of Vitamin C in Orange Juice: 1–8, Section 12.4, pp. 470–471

General Outcome 2: Students will apply the principles of oxidation–reduction to electrochemical cells.

	Student Textbook	Assessment Options
Outcomes for Knowledge		
30–B2.1k define anode, cathode, anion, cation, salt bridge/porous cup, electrolyte, external circuit, power supply, voltaic cell and electrolytic cell	**The Voltaic Cell, Section 13.1, p. 478** **Electrolytic Cells, Section 13.3, p. 502**	**Questions for Comprehension: 2, 3, Section 13.1, p. 480** **Questions for Comprehension: 9, Section 13.2, p. 493** **Chapter 13 Review: 1, pp. 522–523** **Chapter 13 Test**

	Student Textbook	Assessment Options
30–B2.2k identify the similarities and differences between the operation of a voltaic cell and that of an electrolytic cell	**The Voltaic Cell, Section 13.1, p. 478** **Electrolytic Cells, Section 13.3, p. 502**	**Questions for Comprehension: 2, 3, Section 13.1, p. 480** **Questions for Comprehension: 8, 9, Section 13.2, p. 493** **Section 13.2 Review: 9, p. 501** **Chapter 13 Review: 1, 13, pp. 522–523** **Chapter 13 Test** Unit 6 Review: 13, 23, 24, 28, 32, pp. 526–529
30–B2.3k predict and write the half–reaction equation that occurs at each electrode in an electrochemical cell	**Electrolytic Cells, Section 13.3, p. 502**	**Practice Problems: 5–8, Section 13.3, p. 508** **Questions for Comprehension: 17–20, Section 13.3, p. 510** **Chapter 13 Review: 13, 19, 20, pp. 522–523** **Chapter 13 Test** Unit 6 Review: 11, 31, 36, 44, pp. 526–529
30–B2.4k recognize that predicted reactions do not always occur, e.g., the production of chlorine gas from the electrolysis of brine	**Predicting the Products of Electrolysis of Aqueous Solutions, Section 13.3, p. 507** **Sample Problem: Electrolysis of an Aqueous Solution, Section 13.3, p. 507**	**Practice Problems 5–8, Section 13.3, p. 508** **Questions for Comprehension: 17, 18, Section 13.3, p. 510** **Chapter 13 Review: 5, 14, p. 513** **Chapter 13 Test** Unit 6 Review: 9, 17, 20, pp. 526–529
30–B2.5k explain that the values of standard reduction potential are all relative to zero volts set for the hydrogen electrode at standard conditions	**Cell Potentials, Section 13.1, pp. 482–483**	**Thought Lab 13.1: Assigning Reference Values: 1–5, Section 13.1, p. 487** **Section 13.1 Review: 5–7, p. 490** **Questions for Comprehension: 13–16, Section 13.2, p. 500** **Chapter 13 Review: 15, pp. 522–523** **Chapter 13 Test** Unit 6 Review: 9, 13, 23, 25, 31, pp. 526–529
30–B2.6k calculate the standard cell potential for electrochemical cells	**Calculating Standard Cell Potentials, Section 13.1, pp. 485–486** **Thought Lab 13.1: Assigning Reference Values: 1–5, Section 13.1, p. 487** **Sample Problem: Calculating a Standard Cell Potential, Given a Net Ionic Equations, Section 13.1, p. 486** **Sample Problem: Calculating a Standard Cell Potential, Given a Chemical Reaction, Section 13.1, pp. 486–487**	**Thought Lab 13.1: Assigning Reference Values: 1–5, Section 13.1, p. 487** **Practice Problems: 1–4, Section 13.1, p. 487** **Section 13.1 Review: 3–7, p. 490** **Questions for Comprehension: 13, Section 13.2, p. 500** **Chapter 13 Review: 6, 11, 12, 18, 21, pp. 522–523** **Chapter 13 Test** Unit 6 Review: 23, 28, 36, 38, pp. 526–529

	Student Textbook	Assessment Options
30–B2.7k predict the spontaneity or non–spontaneity of redox reactions based on standard cell potential and the relative positions of half–reaction equations on a standard reduction potential table	**Predicting the Products of Electrolysis of Aqueous Solutions, Section 13.3, p. 507** **Sample Problem: Electrolysis of an Aqueous Solution, Section 13.3, p. 507**	Questions for Comprehension: 5–8, Section 12.1, p. 440 Section 12.1 Review; 3, 5, 6, p. 440 **Practice Problems: 5-8, Section 13.3, p. 508** **Section 13.3 Review: 3, p. 513** **Chapter 13 Review: 11, pp. 522–523** **Chapter 13 Test** Unit 6 Review: 3, 38, pp. 526–529
30–B2.8k calculate mass, amounts, current and time in single voltaic and electrolytic cells by applying Faraday's law and stoichiometry.	**Stoichiometry and Faraday's Law, Section 13.4, p. 514** **Sample Problem: Calculating the Mass of an Electrolysis Product, Section 13.4, pp. 515–516** **Faraday's law, Section 13.4, pp. 516–517**	**Practice Problems: 9–12, Section 13.4, p. 516** **Section 13.4 Review: 1–6, p. 520** **Chapter 13 Review: 16, 21–26, pp. 522–523** **Chapter 13 Test** Unit 6 Review: 41–43,, pp. 526–529
Outcomes for Science, Technology and Society (Emphasis on science and technology)		
30–B2.1sts describe the ways in which scientific knowledge may lead to the development of new technologies and new technologies may lead to scientific discoveries by ■ *analyzing the relationship of scientific knowledge and technological development in the applications of voltaic and electrolytic cells for, e.g., batteries, electroplating, refining metals from ores, electrowinning, sanitizing swimming pools with chlorine compounds*	Connections: How Green Is White Paper? Section 12.3, p. 465 **Chapter 13 Launch Lab: What Determines Voltage? p. 477**	Connections: How Green Is White Paper? 1–3, Section 12.3, p. 465 **Chapter 13 Launch Lab: What Determines Voltage? 1–5, p. 477** Unit 6 Review: 45–48, pp. 526–529
30–B2.2sts describe applications of science and technology that have developed in response to human and environmental needs by ■ *investigating the use of technology to solve practical problems related to corrosion; e.g., galvanizing, metallurgy, magnesium coupling, painting*	Chapter 12 Launch Lab: Penny Chemistry, p. 433 Connections: How Green Is White Paper? Section 12.3, p. 465 **Chapter 13 Launch Lab: What Determines Voltage? p. 477**	Chapter 12 Launch Lab: Penny Chemistry: 1–6, p. 433 Connections: How Green Is White Paper? 1–3, Section 12.3, p. 465 **Chapter 13 Launch Lab: What Determines Voltage? 1–5, p. 477** Unit 6 Review: 45–48, pp. 526–529
30–B2.3sts illustrate how science and technology are influenced and supported by society and have influenced and been influenced by historical development and societal needs by ■ *assessing the economic importance of electrochemical cells, particularly fuel cells, to modern society, predicting their future importance in transportation, the recycling of metals and in reducing emissions from smokestacks.*	Connections: How Green Is White Paper? Section 12.3, p. 465	Connections: How Green Is White Paper? 1–3, Section 12.3, p. 465 Unit 6 Review: 45–48, pp. 526–529

	Student Textbook	Assessment Options
Skill Outcomes (Focus on problem solving)		
Initiating and Planning		
30–B2.1s ask questions about observed relationships and plan investigations of questions, ideas, problems and issues by ■ designing an experiment, including a labelled diagram, to test predictions regarding spontaneity, products and the standard cell potential for reactions occurring in electrochemical cells ■ describing procedures for safe handling, storage and disposal of materials used in the laboratory, with reference to WHMIS and consumer product labelling information ■ *developing a plan to build a battery, seeking feedback, testing and reviewing the plan, and making revisions to the plan.*	**Investigation 13.C: Electroplating, Section 13.4, pp. 517–518**	**Investigation 13.C: Electroplating: 1–8, Section 13.4, pp. 517–518** Unit 6 Review: 22, 44, pp. 526–529
Performing and Recording		
30–B2.2s conduct investigations into relationships between and among observable variables and use a broad range of tools and techniques to gather and record data and information by ■ constructing and observing electrochemical cells ■ *investigating the issue of the disposal of used batteries and proposing alternative solutions to this problem* ■ *compiling and displaying evidence and information about voltaic and electrochemical cells, by hand or using technology, in a variety of formats, including diagrams, flow charts, tables, graphs and scatterplots.*	**Chapter 13 Launch Lab: What Determines Voltage? p. 477** **Investigation 13.A: Measuring Cell Potentials of Voltaic Cells, Section 13.1, pp. 488–489**	**Chapter 13 Launch Lab: What Determines Voltage? 1–5, p. 477** **Investigation 13.A: Measuring Cell Potentials of Voltaic Cells: 1–13, Section 13.1, pp. 488–489** Unit 6 Review: 22, 44, pp. 526–529
Analyzing and Interpreting		
30–B2.3s analyze data and apply mathematical and conceptual models to develop and assess possible solutions by ■ identifying the products of electrochemical cells ■ comparing predictions with observations of electrochemical cells ■ identifying the limitations of data collected on a electrochemical cell ■ explaining the discrepancies between theoretical and the actual cell potential ■ *assessing the efficiencies and practicalities of various electrochemical configurations as batteries* ■ *evaluating experimental designs for voltaic and electrolytic cells and suggesting improvements and alternatives*	**Chapter 13 Launch Lab: What Determines Voltage? p. 477** **Thought Lab 13.1: Assigning Reference Values, Section 13.1, p. 487** **Investigation 13.A: Measuring Cell Potentials of Voltaic Cells, Section 13.1, pp. 488–489** **Investigation 13.B: Electrolysis of Aqueous Potassium Iodide, Section 13.3, pp. 508–509**	**Chapter 13 Launch Lab: What Determines Voltage? 1–5, p. 477** **Thought Lab 13.1: Assigning Reference Values: 1–5, Section 13.1, p. 487** **Investigation 13.A: Measuring Cell Potentials of Voltaic Cells: 1–13, Section 13.1, pp. 488–489** **Investigation 13.B: Electrolysis of Aqueous Potassium Iodide: 1–9, Section 13.3, pp. 508–509** Unit 6 Review: 10, 39, 40–43, pp. 526–529

	Student Textbook	Assessment Options
Communication and Teamwork		
30–B2.4s work collaboratively in addressing problems and apply the skills and conventions of science in communicating information and ideas and in assessing results by ■ using appropriate SI notation, fundamental and derived units to communicate answers to problems related to functioning electrolytic cells. ■ creating multiple–linked documents, selecting and integrating information from various print and electronic sources or from several parts of the same source, to prepare a presentation on the use of hydrogen fuel cells for transportation and heating.	**Investigation 13.B: Electrolysis of Aqueous Potassium Iodide, Section 13.3, pp. 508–509**	**Investigation 13.B: Electrolysis of Aqueous Potassium Iodide: 1–9, Section 13.3, pp. 508–509**

Chapter 13

Cells and Batteries

Student Textbook pages 476–523

Chapter Concepts

Section 13.1 Voltaic Cells

- Voltaic cells transform chemical energy into electrical energy.
- The chemical reaction determines the cell voltage.

Section 13.2 Applications of Voltaic Cells

- The common flashlight battery is a dry cell.
- Fuel cells offer a potentially clean form of energy for the future if they can be produced economically.

Section 13.3 Electrolytic Cells

- Electrolysis is the process in which electrical energy is used to drive non-spontaneous redox reactions.
- You can use standard reduction potentials to calculate the potential difference needed to drive a specific reaction.
- Electrolysis is used to produce many industrially important products such as chlorine, sodium hydroxide, and aluminium.

Section 13.4 Stoichiometry and Faraday's Law

- You can apply stoichiometric calculations to redox reactions after determining the charge on a mole of electrons.
- You can calculate the amount of time that a specific current is needed in an electrochemical cell to deposit a known amount of mass on an electrode.
- Some metals are refined by electrolysis of molten salts or salt solutions.

Common Misconceptions

- A common misconception in drawing diagrams of voltaic cells is that the anode must appear on the left. The convention of "anode on the left" applies only to shorthand notation introduced on page 481 of the student textbook. In a diagram of a voltaic cell, the anode can appear on the left or the right, as it can in a lab setting.
- The false idea that free electrons can travel through an electrolyte solution is a major student misconception. A current is usually defined as a flow of electrons, but free electrons do not flow through the electrolyte in a cell. The current flow in a cell involves the flow of anions and cations and a redox reaction (reduction at the cathode and oxidation at the anode). Address this misconception by ensuring that any diagrams you or the students create show the movement of ions, not electrons, though the electrolyte.
- Students may wrongly believe that the quantity (volume) of solution in a cell affects cell potential. The type of substances present and solution concentration affects standard cell potentials, but not the volume of solution present. This can be checked during Investigation 13.A (pp. 488–489) when students construct voltaic cells.
- Students use cell potentials to decide whether reactions will be "spontaneous." The concept of spontaneity is a thermodynamic concept, and it means that a reaction is energetically favourable. A spontaneous reaction may not actually take place. There may be another favourable reaction between the same reactants that takes place instead, or the reaction may require a high activation energy.
- Many students have misconceptions about what is taking place in a voltaic cell, yet are still able to calculate cell potentials correctly. This is consistent with research suggesting that students capable of solving quantitative problems often lack an understanding of the underlying concepts. For example, they may believe that the value of zero for the standard potential of the hydrogen half-cell is based on the chemistry of $H^+(aq)$ and $H_2(g)$. As Thought Lab 13.1 (p. 487) is intended to show, another half-cell could have been chosen as the reference half-cell.
- Students may incorrectly believe that current flows in a voltaic cell because there is a difference in charge at the anode and cathode. Each half-reaction has a different reduction potential, and when the cell circuit is completed, electrons are forced through the external circuit.
- A related misconception involves the salt bridge, when students believe that it allows electrons to flow and complete the circuit. Ask students what would happen to a voltaic cell if the salt bridge were replaced with a copper wire. Students can check their prediction during Investigation 13.A (pp. 488–489). The ions in a salt bridge play the important role of carrying electrical charges to complete the electrical circuit. Although copper is a conductor, it conducts electrons, not ions.
- Some students believe that half-cells need not be electrically neutral; one half-cell can be positive with cations, and the other negative with an equal number of anions. This is not true. Ions migrate in the internal circuit of the cell while electrons flow in the external circuit, maintaining electrically neutral half-cell compartments.

Helpful Resources

Books and Journal Articles

- Sanger, M.J. and Greebowe, T.J. "An Analysis of College Chemistry Textbooks as Sources of Misconceptions and Errors in Electrochemistry," *Journal of Chemical Education.* Vol. 76, 1999, p. 853.
- Armstrong, R. and Robertson, A. "Time to Recharge," *Chemistry in Britain.* 2002, p. 38.
- Brett, C.M.A. and Brett, A.M.O. Electrochemistry: *Principles, Methods, and Applications.* Oxford University, Oxford, 1993.
- Koppel, T. *Powering the Future: The Ballard Fuel Cell and The Race To Change The World.* Wiley, New York, 1999.

- Larminie, J. and Dicks, A., *Fuel Cell Systems Explained.* Wiley, New York, 2000.
- Wright, S.W. "Trusty or Rusty? Oxidation Rate of Nails," *Journal of Chemical Education.* Vol. 82, 2005, p. 1648A.

Web Sites

Web links related to cells and batteries 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

13.0.1 (HAND) Launch Lab: What Determines Voltage?
13.0.1A (ANS) Launch Lab: What Determines Voltage? Answer Key

13.1.1 (OH) The Motion of Ions and Electrons
13.1.2 (OH) The Hydrogen Half-Cell with Zinc
13.1.3 (HAND) Thought Lab 13.1: Assigning References Values
13.1.3A (ANS) Thought Lab 13.1: Assigning References Values Answer Key
13.1.4 (HAND) Investigation 13.A: Measuring Cell Potentials of Voltaic Cells
13.1.4A (ANS) Investigation 13.A: Measuring Cell Potentials of Voltaic Cells Answer Key
13.1.5 (AST) Voltaic Cell Notation
13.1.5A (ANS) Voltaic Cell Notation Answer Key

13.2.1 (OH) Fuel Cell Technology

13.3.1 (OH) Electrolytic Cells

13.3.1A (ANS) Electrolytic Cells Answer Key
13.3.2 (OH) The Chlor-alkali Cell
13.3.3 (HAND) Investigation 13.B: Electrolysis of Aqueous Potassium Iodide
13.3.3A (ANS) Investigation 13.B: Electrolysis of Aqueous Potassium Iodide Answer Key

13.4.1 (OH) Calculating the Mass of an Electrolysis Product
13.4.2 (HAND) Investigation 13.C: Electroplating
13.4.2A (ANS) Investigation 13.C: Electroplating Answer Key

13.5.1 (AST) Chapter 13 Test
13.5.1A (ANS) Chapter 13 Test Answer Key

Using the Chapter 13 Opener

Student Textbook pages 476–477

Teaching Strategies

- Bring in samples of batteries to help illustrate the text in the opener.

Launch Lab What Determines Voltage?

Student Textbook page 477

Purpose

The purpose of this lab is to introduce students to a practical application of redox chemistry. Students will be encouraged to identify the relationship between redox reactions and how batteries work.

Outcomes

- 30-B2.1s
- 30-B2.2s

Advance Preparation

When to Begin	What to Do
2–3 weeks before	■ Order the required supplies.
1 day before	■ Acquire lemons. ■ Photocopy **BLM 13.0.1: Launch Lab.**

Materials
■ 4 zinc strips (1 cm × 5 cm) ■ 4 copper strips (1 cm × 5 cm) ■ 4 lemons ■ fine sandpaper ■ 8 electrical leads with alligator clips ■ voltmeter set to scale of 0 V to 5 V ■ small flashlight bulb

Time Required

- 60 minutes

Helpful Tips

- Remind students to use the sandpaper to clean the contact points of the metal strips and alligator clips.
- Before making cuts in the lemon, roll it around on the bench while pressing down on it. This will release as much juice as possible. Make small cuts in the lemon and insert the zinc and copper strips. (Use a penny and a galvanized nail if strips are not available.) Attach the strips to a galvanometer using the alligator clips, as shown in the

student textbook. You can try powering a digital watch with the lemon cell. Remove the button battery, and then attach the leads to the connections. (Try switching the leads if the watch does not work. You may have to connect several lemon cells in series to get enough current.) You and your students can experiment with other fruits and vegetables. Grapefruits will work, and so will potatoes.

- Use **BLM 13.0.1 (HAND) Launch Lab: What Determines Voltage?** to support this activity. Remove sections as appropriate to meet the needs of students in your class.
- ***Expected Results:*** The series lemons will produce the highest voltage.

Safety Precautions

- Since lemon juice is acidic, remind students to wear their goggles.
- Ensure that students have read "Safety in Your Chemistry Laboratory and Classroom" on pages xii-xv in the student textbook.

Answers to Analysis Questions

1. (a) The connection in series provides a greater voltage, as the effects of the cells are cumulative.

(b) Again, connecting the lemons in series should provide the most light.

3. The main difference between a cell and a battery is that a battery is made up of several cells connected together.

4. Batteries of different size may produce the same voltage since they may contain the same components in the same concentrations. However, larger batteries are used for purposes that have a higher electrical demand and so are designed to last longer.

Assessment Options

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

13.1 Voltaic Cells

Student Textbook pages 478–490

Section Outcomes

Students will:

- predict the spontaneity of redox reactions
- evaluate data from an experiment to derive a simple reduction table
- define anode, cathode, salt bridge/porous cup, electrolyte, external circuit, and voltaic cell
- predict and write half-reactions that occur at each electrode in an electrochemical cell
- explain that the values of standard reduction potential are all relative to the hydrogen electrode at standard conditions
- construct and observe electrochemical cells
- compare predictions about and observations of electrochemical cells
- evaluate experimental designs for voltaic cells and suggest improvements
- explain that the hydrogen electrode is set at 0 V as a reference for all half cells at standard conditions
- calculate standard cell potentials for electrochemical cells

Key Terms

electric current
electrochemistry
voltaic cell
external circuit
electrodes
electrolytes
anode
cathode
salt bridge
inert electrode
cell notation
electrical potential difference
voltage
cell potential
standard cell potential
standard reduction potential

Chemistry Background

- A voltaic cell involves spontaneous oxidation and reduction half-reactions that occur at two electrodes—the anode and the cathode. Oxidation occurs at the anode, which is negatively charged, and reduction occurs at the cathode, which is positively charged. Since the reactants in the redox reaction do not come in contact with one another, electrons must flow through an external circuit from the anode to the cathode, creating an electric current.
- Standard reduction potentials are used to calculate cell potentials for reactions that take place under standard state conditions (101.3 kPa and 25 °C for 1.0 mol/L solution of the chemical involved). Standard half-cell potentials are measured relative to the standard hydrogen electrode, which has been assigned the value of 0.000 V. This scale is similar to enthalpy, another relative scale, in which enthalpy changes are defined based on zero values assigned for convenience for elements in their standard states.
- Cell voltage is dependent on the nature of the reactants and the concentration of the electrolytes, but not on the *quantity* of electrolyte present (i.e., cells using 10 mL or 1 L of identical electrolyte solutions should produce the same voltage).
- The Sample Problem on page 486 demonstrates one of two methods of calculating standard cell potentials. The other method involves changing the positive or negative sign of the anode reaction, since the anode reaction is the

reverse of the cathode reaction it is given the reverse value. Next the E°cathode and E°anode are added together. This method works well for weaker students who are uncomfortable subtracting negative numbers. The two methods are technically equivalent, however, the subtraction method is used for the remainder of the chapter for pedagogical reasons.

- Students use cell potentials to decide whether reactions will be "spontaneous." At this stage, students will think of spontaneous as meaning that a reaction happens. More strictly, it means that a reaction is energetically favourable. A "spontaneous" reaction (i.e., a reaction that is energetically favourable) may not actually take place. There may be another favourable reaction between the same reactants that takes place instead. Also, the concept of spontaneity is a thermodynamic concept. Kinetics can come into play as well. For example, the combustion of human bodies in air is spontaneous. However, under normal conditions, we are kinetically stable in air. Fortunately for us, the activation energy is high, and the spontaneous reaction does not start under normal conditions. (The fact that it does not happen does not make it any less "spontaneous," within the technical meaning of the term).

Teaching Strategies

- Discuss the different types, uses, and prevalence of batteries in Canadian society. Have each student estimate the number of used batteries they discard in a year.
- Electrodes are not named according to their charge but rather according to the chemical reaction in which they take part. A good mnemonic for naming electrodes is: RED CAT (reduction at the cathode) and AN OX (oxidation at the anode). This mnemonic works because "an cat" does not make sense. Alternatives include CCC (cations consumed at cathode) and APA (anions produced at anode). A creative class may enjoy coming up with others of their own.
- If students are having difficulty with the arbitrary nature of the standard reference cell, have them create a horizontal line on a piece of paper or on the chalkboard. Beneath the line, label one end 0 cm and the other 100 cm. Above the line, label the middle 0 cm, with –50 cm on one end and 50 cm on the other. A third set of values could be chosen to represent –30 cm, 0 cm, and 70 cm.
- Although the textbook presents two methods of calculating standard cell potentials, the subtraction method is preferred. Using the addition method may perpetuate the misconception that half-cell potentials are measurable in isolation and have meaning in an absolute sense. The subtraction method emphasizes that only the difference has meaning. Using the subtraction method also avoids the need to explain why you reverse the sign of a half-cell potential when you reverse the equation, but do not change the potential when you multiply an equation by a numerical coefficient. This will seem confusing to a student who may feel that potential difference should also be multiplied, according to Hess's Law. Unlike enthalpy, however, half-cell potentials are intensive properties like density, not extensive properties like enthalpy. The potential difference between two points in a circuit is the amount of work done in moving a unit quantity of charge between the two points. If the quantity of charge is doubled, the work done to move the charge will also double, and the potential difference remains the same.
- As you work through sample problems on the chalkboard, illustrate your answers using potential ladder diagrams, similar to the one shown in Table 13.1 on page 485 of the student textbook. This will help those of your students who are visual learners.
- Overhead masters and quizzes have been prepared for this section. You will find them with the Chapter 13 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
  13.1.1 (OH) The Motion of Ions and Electrons
  13.1.2 (OH) The Hydrogen Half-Cell with Zinc
  13.1.5 (AST) Voltaic Cell Notation
  13.1.5A (ANS) Voltaic Cell Notation Answer Key

SUPPORTING DIVERSE STUDENT NEEDS

- Students with weaker math skills may need a reminder in working with positive and negative numbers. Pair them up with more mathematically advanced students.

Chemistry File: Web Link

Student Textbook page 479

Galvani discovered that, when placed in a container of saline solution, the frog's legs would twitch when he touched them with his metal instrument (completing the circuit). Galvani thought that he had discovered bio-electricity, a form of electricity present in living things. However, Volta suggested that it was a current being transferred that caused the frog's muscles to contract.

Answers to Questions for Comprehension

Student Textbook page 480

Q1. The zinc atoms can reduce the copper(II) ions, even though the two metals are not in contact, because the electrons from the zinc travel through the electrodes via the connecting wire, before coming into contact with the copper(II) ions in solution.

Q2. The function of the salt bridge or porous cup in the voltaic cell is to allow movement of the excess anions and cations, preventing a build-up of charge in each half-cell. In the zinc-copper example, copper(II) ions are turned into copper atoms, leaving the sulfate ions behind. Over time, the negatively charged sulfate ions would build-up around the copper electrode. However, this is prevented by positively charged potassium ions in the salt bridge being attracted to the sulfate ions. Similarly, the chloride ions of the salt bridge are attracted to the increasing number of zinc ions.

Q3. The main difference between the two half-cells is what happens at the electrodes. In the cathode, cations are converted into atoms, causing the electrode to gain in mass. Conversely, the electrode at the anode produces cations and consequently loses mass over time.

Figure 13.5

Student Textbook page 481

An inert electrode like platinum is often used when the substance at either the anode or cathode produces a liquid or a gas. However, platinum has the added quality of being even less reactive than gold.

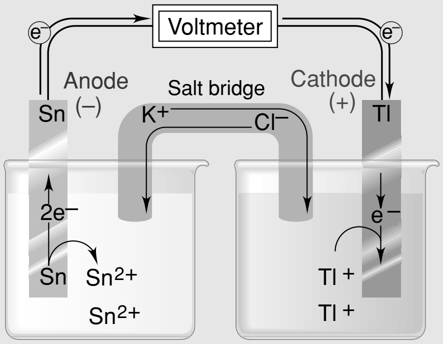
Answers to Questions for Comprehension

Student Textbook page 482

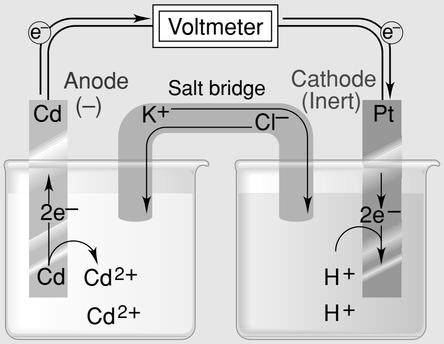
Q4. (a) In a test tube, the zinc would be the reducing agent and the copper(II) ions would be the oxidizing agent.

(b) In a Daniel cell, the copper(II) ions and the copper metal would be at the cathode while the zinc and zinc ions would be at the anode. Again students can determine this by remembering that electron flow is from anode to cathode, from reducing agent to oxidizing agent.

Q5. (a)



(b)



(a) anode: $Sn(s) \rightarrow Sn^{2+}(aq) + 2e^-$
cathode: $Tl^+(aq) + 1e^- \rightarrow Tl(s)$
Net: $Sn(s) + 2Tl^+(aq) \rightarrow Sn^{2+}(aq) + 2Tl(s)$

(b) anode: $Cd(s) \rightarrow Cd^{2+}(aq) + 2e^-$
cathode: $2H^+(aq) + 2e^- \rightarrow H_2(g)$
Net: $Cd(s) + 2H^+(aq) \rightarrow Cd^{2+}(aq) + H_2(g)$

Platinum is an inert electrode. A comma indicates that both species are in the same phase. There is a line instead of a comma to indicate that hydrogen gas is produced.

Q6. (a) The two half-reactions are:

$$MnO_4^-(aq) + 8H^+(aq) + 5e^- \rightarrow Mn^{2+}(aq) + 4H_2O(\ell)$$

$$2I^-(aq) \rightarrow I_2(s) + 2e^-$$

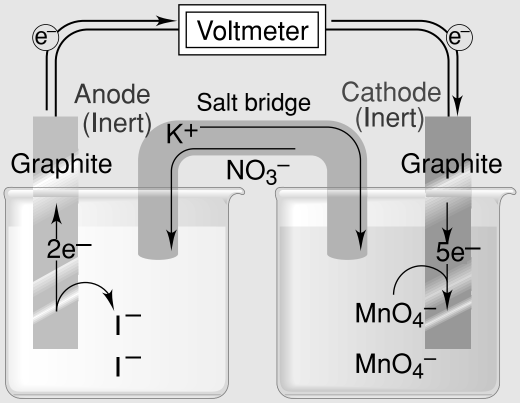
The overall balanced reaction is:

$$2MnO_4^-(aq) + 16H^+(aq) + 10I^-(aq) \rightarrow 5I_2(s) + 2Mn^{2+}(aq) + 8H_2O(\ell)$$

(b) The oxidizing agent is $MnO_4^-(aq)$ and the reducing agent is $I^-(aq)$.

(c) To form solid iodine from iodide ions, electrons must be lost (i.e. the ions must undergo oxidation). Since oxidation occurs at the anode, the solid iodine forms at the anode.

(d)



Q7. There is no double vertical line in the notation because there is no salt bridge or porous cup in the lemon.

Answers to Practice Problems 1–4

Student Textbook page 487

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

1. +0.29 V
2. +3.17 V
3. +0.14 V
4. +1.54 V

Thought Lab 13.1: Assigning Reference Values

Student Textbook page 487

Purpose

This activity is designed to help students understand the central, yet arbitrary, role that the hydrogen half-reaction plays in calculating reference cell values.

Outcomes

- 30-B2.5k
- 30-B2.3s

Time Required

- 30 – 45 minutes

Helpful Tips

- The intent of this exercise is to reinforce the idea that zero values are often arbitrarily set and usually involve clearly defined standards or conditions. Let students know that standard molar enthalpies of formation are another example of a relative scale. The standard enthalpy of formation of an element in its standard state is arbitrarily defined as 0 kJ/mol. All other enthalpies of formation are measured relative to those zero values.
- Pair stronger students with students who are having difficulty understanding the principles behind this activity.
- Use **BLM 13.1.3 (HAND) Thought Lab 13.1: Assigning Reference Values** to support this activity. Remove sections as appropriate to meet the needs of students in your class.

Answers to Procedure Questions

1.

Reduction half-reaction	$E°$ (V)	$E°$ + 1.662 (V)
$F_2(g) + 2e–$ $2F^-(aq)$	2.866	4.528
$Fe^{3+}(aq) + e^-$ $Fe^{2+}(aq)$	0.771	2.433
$2H^+(aQ) + 2e^-$ $H_2(g)$	0.000	1.662
$Al^{3+}(aq) + 3e^-$ $Al(s)$	–1.662	0.000
$Li^+(aq) + e^-$ $Li(s)$	–3.040	–1.378

2. (a) $E^o_{cell} = E^o_{cathode} - E^o_{anode}$
$= 0.00 \text{ V} - (-3.04 \text{ V}) = +3.04 \text{ V}$

(b) $E^o_{cell} = E^o_{cathode} - E^o_{anode}$
$= 2.87 \text{ V} - (-1.66 \text{ V}) = +4.53 \text{ V}$

(c) $E^o_{cell} = E^o_{cathode} - E^o_{anode}$
$= 0.77 \text{ V} - 0.000 \text{ V} = +0.77 \text{ V}$

(d) $E^o_{cell} = E^o_{cathode} - E^o_{anode}$
$= -1.66 \text{ V} - (-3.04 \text{ V}) = +1.38 \text{ V}$

3. (a) $E^o_{cell} = E^o_{cathode} - E^o_{anode}$
$= 1.66 \text{ V} - (-1.38 \text{ V}) = +3.04 \text{ V}$

(b) $E^o_{cell} = E^o_{cathode} - E^o_{anode}$
$= 4.53 \text{ V} - 0.00 \text{ V} = +4.53 \text{ V}$

(c) $E^o_{cell} = E^o_{cathode} - E^o_{anode}$
$= 2.43 \text{ V} - 1.66 \text{ V} = +0.77 \text{ V}$

(d) $E^o_{cell} = E^o_{cathode} - E^o_{anode}$
$= 0.00 \text{ V} - (-1.38 \text{ V}) = +1.38 \text{ V}$

Answers to Analysis Questions

1. (a) Shifting the zero value by +1.66 V increased each reduction potential by +1.66 V.

(b) Changing the zero on the scale of reduction potentials has no effect on the calculated cell potentials.

2. (a) On the Celsius temperature scale the difference between the temperature at which water boils and the temperature at which water freezes is
100.0 °C – 0.0 °C = 100 °C.

(b) On the Kelvin temperature scale the difference between the temperature at which water boils and the temperature at which water freezes is
373.15 K – 273.15 K = 100 K.

3. The Celsius and Kelvin temperature scales have two different zero values, but a change of 1 degree is the same on both scales. Zero on the Celsius scale is 273.15 K on the Kelvin scale. Zero on the Kelvin scale is –273.15 °C. The zero value of the Celsius scale is defined based on the freezing point of water under standard conditions. For the Kelvin scale, the zero is shifted to occur at absolute zero – the temperature at which particles of matter have zero energy under standard conditions.

4. The Celsius scale and reduction potentials have a zero value that is arbitrarily set.

5. Mass is a measure of the quantity of matter in an object. The zero value means that there is no matter present at all. This reference point cannot change, so the zero value of a mass scale is not arbitrary.

Assessment Options

- Collect and assess student tables and answers to Procedure and Analysis questions.

Investigation 13.A: Measuring Cell Potential of Voltaic Cells

Student Textbook pages 488–489

Purpose

The purpose of this investigation is for students to measure the cell potentials of a variety of voltaic cells that they have assembled. Students will use this data to assemble a table of reduction potentials from which they can draw conclusions about periodic trends in reduction.

Outcomes

- 30-B1.3s
- 30-B2.2s
- 30-B2.1s
- 30-B2.3s

Advance Preparation

When to Begin	What to Do
1–2 weeks before	- Order materials.
1–2 days before	- Prepare solutions and graphite. See **BLM 5.4.4** for tips on preparing solutions. - Cut metals into strips. - Photocopy **BLM 13.1.4 Investigation 13.A.**

Materials

- 1 Styrofoam™ or clear plastic egg carton with 12 wells
- 5 mL of 0.1 mol/L solution of each of the following:
  - $Mg(NO_3)_2(aq)$
  - $Cu(NO_3)_2(aq)$
  - $Al(NO_3)_3(aq)$
  - $Ni(NO_3)_2(aq)$
  - $Zn(NO_3)_2(aq)$
  - $SnSO_4(aq)$
  - $Fe(NO_3)_3(aq)$
  - $AgNO_3(aq)$
  - $HNO_3(aq)$
- 5 mL of saturated NaCl(aq) solution
- 5 cm strip of Mg ribbon
- 1 cm × 5 cm strip of each of the following: Cu(s), Al(s), Ni(s), Zn(s), Sn(s), Fe(s), and Ag(s)
- 15 mL of 1.0 mol/L KNO_3
- 5 cm of thick graphite pencil lead or a graphite rod
- sandpaper
- 25 cm clear aquarium rubber tubing (Tygon®; internal diameter: 4–6 mm)
- cotton batting
- disposable pipette
- black and red electrical leads with alligator clips
- voltmeter set to a scale of 0 V to 20 V
- paper towel
- marker

Time Required

- 120 – 180 minutes

Helpful Tips

- Consider using a well plate instead of an egg carton. Add solution to each well so that they are 1/4 full. (The reference cell should be in the centre of the other cells.)
- Create a simple salt bridge by cutting a thin strip of paper towel (10 cm × 0.3 cm). Soak the strip in the salt solution. Place one end in each cell being tested.
- Consider preparing a template for the 9 × 9 grid. This will help ensure consistency amongst students and reduce the time required for completing the experiment.
- Since graphite tends to be somewhat brittle, pencil lead may not last very long as an electrode. If possible, use graphite electrodes. If you do use pencil lead, test it to be sure it conducts. HB lead usually works fine.
- In addition to wiping the ends of the salt bridge to remove excess salt solution, you may want to instruct students to rinse the ends with a little distilled water. You may also want to change the cotton plugs to ensure there is no contamination when transferring the salt bridge from cell to cell.
- Consider providing students with the following bonus question: If you built a voltaic cell with inert electrodes, would you predict that replacing them with inert electrodes made from another material would affect the cell potential? Explain your answer.
  Answer: If you built the galvanic cell from inert electrodes, it would change the substance that functioned as the anode. This would change the cell potential, and in many cases cause the cell to stop functioning.
- Use **BLM 13.1.4 (HAND) Investigation 13.A: Measuring Cell Potential of Voltaic Cells** to support this activity. Remove sections as appropriate to meet the needs of students in your class.
- ***Expected Results:*** Because the investigation calls for 0.10 mol/L solutions instead of the standard 1.00 mol/L concentrations, the students' results will not be the same as the values that they calculate using the standard reductions potentials in the table in Appendix G. As well, the temperature and pressure are not maintained at exactly SATP conditions. Nevertheless, the measured values should be relatively close to the expected values.

Safety Precautions

- Provide proper containers for chemical disposal so that students do not pour the solutions down the drain or throw the electrodes into the garbage.
- 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.
- Remind students to handle the nitric acid solution with care. In addition, remind them to immediately wash off

any solution that they inadvertently spill on their skin with copious amounts of water.

- Remind students to wash their hands when they have completed the investigation.

Answers to Prediction Question

Students will likely use the Alberta Chemistry Data Booklet to calculate cell potentials as follows:

Mg(s) anode, Cu(s) cathode; $E^o_{\text{cell}} = 2.71$ V
Mg(s) anode, Al(s) cathode; $E^o_{\text{cell}} = 0.71$ V
Al(s) anode, Cu(s) cathode; $E^o_{\text{cell}} = 2.00$ V
Mg(s) anode, Ni(s) cathode ; $E^o_{\text{cell}} = 2.11$ V
Ni(s) anode, Cu(s) cathode ; $E^o_{\text{cell}} = 0.60$
Al(s) anode, Ni(s) cathode; $E^o_{\text{cell}} = 1.40$ V
Mg(s) anode, Zn(s) cathode; $E^o_{\text{cell}} = 1.61$ V
Zn(s) anode, Cu(s) cathode; $E^o\text{cell} = 0.416$ V
Al(s) anode, Zn(s) cathode; $E^o_{\text{cell}} = 0.90$ V
Zn(s) anode, Sn(s) cathode; $E^o_{\text{cell}} = 0.62$ V
Mg(s) anode, Sn(s) cathode; $E^o_{\text{cell}} = 2.23$ V
Sn(s) anode, Cu(s) cathode; $E^o_{\text{cell}} = 0.48$ V
Al(s) anode, Sn(s) cathode; $E^o_{\text{cell}} = 1.52$ V
Ni(s) anode, Sn(s) cathode; $E^o_{\text{cell}} = 0.12$ V
Mg(s) anode, Fe(s) cathode; $E^o_{\text{cell}} = 1.92$ V
Fe(s) anode, Cu(s) cathode; $E^o_{\text{cell}} = 0.79$ V
Al(s) anode, Fe(s) cathode; $E^o_{\text{cell}} = 1.21$ V
Ni(s) anode, Fe(s) cathode ; $E^o_{\text{cell}} = -0.19$ V
Zn(s) anode, Fe(s) cathode; $E^o_{\text{cell}} = 0.31$ V
Sn(s) anode, Fe(s) cathode; $E^o_{\text{cell}} = -0.31$ V
Mg(s) anode, Ag(s) cathode; $E^o_{\text{cell}} = 3.17$ V
Cu(s) anode, Ag(s) cathode; $E^o_{\text{cell}} = 0.46$ V
Al(s) anode, Ag(s) cathode; $E^o_{\text{cell}} = 2.46$ V
Ni(s) anode, Ag(s) cathode; $E^o_{\text{cell}} = 1.06$ V
Zn(s) anode, Ag(s) cathode; $E^o_{\text{cell}} = 1.56$ V
Sn(s) anode, Ag(s) cathode; $E^o_{\text{cell}} = 0.94$ V
Fe(s) anode, Ag(s) cathode; $E^o_{\text{cell}} = 1.25$ V
Mg(s) anode, H^+(aq) cathode; $E^o_{\text{cell}} = 2.37$ V
H_2(g) anode, Cu(s) cathode; $E^o_{\text{cell}} = 0.34$ V
Al anode, H^+(aq) cathode; $E^o_{\text{cell}} = 1.66$ V
Ni anode, H^+(aq) cathode; $E^o_{\text{cell}} = 0.26$ V
Zn anode, H+(aq) cathode; $E^o_{\text{cell}} = 0.76$ V
Sn anode, H^+(aq) cathode; $E^o_{\text{cell}} = 0.14$ V
Fe anode, H^+(aq) cathode; $E^o_{\text{cell}} = 0.45$ V
H_2(g) anode, Ag(s) cathode; $E^o_{\text{cell}} = 0.80$ V

Student values will likely be less than predicted, which they will likely blame on faulty equipment or resistance in the wire. While these may be valid points, the other issues to point out are the conditions listed in the Data Booklet under which the standard values are calculated.

Answers to Analysis Questions

1.

	Mg/ Mg^{2+}	**Cu/ Cu^{2+}**	**Al/ Al^{3+}**	**Ni/ Ni^{2+}**	**Zn/ Zn^{2+}**	**Sn/ Sn^{2+}**	**Fe/ Fe^{3+}**	**Ag/ Ag^+**	**H^+/ H_2**
Mg/ Mg^{2+}									
Cu/ Cu^{2+}	(1)								
Al/ Al^{3+}	(2)	(3)							
Ni/ Ni^{2+}	(4)	(5)	(6)						
Zn/ Zn^{2+}	(7)	(8)	(9)	(10)					
Sn/ Sn^{2+}	(11)	(12)	(13)	(14)	(15)				
Fe/ Fe^{3+}	(16)	(17)	(18)	(19)	(20)	(21)			
Ag/ Ag^+	(22)	(23)	(24)	(25)	(26)	(27)	(28)		
H^+/ H_2	(29)	(30)	(31)	(32)	(33)	(34)	(35)	(36)	

(a) (1) Mg(s) anode, Cu(s) cathode; (2) Mg(s) anode, Al(s) cathode; (3) Al(s) anode, Cu(s) cathode; (4) Mg(s) anode, Ni(s) cathode; (5) Ni(s) anode, Cu(s) cathode; (6) Al(s) anode, Ni(s) cathode; (7) Mg(s) anode, Zn(s) cathode; (8) Zn(s) anode, Cu(s) cathode; (9) Al(s) anode, Zn(s) cathode; (10) Zn(s) anode, Sn(s) cathode; (11) Mg(s) anode, Sn(s) cathode; (12) Sn(s) anode, Cu(s) cathode; (13) Al(s) anode, Sn(s) cathode; (14) Ni(s) anode, Sn(s) cathode; (15) Zn(s) anode, Sn(s) cathode; (16) Mg(s) anode, Fe(s) cathode; (17) Fe(s) anode, Cu(s) cathode; (18) Al(s) anode, Fe(s) cathode; (19) Ni(s) anode, Fe(s) cathode; (20) Zn(s) anode, Fe(s) cathode; (21) Sn(s) anode, Fe(s) cathode; (22) Mg(s) anode, Ag(s) cathode; (23) Cu(s) anode, Ag(s) cathode; (24) Al(s) anode, Ag(s) cathode; (25) Ni(s) anode, Ag(s) cathode; (26) Zn(s) anode, Ag(s) cathode; (27) Sn(s) anode, Ag(s) cathode; (28) Fe(s) anode, Ag(s) cathode; (29) Mg(s) anode, H^+(aq) cathode; (30) H_2(g) anode, Cu(s) cathode; (31) Al(s) anode, H^+(aq) cathode; (32) Ni(s) anode, H^+(aq) cathode; (33) Zn(s) anode, H^+(aq) cathode; (34) Sn(s) anode, H^+(aq) cathode; (35) Fe(s) anode, H^+(aq) cathode; (36) H_2(g) anode, Ag(s) cathode

(b) For each of the cells in this investigation, all anodes are negative electrodes and all cathodes are positive electrodes.

2. An oxidation half-reaction occurs at each anode and a reduction half-reaction occurs at each cathode. For example, for cell (1), the Mg(s)/Mg^{2+}(aq) half-cell is the anode and the Cu(s)/Cu^{2+}(aq) half-cell is the cathode.

Oxidation occurs at the anode. The oxidation half-reaction is:

$Mg(s) \rightarrow Mg^{2+}(aq) + 2e^-$

Reduction occurs at the cathode. The reduction half-reaction is:

$Cu^{2+}(aq) + 2e^- \rightarrow Cu(s)$

The overall cell reaction is:

$Mg(s) + Cu^{2+}(aq) \rightarrow Mg^{2+}(aq) + Cu(s)$

3. **(a)** The direction of electron flow for a voltaic cell is always from the anode to the cathode. For example, in cell (1), the electrons flow from the magnesium anode to the copper cathode.

(b) Positive ions (cations) move towards the cathode where reduction occurs and negatively charged ions (anions) move towards the anode where oxidation occurs. For example, in cell (1), potassium ions and copper ions move towards the cathode (copper). Nitrate ions move towards the anode (magnesium).

4. Magnesium is the most effective reducing agent since the magnesium half-cell always acted as the anode when coupled with other half-cells.

5. The reduction half-reactions, in order of least to most effective reducing agents, are as follows:

$Ag^+(aq) + e^- \rightarrow Ag(s)$
$Cu^{2+}(aq) + 2e^- \rightarrow Cu(s)$
$2H^+(aq) + 2e^- \rightarrow H_2(g)$
$Fe^{3+}(aq) + 3e^- \rightarrow Fe(s)$
$Sn^{2+}(aq) + 2e^- \rightarrow Sn(s)$
$Ni^{2+}(aq) + 2e^- \rightarrow Ni(s)$
$Zn^{2+}(aq) + 2e^- \rightarrow Zn(s)$
$Al^{3+}(aq) + 3e^- \rightarrow Al(s)$
$Mg^{2+}(aq) + 2e^- \rightarrow Mg(s)$

6. The metal ions that are the best oxidizing agents are at the top of the list. The least effective reducing agents tend to be the most effective oxidizing agents.

7. This could vary amongst students. However, some might make the link between the relative strength of oxidizing agents discussed in Chapter 12 and the higher reductive potentials.

8. Typically, most measured cell potentials ought to be lower than predicted values. In part this is due to the lower concentration of the solutions, different operating conditions, and resistance in the wiring leading to the voltmeter.

9. Two factors that affect the cell potential of the voltaic cell are the concentration of the electrolyte solution and the identity of the substances used in each half-cell.

10. **(a)** $AgNO_3(aq) + NaCl(aq) \rightarrow AgCl(s) + NaNO_3(aq)$

Silver nitrate and sodium chloride are both soluble in water. A double displacement reaction takes place when their solutions are mixed, and insoluble silver chloride precipitates.

(b) and **(c)** When sodium chloride was added to the $Ag(s)/Ag^+(aq)$ well, students should have observed a change (lowering) in the overall cell voltage. Since the reaction of silver nitrate with silver chloride effectively removes silver ions from solution, it reduces the concentration of the silver ions in solution at the silver electrode. Students should conclude, then, that the concentration of the electrolyte solution does affect the cell potential of a voltaic cell.

Answers to Conclusion Questions

11. Students may suggest that temperature or the electrical resistance of the external circuit would affect the voltage of a voltaic cell. Students could suggest a procedure in which they compare the voltage of a voltaic cell at different temperatures. In addition, using leads with different resistances could be used to test if electrical resistance affects cell voltage.

12. The term "standard cell potential" refers to the cell potential using solutions of 1.0 mol/L at SATP conditions, while the term "cell potential" is used for a cell potential measured at any other set of conditions.

Answer to Application Question

13. Students may choose to investigate such factors as the type or length of the wire.

Assessment Options

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

Section 13.1 Review Answers

Student Textbook page 490

1. It is important to keep each half-cell separate so that an instantaneous reaction does not occur, "short-circuiting" the voltmeter.

2. The standard half-cell potentials are measured using the hydrogen half-cell as a reference. In the equation $E^o_{cell} = E^o_{cathode} - E^o_{anode}$, the hydrogen half-cell is given a value of zero. Any measured voltage would therefore belong to the non-hydrogen half of the cell.

3. **(a)** $E^o_{cell} = E^o_{cathode} - E^o_{anode}$
$= -1.66\text{ V} - (-2.37\text{ V}) = +0.71\text{V}$

(b) $E^o_{cell} = E^o_{cathode} - E^o_{anode}$
$= +2.87\text{ V} - (-2.93\text{ V}) = +5.80\text{ V}$

(c) $E^o_{cell} = E^o_{cathode} - E^o_{anode}$
$= +1.23\ V - (0.80\ V) = +0.43\ V$

4. (a) $E^o_{cell} = E^o_{cathode} - E^o_{anode}$
$= +0.34\ V - (-0.26\ V) = +0.60V$

(b) $E^o_{cell} = E^o_{cathode} - E^o_{anode}$
$= +1.50\ V - (-0.40\ V) = +1.90\ V$

(c) $E^o_{cell} = E^o_{cathode} - E^o_{anode}$
$= -0.04\ V - (-0.46\ V) = +0.50\ V$

5. No, a student using the Alberta Chemistry Data Booklet could not build a voltaic cell with a standard cell potential of 7.0 V. The strongest oxidizing agent $F_2(g)$ and the strongest reducing agent Li(s) would only yield 5.91 V. By building a gold-aluminium cell in series with two cells, a total of 7.32 V could be achieved.

6. The cell potential describes the potential difference between two electrodes of a cell, or the amount of energy on a charge as it moves between two electrodes. The cell potential is dependent on both the anode and the cathode used. The standard reduction potential is a measure of the amount of energy for only the reduction half of cell. Since reduction cannot happen without oxidation, reduction potentials are measured against a standard reference, the hydrogen half-cell, which is set at a reduction potential of 0.00V.

7. $E^o_{cell} = E^o_{cathode} - E^o_{anode}$
$= \text{“X” } V - (-0.76\ V) = +1.75\ V$
$X = 0.99\ V$

13.2 Applications of Voltaic Cells

Section Outcomes

Students will:

- analyze the relationship between scientific knowledge and the technological development of batteries and fuel cells
- investigate the use of technology to solve problems related to corrosion
- predict the future of fuel cells in transportation

Key Terms

dry cell
battery
primary battery
secondary battery
alkaline batteries
button battery
fuel cell
corrosion
galvanizing
sacrificial anode
cathodic protection

Chemistry Background

- There are a variety of types of batteries that differ both in their applications and their environmental impacts. The most common battery, the dry cell, offers size and portability but generates toxic waste products. The fuel cell, on the other hand, is a current example of a “green” technology that offers environmentally friendlier waste products, at a greater economic cost.
- The corrosion of metals involves a redox reaction, typically between the metal, oxygen, and air. A proper understanding of redox spontaneity rules can allow for better protection against corrosion.

Teaching Strategies

- Have students do research on new types of batteries and fuel cells that are safer for the environment. They can present their findings in the form of a report, poster, or brochure.
- Have students identify household items that commonly rust and useful ways to prevent these objects from rusting.
- Invite a tradesperson or engineer to speak to your class about the impact of corrosion in their field.
- **BLM 13.2.1 (OH) Fuel Cell Technology** has been prepared for this section. You will find it with the Chapter 13 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 more creative or artistically inclined students could involve the creation of an advertising campaign for a new type of battery. Before beginning work on their batteries, have students, as a class, devise a set of criteria that the new battery would need to meet in order to be a success. Students can then make presentations that are assessed based on these criteria by other class members. Alternatively, instead of using new battery designs, students could use this activity with existing battery types.

Answers to Questions for Comprehension

Student Textbook page 493

Q8. The dry cell is similar to the Daniel cell in that both use a zinc anode. The dry cell differs from a Daniell cell in that the electrolyte has been thickened into a paste.

Q9. A battery consists of several cells connected in series.

Student Textbook page 496

Q10. The cell potential for the hydrogen fuel cell is:
$E^o_{cathode} - E^o_{anode} = E^o_{cell}$
$(1.23\ V) - (0.00) = 1.23\ V$

Q11. Typically a combustion reaction requires the presence of a flame, however, no flames are present in these reactions.

Q12. Currently, the source of hydrogen for the fuel cell comes from the conversion of ethane to ethene in the petrochemical industry. The cost of this process is one of the leading criticisms against the common use of hydrogen fuel cells.

Student Textbook page 500

Q13. (a) For a corrosion reaction:

$E^o_{\text{cathode}} - E^o_{\text{anode}} = E^o_{\text{cell}}$

$(0.40 \text{ V}) - (-0.45) = 0.85 \text{ V}$

(b) Students should be aware that the standard cell potentials are based on 1.0 mol/L concentrations and SATP conditions. If the concentrations or conditions of their rust "scenarios" differ, then they will get different (typically lower) cell potentials.

Q14. Aluminium provides cathodic protection on an iron object because it is a better reducing agent than iron. This means that the aluminium reacts first, saving the iron from rusting.

Q15. The corrosion problems most likely showed up in the Atlantic provinces due to high volumes of precipitation throughout the year (both rain and snow), coupled with the use of salt on winter roads. Taken together, these factors encourage corrosion.

Q16. (a) Given the cell potentials of the following half-reactions:

$O_2(g) + 4H^+(aq) = E^o_{\text{cell}} = 1.23 \text{ V}$

$O_2(g) + 2H^+(aq) = E^o_{\text{cell}} = 0.70 \text{ V}$

$O_2(g) + 2H_2O(aq) = E^o_{\text{cell}} = 0.40 \text{ V}$

Elemental oxygen is a much better oxidizing agent under acidic conditions, providing an overall reaction of:

$2Fe(s) \rightarrow 2Fe^{2+}(aq) + 4e^-$

$O_2(g) + 4H^+(aq) + 4e^- \rightarrow 2H_2O(\ell)$

$E^o_{\text{cathode}} - E^o_{\text{anode}} = E^o_{\text{cell}}$

$(1.23 \text{ V}) - (-0.45) = 1.68 \text{ V}$

(b) Acid rain greatly increases oxygen's ability to act as an oxidizing agent and contributes to rusting.

Chemistry File: Web Link

Student Textbook page 500

The combinatorial materials synthesis functions by depositing a film of new materials only millimeters in area. This allows for many combinations to be spread over a small area and chemical analysis to be done quickly.

Section 13.2 Review Answers

Student Textbook page 501

1. The oxidizing agent is manganese dioxide. The reducing agent is zinc.
2. The top of a commercial 1.5-V dry cell is marked with a plus sign because the graphite cathode protrudes there.
3. **(a)** oxidation:

   $Zn(s) + 2OH^-(aq) \rightarrow ZnO(s) + H_2O(\ell) + 2e^-$

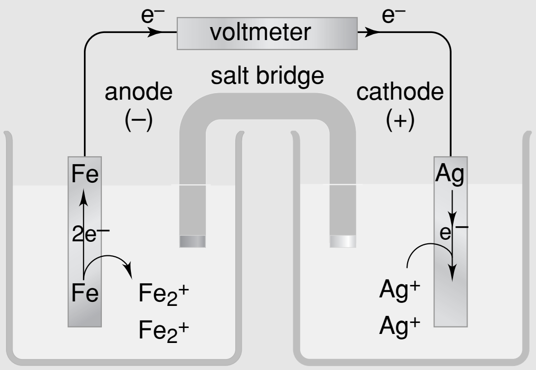
   reduction:

   $Ag_2O(s) + H_2O(\ell) + 2e^- \rightarrow 2Ag(s) + 2OH^-(aq)$

   overall cell reaction:

   $Zn(s) + Ag_2O(s) \rightarrow ZnO(s) + Ag(s)$

   (b) The anode is made of zinc, while the cathode is made of stainless steel.
4. In a flashlight, the two 1.5-V D batteries are connected in series and the voltages are additive; therefore, the total voltage is 3.0 V.
5. A 6-V dry cell battery contains four 1.5-V dry cells connected in series. When dry cells are connected in series, the voltages are additive, and 4 × 1.5 V = 6 V.
6. The potential of a cell depends only on the substances used for the anode and the cathode, not on the amounts of the substances used. A cell with a gold anode and an iron cathode would produce a cell potential of 1.95 V. If three of these cells were combined in series, a potential of almost 6 V would be produced. If only small quantities of these substances were used, the battery would be small. A large cell that produces only 1.5 V of potential contains larger quantities of the reactants required for the redox reaction. A large cell would be able to provide an electrical current over a longer period of time.
7. The zinc container of the dry cell is involved with the oxidation half-reaction. As zinc is oxidized, it forms ions and dissolves into the electrolyte paste. Therefore, the container has less and less zinc as the reaction progresses.
8. Assume that the cell represented by the shorthand notation is a typical galvanic cell.



Oxidation: $Fe(s) \rightarrow Fe^{2+}(aq) + 2e^-$

Reduction: $Ag^+(aq) + e^- \rightarrow Ag(s)$

Overall cell reaction:

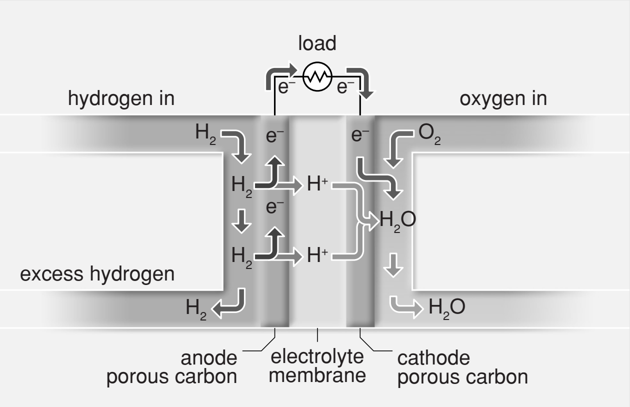
$Fe(s) + 2Ag^+(aq) \rightarrow Fe^{2+}(aq) + 2Ag(s)$

9. Both fuel cells and dry cells are alike in that they dispense with the need for an aqueous solution. In a dry cell, the

electrolytes are in a paste, while fuel cells tend to rely on solid oxides, though the hydrogen fuel cell uses oxygen and hydrogen gases. One of the chief differences of fuel cells and dry cells, however, is that the main waste of fuel cells can be recycled, such as water in the case of the hydrogen cell, whereas the waste of many depleted batteries involve toxic heavy metal cations.

10. The main benefit of fuel cells for space use is the possibility of recycling waste to re-start, reducing the need to re-supply with fuel or the store and remove fuel waste. Second, many fuel cells operate at higher efficiencies, allowing for more energy to be produced overall. Finally, the production of water helps to reduce risks of fuel combustion.

11.



The anode reaction is: $H_2(g) \rightarrow 2H^+(aq) + 2e^-$

The cathode reaction is: $O_2(g) + 4H^+(aq) + 4e^- \rightarrow 2H_2O(\ell)$

The electricity is generated through the passing of electrons from the oxygen to the hydrogen in the presence of the platinum-coated carbon catalyst.

12. The use of salt lowers the melting point of snow, causing more water to be mobile and reactive. Furthermore, the salt also provides electrolytes required for the salt bridge.

13. When the steel cans corrode, they produce aqueous ions. However, the aluminum cans produce solid aluminium oxide, which slows down the corrosion process.

14. Zinc is a more reactive metal than iron, as noted by its cell potential, so it will react before iron does.

15. **(a)** Two metals that do not react in the presence of oxygen and water are silver and gold, as noted by their cell potentials being higher than that of the oxygen and water half-reaction.

(b) Since these metals are quite resistant to corrosion they are useful for industrial processes that take place under extremely adverse conditions that promote corrosion, such as high temperatures or acidic environments, like the aerospace or petroleum industry.

16. Silver sulfide forms when silver is exposed to small quantities of hydrogen sulfide found in the air.

$$4Ag(s) + 2H_2S(g) + O_2(g) \rightarrow 2Ag_2S(s) + 2H_2O(g)$$

The tarnish may be removed by an abrasive polish or, more easily, by chemically reducing the silver ions to silver atoms. To do this, place the tarnished piece of silver in a conducting solution, such as baking soda, in an aluminium pan. When the piece of silver comes in contact with the aluminium, the silver ions in the tarnish are reduced to silver.

$$3Ag_2S(s) + 2Al(s) \rightarrow 6Ag(s) + Al_2S_3(s)$$

The oxidation half-reaction is:

$$Al(s) \rightarrow Al^{3+}(aq) + 3e^-$$

The reduction half-reaction is:

$$Ag^+(aq) + e^- \rightarrow Ag(s)$$

13.3 Electrolytic Cells

Student Textbook pages 502–513

Section Outcomes

Students will:

- define electrolytic cell
- identify similarities and differences between the operation of voltaic cell and that of an electrolytic cell
- recognize that predicted reactions do not always occur
- describe applications of electrolytic cells for rechargeable batteries
- construct and observe an electrolytic cell

Key Terms

electrolyte
electrolysis
overpotential
chlor-alkali process

Chemistry Background

- Electrolytic cells essentially operate in the reverse direction compared with voltaic cells. An external source of current induces a non-spontaneous redox reaction. In an electrolytic cell, oxidation occurs at the anode, which is positively charged, and reduction occurs at the cathode, which is negatively charged.
- Predicting which reaction will occur when an aqueous solution is electrolyzed involves considering the reduction and oxidation of water. The student textbook has students use nonstandard reduction potentials for water when predicting which reaction will occur in aqueous solutions. Students know that the concentration of hydrogen ions and hydroxide ions in water is not 1 mol/L (standard conditions), but 1.0×10^{-7} mol/L. However, using nonstandard reduction potentials is not optimum. For example, many solutions of ionic compounds are not neutral, and the $H^+(aq)$ and $OH^-(aq)$ ion concentrations do not equal 1.0×10^{-7} mol/L. Nonetheless, using nonstandard reduction potentials for water will usually

result in more accurate predictions than using standard reduction potentials.
- Many industrial processes, such as the production of chlorine, sodium hydroxide, and aluminium, involve some form of electrolysis.
- Rechargeable batteries, also called secondary batteries, operate on the basis of electrolysis. Placing the batteries into the electric charge completes the electrolytic cell, and the external electricity provides the voltage needed for the reverse reaction.

Teaching Strategies

- In this section, students will learn how to use standard reduction potentials to predict the products of the electrolysis of aqueous solutions, and to predict the spontaneity of chemical reactions. Students should realize that although predictions made using standard reduction potentials are often correct, they may be unreliable when potentials are similar or when gases are produced. One example is aluminium and water. Using the prediction method laid out in the text, students will predict that aluminium displaces hydrogen from water. In practice it does not. Practice Problems 7 and 8 on page 508 of the student textbook have students attempt to deal with this discrepancy. Consider using these questions in class to illustrate that predictions based on simplified models are not always correct.
- Work through the Sample Problem on page 507 of the student textbook on the chalkboard. When you work through problems predicting the electrolysis of an aqueous solution, be careful to label the nonstandard reduction potentials for water correctly (E instead of E°). Also, the cell voltage must be labelled E_{cell} instead of E°_{cell} if nonstandard potentials are used to determine it.
- Students will learn the term overvoltage in this chapter, but it is not explained in depth. Overvoltage is a complex kinetic phenomenon related to polarization and irreversibility. All that students need to know is that predicted products of electrolysis are not always observed because of the overvoltage required in many cases, especially when gases are produced. Overvoltage has applications in discussing rechargeable car batteries.
- The following overhead masters have been prepared for this section. You will find them with the Chapter 13 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
  13.3.1 (OH) Electrolytic Cells
  13.3.2 (OH) The Chlor-Alkali Cell

SUPPORTING DIVERSE STUDENT NEEDS

Have advanced students research the connection between cell potentials and thermodynamics by investigating the Nernst equation for equilibrium conditions:

$$E^\circ_{cell} = \frac{0.0592\text{ V}}{n} \times \log K_c$$

Figure 13.24

Student Textbook page 504

According to Avogadro's Law, when a gas has twice the volume of another at the same temperature and pressure, it means that twice as many moles are present. If two moles of water decompose into hydrogen and oxygen, they will give two moles of hydrogen and one mole of oxygen.

Chemistry File: Web Link

Student Textbook page 506

Dioxins are chlorine-containing organic compounds that are created through the combustion of hydrocarbons. They have been linked to cancer and reproductive defects in humans. One of the main ways to reduce exposure to dioxins has been prohibiting the burning of garbage, particularly natural waste like fruits, vegetables, and leaves.

Solutions to Practice Problems 5–8

Student Textbook page 508

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

5. **(a)** +0.15 V
  (b) Anode: Cd(s); Cathode: Ni(s)
  (c) Oxidation: $Cd(s) \rightarrow Cd^{2+}(aq) + 2e^-$
  Reduction: $Ni^{2+}(aq) + 2e^- \rightarrow Ni(s)$
  Cell reaction: $Cd(s) + Ni^{2+}(aq) \rightarrow Ni(s) + Cd^{2+}(aq)$
6. **(a)** –0.15 V
  (b) Anode: Ni(s) ; Cathode: Cd(s)
  (c) Oxidation: $Ni(s) \rightarrow Ni^{2+}(aq) + 2e^-$
  Reduction: $Cd^{2+}(aq) + 2e^- \rightarrow Cd(s)$
  Cell reaction: $Ni(s) + Cd^{2+}(aq) \rightarrow Ni^{2+}(aq) + Cd(s)$
7. Co(s) and $Cl_2(g)$
8. Ag(s) and $O_2(g)$

Investigation 13.B: Electrolysis of Aqueous Potassium Iodide

Student Textbook pages 508–509

Purpose

The purpose of this investigation is to provide students with the opportunity to build an electrolytic cell and predict the products of electrolysis. This investigation will also give students the background to evaluate electrolytic cell design and suggest alternatives.

Outcomes

- 30-B2.3s
- 30-B2.4s

Advance Preparation

When to Begin	What to Do
1–2 weeks before	- Order materials.
1–2 days before	- Prepare solutions and materials. - Photocopy **BLM 13.3.3: Investigation 13.B.**

Materials

- 10 mL 1 mol/L KI
- 20 gauge platinum wire (2 cm)
- 1 graphite pencil lead, 2 cm long
- 1 drop 1% starch solution
- 1 drop 1% phenolphthalein
- sheet of white paper
- 1 beaker (600 mL or 400 mL)
- 1 elastic band
- 25 cm clear aquarium rubber tubing (Tygon®; internal diameter: 4–6 mm)
- 3 disposable pipettes
- 3 toothpicks
- 2 wire leads (black and red) with alligator clips
- 9-V battery or variable power source set to 9 V

Time Required

- 120 minutes

Helpful Tips

- Before beginning the investigation, ensure students have a good grasp of the material on pages 503 and 507 of the student textbook, electrolysis in aqueous solutions. Students will predict the products of the electrolysis based on the discussion and examples shown on these pages.
- Consider having students redraw the setup diagram before beginning the investigation. Have students label the battery terminals. The black lead, connected to the copper wire electrode (the cathode), should be attached to the negative terminal of the battery. The red lead, connected to the carbon electrode (the anode), should be connected to the positive terminal of the battery.
- If possible, use a commercial carbon electrode rather than pencil lead. Pencil lead tends to be thin and breaks more easily than the purchased carbon electrodes.
- Caution students that gases dissolved in solution sometimes will come out of solution and appear on the electrodes as the solution warms up. This should not be confused with the production of gases at the electrodes through redox reactions. Indicators, such as starch and phenolphthalein, help suggest the identity of the products of electrolysis which appear at the electrodes. In this case, students will see hydrogen bubbles appearing at the cathode.
- 20 gauge wire is suggested, but 14 gauge wire will also work.
- Use **BLM 13.3.3 (HAND) Investigation 13.B: Electrolysis of Aqueous Potassium Iodide** to support this activity. Remove sections as needed to meet the needs of students in your class.
- ***Expected Results:*** Bubbles should form on the wire lead that is attached to the negative end of the battery. When students drop the starch solution in beside the anode, the graphite electrode connected to the positive terminal of the battery, the solution should turn purple which is the test for iodine. When dropped into the solution beside the cathode, the phenolphthalein should become colourless, indicating that the concentration of hydrogen ions is high.

Safety Precautions

- 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.
- Make sure the lab benches are dry before carrying out this investigation.

Answer to Prediction Question

The $K^+(aq)$ and $I^-(aq)$ concentrations are 1 mol/L, so the standard reduction potentials for the half-reactions that involve these ions should be used. The nonstandard values for water should be used.

$$I_2(s) + 2e^- \rightarrow 2I^-(aq) \quad E^o = 0.54\ V$$

$$O_2(g) + 4H^+(aq) + 4e^- \rightarrow 2H_2O(\ell) \quad E^o = 0.81\ V$$

$$2H_2O(\ell) + 2e^- \rightarrow H_2(g) + 2OH^-(aq) \quad E^o = -0.41\ V$$

$$K^+(aq) + e^- \rightarrow K(s) \quad E^o = -3.04\ V$$

The four possible overall reactions are:

$$2K^+(aq) + 2I^-(aq) \rightarrow 2K(s) + 2I_2(s); \ E^o_{cell} = -3.58\ V$$

$$2H_2O(\ell) \rightarrow 2H_2(g) + O_2(g); \ E\,cell = -1.23\ V$$

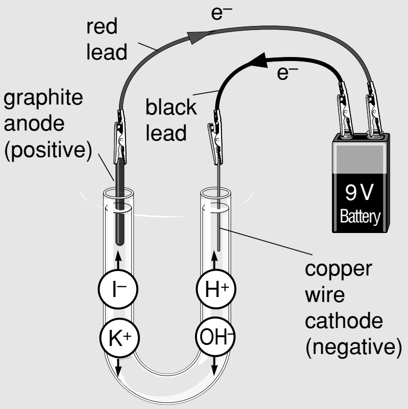
$4K^+(aq) + 2H_2O(\ell) \rightarrow$
$4K(s) + O_2(g) + 4H^+(aq)$; E^o cell = –3.85 V

$2H_2O(\ell) + 2I^-(aq) \rightarrow$
$H_2(g) + 2OH^-(aq) + I_2(s)$; E^o cell= –0.95 V

The cell reaction that produces hydrogen gas, hydroxide ions, and iodine requires the lowest external voltage. Therefore, students should predict the formation of these products. The hydrogen gas and hydroxide ions should form at the cathode, and the iodine should form at the anode.

Answers to Analysis Questions

1.



2. Adding the starch solution to the electrolyte at the anode results in a purple solution, indicating the presence of molecular iodine. At the cathode, the clear, colourless phenolphthalein solution will turn pink, indicating that the solution is basic and suggesting the presence of hydroxide ions. (Hydrogen gas is also formed at the cathode, but students do not test for the gas.)

3. $2I^-(aq) \rightarrow I_2(s) + 2e^-$

4. $2H_2O(\ell) + 2e^- \rightarrow H_2(g) + 2OH^-(aq)$

5. $2I^-(aq) + 2H_2O(\ell) \rightarrow I_2(aq) + H_2(g) + 2OH^-(aq)$

6. $E^o_{cell} = E^o_{cathode} - E^o_{anode}$
$= -0.83\text{ V} - (+0.54\text{ V}) = -1.37\text{ V}$

The calculated external voltage needed is 1.37 V. The actual voltage needed to carry out this electrolytic process is greater than this. As students learned on page 505 of the student textbook, this excess is called overpotential. This is the reason that the external voltage used in this investigation was significantly higher than the calculated value.

Answer to Conclusion Question

7. The products from the electrolysis of 1 mol/L KI(aq) are iodine, hydrogen gas, and hydroxide ions. These are the products that students should have predicted using the method shown in the Sample Problem on page 507 of the student textbook.

Answers to Applications Questions

8. The observations would probably not change. Consider the four possible reactions. The electrolysis of aqueous sodium iodide requires greater external voltage than the electrolysis of water. The reaction producing sodium metal and hydrogen ions also requires greater external voltage than the electrolysis of water. However, the reaction producing iodine, hydroxyl ions, and hydrogen gas requires less external voltage than the electrolysis of water. These are the same products that were observed in the investigation.

9. The procedure would need to be changed. The investigation would need to be carried out at a temperature higher than the melting point of potassium iodide. Also, inert electrodes would need to be used. Refer to the discussion of the electrolysis of molten salts on page 509 of the student textbook.

Assessment Options

- Collect and assess students' Predictions and Observations, as well as answers to Analysis, Conclusion, and Applications questions.
- Use Assessment Checklist 2: Laboratory Report from Appendix A.

Answers to Questions for Comprehension

Student Textbook page 510

Q17. (a) $2Cl^-(\ell) \rightarrow Cl_2(g) + 2e^-$

(b) $Ca^{2+}(\ell) + 2e^- \rightarrow Ca(s)$

(c) $Ca^{2+}(\ell) + 2Cl^-(\ell) \rightarrow Cl_2(g) + Ca(s)$

Q18. (a) $Li^+(el) + e^- \rightarrow Li(s)$

(b) $2Br^-(el) \rightarrow Br_2(\ell) + 2e^-$

(c) $2Li^+(el) + 2Br^-(el) \rightarrow Br_2(\ell) + 2Li(s)$

Q19. If alternating current were to supply the power to an electrolytic cell, then power would be supplied to both anode and cathode, causing the reverse reactions to occur as the power alternated.

Q20. Electrons leave the battery at the negative terminal and the cathode needs to accept electrons so that they can enter the solution containing the cations.

Section 13.3 Review Answers

Student Textbook page 513

1. The four half-reactions involved are:

(1) $Br_2(\ell) + 2e^- \rightarrow 2Br^-(aq)$ $E^o = 1.07$ V

(2) $O_2(g) + 4H^+(aq) + 4e^- \rightarrow 2H_2O(\ell)$ $E^o = 0.81$ V (nonstandard reduction potential)

(3) $H_2O(\ell) + 2e^- \rightarrow H_2(g) + 2OH^-(aq)$ $E^o = -0.41$ V (nonstandard reduction potential)

(4) $Cu^+(aq) + e^- \rightarrow Cu(s)$ $E^o = 0.52$ V

The possible anode half-reactions are the reverse of (1) and (2). The possible cathode half-reactions are (3) and (4).

Using $E^\circ_{cell} = E^\circ_{cathode} - E^\circ_{anode}$ the most positive (least negative) cell potential is obtained using half-reactions (2) and (4):

$E^\circ_{cell} = 0.52\text{ V} - 0.81\text{ V} = -0.29\text{ V}$

The overall reaction is:

$4Cu^+(aq) + 2H_2O(\ell) \rightarrow 4Cu(s) + O_2(g) + 4H^+(aq)$

The predicted products are copper, oxygen, and hydrogen ions.

2. (a) Since the electrolysis takes place in aqueous solution the relevant reduction potentials are:

(1) $Zn^{2+}(aq) + 2e^- \rightarrow Zn(s)\ E^o = -0.76\text{ V}$

(2) $2H_2O(\ell) + 2e^- \rightarrow H_2(g) + 2OH^-(aq)\ E = -0.41\text{ V}$

(3) $O_2(g) + 4H^+(aq) + 4e^- \rightarrow 2H_2O(\ell)\ E = 0.81\text{ V}$

(4) $Cu^{2+}(aq) + 2e^- \rightarrow Cu(s)\ E^\circ = 0.34\text{ V}$

(1) and (2) are the possible half-reactions at the cathode. The reverse of (3) and (4) are the possible half-reactions at the anode. The lowest external voltage is required by the reaction including half-reaction (2) and the reverse of (4):

$2H_2O(\ell) + Cu(s) \rightarrow Cu^{2+}(aq) + H_2(g) + 2OH^-(aq)\ E_{cell} = -0.76\text{ V}$

The predicted products are hydrogen gas and copper (II) ions.

(b) Hydrogen gas requires overvoltage. Therefore, zinc deposits on the cathode.

3. (a) The half-reactions are:

Reduction (occurs at the cathode):

$Fe^{3+}(aq) + 3e^- \rightarrow Fe(s)$

Oxidation (occurs at the anode):

$2I^-(aq) \rightarrow I_2(s) + 2e^-$

The standard reduction potentials are:

$Fe^{3+}(aq) + 3e^- \rightarrow Fe(s)\ E^\circ = -0.04\text{ V}$

$I_2(s) + 2e^- \rightarrow 2I^-(aq)\ E^o = 0.54\text{ V}$

$E^o_{cell} = E^o_{cathode} - E^o_{anode}$

$= -0.04\text{ V} - (0.54\text{ V})$

$= -0.58\text{ V}$

The reaction has a negative standard cell potential; therefore, it is non-spontaneous under standard conditions.

(b) The half-reactions are:

Reduction (occurs at the cathode):

$Ag^+(aq) + e^- \rightarrow Ag(s)$

Oxidation (occurs at the anode):

$H_2SO_3(aq) + H_2O(\ell) \rightarrow SO_4^{2-}(aq) + 4H^+(aq) + 2e^-$

The standard reduction potentials are:

$SO_4^{2-}(aq) + 4H^+(aq) + 2e^- \rightarrow H_2SO_3(aq) + H_2O(\ell)\ E^o = +0.17\text{ V}$

$Ag^+(aq) + e^- \rightarrow Ag(s)\ E^o = +0.80\text{ V}$

$E^o_{cell} = E^o_{cathode} - E^o_{anode}$

$= 0.80\text{ V} - (0.17\text{ V}) = +0.63\text{ V}$

This reaction has a positive standard cell potential; therefore, it is spontaneous under standard conditions.

4. On discharging, the half-reactions are:

$Cd(s) + 2OH^-(aq) \rightarrow Cd(OH)_2(s) + 2e^-$

$NiO(OH)(s) + H_2O(\ell) + e^- \rightarrow Ni(OH)_2(s) + OH^-(aq)$

On recharging, therefore, the half-reactions are:

$Cd(OH)_2(s) + 2e^- \rightarrow Cd(s) + 2OH^-(aq)$

$Ni(OH)_2(s) + OH^-(aq) \rightarrow NiO(OH)(s) + H_2O(\ell) + e^-$

To balance charge, multiply the oxidation equation by 2:

$2Ni(OH)_2(s) + 2OH^-(aq) \rightarrow 2NiO(OH)(s) + 2H_2O(\ell) + 2e^-$

The overall reaction for recharging a nicad battery can then be written as:

$Cd(OH)_2(s) + 2Ni(OH)_2(s) \rightarrow Cd(s) + 2NiO(OH)(s) + 2H_2O(\ell)$

5. According to the caption for Figure 13.27 on page 511, a lead-acid car battery consists of six 2-V cells. The total potential is 12 V. Therefore, an external voltage of over 12 V is required to recharge the battery.

6. The reaction in an electrolytic cell is non-spontaneous; it requires an input of energy to occur. The energy is in the form of electrical energy. The external source of electrons forces electrons onto one electrode, which becomes negative relative to the other electrode. This drives the movement of ions and non-spontaneous redox reactions.

7. (a) The question is asking whether the following reaction is spontaneous:

$Al(s) + 3H_2O(\ell) \rightarrow H_2(g) + Al(OH)_3(s)$

The oxidation of aluminium is considered to occur at the anode. The reduction of water is considered to occur at the cathode.

$E^o_{cell} = E^o_{cathode} - E^o_{anode}$

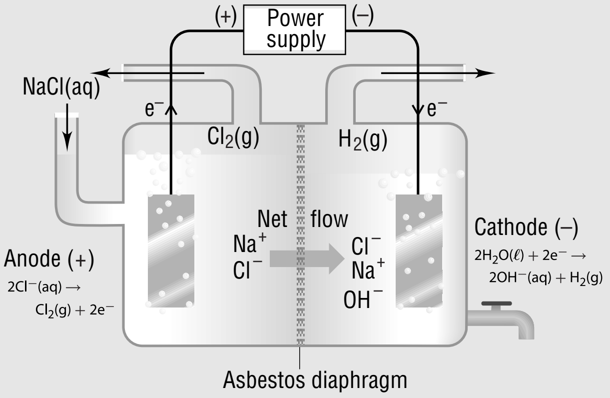
$= -0.83\text{ V} - (-1.66\text{ V}) = +0.83\text{ V}$

The reaction should be spontaneous under standard conditions.

(b) The boiling point of water is 100 °C, which is not the standard temperature. The aluminium vessel is likely covered by a thin but durable coating of unreactive aluminium oxide. Also, predictions are not reliable when reduction potentials are similar (aluminium is just below water on the chart of standard reduction potentials), and when gaseous products are expected, which is the case in this example.

8. There are advantages and disadvantages to lead-acid batteries. Lead-acid batteries have been available for almost 150 years, so the technology is stable and has been improved on over the years. Modern batteries are completely sealed to minimize one of the potential risks, which is the possibility of a chemical burn if someone comes in contact with the acid in the battery. Although the charging process is never perfect, lead-acid batteries are long-lasting and if handled properly, much of the lead (up to 97%) can be recycled during disposal. On the negative side, lead is heavy, which decreases the energy efficiency of the car, and if disposed of improperly, lead will contaminate soil and water. If lead-acid batteries are cracked, the sulfuric acid inside will leak and may cause burns. In addition, if improperly charged or if an electrical short-circuit takes place, hydrogen gas that is produced when the battery is charged could be ignited if sparked, leading to a fire or an explosion.

9.



The chlor-alkali cell is an electrolytic cell that oxidizes chloride ions, $Cl^-(aq)$ from brine (salt water) and reduces water, forming hydrogen gas and hydroxide ions.

Oxidation: $2Cl^-(aq) \rightarrow Cl_2(g) + 2e^-$
Reduction: $2H_2O(l) + 2e^- \rightarrow H_2(g) + 2OH^-(aq)$

The salt water enters the cell, and a battery provides the potential required to electrolyze the salt water. The products of the electrolysis are removed from the cell – the chlorine gas from the anode, the hydrogen gas from the cathode, and the aqueous sodium hydroxide is isolated from the electrolyte.

13.4 Stoichiometry and Faraday's Law

Student Textbook pages 514–523

Section Outcomes

Students will:

- calculate mass, amounts, current, and time in electrolytic cells by applying Faraday's law and stoichiometry
- describe the ways in which scientific knowledge led to the development of electroplating and refining metals from ores

Key Terms

Faraday's law
extraction
refining

Chemistry Background

- Faraday's law helps to predict how many electrons will be generated over a given amount of time, which can be applied to such processes as electroplating and metal refining.
- During electroplating, the object to be plated is placed at the cathode where metallic cations deposit on it.

Teaching Strategies

- Introduce this section by asking students what information they would need to determine the cost of refining a gram of copper. In your discussion, highlight the need for a way to relate current, charge, and amount of chemical products produced in an electrolytic cell.
- Discuss the units for current and quantity of electricity to help students who like to use unit analysis. One coulomb per second is equivalent to one ampere. In physics, students may have learned that the symbol for current is I (measured in A, amperes), and the symbol for quantity of electricity is Q (measured in C, coulombs). The relationship between the two is expressed by:

$$I = \frac{Q}{\Delta t}$$

$$\text{Unit analysis: A} = \frac{\text{coulomb}}{\text{second}} = \frac{\text{C}}{\text{s}}$$

- **BLM 13.4.1 (OH) Calculating the Mass of an Electrolysis Product** has been prepared for this section. You will find it with the Chapter 13 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

Many of the calculations in this section involve the production of moles of electrons. Turn Faraday's law into a relay race as an activity for kinesthetic learners. Have students fill a bucket (the cathode) with bean bags (moles of electrons) that they must carry a different number of laps (amps and time). Each student can represent a different metal, changing the number of bean bags carried during each lap.

Answers to Practice Problems 9–12

Student Textbook page 516

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

9. 2.97 g

10. 8.90 min

11. 1.85 A

12. **(a)** 1.09 g increase

(b) Because the overall balanced equation can be determined from the given information, the mole ratio is sufficient information to calculate the mass change.

Investigation 13.C: Electroplating

Student Textbook pages 517–518

Purpose

This activity is designed to give students some experience in electroplating an object. However, it is also a good activity to get students thinking about sources of error and the limitations of their collected data because the actual amount of material plated is often considerably less than predicted.

Outcomes

- 30-B2.8k
- 30-B2.3s
- 30-B2.1s

Advance Preparation

When to Begin	What to Do
1 – 2 weeks before	■ Order materials.
1 – 2 days before	■ Prepare solutions. See **BLM 15.4.4** for tips. ■ Cut wire and metal strips. ■ Photocopy **BLM 13.4.2: Investigation 13.C.**

Materials
■ 3 cm × 12 cm × 1 mm Cu strip
■ 150 mL 1.0 mol/L HNO_3 in a 250 mL beaker
■ deionized water in a wash bottle
■ 50 cm 16-gauge bare solid copper wire
■ 120 mL acidified 0.50 mol/L $CuSO_4$ solution (with 5 mL of 6 mol/L H_2SO_4 and 3 mL of 0.1 mol/L HCl added)
■ fine sandpaper
■ 250 mL beaker
■ 2 electrical leads with alligator clips
■ adjustable D.C. power supply with ammeter
■ drying oven
■ electronic balance

Time Required

- 20–180 minutes

Helpful Tips

- You must use DC power supplies to control current and voltage. The physics department may have DC power supplies.
- Analyze the results of one or several classes, rather than having students look at their individual experimental results. Compare the class median to the theoretical number calculated using Faraday's law. (Using the median eliminates high and low end errors and provides a representative number that reflects the conditions, materials, and expertise of the students.)
- Use **BLM 13.4.2 (HAND) Investigation 13.C: Electroplating** to support this activity. Remove sections as appropriate to meet the needs of students in your class.
- ***Expected Results:*** Solid copper will be deposited on the cathode, making its mass greater. The anode will lose mass because the copper will be converted into copper ions and dissolve in the solution.

Safety Precautions

- 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.
- Provide proper disposal vessels and ensure students know how to dispose of materials properly. (Follow your board's specific guidelines for disposal procedures.)
- Nitric acid is corrosive. Note that the $CuSO_4$(aq) solution contains sulfuric acid and hydrochloric acid. Wash any spills on skin with plenty of cold water. Remind students to avoid touching the parts of the electrodes that have been washed with nitric acid.
- Acetone must be used in a fume hood. If your classroom does not have a fume hood, arrange to switch classrooms with another class or to use these facilities if they are available in another area.
- Remind students to make sure their hands and lab bench are dry before handling any electrical equipment.

Answers to Prediction

Students will likely predict that the mass of the copper plated onto the cathode will be less than predicted using Faraday's law. A prediction made using Faraday's law involves a theoretical calculation, which does not take into account the experimental conditions, variations in materials used, and experimental expertise of the students.

Answers to Analysis Questions

1. The half-reaction occurring at the cathode (reduction of copper) is:

$Cu^{2+}(aq) + 2e^- \rightarrow Cu(s)$

2. Students should use their measured results to calculate quantity of electricity, as done in the Sample Problem on page 515 of the student textbook. Assuming a current of 1.00 A and 1200 s:

$$\begin{aligned}\text{Quantity of electricity} &= 1.00\text{ A} \times 1200\text{ s}\\ &= 1.20 \times 10^3\text{ C}\end{aligned}$$

3. $$\begin{aligned}\text{Amount of electrons} &= 1200\text{ C} \times \frac{1\text{ mol e}^-}{96\ 500\text{ C}}\\ &= 0.0124\text{ mol e}^-\end{aligned}$$

The half reactions is: $Cu^{2+}(aq) + 2e^- \rightarrow Cu(s)$

$$\begin{aligned}\text{Amount of copper formed} &= 0.0124\text{ mol e}^- \times \frac{1\text{ mol Cu}}{2\text{ mol e}^-}\\ &= 0.0062\text{ mol Cu}\end{aligned}$$

$$\begin{aligned}\text{Mass of copper formed} &= 0.0062\text{ mol Cu} \times \frac{63.6\text{ g}}{\text{mol Cu}}\\ &= 0.40\text{ g}\end{aligned}$$

4. Assuming no trivial errors, the differences between the calculated mass and the actual mass will likely be the result of improper drying or differences in the actual current flow (quantity of electrons) moving through the circuit.

Answer to Conclusion Question

5. Typically the mass of the copper electroplated onto the cathode will be less than predicted, assuming students have completely dried the anode. The answer for the second part of this question will be based on student predictions at the start of the investigation.

Answers to Applications Questions

6. Assuming that the iron(II) ions will plate out in the reaction, compare the mass of iron that should theoretically plate out to the mass of the copper plated out. In this case, since the mass of 1 mol of iron is less than that of 1 mol of copper, and since the number of moles of electrons required per mole of iron(II) ions is the same as required for copper(II) ions, the total mass of iron plated out should be less than the mass of copper plated out.

7. With the current running for half the time, you would expect the mass of copper plated to be half the original amount. If the current runs for half the time, then only half the number of electrons are provided, and only half of the reduction reactions can take place.

8. The simple answer is no because there would be no difference between the standard reduction potentials of the electrodes. The more complicated answer is yes, if a porous barrier separates the anode and cathode and the concentrations of the copper sulfate solution (same electrolyte) are different on each side of the porous barrier. This voltaic cell, called a concentration cell, uses the fact that the two solutions will have a tendency to reach the same concentration level as the basis for the movement of electrons through the external circuit. Students may also suggest that the slight differences in the purity of the electrodes could result in a very small cell potential.

Assessment Options

- Collect and assess students' Predictions and Observations, as well as their answers to Analysis, Conclusion, and Application questions.
- Use Assessment Checklist 2: Laboratory Report from Appendix A.
- Answers to Applications questions could be used to launch a poster activity. Use Assessment Checklists 3: Performance Task Self-Assessment or 4: Performance Task Group Assessment could be adapted to assess the posters.

Section 13.4 Review Answers

Student Textbook page 520

1. (a) Nickel is electroplated onto the silver master disc. The master disc must provide electrons to reduce the nickel ions to nickel atoms. Reduction reactions take place at the cathode of the electrolytic cell. Therefore, the master disc is the cathode.

(b) Number of moles of Ni formed

$$\begin{aligned}&= \frac{1\text{ g Ni}}{58.69\text{ g/mol}}\\ &= 1.704 \times 10^{-2}\text{ mol}\end{aligned}$$

To reduce each mole of nickel, 2 mol of electrons are required. Therefore, 0.034 08 mol of electrons are required for 1.704×10^{-2} mol of nickel.

$$\begin{aligned}\text{Quantity of electricity} &= 0.034\ 08\text{ mol e}^- \times \frac{96\ 500\text{ C}}{1\text{ mol e}^-}\\ &= 3.289 \times 10^3\text{ C}\end{aligned}$$

2. $$\begin{aligned}\text{Quantity of electricity} &= 3.00 \times 10^4\text{ A} \times 8.64 \times 10^4\text{ s}\\ &= 2.59 \times 10^9\text{ C}\end{aligned}$$

$$\begin{aligned}\text{Amount of electrons} &= 2.59 \times 10^9\text{ C} \times \frac{1\text{ mol e}^-}{96\ 500\text{ C}}\\ &= 2.69 \times 10^4\text{ mol e}^-\end{aligned}$$

The half-reactions are:

$Na^+(\ell) + 1e^- \rightarrow Na(s)$

$2Cl^-(\ell) \rightarrow Cl_2(g) + 2e^-$

Amount of sodium formed

$$\begin{aligned}&= 2.69 \times 10^4\text{ mol e}^- \times \frac{1\text{ mol Na}}{1\text{ mol e}^-}\\ &= 2.69 \times 10^4\text{ mol Na}\end{aligned}$$

Mass of sodium formed

$$= 2.69 \times 10^4 \text{ mol Na} \times \frac{22.99 \text{ g}}{\text{mol Na}}$$
$$= 6.18 \times 10^5 \text{ g}$$
$$= 618 \text{ kg}$$

Amount of chlorine formed

$$= 2.69 \times 10^4 \text{ mol e}^- \times \frac{1\text{mol Cl}_2}{2 \text{ mol e}^-}$$
$$= 1.34 \times 10^4 \text{ mol Cl}_2$$

Mass of chlorine formed

$$= 1.34 \times 10^4 \text{ mol Cl}_2 \times \frac{70.90 \text{ g}}{\text{mol Cl}_2}$$
$$= 9.50 \times 10^5 \text{ g}$$
$$= 950 \text{ kg}$$

3. Time in seconds

$$= 365 \frac{\text{days}}{\text{year}} \times 24 \frac{\text{hours}}{\text{day}} \times \frac{60 \text{ min}}{\text{hour}} \times \frac{60 \text{ s}}{\text{min}}$$
$$= 3.154 \times 10^7 \text{ s}$$

Quantity of electricity

$$= 2.00 \times 10^2 \text{ A} \times 3.154 \times 10^7 \text{ s}$$
$$= 6.31 \times 10^9 \text{ C}$$

Amount of electrons

$$= 6.31 \times 10^9 \text{ C} \times \frac{1 \text{ mol e}^-}{96\,500 \text{ C}}$$
$$= 6.54 \times 10^4 \text{ mol e}^-$$

Assume the reduction reaction is:

$Cu^{2+}(aq) + 2e^- \rightarrow Cu(s)$

Amount of copper formed

$$= 6.54 \times 10^4 \text{ mol e}^- \times \frac{1 \text{ mol Cu}}{2 \text{ mol e}^-}$$
$$= 3.27 \times 10^4 \text{ mol Cu}$$

Mass of copper formed

$$= 3.27 \times 10^4 \text{ mol Cu} \times \frac{63.6 \text{ g}}{\text{mol Cu}} = 2.08 \times 10^6 \text{ g}$$
$$= 208 \text{ t}$$

4. Even with the cost of transporting the bauxite to Canada, it is still cheaper to extract the aluminium by electrolytic processes because of the abundance of relatively cheap electricity in Canada.

5. $$m\text{Al(s)} = 22.7 \text{ t} \times \frac{1000 \text{ kg}}{1\text{t}} \times \frac{1000 \text{ g}}{1 \text{ kg}}$$
$$= 2.27 \times 10^7 \text{ g Al(s)}$$

$$n\text{Al(s)} = \frac{m}{M}$$
$$= \frac{2.27 \times 10^7 \text{ g Al(s)}}{\frac{26.98 \text{ g Al(s)}}{\text{mol Al(s)}}}$$

8.41×10^5 mol Al(s)

$Al(s) \rightarrow Al^{3+}(aq) + 3e^-$

$$n_e = 8.41 \times 10^5 \text{ mol Al(s)} \times \frac{3 \text{ mol e}^-}{1 \text{ mol Ag(s)}}$$
$$= 2.52 \times 10^6 \text{ mol e}^-$$

$$\text{amount of electrons} = 2.52 \times 10^6 \text{ mol e}^- \times \frac{9.65 \times 10^4 \text{ C}}{\text{mol e}^-}$$
$$= 2.43 \times 10^{11} \text{ C}$$

$$q = I\Delta t$$
$$\Delta t = \frac{2.43 \times 10^{11} \text{ C}}{2.5 \times 10^5 \text{ C/s}}$$
$$= 9.74 \times 10^5 \text{ s}$$
$$= 11 \text{ days}$$

It would take 11 days of continuous cell operation to produce 22.7 t of aluminium.

$$E = qV$$
$$= (2.44 \times 10^{11} \text{ C})(4.5 \text{ V})$$
$$= 1.1 \times 10^{12} \text{ J}$$

The amount of energy used to produce 22.7 t of aluminium is 1.1×10^{12} J.

6. Nickel extraction is analogous to copper extraction. The $Cu_2S(s)$ is roasted in air and some of it is converted to $Cu_2O(s)$:

$2Cu_2S(s) + 3O_2(g) \rightarrow 2Cu_2O(s) + 2SO_2(g)$

The copper compounds react to produce copper:

$Cu_2S(s) + 2Cu_2O(s) \rightarrow 6Cu(s) + SO_2(g)$

The copper is further refined by electrolysis. Sulfur dioxide gas reacts with water to form sulfurous acid, which contributes to acid precipitation. To prevent the release of sulfur dioxide into the atmosphere, various recovery programs have been initiated. One example is the electrostatic precipitation program at INCO's refinery in Sudbury, Ontario.

Chapter 13 Review Answers

Student Textbook pages 522–523

Answers to Understanding Concepts Questions

1. (a) Electrodes are the conductors that carry electrons into and out of the cell. Oxidation and reduction occur at the electrodes.

(b) Electrolytes are substances that dissolve in water to form ions that can move and conduct electricity.

(c) The external voltage is a source of electricity that is included in the external circuit of an electrolytic cell.

2. $Mg(s)\,|Mg^{2+}(aq)||Cd^{2+}(aq)|\,Cd(s)$

3. The oxidation half-reaction is:

$NO(g) + 2H_2O(\ell) \rightarrow NO_3^-(aq) + 4H^+(aq) + 3e^-$

The reduction half-reaction is:

$I_2(s) + 2e^- \rightarrow 2I^-(aq)$

Balance the electrons and add the equations to get the overall cell reaction:

$$2NO(g) + 4H_2O(\ell) + 3I_2(s) \rightarrow 2NO_3^-(aq) + 8H^+(aq) + 6I^-(aq)$$

4. The hydrogen electrode is defined as the reference electrode against which other reduction potentials are measured. The reduction potential of the standard hydrogen electrode is assigned a value of 0.00 V.

5. Electrolyzing aqueous solutions often results in unwanted products due to redox reactions involving water.

6. The two half-reactions are:

$\text{CuI(s)} + 1\text{e}^- \rightarrow \text{Cu(s)} + \text{I}^-\text{(aq)}$

$\text{N}_2\text{O(g)} + 2\text{H}^+\text{(aq)} + 2\text{e}^- \rightarrow \text{N}_2\text{(g)} + \text{H}_2\text{O}(\ell)$

First assume that copper is oxidized at the anode.

Write a balanced equation.

$2\text{Cu(s)} + \text{I}^-\text{(aq)} \rightarrow 2\text{CuI(s)} + 2\text{e}^-$

Add the two equations to get the balanced net ionic equation.

$\text{N}_2\text{O(g)} + 2\text{H}^+\text{(aq)} + 2\text{Cu(s)} + 2\text{I}^-\text{(aq)} \rightarrow \text{N}_2\text{(g)} + \text{H}_2\text{O}(\ell) + 2\text{CuI(s)}$

$E^o_{\text{cell}} = E^o_{\text{cathode}} - E^o_{\text{anode}} = 1.77\text{ V} - (-0.18\text{ V})$

$= +1.95\text{ V}$ (spontaneous reaction)

Now assume that copper ions are reduced at the cathode:

$\text{N}_2\text{(g)} + \text{H}_2\text{O}(\ell) + 2\text{CuI(s)} \rightarrow \text{N}_2\text{O(g)} + 2\text{H}^+\text{(aq)} + 2\text{Cu(s)} + 2\text{I}^-\text{(aq)}$

$E^o_{\text{cell}} = E^o_{\text{cathode}} - E^o_{\text{anode}} = -0.18\text{ V} - (+1.77\text{ V})$

$= -1.95\text{ V}$ (non-spontaneous reaction)

7. (a) The oxidizing agent is $\text{PbO}_2\text{(s)}$ and the reducing agent is Pb(s).

(b) The oxidizing agent is Pb(s) and the reducing agent is $\text{PbO}_2\text{(s)}$.

8. In order of the most effective to least effective oxidizing agents, based on the table of standard reduction potentials and assuming that the reduction potential for cobalt is between that of hydrogen ions and zinc (based on the activity series), the order is:

$\text{Br}_2(\ell)$, $\text{H}^+\text{(aq)}$, $\text{Co}^{3+}\text{(aq)}$, $\text{Zn}^{2+}\text{(aq)}$

9. Based on the table of standard reduction potentials, the order is:

Al(s), $\text{H}_2\text{(g)}$, Ag(s), $\text{Cl}^-\text{(aq)}$

10. The metals would plate-out in order of their standard reduction potentials. Of the three metals, silver has the most positive standard reduction potential (i.e., it is most easily reduced). It will plate out first. Then copper will plate out, followed by iron.

Answers to Applying Concepts Questions

11. (a) $\text{Zn(s)} \rightarrow \text{Zn}^{2+}\text{(aq)} + 2\text{e}^-$ (anode)

$\text{Fe}^{2+}\text{(aq)} + 2\text{e}^- \rightarrow \text{Fe(s)}$ (cathode)

$E^o_{\text{cell}} = E^o_{\text{cathode}} - E^o_{\text{anode}} = -0.45\text{ V} - (-0.76\text{ V})$

$= +0.31\text{ V}$ (spontaneous)

(b) $\text{Cr(s)} \rightarrow \text{Cr}^{3+}\text{(aq)} + 3\text{e}^-$ (anode)

$\text{Al}^{3+}\text{(aq)} + 3\text{e}^- \rightarrow \text{Al(s)}$ (cathode)

$E^o_{\text{cell}} = E^o_{\text{cathode}} - E^o_{\text{anode}} = -1.66\text{ V} - (-0.74\text{ V})$

$= -0.92\text{ V}$ (non-spontaneous)

(c) $2\text{H}_2\text{O}_2(\ell) \rightarrow \text{O}_2\text{(g)} + 2\text{H}^+\text{(aq)} + 2\text{e}^-$ (anode)

$\text{Ag}^+\text{(aq)} + \text{e}^- \rightarrow \text{Ag(s)}$ (cathode)

$E^o_{\text{cell}} = E^o_{\text{cathode}} - E^o_{\text{anode}} = 0.80\text{ V} - (+0.70\text{ V})$

$= +0.10\text{ V}$ (spontaneous)

12. (a) Connect a voltmeter to the external circuit with the negative terminal connected to the tin anode and the positive terminal connected to the lead cathode.

(b) This cell is unlikely to find any practical uses because the cell voltage is so low and the components are heavy. ($E^o_{\text{cell}} = +0.01\text{ V}$)

13. (a) oxidation at anode: $\text{Fe(s)} \rightarrow \text{Fe}^{2+}\text{(aq)} + 2\text{e}^-$

reduction at cathode: $\text{Ag}^+\text{(aq)} + \text{e}^- \rightarrow \text{Ag(s)}$

$E^o_{\text{cell}} = E^o_{\text{cathode}} - E^o_{\text{anode}}$

$= 0.80\text{ V} - (-0.45\text{ V}) = +1.25\text{ V}$

Balance charge and add the two half-reactions to get the cell reaction:

$2\text{Ag}^+\text{(aq)} + \text{Fe(s)} \rightarrow 2\text{Ag(s)} + \text{Fe}^{2+}\text{(aq)}$

(b) oxidation at anode: $\text{Ag(s)} \rightarrow \text{Ag}^+\text{(aq)} + \text{e}^-$

reduction at cathode: $\text{Fe}^{2+}\text{(aq)} + 2\text{e}^- \rightarrow \text{Fe(s)}$

$E^o_{\text{cell}} = E^o_{\text{cathode}} - E^o_{\text{anode}} = 0.80\text{ V} - (-0.45\text{ V})$

$= +\ 1.25\text{ V}$ (spontaneous)

Balance charge and add the two half-reactions to get the cell reaction:

$2\text{Ag(s)} + \text{Fe}^{2+}\text{(aq)} \rightarrow 2\text{Ag}^+\text{(aq)} + \text{Fe(s)}$

(c) For the voltaic cell, the mass of the iron electrode will decrease as iron atoms form iron(II) ions. The mass of the silver electrode will increase as the silver ions form silver atoms.

(d) For the electrolytic cell, the mass of the silver electrode will decrease as silver atoms form silver ions. The mass of the iron electrode will increase as iron(II) ions become iron atoms.

14. (a) Students should suggest using inert electrodes to electrolyze the aqueous zinc bromide and testing the solutions around the electrodes (e.g., with phenolphthalein) to determine what is produced. Bromine in solution would appear orange-yellow.

(b) Based on standard reduction potentials, you would not expect zinc to be oxidized. Instead, you would expect water to be oxidized.

15. Numerous combinations are possible. For example, a cell in which $\text{Al}^{3+}\text{(aq)}$ is reduced to Al(s) and Fe(s) is oxidized to $\text{Fe}^{3+}\text{(aq)}$ has a cell potential of 1.62 V. Five of these cells in series make a battery of 8.10 V.

16. (a) Luigi Galvani's great interest was "animal electricity," which he studied in his post as a teacher of medicine in Bologna. He constructed a crude electric cell with two different metals and the natural fluids from a dissected frog. Alessandro Volta's first important invention was the electrophorus, a device capable of passing an electric charge to another object. Arguably his most important contribution was the development and study of the "voltaic pile," which provided proof that electricity could be conducted through, and be produced by, non-organic matter.

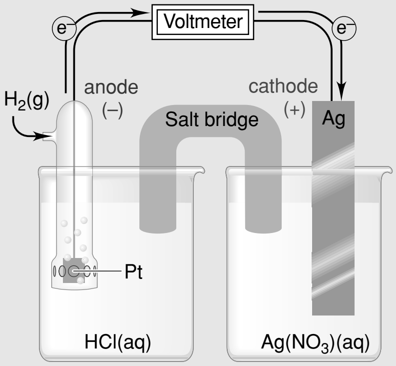
(b) Davy studied Galvani's work with the "voltaic pile," and followed the examples of Nicholson and Carlisle by using electrolysis to extract elemental potassium and sodium from molten potash and table salt, respectively. He continued his work on electrolysis and he discovered that any electrolyte would pass current, as long as it was capable of oxidizing one of the metals. He also determined that the intensity of the effect (the voltage) was directly related to the reactivity of the electrolyte with the metal.

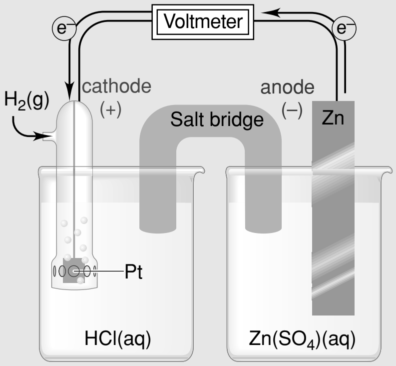
In the 1830s Faraday worked on developing his ideas on electricity. He stated a new theory of electrochemical action. One of the results of this work was that he coined, with William Whewell, many of the words we currently use, including electrode, electrolyte, anode, cathode, and ion. His work led him to reject the traditional theory that electricity was an imponderable fluid or fluids, and he proposed that electricity was a form of force that passed from particle to particle of matter.

17. The surface of the moon does not have an atmosphere or water. Since oxygen is not available for the redox reaction, iron would not corrode.

18. It is simply reversing the direction of the electron flow so that a reduction reaction becomes an oxidation reaction, and visa versa.

19.





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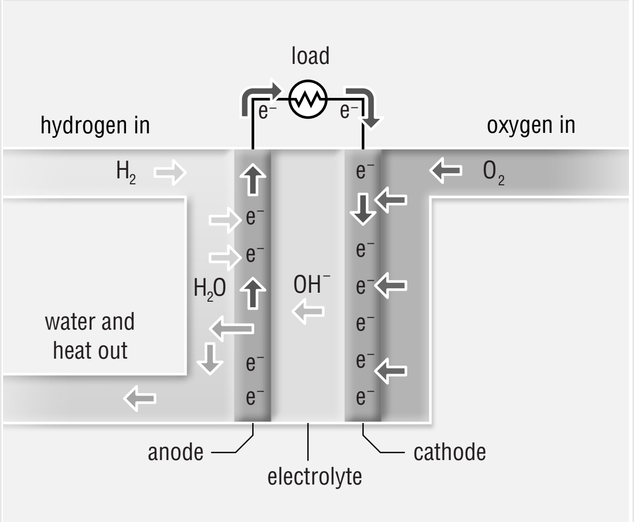


Answers to Solving Problems Questions

21. 2.46 V; Al/Al^{3+} oxidation half cell; Ag/Ag^{+} reduction half cell

22. 1.2 g of copper

23. Time in s $= 55.0 \text{ min} \times 60 \text{ s/min} = 3.300 \times 10^3 \text{ s}$

$$\text{Quantity of electricity} = 3.65 \text{ A} \times 3.300 \times 10^3 \text{ s}$$
$$= 1.204 \times 10^4 \text{ C}$$

$$\text{Amount of electrons} = 1.2048 \times 10^4 \text{ C} \times \frac{1 \text{ mol e}^-}{96\ 500 \text{ C}}$$
$$= 0.125 \text{ mol e}^-$$

Assume the reduction reaction is:

$$Mg^{2+}(aq) + 2e^- \rightarrow Mg(s)$$

$$\text{Amount of Mg formed} = 0.1248 \text{ mol e}^- \times \frac{1 \text{ mol Mg}}{2 \text{ mol e}^-}$$
$$= 0.0624 \text{ mol Mg}$$

$$\text{Mass of Mg formed} = 0.0624 \text{ mol Mg} \times \frac{24.31 \text{ g}}{1 \text{ mol Mg}}$$
$$= 1.52 \text{ g}$$

24. Number of moles Na $= 1000 \text{ g} \times \dfrac{1 \text{ mol}}{22.99 \text{ g}}$

$= 43.50$ mol

Number of moles Al $= 1000 \text{ g} \times \dfrac{1 \text{ mol}}{26.98 \text{ g}}$

$= 37.06$ mol

The amount of each material in terms of moles is very close. However, reducing 1 mole $Al^{3+}(aq)$ requires 3 mole of electrons, while reducing 1 mole $Na^{+}(aq)$ requires only 1 mole of electrons. Producing 1 kg of aluminium would require about 2.6 times more electricity than producing 1 kg of Na(s).

25. 1.98 kg/h

26. 0.133 A

Answers to Making Connections Questions

27. The quantity of electricity depends on the current generated. The current generated depends on the quantity and identity of substances involved in the electrochemical reaction. Because the batteries are different sizes, they probably contain different materials. Also, the larger battery is likely to last longer. Even if the batteries produced the same current, the larger battery would produce a greater overall quantity of electricity.

28. **(a)** Yes, because aluminium is reduced more readily than iron. The aluminium would be oxidized, forming a relatively inert oxide coating, and thus hold the iron gutter in place while protecting the iron from oxidation.

(b) No, because after the aluminium siding formes the surface oxide, the iron nails would rust and flake to pieces.

29. A cell fuelled by aluminium generates power through an electrochemical reaction between the metal, placed in a saline or alkaline solution, and oxygen from the air. Electricity is produced as the aluminium oxidizes (at the anode), so it acts as the reducing agent. Oxygen from the air acts as a cathode. To recharge the cell when the aluminium is consumed, the plates are replaced and more electrolyte is added.

30. About 95% less energy is required to process used aluminium than is to process the aluminium from the ore. The aluminium in most products tends to be relatively pure, and there is little surface oxide to remove during the recycling process.

31. Students may say they are against the smelter because the waste products that are formed, such as sulfur dioxide, can contribute to acid precipitation. Students may say they are in favour of the plant because of the economic situation and the fact that the modern mechanisms for removing waste (sulfur dioxide) from the roasting process results in a reduction in the quantities of these pollutants. Accept all reasoned and reasonable answers, but students should demonstrate some understanding of the processes involved.

32. The iron "ribs" of the statue, which supported the copper "skin," corroded due to the effects of wind, salty air, pollution, and time. Because the iron was in contact with the copper, which acted as a cathode, the ribs rusted much faster than if they had not been in contact with the copper. During the restoration process, the iron ribs were removed and replaced with copper ribs.

33. **(a)** Estimates will vary based on your classroom. Ensure students consider all uses of batteries, including watches, cellular phones, and portable CD players.

(b) Students may suggest purchasing rechargeable or longer-life batteries as a partial solution. They could suggest using an AC adapter for battery powered devices whenever possible. Or, students could simply find ways to reduce their use of battery-powered devices.

Career Focus: Ask a Pipeline Corrosion Expert

Student Textbook pages 524–525

Teaching Strategies

- If possible, invite an individual whose profession involves dealing with corrosion. This may be someone who works with the chemistry involved, who works at developing preventative measures, or someone who must deal with the consequences of corrosion.
- Discuss with students some everyday examples of corrosion that they have heard about in the town or city that you live in, or that they may have experienced first-hand.
- Review the corrosion process and reactions that were covered in this unit with students.
- The following web sites may be of interest to students interested in pursuing occupations mentioned in this Career Focus.
  - The National Association of Corrosion Engineers (NACE) at http://www.nace.org
  - Canadian Association of Home and Property Inspectors (CAHPI) at http://www.cahi.ca
  - The Alberta Energy and Utilities Board http://www.eub.gov.ab.ca

Answers to Go Further Questions

1. Corrosion of iron:

$$4Fe^{2+}(aq) + O_2(g) + 8OH^{-}(aq) \rightarrow 2Fe_2O_3{\bullet}H_2O(s) + 2H_2O(\ell)$$

2. Corrosion of copper:

$$2Cu(s) + O_2(g) + CO_2(g) + H_2O(\ell) \rightarrow Cu_2(OH)_2CO_3(s)$$

3. Hydrogen sulfide acts as a reducing agent forming metal sulfides. The metal in the pipeline is reduced by the hydrogen sulfide, which leads to corrosion of the metal.
4. Infrastructure refers to the foundation elements necessary to carry something out. For example, educational infrastructure includes school buildings, administration, and finances; municipal infrastructures include roads, sewers, water systems, power supplies and delivery. In the oil industry, infrastructure refers to pipeline networks, roads, and utilities, as well as groups such as the Alberta Energy and Utilities Board to supply expertise and support for common elements required by the industry.

Unit 6 Review Answers

Student Textbook pages 526–529

Answers to Understanding Concepts Questions

1. Oxidation was originally defined as the reaction of an atom or compound with oxygen. Since then, oxidation has been redefined as a reaction in which electrons are lost.

   $2Na(s) + O_2(g) \rightarrow Na_2O(s)$

   ionic equation: $2Na(s) + O_2(g) \rightarrow 2Na+(s) + O^{2-}(s)$

   $2Na(s) + Cl_2(g) \rightarrow 2NaCl(s)$

   ionic equation: $2Na(s) + Cl_2(g) \rightarrow 2Na^+(s) + 2Cl^-(s)$

   The modern theoretical definition of oxidation describes chemical reactions that resemble the reaction with oxygen, but with compounds other than oxygen. In both cases, electrons are lost by the sodium atoms.
2. ionic equation: $Sn(s) + Cd^{2+}(aq) + 2NO_3^-(aq) \rightarrow Sn^{2+}(aq) + 2NO_3^-(aq) + Cd(s)$

   net ionic equation: $Sn(s) + Cd^{2+}(aq) \rightarrow Sn^{2+}(aq) + Cd(s)$

   (a) The spectator ion, $NO_3^-(aq)$, appears unchanged on both the reactant and product side of the equation. Just as a spectator watches a sporting event but does not take part, a spectator ion "watches" a chemical reaction as it does not change during the reaction.

   (b) The reducing agent is Sn(s), as it is oxidized in the reaction.

   (c) The oxidizing agent is $Cd(NO_3)_2(aq)$, or $Cd^{2+}(aq)$, as it is reduced in the reaction.

   (d) Sn(s) is oxidized.

   (e) $Cd^{2+}(aq)$ is reduced.
3. The reaction in b) will occur spontaneously.
4. Smelting reduces the iron from iron(II) ions to metallic iron, also called pig iron. The pig iron is then refined in a series of steps that involves the oxidation of the impurities to remove them from the purified steel.
5. **(a)** N = +3

   (b) P = +5

   (c) Si = +4

   (d) N = –3, H = +1, S = +6, O = –2
6. No. An oxidation number of –2 for oxygen in water means that the oxygen atom attracts electrons more strongly than the hydrogen atom in each O—H bond. Oxygen is considered to have "gained" the two electrons it shares with the hydrogen atoms.
7. For reactions that use only molecular compounds, the oxidation number method is easier.
8. **(a)** $Ca(OH)_2(aq) + CO_2(g) \rightarrow CaCO_3(s) + H_2O(\ell)$

   (b) The reaction is not a redox reaction because the oxidation numbers do not change.
9. Sodium chloride melts at temperatures above 800 °C. Therefore, the electrolysis is not carried out under standard conditions.

Answers to Applying Concepts Questions

10. Students should use either the oxidation number method or the half-reaction method (preferable) to write the reaction between acidified hydrogen peroxide, $H_2O_2(aq)$ and aqueous potassium permanganate. The balanced net ionic equation is shown below:

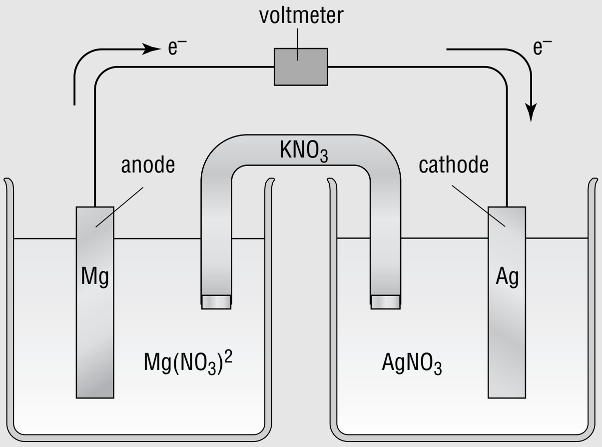
    $$5H_2O_2(aq) + 2MnO_4^-(aq) + 6H^+(aq) \rightarrow 5O_2(g) + 2Mn^{2+}(aq) + 8H_2O(\ell)$$

    Volume of $MnO_4^-(aq)$ added =

    21.54 mL – 1.34 mL = 20.20 mL

    $$C_{H_2O_2} = \frac{0.852\text{ g }KMnO_4(s)}{1} \times \frac{1\text{ mol }KMnO_4(s)}{158.04\text{ g }KMnO_4(s)} \times \frac{1}{0.250\text{ L}} \times \frac{0.0202\text{ L}}{1} \times \frac{1\text{ mol }MnO_4^-(aq)}{1\text{ mol }KMnO_4(s)} \times \frac{5\text{ mol }H_2O_2}{2\text{ mol }MnO_4} \times \frac{1}{0.025\text{ L}}$$

    $$= 4.36 \times 10^{-2}\text{ mol/L }H_2O_2(aq)$$
11. **(a)** to **(c)**



    (d) anode: $Mg(s) \rightarrow Mg^{2+}(aq) + 2e^-$

    cathode: $Ag^+(aq) + e^- \rightarrow Ag(s)$

(e) $Mg(s) + 2Ag^+(aq) \rightarrow Mg^{2+}(aq) + 2Ag(s)$

(f) $E^\circ_{cell} = E^\circ_{cathode} - E^\circ_{anode}$

$E^\circ_{cell} = 0.80\ V - (-2.37\ V)$

$E^\circ_{cell} = +3.17\ V$

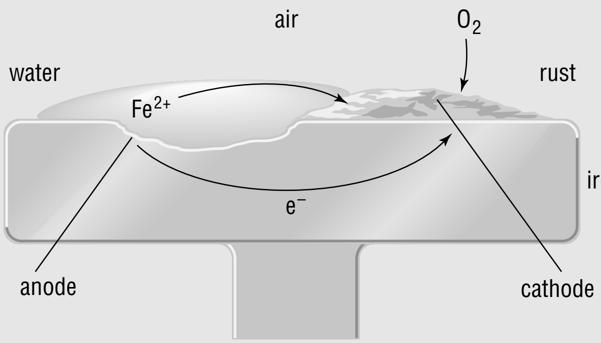
(g) $Mg(s) \mid Mg^{2+}(aq) \parallel Ag^+(aq) \mid Ag(s)$

12. An inert electrode is made from a material that is neither a reactant nor a product of the redox reaction but can carry a current and provide a surface on which redox reactions can occur. Carbon and platinum are commonly used as inert electrodes.

13. "Standard reduction potential" refers to the amount of energy conferred on a unit of charge as the material listed is reduced. Since potential must be measured in reference to another half-cell, the hydrogen half-cell is used as the reference point; its reduction potential is set at 0.00 V and all other half-cells are measured relative to this.

14. A primary battery cannot be recharged while a secondary battery can be recharged.

15.



Anode: $Fe(s) \rightarrow Fe^{2+}(aq) + 2e^-$

Cathode: $O_2(g) + 2H_2O(\ell) + 4e^- \rightarrow 4OH^-(aq)$

Overall reaction: $2Fe(s) + O_2(g) + 2H_2O(\ell) \rightarrow 2Fe^{2+}(aq) + 4OH^-(aq)$

16. Protect from exposure to oxygen and water: Cover the material with paint, oil, or plastic to prevent exposure to the reactants involved in the corrosion reaction.

Alloys: Mix iron with other metals to reduce corrosion.

Galvanization: Coat with zinc. As the zinc corrodes, a protective layer is formed. Also, zinc is a stronger reducing agent than iron, so it will corrode preferentially.

Cathodic protection: Attach a stronger reducing agent such as magnesium, which will corrode preferentially, protecting the iron.

17. In most cases, electrolysis will lead to the reduction of the strongest oxidizing agent. In an aqueous solution, water will also behave as an oxidizing agent. Water is a stronger oxidizing agent than many reactive metal ions including $Cr^{2+}(aq)$, $Al^{3+}(aq)$, $Mg^{2+}(aq)$, $Na^+(aq)$, $Ca^{2+}(aq)$, $Ba^{2+}(aq)$, $K^+(aq)$ and $Li^+(aq)$, and will therefore be reduced before any of these metal ions.

18. The impure metal is placed at the anode. As a current is applied, the copper on the anode dissolved, and pure copper can be plated on the cathode.

19. The electrolyte is sulfuric acid. As the battery discharges, sulfate ions combine with the lead in the battery to form solid lead(II) sulfate and water. Sulfuric acid solution is denser than water. Therefore, the electrolyte solution is most dense when it is fully charged.

20. The following half-reactions are involved:

(1) $2H^+(aq) + 2e^- \rightarrow H_2(g)$

(reduction at cathode, $E^o_{cell} = 0.00\ V$)

(2) $2H_2O(\ell) \rightarrow O_2(g) + 4H^+(aq) + 4e^-$

(oxidation at anode, E°_{cell} for reverse reaction $= 1.23\ V$)

(3) $2H_2O(\ell) + 2e^- \rightarrow H_2(g) + 2OH^-(aq)$

(reduction at cathode, $E^\circ_{cell} = -0.83\ V$)

(4) $2Cl^-(aq) \rightarrow Cl_2(g) + 2e^-$

(oxidation at anode, E°_{cell} for reverse reaction $= 1.36\ V$)

Using $E^\circ_{cell} = E^\circ_{cathode} - E^\circ_{anode}$ the most positive (least negative) cell potential is obtained using half-reactions (1) and (2):

$E^\circ_{cell} = 0.000\ V - (1.23\ V) = -1.23\ V$

The overall reaction is:

$2H_2O(\ell) \rightarrow O_2(g) + 2H_2(g)$

The predicted products are oxygen and hydrogen. However, the production of hydrogen requires overvoltage. Therefore, chlorine may form instead.

21. (a) If the lead completely covers the iron, no oxygen can contact the iron as long as the coating is intact. Lead has a more positive reduction potential than iron, and will oxidize more slowly than iron.

(b) If any iron is exposed, it acts as an anode and is oxidized (rusts).

22. Students could choose to construct an electrolytic cell with an iron nail at the cathode in a solution of any cation listed below it on the reactivity series (such as zinc or aluminium).

23. (a) The half-cell that has the most positive reduction potential acts as an anode with the three other half-cells. The half-cell that has the second most positive reduction potential acts as an anode with the two half-cells with more negative reduction potentials. The half-cell that has the third most positive reduction potential acts as an anode with the half-cell with the most negative reduction potential. In total, you could make six different voltaic cells.

(b) You only need to determine three cell potentials. Choose one half-cell as your reference, and measure the cell potential of the reference half-cell paired with each of the other three half-cells. This is analogous to determining standard reduction potentials based on the hydrogen electrode.

24. Refer to Figure 13.23 on page 502 of the student textbook. Replace the zinc electrode with lead and the copper electrode with silver. Use the electrolytes lead nitrate and silver nitrate.

25. Reduction of hydrogen ions:

$2H^+(aq) + 2e^- \rightarrow H_2(g)\ E^o = 0.00\ V$

Reduction of water molecules:

$2H_2O(\ell) + 2e^- \rightarrow H_2(g) + 2OH^-(aq)\ E^o = -0.83\ V$

Reverse the first equation and add the two equations:

$2H_2O(\ell) \rightarrow 2H^+(aq) + 2OH^-(aq)$

$E^o_{cell} = E^o_{cathode} - E^o_{anode}$

$= -0.83\ V - (0.00) = -0.83\ V$

The negative sign indicates that the reaction is non-spontaneous under standard conditions. This is consistent with the extremely low concentration of hydrogen and hydroxide ions in pure water.

26. A balanced half-reaction equation indicates whether oxidation or reduction is taking place, how many electrons are involved, and identifies oxidizing and reducing agents and products. For example, consider the following equations:

$Zn(s) \rightarrow Zn^{2+}(aq) + 2e^-$

Because the electrons are on the products side, the half-reaction is an oxidation. Zinc is a reducing agent. Each zinc atom loses two electrons, and Zn^{2+} is produced.

$Cu^{2+} + 2e^- \rightarrow Cu(s)$

Because the electrons are on the reactants side, the half-reaction is a reduction. Copper(II) ions are an oxidizing agent. Each copper ion gains two electrons, and copper metal is produced.

27. (a) Metal A is the more effective reducing agent. It is oxidized and facilitates the reduction of the other metal. For example, in the following reaction, magnesium is the more effective reducing agent.

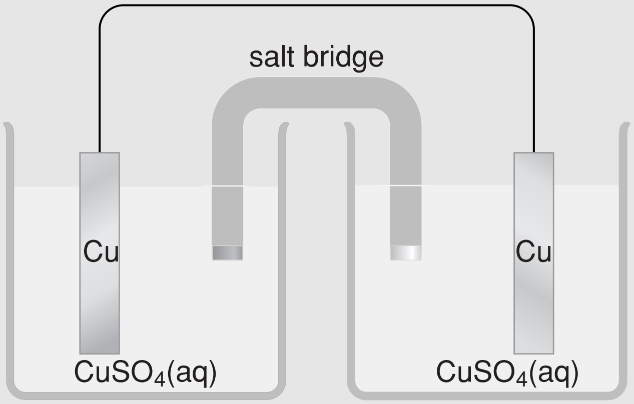
$Mg(s) + Zn(NO_3)_2(aq) \rightarrow Zn(s) + Mg(NO_3)_2(aq)$

(b) For one nonmetal to replace another during a single displacement reaction, it must gain electrons more easily than the nonmetal it is displacing. Therefore, it must be a better oxidizing agent. An example is:

$F_2(g) + 2NaCl(aq) \rightarrow 2NaF(aq) + Cl_2(g)$

28. (a) Since the electrodes are both made of copper, and neither water nor sulfate ions will create a potential difference in the cell, the cell potential is zero.

(b) By adding an external source of electrons to create an anode and a cathode, the cell can be operated as an electrolytic cell. The copper metal anode will decrease in mass and the copper metal cathode will increase in mass.



29. Half-reactions represent oxidations or reductions, which never occur independently. For each electron "lost" in an oxidation reaction, an electron must be "gained" in a reduction reaction. Electrons never appear in the overall equation for a redox reaction, which is consistent with the fact that electrons are not created or destroyed in chemical reactions.

30. (a) A disproportionation reaction occurs when one element is both oxidized and reduced in the same reaction.

(b) $3ClO^-(aq) \rightarrow 2Cl^-(aq) + ClO_3^-(aq)$

31. (a) An electrolytic cell drives a non-spontaneous reaction, meaning that E^o_{cell} is negative. Since

$E^o_{cell} = E^o_{cathode} - E^o_{anode}$ and E^o_{anode} is 0 V, E^o cathode must be negative.

(b) Since $E^o_{cell} = E^o_{cathode} - E^o_{anode}$ and $E^o_{cathode}$ is 0 V, E^o_{anode} must be positive for $E°_{cell}$ to be negative.

32. The Downs cell and the chlor-alkali process are similar in that they both produce chlorine gas through oxidation of chloride ions. In the chlor-alkali process, a saturated aqueous solution of sodium chloride (brine) is electrolyzed. The products are chlorine gas, hydrogen gas, and sodium hydroxide. A Downs cell, in contrast, electrolyzes molten sodium chloride. The products are sodium metal and chlorine gas. To maintain the sodium chloride as a liquid, the reaction in a Downs cell must be carried out at a very high temperature (about 600 °C). This is not necessary in the chlor-alkali process.

33. (a) Students should emphasize that oxidation occurs when an atom "loses" electrons and reduction occurs when an atom "gains" electrons, and that neither process occurs independently of the other. They may include sample half-reactions and redox equations showing oxidation numbers.

(b) Voltaic cells convert chemical energy to electrical energy through spontaneous redox reactions. Electrolytic cells convert electrical energy to chemical energy by providing a source of electrons that allow a non-spontaneous redox reaction to occur. Students should include the idea that voltaic cells are essentially the reverse of electrolytic cells. Students should provide an example, including a diagram, of each type of cell.

Answers to Solving Problems Questions

34. (a) $Hg_2^{2+}(aq) + 2e^- \rightarrow 2Hg$

(b) $TiO_2(s) \rightarrow Ti^{2+}(aq)$ (acidic conditions)

Titanium is reduced (+4 to +2):

$TiO_2(s) \rightarrow Ti^{2+}(aq)$

Balance the number of oxygen atoms:

$TiO_2(s) \rightarrow Ti^{2+}(aq) + 2H_2O$

Balance the number of hydrogen atoms (acidic conditions):

$4H^+(aq) + TiO_2(s) \rightarrow Ti^{2+}(aq) + 2H_2O$

Balance charge by adding electrons:

$4H^+(aq) + TiO_2(s) + 2e^- \rightarrow Ti^{2+}(aq) + 2H_2O$

(c) $I_2 \rightarrow H_3IO_6^{3-}$ (basic conditions)

Iodine is oxidized (0 to +6):

$I_2(s) \rightarrow H_3IO_6^{3-}(aq)$

Balance the number of iodine atoms:

$I_2(s) \rightarrow 2H_3IO_6^{3-}(aq)$

Balance the number of oxygen atoms:

$12H_2O(\ell) + I_2(s) \rightarrow 2H_3IO_6^{3-}(aq)$

Balance the number of hydrogen atoms:

$12H_2O(\ell) + I_2(s) \rightarrow 2H_3IO_6^{3-} + 18H^+(aq)$

Add $OH^-(aq)$ to neutralize the hydrogen ions:

$18OH^-(aq) + 12H_2O(\ell) + I_2(s) \rightarrow$
$2H_3IO_6^{3-}(aq) + 18H^+(aq) + 18OH^-(aq)$

Combine hydrogen and hydroxide ions and remove water common to both sides. Balance charge by adding electrons:

$18OH^-(aq) + I_2(s) \rightarrow$
$2H_3IO_6^{3-}(aq) + 6H_2O(\ell) + 12e^-$

35. (a) Oxidation

$Fe(s) \rightarrow Fe^{2+}(aq) + 2e^-$

Reduction

$4H^+(aq) + NO_3^-(aq) + 3e^- \rightarrow NO(g) + 2H_2O(\ell)$

Write the half-reactions using the LCM

$3Fe(s) \rightarrow 3Fe^{2+}(aq) + 6e^-$

$8H^+(aq) + 2NO_3^-(aq) + 6e^- \rightarrow 2NO(g) + H_2O(\ell)$

Add the two half-reactions, replace the spectator nitrate ions, and combine the nitrate ions with the hydrogen ions and the iron(II) ions:

$3Fe(s) + 8HNO_3(aq) \rightarrow$
$3Fe(NO_3)_2(aq) + 2NO(g) + 4H_2O(\ell)$

(b) Oxidation

$Fe^{2+}(aq) \rightarrow Fe^{3+}(aq) + e^-$

Reduction

$4H^+(aq) + NO_3^-(aq) + 3e^- \rightarrow NO(g) + H_2O(\ell)$

Write the half-reactions using the LCM

$3Fe^{2+}(aq) \rightarrow 3Fe^{3+}(aq) + 3e^-$
$4H^+(aq) + NO_3^-(aq) + 3e^- \rightarrow NO(g) + 2H_2O(\ell)$

Add the two half-reactions, replace the spectator nitrate ions, and combine the nitrate ions with the hydrogen ions and the iron(III) ions:

$3Fe(NO_3)_2(aq) + 4HNO_3(aq) \rightarrow$
$3Fe(NO_3)_3(aq) + NO(g) + 2H_2O(\ell)$

(c) Oxidation

$Fe(s) \rightarrow Fe^{3+}(aq) + 3e^-$

Reduction

$4H^+(aq) + NO_3^-(aq) + 3e^- \rightarrow NO(g) + 2H_2O(\ell)$

The balanced reaction is:

$Fe(s) + 4H^+(aq) + NO_3^-(aq) \rightarrow$
$Fe^{3+}(aq) + NO(g) + 2H_2O(\ell)$

(d) These all represent corrosion reactions.

36. (a) The anode is an inert carbon electrode, and the cathode is a silver electrode. The positive electrode is the cathode, the negative electrode is the anode.

(b) oxidation at anode: $2I^-(aq) \rightarrow I_2(s) + 2e^-$

reduction at cathode: $Ag^+(aq) + 1e^- \rightarrow Ag(s)$

overall cell reaction: $2I^-(aq) + 2Ag^+(aq) \rightarrow$
$I_2(s) + 2Ag(s)$

(c) Silver is the oxidizing agent and the iodide ion is the reducing agent.

(d) $E^o_{cell} = E^o_{cathode} - E^o_{anode}$

$= 0.80\ V - (+0.54\ V) = +0.26\ V$

37. (a) The reaction is not a redox reaction because the oxidation numbers do not change.

$Cl_2O_7(aq) + H_2O(\ell) \rightarrow 2HClO_4(aq)$

(b) The oxidizing agent is $ClO_3^-(aq)$ and the reducing agent is $I_2(s)$. Students should use the half-reaction method to balance the equation for acidic conditions.

$3I_2(s) + 3H_2O(\ell) + 5ClO_3^-(aq) \rightarrow$
$6IO_3^-(aq) + 5Cl^-(aq) + 6H^+(aq)$

(c) Br_2 is the oxidizing agent and S^{2-} is the reducing agent. Students should use the half-reaction method to balance the equation for basic conditions.

$8OH^-(aq) + S^{2-}(aq) + 4Br_2(\ell) \rightarrow$
$SO_4^{2-}(aq) + 8Br_2(\ell) + 4H_2O(\ell)$

(d) The oxidizing agent is HNO_3 and the reducing agent is H_2S. Students can balance this equation using the half-reactions method or the oxidation number method.

$3H_2S(aq) + 2HNO_3(aq) \rightarrow$
$3S(s) + 2NO(g) + 4H_2O(\ell)$

38. (a) $E^o_{\text{cell}} = E^o_{\text{cathode}} - E^o_{\text{anode}} = -0.23$ V; non–spontaneous

(b) $E^o_{\text{cell}} = E^o_{\text{cathode}} - E^o_{\text{anode}} = +0.70$ V; spontaneous

(c) $E^o_{\text{cell}} = E^o_{\text{cathode}} - E^o_{\text{anode}} = +0.42$ V; spontaneous

39. (a) Students should balance the equation for acidic conditions using the half-reaction method:

$$2MnO_4^-(aq) + 5C_2O_4^{2-}(aq) + 16H^+(aq) \rightarrow 2Mn^{2+}(aq) + 10CO_2(g) + 8H_2O(\ell)$$

(b) Amount of $KMnO_4$ = 14.28 ml × 0.1575 mol/1000 mL

= 2.249×10^{-3} mol

Amount $(NH_4)_2C_2O_4$ = 5 mol $Na_2C_2O_4$/2 $KMnO_4$ × 2.249×10^{-3} mol $KMnO_4$

= 5.623×10^{-3} mol $Na_2C_2O_4$

Concentration of oxalate solution = 5.623×10^{-3} mol $Na_2C_2O_4$/0.02500L

= 0.225 mol/L

40. Cerium(IV) ions will be reduced and tin(II) ions will be oxidized according to the following equations:

Reduction: $Ce^{4+}(aq) + e^- \rightarrow Ce^{3+}(aq)$

Oxidation: $Sn^{2+}(aq) \rightarrow Sn^{4+}(aq) + 2e^-$

Based on these equations, 1 mol of cerium(IV) ions are reduced by 0.5 mol of tin(II) ions.

Amount of cerium(IV) ions = 0.024 38 L × 0.2113 mol/L

= 5.151×10^{-3} mol cerium(IV) ions

Amount of Sn(II) ions = 2.576×10^{-3} mol

Mass of Sn = 2.576×10^{-3} mol × 118.71 g/mol

= 0.305 8 g Sn

Mass % of tin = 20.73%

41. (a) Amount of magnesium ions = $\dfrac{3.38 \text{ g}}{24.31 \text{ g/mol}}$

= 0.139 mol

Amount of electrons = 0.139 mol Mg^{2+} × $\dfrac{2 \text{ mol } e^-}{1 \text{ mol } Mg^{2+}}$

= 0.278 mol e^-

Quantity of electricity = $\dfrac{96\ 500 \text{ C}}{1 \text{ mol } e^-} \times 0.278 \text{ mol } e^-$

= 2.68×10^4 C

(b) 2.68×10^4 C = 100×10^{-3} A × t

t = 2.68×10^5 s ÷ 3600 s/h

= 74.5 h

42. Quantity of electricity = 2.92 A × 71.0 min × $\dfrac{60 \text{ s}}{1 \text{ min}}$

= 1.24×10^4 C

Amount of electrons = $1.244 \times 10^4 \text{ C} \times \dfrac{1 \text{ mol } e^-}{96\ 500 \text{ C}}$

= 0.1289 mol e^-

Amount of aluminium = $0.1289 \text{ mol } e^- \times \dfrac{1 \text{ mol Al}}{3 \text{ mol } e^-}$

= 0.042 97 mol Al

Mass of aluminium = $0.0429 \text{ mol Al} \times \dfrac{26.98 \text{ g}}{\text{mol Al}}$

= 1.16 g

43. Quantity of electricity = $5000 \text{ A} \times 7 \text{ days} \times \dfrac{24 \text{ h}}{1 \text{ day}} \times \dfrac{3600 \text{ s}}{1 \text{ h}}$

= 3.024×10^9 C

Amount of electrons = $3.024 \times 10^9 \text{ C} \times \dfrac{1 \text{ mol } e^-}{96\ 500 \text{ C}}$

= 3.134×10^4 mol e^-

Amount of fluorine = $3.134 \times 10^4 \text{ mol } e^- \times \dfrac{1 \text{ mol } F_2}{2 \text{ mol } e^-}$

= 1.567×10^4 mol F_2

Mass of fluorine = $1.567 \times 10^4 \text{ mol } F_2 \times \dfrac{38.00 \text{ g}}{\text{mol } F_2}$

= 5.954×10^5 g or 0.5954 t

44. (a) (i) The gas at the anode would be chlorine gas, $Cl_2(g)$, as long as sufficient voltage is supplied (this is the chlor-alkali cell).

anode: $2Cl^-(aq) \rightarrow Cl_2(g) + 2e^-$

(ii) A "pop" test could be used to identify the gas in tube C. A lit splint would ignite the hydrogen gas, $H_2(g)$, produced, resulting in a pop.

cathode: $2H_2O(\ell) + 2e^- \rightarrow H_2(g) + 2OH^-(aq)$

(iii) Platinum can also be used as an inert electrode.

(iv) $E^o_{\text{cell}} = E^o_{\text{cathode}} - E^o_{\text{anode}}$

$E^o_{\text{cell}} = 0.83 \text{ V} - (-1.36 \text{ V})$

$E^o_{\text{cell}} = 2.19$ V

(v) Collector A contains chlorine gas, which is made at the same rate as hydrogen gas, therefore equal volumes will be made. Chlorine is very soluble in water, however, so you will not see as much gas collected in the tube.

(vi) Litmus will turn blue at the cathode since hydroxide ions are produced. At the anode, the litmus will be bleached by the presence of chlorine gas dissolved in the water.

(b) (i) Chlorine gas will still be produced at the anode, but at the cathode, solid sodium will be deposited:

anode: $2Cl^-(\ell) \rightarrow Cl_2(g) + 2e^-$

cathode: $Na^+(\ell) + e^- \rightarrow Na(s)$

(ii) $E^o_{\text{cell}} = E^o_{\text{cathode}} - E^o_{\text{anode}}$

$E^o_{\text{cell}} = 2.71 \text{ V} - (-1.36 \text{ V})$

$E^o_{\text{cell}} = 4.07$ V

The calculated minimum voltage uses "standard" reduction potentials, which are recorded based on 1.0 mol/L aqueous solutions at 0 °C and 101.325 kPa. In order to operate this cell, the temperature is raised to 600 °C, and calcium chloride is added to the molten salt. This calculation assumes that the change in temperature and the use of molten salts instead of aqueous solutions will not affect the cell potential.

(iii) Cell 1: $2H_2O(\ell) + 2Cl^-(aq) \rightarrow Cl_2(g) + H_2(g) + 2OH^-(aq)$

Cell 2: $2Cl^-(\ell) + 2Na^+(\ell) \rightarrow Cl_2(g) + 2Na(s)$

In cell 1, water is reduced while in cell 2, sodium ions are reduced.

(c) (i) When water is present, it will be preferentially reduced compared to sodium ions because it is a much stronger oxidizing agent than sodium ions. To reduce sodium ions, stronger oxidizing agents cannot be present.

(ii) Aqueous salts such as sodium nitrate, $NaNO_3(aq)$, or sodium sulfate, $Na_2SO_4(aq)$, would lead to the disproportionation of water and lead to oxygen gas production at the anode and hydrogen gas production at the cathode.

(iii) The flow of electrons is always from the anode to the cathode.

Answers to Making Connections Questions

45. Since the speed of developing is related to the rate of decomposition of the silver halide, and since silver iodide is used for high-speed film, silver iodide is more sensitive to light than silver bromide.

46. (a) $8Al(s) + 3NH_4ClO_4(aq) \rightarrow 4Al_2O_3(s) + 3NH_4Cl(aq)$

The oxidizing agent is ammonium perchlorate. The reducing agent is aluminium.

(b) $2H_2(g) + O_2(g) \rightarrow 2H_2O(\ell)$

(c) In the first reaction, ammonium chloride is produced, which could participate in harmful reactions in the environment. The second reaction produces only water. Therefore, the second reaction is more environmentally friendly.

47. Corroding iron mains must be repaired or replaced, which costs money. Corroding mains may also leak and their hydraulic capacity may decrease, which decreases their efficiency. Leaking and corroding water mains may allow contaminants into drinking water. Several ways to combat these problems include replacing iron water mains with polymer or cement mains, which do not corrode, or using sacrificial anodes to prevent corrosion.

48. Iron is abundant (the second most abundant metal after aluminium), relatively inexpensive to mine and process, and can be used to create a number of different alloys with a variety of useful physical properties.