Q14. Copper is plated onto the cathode of an electrolytic cell containing $CuCl_2(aq)$. How long does it take to plate 111 mg of copper with a current of 3.8 A? **MISSED THIS?** *Read Section 20.8*

a) 1.3×10^3 s b) 44 s c) 89 s d) 22 s

Q15. Which metal can be used as a sacrificial electrode to prevent the rusting of an iron pipe? **MISSED THIS?** *Read Section 20.9*

a) Au b) Ag c) Cu d) Mn

Answers: **1.** (a) **2.** (d) **3.** (d) **4.** (a) **5.** (c) **6.** (a) **7.** (c) **8.** (a) **9.** (b) **10.** (d) **11.** (c) **12.** (a) **13.** (c) **14.** (c) **15.** (d)

CHAPTER 20 IN REVIEW

TERMS

Section 20.3

electrical current (901)
electrochemical cell (902)
voltaic (galvanic) cell (902)
electrolytic cell (902)
half-cell (902)
electrode (902)
ampere (A) (903)
potential difference (903)
volt (V) (903)
electromotive force (emf) (903)
cell potential (cell emf) (E_{cell}) (904)
standard cell potential (standard emf) (E°_{cell}) (904)
anode (904)
cathode (904)
salt bridge (904)

Section 20.4

standard electrode potential (905)
standard hydrogen electrode (SHE) (906)

Section 20.5

Faraday's constant (F) (914)

Section 20.6

Nernst equation (918)

Section 20.7

dry-cell battery (922)
alkaline battery (923)
lead–acid storage battery (923)
nickel–cadmium (NiCad) battery (924)
nickel–metal hydride (NiMH) battery (924)
lithium ion battery (924)
fuel cell (925)

Section 20.8

electrolysis (926)

Section 20.9

corrosion (933)

CONCEPTS

Pulling the Plug on the Power Grid (20.1)

- Oxidation–reduction reactions are reactions in which electrons are transferred from one reactant to another.
- In the most common form of fuel cell, an electrical current is created as hydrogen is oxidized and oxygen is reduced; water is the only product.

Balancing Oxidation–Reduction Equations (20.2)

- Oxidation is the loss of electrons and corresponds to an increase in oxidation state; reduction is the gain of electrons and corresponds to a decrease in oxidation state.
- We can balance redox reactions using the half-reaction method, in which the oxidation and reduction reactions are balanced separately and then added. This method differs slightly for redox reactions in acidic and in basic solutions.

Voltaic (or Galvanic) Cells: Generating Electricity from Spontaneous Chemical Reactions (20.3)

- A voltaic electrochemical cell separates the reactants of a spontaneous redox reaction into two half-cells that are connected by a wire and a means to exchange ions so that electricity is generated.
- In an electrochemical cell, the electrode where oxidation occurs is the anode and the electrode where reduction occurs is the cathode; electrons flow from the anode to the cathode.
- The rate of electrons flowing through a wire is measured in amperes (A), and the cell potential is measured in volts (V).
- A salt bridge allows ions to flow between the half-cell solutions and prevent the buildup of charge.
- The cell diagram or line notation symbolizes electrochemical cells concisely by separating the components of the reaction using lines or commas.

Standard Electrode Potentials (20.4)

- The electrode potentials of half-cells are measured in relation to that of a standard hydrogen electrode, which is assigned an electrode potential of zero under standard conditions (solution concentrations of 1 M, gas pressures of 1 atm, and a temperature of 25 °C).
- A species with a highly positive E° has a strong tendency to attract electrons and undergo reduction (and is therefore an excellent oxidizing agent).
- A species with a highly negative E° has a strong tendency to repel electrons and undergo oxidation (and is therefore an excellent reducing agent).

Cell Potential, Free Energy, and the Equilibrium Constant (20.5)

- In a spontaneous reaction, E°_{cell} is positive, the change in free energy (ΔG°) is negative, and the equilibrium constant (K) is greater than 1.
- In a nonspontaneous reaction, E°_{cell} is negative, ΔG° is positive, and K is less than 1.
- Because E°_{cell}, ΔG°, and K all relate to spontaneity, we can derive equations relating all three quantities.

Cell Potential and Concentration (20.6)

- The standard cell potential (E°_{cell}) is related to the cell potential (E_{cell}) by the Nernst equation, $E_{cell} = E^\circ_{cell} - (0.0592\ V/n) \log Q$.
- As shown by the Nernst equation, E_{cell} is related to the reaction quotient (Q); E_{cell} equals zero when Q equals K.
- In a concentration cell, the reactions at both electrodes are identical and electrons flow because of a difference in concentration. Nerve cells are a biological example of concentration cells.

Batteries: Using Chemistry to Generate Electricity (20.7)

- Batteries are packaged voltaic cells.
- Dry-cell batteries, including alkaline batteries, do not contain large amounts of water.
- The reactions in rechargeable batteries, such as lead–acid storage, nickel–cadmium, nickel–metal hydride, and lithium ion batteries, can be reversed.
- Fuel cells are similar to batteries except that fuel-cell reactants must be continually replenished from an external source.

Electrolysis: Driving Nonspontaneous Chemical Reactions with Electricity (20.8)

- An electrolytic electrochemical cell differs from a voltaic cell in that (1) an electrical charge is used to drive the reaction, and (2) although the anode is still the site of oxidation and the cathode the site of reduction, they are represented with signs opposite those of a voltaic cell (anode +, cathode −).
- In electrolysis reactions, the anion is oxidized; if there is more than one anion, the anion with the more negative *E*° is oxidized.
- We can use stoichiometry to calculate the quantity of reactants consumed or products produced in an electrolytic cell.

Corrosion: Undesirable Redox Reactions (20.9)

- Corrosion is the undesired oxidation of metal by environmental oxidizing agents.
- When some metals, such as aluminum, oxidize, they form a stable compound that prevents further oxidation. Iron, in contrast, does not form a structurally stable compound when oxidized, and therefore, rust flakes off and exposes more iron to corrosion.
- Iron corrosion can be prevented by protecting iron from water, minimizing the presence of electrolytes and acids, or using a sacrificial electrode.

EQUATIONS AND RELATIONSHIPS

Definition of an Ampere (**20.3**)

$$1\ \text{A} = 1\ \text{C/s}$$

Definition of a Volt (**20.3**)

$$1\ \text{V} = 1\ \text{J/s}$$

Standard Hydrogen Electrode (**20.4**)

$$2\ \text{H}^+(aq) + 2\ \text{e}^- \longrightarrow \text{H}_2(g) \quad E^\circ = 0.00\ \text{V}$$

Equation for Cell Potential (**20.4**)

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

Relating ΔG° and E°_{cell} (**20.5**)

$$\Delta G^\circ = -nFE^\circ_{\text{cell}} \qquad F = \frac{96{,}485\ \text{C}}{\text{mol e}^-}$$

Relating E°_{cell} and K (**20.5**)

$$E^\circ_{\text{cell}} = \frac{0.0592\ \text{V}}{n}\log K \qquad (\text{at } 25\ ^\circ\text{C})$$

The Nernst Equation (**20.6**)

$$E_{\text{cell}} = E^\circ_{\text{cell}} - \frac{0.0592\ \text{V}}{n}\log Q \qquad (\text{at } 25\ ^\circ\text{C})$$

LEARNING OUTCOMES

Chapter Objectives	Assessment
Balance acidic solution redox equations using the half-reaction method (**20.2**)	Examples 20.1, 20.2 **For Practice** 20.1, 20.2 **Exercises** 37–40
Balance basic solution redox equations using the half-reaction method (**20.2**)	Example 20.3 **For Practice** 20.3 **Exercises** 41–42
Identify the parts of a voltaic cell (**20.3**)	**Exercises** 43–44
Represent electrochemical cells using cell diagram notation (**20.3**)	**Exercises** 49–52
Calculate standard cell potentials (E°_{cell}) for electrochemical cells (**20.4**)	Example 20.4 **For Practice** 20.4 **Exercises** 45–48, 61–62
Predict the spontaneity of redox reactions (**20.4**)	Example 20.5 **For Practice** 20.5 **Exercises** 53–54, 61–62
Identify solutions that can selectively oxidize metals and metal ions (**20.4**)	**Exercises** 55–60, 63–64
Calculate standard free energy changes (ΔG°) for electrochemical reactions from standard cell potentials (E°_{cell}) (**20.5**)	Example 20.6 **For Practice** 20.6 **Exercises** 65–66
Perform calculations involving ΔG°, E°_{cell}, and K for electrochemical reactions (**20.5**)	Example 20.7 **For Practice** 20.7 **Exercises** 67–72, 103–104
Analyze cell potentials under nonstandard conditions (**20.6**)	Example 20.8 **For Practice** 20.8 **Exercises** 73–82
Analyze batteries in terms of mass components, redox reactions, and standard cell potentials (**20.7**)	**Exercises** 83–86
Analyze electrolytic cells in terms of component parts, reactions, voltages, and direction of electron flow (**20.8**)	**Exercises** 89–90, 97–98
Analyze molten-salt electrolysis reactions in terms of products and half-reactions (**20.8**)	**Exercises** 91–94
Analyze aqueous-solution electrolysis reactions in terms of products and half-reactions (**20.8**)	Example 20.9 **For Practice** 20.9 **Exercises** 95–96
Perform stoichiometric calculations for electrolysis reactions (**20.8**)	Example 20.10 **For Practice** 20.10 **Exercises** 99–102
Predict metals that act as sacrificial electrodes for iron (**20.9**)	**Exercises** 87–88

EXERCISES

Mastering Chemistry provides end-of-chapter exercises, feedback-enriched tutorial problems, animations, and interactive activities to encourage problem-solving practice and deeper understanding of key concepts and topics.

REVIEW QUESTIONS

1. In electrochemistry, spontaneous redox reactions are used for what purpose?

2. In electrochemistry, what kind of reaction can be driven by electricity?

3. Give the basic definitions of oxidation and reduction and explain the basic procedure for balancing redox reactions.

4. Explain the difference between a voltaic (or galvanic) electrochemical cell and an electrolytic cell.

5. What reaction (oxidation or reduction) occurs at the anode of a voltaic cell? What is the sign of the anode? Do electrons flow toward or away from the anode?

6. What reaction (oxidation or reduction) occurs at the cathode of a voltaic cell? What is the sign of the cathode? Do electrons flow toward or away from the cathode?

7. Explain the purpose of a salt bridge in an electrochemical cell.

8. What unit is used to measure the magnitude of electrical current? What unit is used to measure the magnitude of a potential difference? Explain how electrical current and potential difference differ.

9. What is the definition of the standard cell potential (E°_{cell})? What does a large positive standard cell potential imply about the spontaneity of the redox reaction occurring in the cell? What does a negative standard cell potential imply about the reaction?

10. Describe the basic features of a cell diagram (or line notation) for an electrochemical cell.

11. Why do some electrochemical cells employ inert electrodes such as platinum?

12. Describe the standard hydrogen electrode (SHE) and explain its use in determining standard electrode potentials.

13. How is the cell potential of an electrochemical cell (E°_{cell}) related to the potentials of the half-cells?

14. Does a large positive electrode potential indicate a strong oxidizing agent or a strong reducing agent? What about a large negative electrode potential?

15. Is a spontaneous redox reaction obtained by pairing any reduction half-reaction with one listed above it or with one listed below it in Table 20.1?

16. How can Table 20.1 be used to predict whether or not a metal will dissolve in HCl? In HNO_3?

17. Explain why E°_{cell}, $\Delta G^\circ_{\text{rxn}}$, and K are all interrelated.

18. Does a redox reaction with a small equilibrium constant ($K < 1$) have a positive or a negative E°_{cell}? Does it have a positive or a negative $\Delta G^\circ_{\text{rxn}}$?

19. How does E_{cell} depend on the concentrations of the reactants and products in the redox reaction occurring in the cell? What effect does increasing the concentration of a reactant have on E_{cell}? Increasing the concentration of a product?

20. Use the Nernst equation to show that $E_{\text{cell}} = E^\circ_{\text{cell}}$ under standard conditions.

21. What is a concentration electrochemical cell?

22. What are the anode and cathode reactions in a common dry-cell battery? In an alkaline battery?

23. What are the anode and cathode reactions in a lead–acid storage battery? What happens when the battery is recharged?

24. What are the three common types of portable rechargeable batteries, and how does each one work?

25. What is a fuel cell? What is the most common type of fuel cell, and what reactions occur at its anode and cathode?

26. Explain how a fuel-cell breathalyzer works.

27. List some applications of electrolysis.

28. The anode of an electrolytic cell must be connected to which terminal—positive or negative—of the power source?

29. What species is oxidized, and what species is reduced in the electrolysis of a pure molten salt?

30. If an electrolytic cell contains a mixture of species that can be oxidized, how do you determine which species will actually be oxidized? If it contains a mixture of species that can be reduced, how do you determine which one will actually be reduced?

31. Why does the electrolysis of an aqueous sodium chloride solution produce hydrogen gas at the cathode?

32. What is overvoltage in an electrochemical cell? Why is it important?

33. How is the amount of current flowing through an electrolytic cell related to the amount of product produced in the redox reaction?

34. What is corrosion? Why is corrosion only a problem for some metals (such as iron)?

35. Explain the role of each of the following in promoting corrosion: moisture, electrolytes, and acids.

36. How can the corrosion of iron be prevented?

PROBLEMS BY TOPIC

Balancing Redox Reactions

37. Balance each redox reaction occurring in acidic aqueous solution.
MISSED THIS? *Read Section 20.2; Watch KCV 20.2, IWE 20.2*
a. $K(s) + Cr^{3+}(aq) \longrightarrow Cr(s) + K^{+}(aq)$
b. $Al(s) + Fe^{2+}(aq) \longrightarrow Al^{3+}(aq) + Fe(s)$
c. $BrO_3^{-}(aq) + N_2H_4(g) \longrightarrow Br^{-}(aq) + N_2(g)$

38. Balance each redox reaction occurring in acidic aqueous solution.
a. $Zn(s) + Sn^{2+}(aq) \longrightarrow Zn^{2+}(aq) + Sn(s)$
b. $Mg(s) + Cr^{3+}(aq) \longrightarrow Mg^{2+}(aq) + Cr(s)$
c. $MnO_4^{-}(aq) + Al(s) \longrightarrow Mn^{2+}(aq) + Al^{3+}(aq)$

39. Balance each redox reaction occurring in acidic aqueous solution.
MISSED THIS? *Read Section 20.2; Watch KCV 20.2, IWE 20.2*
a. $PbO_2(s) + I^{-}(aq) \longrightarrow Pb^{2+}(aq) + I_2(s)$
b. $SO_3^{2-}(aq) + MnO_4^{-}(aq) \longrightarrow SO_4^{2-}(aq) + Mn^{2+}(aq)$
c. $S_2O_3^{2-}(aq) + Cl_2(g) \longrightarrow SO_4^{2-}(aq) + Cl^{-}(aq)$

40. Balance each redox reaction occurring in acidic aqueous solution.
a. $I^{-}(aq) + NO_2^{-}(aq) \longrightarrow I_2(s) + NO(g)$
b. $ClO_4^{-}(aq) + Cl^{-}(aq) \longrightarrow ClO_3^{-}(aq) + Cl_2(g)$
c. $NO_3^{-}(aq) + Sn^{2+}(aq) \longrightarrow Sn^{4+}(aq) + NO(g)$

41. Balance each redox reaction occurring in basic aqueous solution.
MISSED THIS? *Read Section 20.2; Watch KCV 20.2, IWE 20.3*

a. $H_2O_2(aq) + ClO_2(aq) \longrightarrow ClO_2^-(aq) + O_2(g)$
b. $Al(s) + MnO_4^-(aq) \longrightarrow MnO_2(s) + Al(OH)_4^-(aq)$
c. $Cl_2(g) \longrightarrow Cl^-(aq) + ClO^-(aq)$

42. Balance each redox reaction occurring in basic aqueous solution.

a. $MnO_4^-(aq) + Br^-(aq) \longrightarrow MnO_2(s) + BrO_3^-(aq)$
b. $Ag(s) + CN^-(aq) + O_2(g) \longrightarrow Ag(CN)_2^-(aq)$
c. $NO_2^-(aq) + Al(s) \longrightarrow NH_3(g) + AlO_2^-(aq)$

Voltaic Cells, Standard Cell Potentials, and Direction of Spontaneity

43. Sketch a voltaic cell for each redox reaction. Label the anode and cathode and indicate the half-reaction that occurs at each electrode and the species present in each solution. Also indicate the direction of electron flow. **MISSED THIS?** *Read Section 20.3*

a. $2\,Ag^+(aq) + Pb(s) \longrightarrow 2\,Ag(s) + Pb^{2+}(aq)$
b. $2\,ClO_2(g) + 2\,I^-(aq) \longrightarrow 2\,ClO_2^-(aq) + I_2(s)$
c. $O_2(g) + 4\,H^+(aq) + 2\,Zn(s) \longrightarrow 2\,H_2O(l) + 2\,Zn^{2+}(aq)$

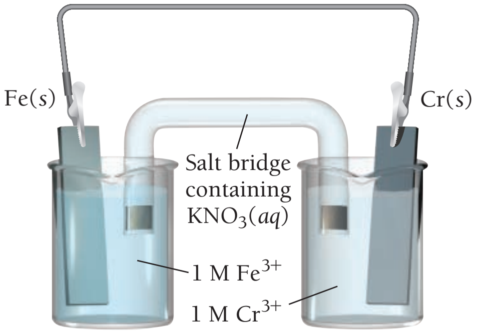
44. Sketch a voltaic cell for each redox reaction. Label the anode and cathode and indicate the half-reaction that occurs at each electrode and the species present in each solution. Also indicate the direction of electron flow.

a. $Ni^{2+}(aq) + Mg(s) \longrightarrow Ni(s) + Mg^{2+}(aq)$
b. $2\,H^+(aq) + Fe(s) \longrightarrow H_2(g) + Fe^{2+}(aq)$
c. $2\,NO_3^-(aq) + 8\,H^+(aq) + 3\,Cu(s) \longrightarrow 2\,NO(g) + 4\,H_2O(l) + 3\,Cu^{2+}(aq)$

45. Calculate the standard cell potential for each of the electrochemical cells in Problem 43.
MISSED THIS? *Read Section 20.4; Watch KCV 20.4, IWE 20.4*

46. Calculate the standard cell potential for each of the electrochemical cells in Problem 44.

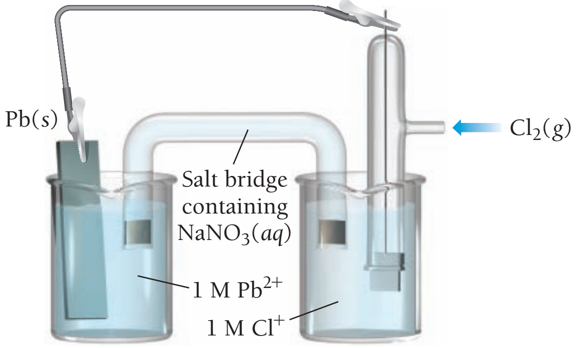
47. Consider the voltaic cell:
MISSED THIS? *Read Sections 20.3, 20.4; Watch KCV 20.4, IWE 20.4*



a. Determine the direction of electron flow and label the anode and the cathode.
b. Write a balanced equation for the overall reaction and calculate E°_{cell}.
c. Label each electrode as negative or positive.
d. Indicate the direction of anion and cation flow in the salt bridge.

48. Consider the voltaic cell:

a. Determine the direction of electron flow and label the anode and the cathode.
b. Write a balanced equation for the overall reaction and calculate E°_{cell}.
c. Label each electrode as negative or positive.
d. Indicate the direction of anion and cation flow in the salt bridge.



49. Use line notation to represent each electrochemical cell in Problem 43. **MISSED THIS?** *Read Section 20.3*

50. Use line notation to represent each electrochemical cell in Problem 44.

51. Make a sketch of the voltaic cell represented by the line notation. Write the overall balanced equation for the reaction and calculate E°_{cell}. **MISSED THIS?** *Read Section 20.3*

$$Sn(s)\,|\,Sn^{2+}(aq)\,\|\,NO(g)\,|\,NO_3^-(aq),\,H^+(aq)\,|\,Pt(s)$$

52. Make a sketch of the voltaic cell represented by the line notation. Write the overall balanced equation for the reaction and calculate E°_{cell}.

$$Mn(s)\,|\,Mn^{2+}(aq)\,\|\,ClO_2^-(aq)\,|\,ClO_2(g)\,|\,Pt(s)$$

53. Determine whether or not each redox reaction occurs spontaneously in the forward direction.
MISSED THIS? *Read Section 20.4; Watch KCV 20.4*

a. $Ni(s) + Zn^{2+}(aq) \longrightarrow Ni^{2+}(aq) + Zn(s)$
b. $Ni(s) + Pb^{2+}(aq) \longrightarrow Ni^{2+}(aq) + Pb(s)$
c. $Al(s) + 3\,Ag^+(aq) \longrightarrow Al^{3+}(aq) + 3\,Ag(s)$
d. $Pb(s) + Mn^{2+}(aq) \longrightarrow Pb^{2+}(aq) + Mn(s)$

54. Determine whether or not each redox reaction occurs spontaneously in the forward direction.

a. $Ca^{2+}(aq) + Zn(s) \longrightarrow Ca(s) + Zn^{2+}(aq)$
b. $2\,Ag^+(aq) + Ni(s) \longrightarrow 2\,Ag(s) + Ni^{2+}(aq)$
c. $Fe(s) + Mn^{2+}(aq) \longrightarrow Fe^{2+}(aq) + Mn(s)$
d. $2\,Al(s) + 3\,Pb^{2+}(aq) \longrightarrow 2\,Al^{3+}(aq) + 3\,Pb(s)$

55. Which metal could you use to reduce Mn^{2+} ions but not Mg^{2+} ions? **MISSED THIS?** *Read Section 20.4; Watch KCV 20.4*

56. Which metal can be oxidized with an Sn^{2+} solution but not with an Fe^{2+} solution?

57. Determine whether or not each metal dissolves in 1 M HCl. For those metals that do dissolve, write a balanced redox reaction showing what happens when the metal dissolves.
MISSED THIS? *Read Section 20.4; Watch KCV 20.4*

a. Al **b.** Ag **c.** Pb

58. Determine whether or not each metal dissolves in 1 M HCl. For those metals that do dissolve, write a balanced redox reaction showing what happens when the metal dissolves.

a. Cu **b.** Fe **c.** Au

59. Determine whether or not each metal dissolves in 1 M HNO_3. For those metals that do dissolve, write a balanced redox reaction showing what happens when the metal dissolves.
MISSED THIS? *Read Section 20.4; Watch KCV 20.4*

a. Cu
b. Au

60. Determine whether or not each metal dissolves in 1 M HIO_3. For those metals that do dissolve, write a balanced redox equation for the reaction that occurs.

a. Au

b. Cr

61. Calculate E°_{cell} for each balanced redox reaction and determine if the reaction is spontaneous as written.

MISSED THIS? *Read Section 20.4; Watch KCV 20.4, IWE 20.4*

a. $2\,Cu(s) + Mn^{2+}(aq) \longrightarrow 2\,Cu^{+}(aq) + Mn(s)$

b. $MnO_2(aq) + 4\,H^{+}(aq) + Zn(s) \longrightarrow Mn^{2+}(aq) + 2\,H_2O(l) + Zn^{2+}(aq)$

c. $Cl_2(g) + 2\,F^{-}(aq) \longrightarrow F_2(g) + 2\,Cl^{-}(aq)$

62. Calculate E°_{cell} for each balanced redox reaction and determine if the reaction is spontaneous as written.

a. $O_2(g) + 2\,H_2O(l) + 4\,Ag(s) \longrightarrow 4\,OH^{-}(aq) + 4\,Ag^{+}(aq)$

b. $Br_2(l) + 2\,I^{-}(aq) \longrightarrow 2\,Br^{-}(aq) + I_2(s)$

c. $PbO_2(s) + 4\,H^{+}(aq) + Sn(s) \longrightarrow Pb^{2+}(aq) + 2\,H_2O(l) + Sn^{2+}(aq)$

63. Which metal cation is the best oxidizing agent?

MISSED THIS? *Read Section 20.4; Watch KCV 20.4*

a. Pb^{2+} **b.** Cr^{3+} **c.** Fe^{2+} **d.** Sn^{2+}

64. Which metal is the best reducing agent?

a. Mn **b.** Al **c.** Ni **d.** Cr

Cell Potential, Free Energy, and the Equilibrium Constant

65. Use tabulated electrode potentials to calculate ΔG°_{rxn} for each reaction at 25 °C.

MISSED THIS? *Read Section 20.5; Watch KCV 20.5, IWE 20.6*

a. $Pb^{2+}(aq) + Mg(s) \longrightarrow Pb(s) + Mg^{2+}(aq)$

b. $Br_2(l) + 2\,Cl^{-}(aq) \longrightarrow 2\,Br^{-}(aq) + Cl_2(g)$

c. $MnO_2(s) + 4\,H^{+}(aq) + Cu(s) \longrightarrow Mn^{2+}(aq) + 2\,H_2O(l) + Cu^{2+}(aq)$

66. Use tabulated electrode potentials to calculate ΔG°_{rxn} for each reaction at 25 °C.

a. $2\,Fe^{3+}(aq) + 3\,Sn(s) \longrightarrow 2\,Fe(s) + 3\,Sn^{2+}(aq)$

b. $O_2(g) + 2\,H_2O(l) + 2\,Cu(s) \longrightarrow 4\,OH^{-}(aq) + 2\,Cu^{2+}(aq)$

c. $Br_2(l) + 2\,I^{-}(aq) \longrightarrow 2\,Br^{-}(aq) + I_2(s)$

67. Calculate the equilibrium constant for each of the reactions in Problem 65. **MISSED THIS?** *Read Section 20.5; Watch KCV 20.5*

68. Calculate the equilibrium constant for each of the reactions in Problem 66.

69. Calculate the equilibrium constant for the reaction between $Ni^{2+}(aq)$ and $Cd(s)$ (at 25 °C).

MISSED THIS? *Read Section 20.5; Watch KCV 20.5*

70. Calculate the equilibrium constant for the reaction between $Fe^{2+}(aq)$ and $Zn(s)$ (at 25 °C).

71. Calculate ΔG°_{rxn} and E°_{cell} for a redox reaction with $n = 2$ that has an equilibrium constant of $K = 25$ (at 25 °C).

MISSED THIS? *Read Section 20.5; Watch KCV 20.5*

72. Calculate ΔG°_{rxn} and E°_{cell} for a redox reaction with $n = 3$ that has an equilibrium constant of $K = 0.050$ (at 25 °C).

Nonstandard Conditions and the Nernst Equation

73. A voltaic cell employs the following redox reaction:

MISSED THIS? *Read Section 20.6*

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

Calculate the cell potential at 25 °C under each set of conditions.

a. standard conditions

b. $[Sn^{2+}] = 0.0100\ M$; $[Mn^{2+}] = 2.00\ M$

c. $[Sn^{2+}] = 2.00\ M$; $[Mn^{2+}] = 0.0100\ M$

74. A voltaic cell employs the redox reaction:

$$2\,Fe^{3+}(aq) + 3\,Mg(s) \longrightarrow 2\,Fe(s) + 3\,Mg^{2+}(aq)$$

Calculate the cell potential at 25 °C under each set of conditions.

a. standard conditions

b. $[Fe^{3+}] = 1.0 \times 10^{-3}\ M$; $[Mg^{2+}] = 2.50\ M$

c. $[Fe^{3+}] = 2.00\ M$; $[Mg^{2+}] = 1.5 \times 10^{-3}\ M$

75. An electrochemical cell is based on these two half-reactions:

MISSED THIS? *Read Section 20.6*

Ox: $Pb(s) \longrightarrow Pb^{2+}(aq, 0.10\ M) + 2\,e^{-}$

Red: $MnO_4^{-}(aq, 1.50\ M) + 4\,H^{+}(aq, 2.0\ M) + 3\,e^{-} \longrightarrow MnO_2(s) + 2\,H_2O(l)$

Calculate the cell potential at 25 °C.

76. An electrochemical cell is based on these two half-reactions:

Ox: $Sn(s) \longrightarrow Sn^{2+}(aq, 2.00\ M) + 2\,e^{-}$

Red: $ClO_2(g, 0.100\ atm) + e^{-} \longrightarrow ClO_2^{-}(aq, 2.00\ M)$

Calculate the cell potential at 25 °C.

77. A voltaic cell consists of a Zn/Zn^{2+} half-cell and a Ni/Ni^{2+} half-cell at 25 °C. The initial concentrations of Ni^{2+} and Zn^{2+} are 1.50 M and 0.100 M, respectively.

MISSED THIS? *Read Section 20.6*

a. What is the initial cell potential?

b. What is the cell potential when the concentration of Ni^{2+} has fallen to 0.500 M?

c. What are the concentrations of Ni^{2+} and Zn^{2+} when the cell potential falls to 0.45 V?

78. A voltaic cell consists of a Pb/Pb^{2+} half-cell and a Cu/Cu^{2+} half-cell at 25 °C. The initial concentrations of Pb^{2+} and Cu^{2+} are 0.0500 M and 1.50 M, respectively.

a. What is the initial cell potential?

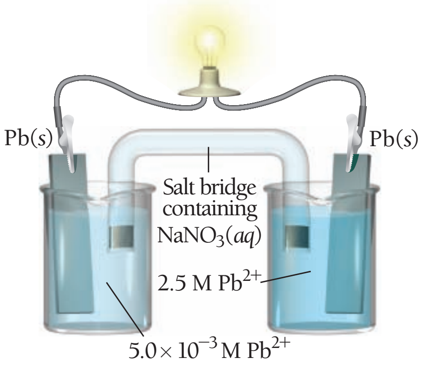
b. What is the cell potential when the concentration of Cu^{2+} has fallen to 0.200 M?

c. What are the concentrations of Pb^{2+} and Cu^{2+} when the cell potential falls to 0.35 V?

79. Make a sketch of a concentration cell employing two Zn/Zn^{2+} half-cells. The concentration of Zn^{2+} in one of the half-cells is 2.0 M, and the concentration in the other half-cell is 1.0×10^{-3} M. Label the anode and the cathode and indicate the half-reaction occurring at each electrode. Also indicate the direction of electron flow.

MISSED THIS? *Read Section 20.6*

80. Consider the concentration cell:



a. Label the anode and cathode.

b. Indicate the direction of electron flow.

c. Indicate what happens to the concentration of Pb^{2+} in each half-cell.

81. A concentration cell consists of two Sn/Sn^{2+} half-cells. The cell has a potential of 0.10 V at 25 °C. What is the ratio of the Sn^{2+} concentrations in the two half-cells?

MISSED THIS? *Read Section 20.6*

82. A Cu/Cu^{2+} concentration cell has a voltage of 0.22 V at 25 °C. The concentration of Cu^{2+} in one of the half-cells is 1.5×10^{-3} M. What is the concentration of Cu^{2+} in the other half-cell? (Assume the concentration in the unknown cell is the *lower* of the two concentrations.)

Batteries, Fuel Cells, and Corrosion

83. Determine the optimum mass ratio of Zn to MnO_2 in an alkaline battery. **MISSED THIS?** *Read Section 20.7*

84. What mass of lead sulfate is formed in a lead–acid storage battery when 1.00 g of Pb undergoes oxidation?

85. Refer to the tabulated values of ΔG°_f in Appendix IIB to calculate E°_{cell} for a fuel cell that employs the reaction between methane gas (CH_4) and oxygen to form carbon dioxide and gaseous water. **MISSED THIS?** *Read Sections 20.5, 20.7; Watch IWE 20.6*

86. Refer to the tabulated values of ΔG°_f in Appendix IIB to calculate E°_{cell} for the fuel-cell breathalyzer, which employs the following reaction. (ΔG°_f for $HC_2H_3O_2(g) = -374.2$ kJ/mol.)

$$CH_3CH_2OH(g) + O_2(g) \longrightarrow HC_2H_3O_2(g) + H_2O(g)$$

87. Determine whether or not each metal, if coated onto iron, would prevent the corrosion of iron.
MISSED THIS? *Read Section 20.9*

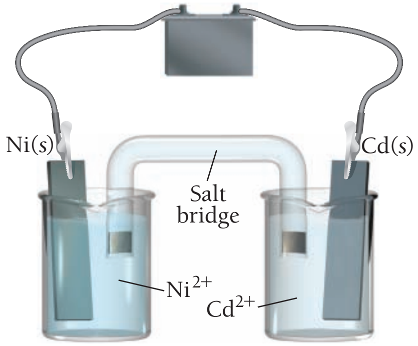
a. Zn **b.** Sn **c.** Mn

88. Determine whether or not each metal, if coated onto iron, would prevent the corrosion of iron.

a. Mg **b.** Cr **c.** Cu

Electrolytic Cells and Electrolysis

89. Consider the electrolytic cell: **MISSED THIS?** *Read Section 20.8*



a. Label the anode and the cathode and indicate the half-reactions occurring at each.
b. Indicate the direction of electron flow.
c. Label the terminals on the battery as positive or negative and calculate the minimum voltage necessary to drive the reaction.

90. Draw an electrolytic cell in which Mn^{2+} is reduced to Mn and Sn is oxidized to Sn^{2+}. Label the anode and cathode, indicate the direction of electron flow, and write an equation for the half-reaction occurring at each electrode. What minimum voltage is necessary to drive the reaction?

91. Write equations for the half-reactions that occur in the electrolysis of molten potassium bromide. **MISSED THIS?** *Read Section 20.8*

92. What products are obtained in the electrolysis of molten NaI?

93. Write equations for the half-reactions that occur in the electrolysis of a mixture of molten potassium bromide and molten lithium bromide. **MISSED THIS?** *Read Section 20.8*

94. What products are obtained in the electrolysis of a molten mixture of KI and KBr?

95. Write equations for the half-reactions that occur at the anode and cathode for the electrolysis of each aqueous solution.
MISSED THIS? *Read Section 20.8*

a. NaBr(*aq*) **b.** $PbI_2(aq)$ **c.** $Na_2SO_4(aq)$

96. Write equations for the half-reactions that occur at the anode and cathode for the electrolysis of each aqueous solution.

a. $Ni(NO_3)_2(aq)$ **b.** KCl(*aq*) **c.** $CuBr_2(aq)$

97. Make a sketch of an electrolysis cell that electroplates copper onto other metal surfaces. Label the anode and the cathode and indicate the reactions that occur at each.
MISSED THIS? *Read Section 20.8*

98. Make a sketch of an electrolysis cell that electroplates nickel onto other metal surfaces. Label the anode and the cathode and indicate the reactions that occur at each.

99. Copper can be electroplated at the cathode of an electrolysis cell by the half-reaction: **MISSED THIS?** *Read Section 20.8*

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

How much time would it take for 325 mg of copper to be plated at a current of 5.6 A?

100. Silver can be electroplated at the cathode of an electrolysis cell by the half-reaction:

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

What mass of silver would plate onto the cathode if a current of 6.8 A flowed through the cell for 72 min?

101. A major source of sodium metal is the electrolysis of molten sodium chloride. What magnitude of current produces 1.0 kg of sodium metal in 1 hour? **MISSED THIS?** *Read Section 20.8*

102. What mass of aluminum metal can be produced per hour in the electrolysis of a molten aluminum salt by a current of 25 A?

103. Consider the reaction shown here occurring at 25 °C.
MISSED THIS? *Read Section 20.5, Watch KCV 20.5, IWE 20.6*

$$A(s) + B^{2+}(aq) \longrightarrow A^{2+}(aq) + B(s) \quad \Delta G^\circ_{rxn} = -14.0 \text{ kJ}$$

Determine the value of E°_{cell} and *K* for the reaction and complete the table.

$[B^{2+}]$	$[A^{2+}]$	Q	E_{cell}	ΔG_{rxn}
1.00	1.00			
1.00	1.00×10^{-4}			
1.0×10^{-4}	1.0			
3.54×10^{-3}	1.0			

104. Consider the reaction shown here occurring at 25 °C.

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

Determine E°_{cell}, *K*, and ΔG°_{rxn} for the reaction and complete the table.

$[Cd^{2+}]$	$[Cr^{2+}]$	Q	E_{cell}	ΔG_{rxn}
1.00	1.00			
1.00	1.00×10^{-5}			
1.00×10^{-5}	1.00			
4.18×10^{-4}	1.00			

CUMULATIVE PROBLEMS

105. Consider the unbalanced redox reaction:

$$MnO_4^-(aq) + Zn(s) \longrightarrow Mn^{2+}(aq) + Zn^{2+}(aq)$$

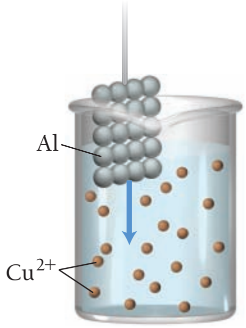
Balance the equation and determine the volume of a 0.500 M $KMnO_4$ solution required to completely react with 2.85 g of Zn.

106. Consider the unbalanced redox reaction:

$$Cr_2O_7^{2-}(aq) + Cu(s) \longrightarrow Cr^{3+}(aq) + Cu^{2+}(aq)$$

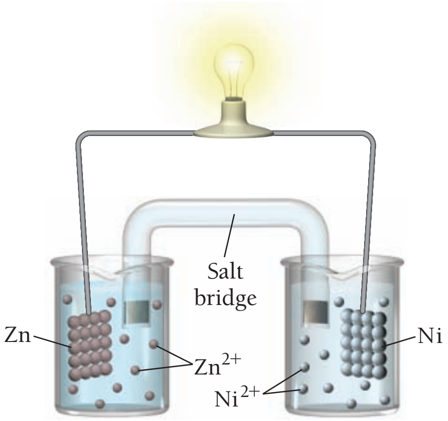
Balance the equation and determine the volume of a 0.850 M $K_2Cr_2O_7$ solution required to completely react with 5.25 g of Cu.

107. Consider the molecular views of an Al strip and Cu^{2+} solution. Draw a similar sketch showing what happens to the atoms and ions after the Al strip is submerged in the solution for a few minutes.



108. Consider the molecular view of an electrochemical cell involving the overall reaction:

$$Zn(s) + Ni^{2+}(aq) \longrightarrow Zn^{2+}(aq) + Ni(s)$$



Draw a similar sketch of the cell after it has generated a substantial amount of electrical current.

109. Determine whether HI can dissolve each metal sample. If it can, write a balanced chemical reaction showing how the metal dissolves in HI and determine the minimum volume of 3.5 M HI required to completely dissolve the sample.

a. 2.15 g Al **b.** 4.85 g Cu **c.** 2.42 g Ag

110. Determine if HNO_3 can dissolve each metal sample. If it can, write a balanced chemical reaction showing how the metal dissolves in HNO_3 and determine the minimum volume of 6.0 M HNO_3 required to completely dissolve the sample.

a. 5.90 g Au **b.** 2.55 g Cu **c.** 4.83 g Sn

111. The cell potential of this electrochemical cell depends on the pH of the solution in the anode half-cell.

$$Pt(s)\,|\,H_2(g, 1\text{ atm})\,|\,H^+(aq, ?\text{ M})\,\|\,Cu^{2+}(aq, 1.0\text{ M})\,|\,Cu(s)$$

What is the pH of the solution if E_{cell} is 355 mV?

112. The cell potential of this electrochemical cell depends on the gold concentration in the cathode half-cell.

$$Pt(s)\,|\,H_2(g, 1.0\text{ atm})\,|\,H^+(aq, 1.0\text{ M})\,\|\,Au^{3+}(aq, ?\text{ M})\,|\,Au(s)$$

What is the concentration of Au^{3+} in the solution if E_{cell} is 1.22 V?

113. A friend wants you to invest in a new battery she has designed that produces 24 V in a single voltaic cell. Why should you be wary of investing in such a battery?

114. What voltage can theoretically be achieved in a battery in which lithium metal is oxidized and fluorine gas is reduced? Why might such a battery be difficult to produce?

115. A battery relies on the oxidation of magnesium and the reduction of Cu^{2+}. The initial concentrations of Mg^{2+} and Cu^{2+} are 1.0×10^{-4} M and 1.5 M, respectively, in 1.0-liter half-cells.

a. What is the initial voltage of the battery?

b. What is the voltage of the battery after delivering 5.0 A for 8.0 h?

c. How long can the battery deliver 5.0 A before going dead?

116. A rechargeable battery is constructed based on a concentration cell constructed of two Ag/Ag^+ half-cells. The volume of each half-cell is 2.0 L, and the concentrations of Ag^+ in the half-cells are 1.25 M and 1.0×10^{-3} M.

a. How long can this battery deliver 2.5 A of current before it goes dead?

b. What mass of silver is plated onto the cathode by running at 3.5 A for 5.5 h?

c. Upon recharging, how long would it take to redissolve 1.00×10^2 g of silver at a charging current of 10.0 amps?

117. If a water electrolysis cell operates at a current of 7.8 A, how long will it take to generate 25.0 L of hydrogen gas at a pressure of 25.0 atm and a temperature of 25 °C?

118. When a suspected drunk driver blows 188 mL of his breath through the fuel-cell breathalyzer described in Section 20.7, the breathalyzer produces an average of 324 mA of current for 10 s. Assuming a pressure of 1.0 atm and a temperature of 25 °C, what percent (by volume) of the driver's breath is ethanol?

119. The K_{sp} of CuI is 1.1×10^{-12}. Find E_{cell} for the cell:

$$Cu(s)\,|\,CuI(s)\,|\,I^-(aq)(1.0\text{ M})\,\|\,Cu^+(aq)(1.0\text{ M})\,|\,Cu(s)$$

120. The K_{sp} of $Zn(OH)_2$ is 1.8×10^{-14}. Find E_{cell} for the half-reaction:

$$Zn(OH)_2(s) + 2\,e^- \rightleftharpoons Zn(s) + 2\,OH^-(aq)$$

121. Calculate ΔG°_{rxn} and K for each reaction.

a. The disproportionation of $Mn^{2+}(aq)$ to $Mn(s)$ and $MnO_2(s)$ in acid solution at 25 °C.

b. The disproportionation of $MnO_2(s)$ to $Mn^{2+}(aq)$ and $MnO_4^-(aq)$ in acid solution at 25 °C.

122. Calculate ΔG°_{rxn} and K for each reaction.

a. The reaction of $Cr^{2+}(aq)$ with $Cr_2O_7^{2-}(aq)$ in acid solution to form $Cr^{3+}(aq)$.

b. The reaction of $Cr^{3+}(aq)$ and $Cr(s)$ to form $Cr^{2+}(aq)$. [The electrode potential of $Cr^{2+}(aq)$ to $Cr(s)$ is -0.91 V.]

123. The molar mass of a metal (M) is 50.9 g/mol; it forms a chloride of unknown composition. Electrolysis of a sample of the molten chloride with a current of 6.42 A for 23.6 minutes produces 1.20 g of M at the cathode. Determine the empirical formula of the chloride.

124. A metal forms the fluoride MF_3. Electrolysis of the molten fluoride by a current of 3.86 A for 16.2 minutes deposits 1.25 g of the metal. Calculate the molar mass of the metal.

125. A sample of impure tin of mass 0.535 g is dissolved in strong acid to give a solution of Sn^{2+}. The solution is then titrated with a 0.0448 M solution of NO_3^-, which is reduced to $NO(g)$. The equivalence point is reached upon the addition of 0.0344 L of the NO_3^- solution. Find the percent by mass of tin in the original sample, assuming that it contains no other reducing agents.

126. A 0.0251-L sample of a solution of Cu^+ requires 0.0322 L of 0.129 M $KMnO_4$ solution to reach the equivalence point. The products of the reaction are Cu^{2+} and Mn^{2+}. What is the concentration of the Cu^{2+} solution?

127. A current of 11.3 A is applied to 1.25 L of a solution of 0.552 M HBr converting some of the H^+ to $H_2(g)$, which bubbles out of solution. What is the pH of the solution after 73 minutes?

128. A 215-mL sample of a 0.500 M NaCl solution with an initial pH of 7.00 is subjected to electrolysis. After 15.0 minutes, a 10.0-mL portion (or aliquot) of the solution was removed from the cell and titrated with 0.100 M HCl solution. The endpoint in the titration was reached upon addition of 22.8 mL of HCl. Assuming constant current, what was the current (in A) running through the cell?

129. An $MnO_2(s)/Mn^{2+}(aq)$ electrode in which the pH is 10.24 is prepared. Find the $[Mn^{2+}]$ necessary to lower the potential of the half-cell to 0.00 V (at 25 °C).

130. To what pH should you adjust a standard hydrogen electrode to get an electrode potential of -0.122 V? (Assume that the partial pressure of hydrogen gas remains at 1 atm.)

CHALLENGE PROBLEMS

131. Suppose a hydrogen–oxygen fuel-cell generator produces electricity for a house. Use the balanced redox reactions and the standard cell potential to predict the volume of hydrogen gas (at STP) required each month to generate the electricity. Assume the home uses 1.2×10^3 kWh of electricity per month.

132. A voltaic cell designed to measure $[Cu^{2+}]$ is constructed of a standard hydrogen electrode and a copper metal electrode in the Cu^{2+} solution of interest. If you want to construct a calibration curve for how the cell potential varies with the concentration of copper(II), what do you plot in order to obtain a straight line? What is the slope of the line?

133. The surface area of an object to be gold plated is 49.8 cm^2, and the density of gold is 19.3 g/cm^3. A current of 3.25 A is applied to a solution that contains gold in the +3 oxidation state. Calculate the time required to deposit an even layer of gold 1.00×10^{-3} cm thick on the object.

134. To electrodeposit all the Cu and Cd from a solution of $CuSO_4$ and $CdSO_4$ required 1.20 F of electricity (1 F = 1 mol e^-). The mixture of Cu and Cd that was deposited had a mass of 50.36 g. What mass of $CuSO_4$ was present in the original mixture?

135. Sodium oxalate, $Na_2C_2O_4$, in solution is oxidized to $CO_2(g)$ by MnO_4^-, which is reduced to Mn^{2+}. A 50.1-mL volume of a solution of MnO_4^- is required to titrate a 0.339-g sample of sodium oxalate. This solution of MnO_4^- is used to analyze uranium-containing samples. A 4.62-g sample of a uranium-containing material requires 32.5 mL of the solution for titration. The oxidation of the uranium can be represented by the change $UO^{2+} \longrightarrow UO_2^{2+}$. Calculate the percentage of uranium in the sample.

136. Three electrolytic cells are connected in a series. The electrolytes in the cells are aqueous copper(II) sulfate, gold(III) sulfate, and silver nitrate. A current of 2.33 A is applied, and after some time 1.74 g Cu is deposited. How long was the current applied? What mass of gold and silver was deposited?

137. The cell $Pt(s)\,|\,Cu^+(1\,M), Cu^{2+}(1\,M)\,\|\,Cu^+(1\,M)\,|\,Cu(s)$ has $E° = 0.364$ V. The cell $Cu(s)\,|\,Cu^{2+}(1\,M)\,\|\,Cu^+(1\,M)\,|\,Cu(s)$ has $E° = 0.182$ V. Write the cell reaction for each cell and explain the differences in $E°$. Calculate $\Delta G°$ for each cell reaction to help explain these differences.

CONCEPTUAL PROBLEMS

138. An electrochemical cell has a positive standard cell potential but a negative cell potential. Which statement is true for the cell?

a. $K > 1; Q > K$ **b.** $K < 1; Q > K$
c. $K > 1; Q < K$ **d.** $K < 1; Q < K$

139. Which oxidizing agent will oxidize Br^- but not Cl^-?

a. $K_2Cr_2O_7$ (in acid) **b.** $KMnO_4$ (in acid)
c. HNO_3

140. A redox reaction employed in an electrochemical cell has a negative $\Delta G°_{rxn}$. Which statement is true?

a. $E°_{cell}$ is positive; $K < 1$
b. $E°_{cell}$ is positive; $K > 1$
c. $E°_{cell}$ is negative; $K > 1$
d. $E°_{cell}$ is negative; $K < 1$

141. A redox reaction has an equilibrium constant of $K = 0.055$. What is true of $\Delta G°_{rxn}$ and $E°_{cell}$ for this reaction?

QUESTIONS FOR GROUP WORK

Active Classroom Learning

Discuss these questions with the group and record your consensus answer.

142. Balance the redox reactions by following the steps in the text. Rotate through the group, having each group member do the next step in the process and explain that step to the rest of the group.

a. $I_2(s) + Fe(s) \longrightarrow FeI_2(s)$
b. $Cl_2(g) + H_2O_2(aq) \longrightarrow Cl^-(aq) + O_2(g)$ (acidic)
c. $Hg^{2+}(aq) + H_2(g) \longrightarrow Hg(l) + H_2O(l)$ (basic)
d. $CH_3OH(l) + O_2(g) \longrightarrow CO_2(g) + H_2O(l)$ (acidic)

143. Have each group member select a half-reaction from Table 20.1. Each member should calculate the standard cell potential of an electrochemical cell formed between each member's half-reaction and the half-reaction of each of the other group members. For each pair of half-reactions, write the overall balanced chemical reaction that will be spontaneous.

144. Calculate $\Delta G°$ and K for each reaction the group created in Question 143. For one of the reactions, explain how the sign or magnitude of each quantity ($E°_{cell}$, $\Delta G°$, and K) is consistent with the fact that the reaction is spontaneous in the direction written.

145. Design a device that uses an electrochemical cell to determine the amount of Cu^{2+} in a sample of water. Describe, in detail, the construction and the theory of operation of your device. If you are able to measure voltage with one-millivolt accuracy, what will the uncertainty in your measured concentration be?

146. Using a library or the Internet, research a fuel cell that uses methanol for fuel. What is the reaction at the anode? What is the reaction at the cathode? What is the overall reaction? What is the standard cell potential? How many kWh can it generate from 1 L (0.792 kg) of methanol?

DATA INTERPRETATION AND ANALYSIS

Measuring Concentration with Voltage

147. In this chapter, you have seen that the voltage of an electrochemical cell is sensitive to the concentrations of the reactants and products in the cell. As a result, electrochemical cells can be used to measure the concentrations of certain species in solution. For example, the voltage of an electrochemical cell based on the reaction $H_2(g) + Cu^{2+}(aq) \longrightarrow 2\,H^+(aq) + Cu(s)$ is sensitive to both the Cu^{2+} concentration and the H^+ concentration in solution. If the H^+ concentration is held constant, then the voltage only depends on the Cu^{2+} concentration, and we can use the cell to measure the Cu^{2+} concentration in an unknown solution. The tabulated data shows the measured voltage in the hydrogen/copper electrochemical cell just discussed for several different Cu^{2+} concentrations. Examine the data and answer the questions that follow.

$[Cu^{2+}]$	Voltage (V)	$[Cu^{2+}]$	Voltage (V)
0.100	0.310	0.500	0.331
0.200	0.319	0.700	0.335
0.300	0.325	1.00	0.340
0.400	0.328		

a. Construct a graph of the measured voltage versus the copper concentration. Is the graph linear?

b. Determine how you might manipulate the data to produce a linear graph. (*Hint*: See the Nernst equation.)

c. Reconstruct a graph of the data using the method to produce a linear graph from part b. Determine the slope and y-intercept of the best-fitting line to the points in your graph. Could you have predicted the slope and intercept from the Nernst equation?

d. The voltage of two unknown solutions are measured and recorded. Use the slope and intercept from part c to determine the Cu^{2+} concentrations of the unknown solutions.

Unknown Cu^{2+} Solution	Voltage (V)
i	0.303
ii	0.338

ANSWERS TO CONCEPTUAL CONNECTIONS

Potential Difference and Electrical Current

20.1 **(a)** The volt is a unit of potential difference, the difference in potential energy and the amp is a unit of electrical current, the flow of electrical charge.

Voltaic Cells

20.2 **(a)** Electrons are negatively charged and therefore flow away from the more negatively charged electrode and toward the more positively charged electrode.

Standard Electrode Potentials

20.3 **(b)** A negative electrode potential indicates that an electron at that electrode has greater potential energy than it has at a standard hydrogen electrode.

Selective Oxidation

20.4 **(d)** The reduction of HNO_3 is below the reduction of Br_2 and above the reduction of I_2 in Table 20.1. Since any reduction half-reaction is spontaneous when paired with the reverse of a half-reaction below it in the table, the reduction of HNO_3 is spontaneous when paired with the oxidation of I^- but is not spontaneous when paired with the oxidation of Br^-.

Metals Dissolving in Acids

20.5 **(c)** Ag falls above the half-reaction for the reduction of H^+ but below the half-reaction for the reduction of NO_3^- in Table 20.1.

Periodic Trends and the Direction of Spontaneity for Redox Reactions

20.6 **(a)** Br is more electronegative than I. If the two atoms were in competition for the electron, the electron would go to the more electronegative atom (Br). Therefore, I_2 does not spontaneously gain electrons from Br^-.

Relating K, $\Delta G^\circ_{\text{rxn}}$, and E°_{cell}

20.7 **(c)** Since $K > 1$, the reaction is spontaneous under standard conditions (when $Q = 1$, the reaction proceeds toward the products). Therefore, E°_{cell} is positive and $\Delta G^\circ_{\text{rxn}}$ is negative.

Relating Q, K, E_{cell}, and E°_{cell}

20.8 **(a)** Since $K > 1$, E°_{cell} is negative (under standard conditions, the reaction is not spontaneous). Since $Q < K$, E_{cell} is positive (the reaction is spontaneous under the nonstandard conditions of the cell).

Voltaic and Electrolytic Cells

20.9 **(b)** Since oxidation always occurs at the anode, electrons flow away from the anode and toward the cathode (where reduction occurs).

Electrolysis

20.10 **(c)** The stoichiometry of the reaction shows that 1 mole of silver plates for every 1 mole of electrons; therefore 3 moles of electrons are required to plate 3 moles of silver.

Sacrificial Electrodes

20.11 **(a)** Cu. The electrode potential for Fe is more negative than that of Cu. Therefore, Fe oxidizes more easily than Cu.