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21 ELECTROCHEMISTRY

SECTION 21.1 ELECTROCHEMICAL CELLS (pages 663–670)

This section describes how redox reactions interconvert electrical energy and chemical energy. It also explains the structure of a dry cell and identifies the substances that are oxidized and reduced.

▶ Electrochemical Processes (pages 663–665)

1. What do the silver plating of tableware and the manufacture of aluminum have in common?

_____

Look at Figure 21.1 on page 663 and the related text to help you answer Questions 2–6.

2. In what form are the reactants when the reaction starts?

_____

3. What kind of reaction occurs? Is it spontaneous?

_____

4. Which substance is oxidized in the reaction? _____

5. Which substance is reduced? _____

6. Which atoms lose electrons and which ions gain electrons during the reaction?

_____

7. Look at Table 21.1 on page 664. What information in this table explains why the reaction in Figure 21.1 occurs spontaneously?

_____

_____

_____

_____

8. What happens when a copper strip is placed in a solution of zinc sulfate? Explain.

_____

_____

9. The flow of _____ from zinc to copper is an electric _____ .

10. Circle the letter of each sentence that is true about electrochemical cells.

a. An electrochemical cell either produces an electric current or uses an electric current to produce a chemical change.

b. Redox reactions occur in electrochemical cells.

c. For an electrochemical cell to be a source of useful electrical energy, the electrons must pass through an external circuit.

d. An electrochemical cell can convert chemical energy to electrical energy, but not electrical energy into chemical energy.

▶ Voltaic Cells (pages 665–667)

For Questions 11–15, match each description with the correct term by writing its letter in the blank.

_____ **11.** Any electrochemical cell used to convert chemical energy into electrical energy

_____ **12.** One part of a voltaic cell in which either reduction or oxidation occurs

_____ **13.** The electrode at which oxidation occurs

_____ **14.** A tube containing a strong electrolyte, which allows transport of ions between the half-cells

_____ **15.** The electrode at which reduction occurs

a. cathode

b. salt bridge

c. voltaic cell

d. half-cell

e. anode

▶ Using Voltaic Cells as Energy Sources (pages 667–670)

16. Look at Figure 21.4 on page 667. How is a common dry cell constructed?

_____

_____

_____

17. Why are alkaline cells better and longer lasting than common cells?

_____

_____

18. Which element is oxidized in a dry cell? Which element is reduced?

_____

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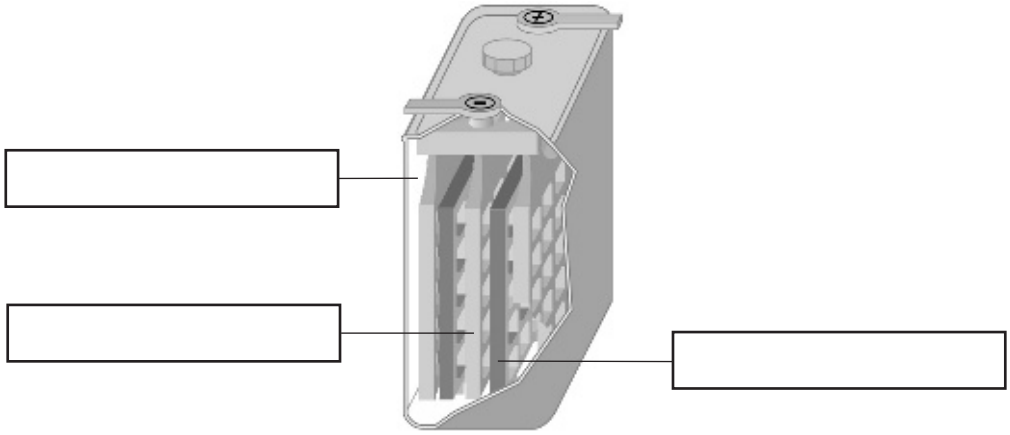
19. What is a battery? _____

20. How many voltaic cells are connected inside a lead storage battery typically found in a car? About how many volts are produced by each cell and what is the total voltage of such a battery?

_____

_____

21. Look at Figure 21.5 on page 668. In the diagram below label the following parts of a lead storage battery: electrolyte, anode, and cathode. Also indicate where oxidation and reduction occur while the battery is discharging.



22. Are the following sentences true or false? As a lead storage battery discharges, lead sulfate builds up on the electrodes. Recharging the battery reverses this process. _____

23. Name two advantages of fuel cells. _____

_____

Reading Skill Practice

Outlining is a way to help you understand and remember what you have read. Write an outline for Section 21.1 *Electrochemical Cells*. Begin your outline by copying the headings in the textbook. Under each heading, write the main idea. Then list details that support, or back up, the main idea. Do your work on a separate sheet of paper.

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SECTION 21.2 HALF-CELLS AND CELL POTENTIALS (pages 671–677)

This section defines standard cell potential and standard reduction potential. It also explains how to use standard reduction potential to calculate standard cell potential.

▶ Electrical Potential (page 671)

1. What unit is usually used to measure electrical potential?

_____

2. What is the equation for cell potential? _____

▶ Standard Cell Potential (page 672)

3. What value have chemists assigned as the standard reduction potential of the hydrogen electrode? _____

4. Describe a standard hydrogen electrode. _____

_____

_____

▶ Standard Reduction Potentials (pages 672–675)

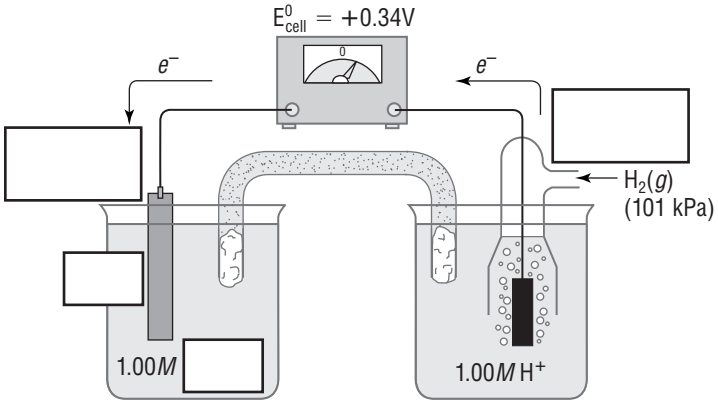
5. Use of a standard hydrogen electrode allows scientists to determine the _____ for many half-cells.

6. Look at Figure 21.10 on page 673. Which substance, zinc metal or hydrogen gas, has a greater potential to be oxidized? How can you tell?

_____

_____

7. In the diagram below use the value given for E^0_{cell} above the voltmeter and Table 21.2 to identify the chemical substances in the left half-cell. Use symbols to label the metal electrode and the ions in the half-cell. Also label the cathode and the anode.



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▶ Calculating Standard Cell Potentials (pages 675–677)

8. If the cell potential for a given redox reaction is _____ , then the reaction is _____ . If the cell potential is _____ , then the reaction is _____ .

SECTION 21.3 ELECTROLYTIC CELLS (pages 678–683)

This section differentiates electrolytic cells from voltaic cells, and lists uses of electrolytic cells. It also identifies the products of the electrolysis of brine, molten sodium chloride, and water.

▶ Electrolytic vs. Voltaic Cells (pages 678–679)

1. An electrochemical cell used to cause a chemical change through the application of electrical energy is called _____ .

2. For each sentence below, fill in *V* if it is true about voltaic cells, *E* if it is true about electrolytic cells, and *B* if it is true about both voltaic and electrolytic cells.

_____ **a.** Electrons are pushed by an outside power source.

_____ **b.** Reduction occurs at the cathode and oxidation occurs at the anode.

_____ **c.** The flow of electrons is the result of a spontaneous redox reaction.

_____ **d.** Electrons flow from the anode to the cathode.

▶ Electrolysis of Water (page 680)

3. Write the net reaction for the electrolysis of water.

_____

▶ Electrolysis of Brine (page 681)

4. Which three important industrial chemicals are produced through the electrolysis of brine?

_____

5. Why are the sodium ions not reduced to sodium metal during the electrolysis of brine?

_____

▶ Other Applications of Electrolytic Cells (page 682)

6. Deposition of a thin layer of metal on an object in an electrolytic cell is called _____ .

7. The object to be plated is made the _____ in the cell.

GUIDED PRACTICE PROBLEM

GUIDED PRACTICE PROBLEM 8 (page 676)

8. A voltaic cell is constructed using the following half-reactions.

$Cu^{2+}(aq) + 2e^- \longrightarrow Cu(s)$ $\quad E^0_{Cu^{2+}} = +0.34\ V$

$Al^{3+}(aq) + 3e^- \longrightarrow Al(s)$ $\quad E^0_{Al^{3+}} = -1.66\ V$

Determine the cell reaction and calculate the standard cell potential.

Analyze

Step 1. What are the known values?

_____

Step 2. Which half-reaction is a reduction? An oxidation?

Reduction: _____

Oxidation: _____

Step 3. Write both half-reactions in the direction they actually occur.

_____

Step 4. What is the expression for the standard cell potential?

E^0_{cell} = _____

Calculate

Step 5. Write the cell reaction by adding the half-reactions, making certain that the number of electrons lost equals the number of electrons gained. The electrons gained and lost will cancel out.

$\square\ [Cu^{2+}(aq) + 2e^- \longrightarrow Cu(s)]$

$\square\ [Al(s) \longrightarrow Al^{3+}(aq) + 3e^-]$

_____

$\square$

Step 6. Calculate the standard cell potential.

E^0_{cell} = _____

Evaluate

Step 7. How do you know that the cell reaction is correct?

_____

Step 8. When a reaction is spontaneous, will the standard cell potential be positive or negative?

_____