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21 Electrochemistry

Big idea MATTER AND ENERGY

21.1 Electrochemical Cells

For students using the Foundation edition, assign problems 1–7, 10–16, 19, 21–23.

Essential Understanding Electrochemical cells produce energy in electrochemical processes and have many practical uses.

Reading Strategy

Cause and Effect A cause and effect chart is a useful tool when you want to describe how, when, or why one event causes another. It provides a pictorial display in order to recognize an effect as something that happens and a cause as the reason why it happens.

As you read Lesson 21.1, use the cause and effect chart below. Fill in the chart with a cause of the listed effect and an effect of the listed cause.

Cause	Effect
1. Atoms lose electrons.	***Other atoms gain electrons.***
2. ***A lead storage battery discharges.***	The electrical energy needed to start a car is produced.
3. Zinc metal is added to a solution containing Fe^{2+} ions.	***Zn^{2+} ions and iron metal form.***
4. A salt bridge is added to a voltaic cell.	***Ions pass from one half-cell to another.***

EXTENSION **In the last two rows of the chart, write your own causes and effects from this lesson.**

Lesson Summary

Electrochemical Processes During an electrochemical process, a conversion occurs between chemical energy and electrical energy.

- Electrochemical processes involve redox reactions because electrons are transferred.
- The activity series of metals helps determine the result of chemical reactions.

Voltaic Cells Voltaic cells use chemical energy from a redox reaction to produce electrical energy.

- A voltaic cell consists of two half-cells connected by a salt bridge.
- Each half-cell contains an electrode and a solution that will react with the electrode.

- Electrons flow from the anode, or negative electrode, to the cathode, or positive electrode.
- Oxidation occurs at the anode; reduction occurs at the cathode.

Using Voltaic Cells as Energy Sources Dry cells, lead-storage batteries, and fuel cells use electrochemical processes to produce energy.

- A common flashlight battery is a dry cell that contains a zinc anode, a graphite cathode, and a paste that acts as an electrolyte.
- A lead-storage battery is composed of several voltaic cells, each of which contains lead and lead(IV) oxide electrodes and concentrated sulfuric acid.
- Fuel cells contain renewable electrodes and a fuel that is oxidized.

After reading Lesson 21.1, answer the following questions.

Electrochemical Processes

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

Both processes use electricity.

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

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

A solid strip of zinc is placed in an aqueous solution of copper sulfate.

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

a redox reaction; yes

4. Which substance is oxidized in the reaction? *zinc*

5. Which substance is reduced? *copper ions*

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

Zinc atoms lose electrons as copper ions gain electrons.

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

Zinc is higher in the table than copper. This means that zinc is more active and more easily oxidized than copper. Thus, the reaction between zinc atoms and copper ions is spontaneous.

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

Nothing happens. Copper metal is not oxidized by zinc ions.

9. The flow of *electrons* from zinc to copper is an electric *current*.

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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 into electrical energy, but not electrical energy into chemical energy.

Voltaic Cells

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

c	**11.** any electrochemical cell used to convert chemical energy into electrical energy	**a.** cathode
d	**12.** one part of a voltaic cell in which either reduction or oxidation occurs	**b.** salt bridge
e	**13.** the electrode at which oxidation occurs	**c.** voltaic cell
b	**14.** a tube containing a strong electrolyte, which allows transport of ions between the half-cells	**d.** half-cell
a	**15.** the electrode at which reduction occurs	**e.** anode

Using Voltaic Cells as Energy Sources

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

A common dry cell consists of a zinc container filled with an electrolyte paste of manganese(IV) oxide, zinc chloride, ammonium chloride, and water. The paste surrounds a central graphite rod.

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

The design eliminates the buildup of ammonia gas and maintains the zinc electrode, which corrodes more slowly under alkaline conditions.

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

zinc; manganese

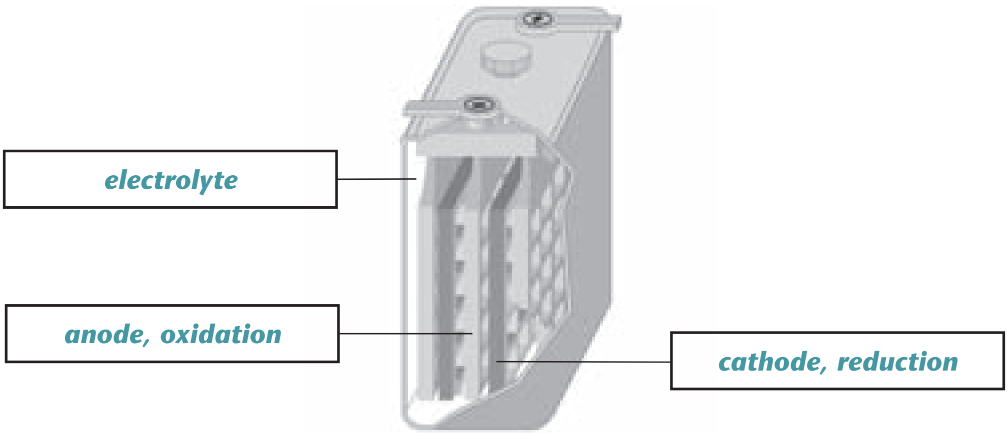
19. What is a battery?

A battery is a group of connected cells.

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?

A 12-V car battery consists of six voltaic cells. Each cell produces about 2 volts, making the total voltage 12 volts.

21. Look at Figure 21.5. Then, 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.

true

23. Name two advantages of fuel cells.

They do not need to be recharged, and they can be designed to emit no pollutants.

21.2 Half-Cells and Cell Potentials

For students using the Foundation edition, assign problems 1–3, 5–8.

Essential Understanding The electrical potential of each half-cell determines how much electrical energy is produced by a voltaic cell.

Lesson Summary

Electrical Potential A cell's electrical potential is a measure of its ability to produce an electric current.

- The reduction potential of a cell describes the likelihood that its half-reaction will be a reduction reaction. The electrical potential of the entire cell is the difference in the reduction potential of its half-cells.
- Standard cell potential is measured at standard conditions.
- The standard hydrogen electrode is used with other electrodes to compare their electrical potential.

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Standard Reduction Potentials A half-cell's standard reduction potential can be found by using a standard hydrogen electrode and applying the equation for standard cell potential.

- The half-cell with the higher tendency to be reduced has a positive standard reduction potential.
- The half-cell with the lower tendency to be reduced has a negative standard reduction potential.

Calculating Standard Cell Potentials Known standard reduction potentials can be used to predict where oxidation and reduction occur and to calculate standard cell potentials.

- The standard cell potential for a spontaneous redox reaction is positive; that for a nonspontaneous redox reaction is negative.
- Once half-cell reactions are written, the standard cell potential can be calculated by subtracting the standard cell potential for the oxidation reaction from the standard cell potential for the reduction reaction.

After reading Lesson 21.2, answer the following questions.

Electrical Potential

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

Electric potential is usually measured in volts.

2. What is the equation for cell potential?

$E^0_{cell} = E^0_{red} - E^0_{oxid}$

3. What value have chemists assigned as the standard reduction potential of the hydrogen electrode? *0.00V*

4. Describe a standard hydrogen electrode.

A standard hydrogen electrode consists of a platinum electrode coated with platinum black, immersed in a 1.0M hydrogen-ion solution with hydrogen gas bubbled into the solution.

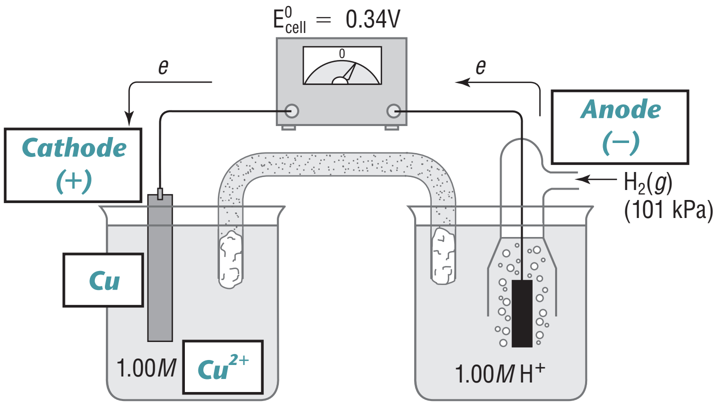
Standard Reduction Potentials

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

6. Look at Figure 21.9. Which substance, zinc metal or hydrogen gas, has a greater potential to be oxidized? How can you tell?

Zinc metal has a greater potential to be oxidized. E^0_{cell} is positive, showing that current is flowing away from zinc as it is oxidized.

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.



Calculating Standard Cell Potentials

8. If the cell potential for a given redox reaction is ***positive***__________, then the reaction is ***spontaneous***__________. If the cell potential is ***negative***__________, then the reaction is ***nonspontaneous***__________.

21.3 Electrolytic Cells

For students using the Foundation edition, assign problems 1–3, 6, 7.

Essential Understanding Electrolytic cells use the process of electrolysis to bring about chemical reactions.

Lesson Summary

Electrolytic vs. Voltaic Cells In contrast to voltaic cells, which use a chemical reaction to produce an electric current, electrolytic cells use electrical energy to produce a chemical change.

- The process that uses electric energy to produce chemical change is called electrolysis and is carried out in an electrolytic cell.
- In a voltaic cell, electrons flow because of a spontaneous chemical reaction; in an electrolytic cell, electrons flow because of an outside power source.

Driving Nonspontaneous Processes Nonspontaneous redox reactions sometimes occur because of electrolysis.

- Passing a current through a solution or a molten ionic compound can break the compound into its component elements.
- Electrolysis is also used in plating, purifying, and refining metals.

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After reading Lesson 21.3, answer the following questions.

Electrolytic vs. Voltaic Cells

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

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.

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

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

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

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

Driving Nonspontaneous Processes

3. Write the net reaction for the electrolysis of water.
$2H_2O(l) \rightarrow 2H_2(g) + O_2(g)$

4. Which three important industrial chemicals are produced through the electrolysis of brine?
chlorine gas, hydrogen gas, and sodium hydroxide

5. Why are the sodium ions not reduced to sodium metal during the electrolysis of brine?
Water molecules are more easily reduced than sodium ions.

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

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

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Guided Practice Problems

Use this information to answer questions about Practice Problem 10.

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

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

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

Determine the cell reaction.

Analyze

Step 1. Which half-reaction is a reduction? An oxidation?

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

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

Step 2. Write both half-reactions in the direction they actually occur.

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

Calculate

Step 3. 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.

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

$\boxed{2}\ Al(s) \rightarrow Al^{3+}(aq) + 3e^-$

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

Evaluate

Step 4. How do you know that the cell reaction is correct?

The number of atoms and ions is the same on both sides of the reaction.

Use this information to answer questions about Practice Problem 12.

Calculate the standard cell potential of a voltaic cell constructed using the half-reactions described in Problem 10.

Analyze

Step 1. What are the known values?

$E^0_{Cu^{2+}} = +0.34\ V$ and $E^0_{Al^{3+}} = -1.66\ V$

Step 2. What is the expression for the standard cell potential?

$E^0_{cell} =$ ***$E^0_{red} - E^0_{oxid}$***

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Calculate

Step 3. Calculate the standard cell potential.

E^0_{cell} = ***+0.34 V − (−1.66 V) = + 2.00 V***

Evaluate

Step 4. When a reaction is spontaneous, will the standard cell potential be positive or negative?

In a spontaneous reaction, the standard cell potential must be positive.

Apply the Big idea

Zinc is sometimes used to coat iron objects to keep them from rusting. The most common way of doing this is called galvanizing. When iron items are galvanized, they are dipped in molten zinc, forming a thin coat of zinc over the iron.

a. Would a thin coat of zinc form on an iron object placed in a solution containing zinc ions? Explain your answer.

According to Table 21.1, zinc is more active than iron. Thus, no reaction would occur. The zinc ions would remain in solution and not react with the iron.

b. Could zinc be electroplated on an iron object? Explain.

Yes. The iron object is placed in a solution containing zinc ions. The iron object acts as the cathode. A piece of zinc is placed in the solution and acts as the anode. When a current is applied, zinc ions are reduced to zinc metal and are deposited on the iron object.

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21 Self-Check Activity

For Questions 1–9, complete each statement by writing the correct word or words. If you need help, you can go online.

21.1 Electrochemical Cells

1. All electrochemical processes, such as those in a battery, use ________________ reactions.

2. In a voltaic cell, electrical energy is produced by a(n) ________________ chemical reaction.

3. Dry cells, lead storage batteries, and ________________ are all examples of applications that use electrochemical processes.

21.2 Half-Cells and Cell Potentials

4. The ________________________ of a cell is determined by the relative attraction of each half-cell for electrons.

5. The standard reduction potential of a half-cell can be determined by using a(n) ________________________ electrode and the equation for ________________________.

6. A(n) ________________ cell potential for a given redox reaction means that the reaction is spontaneous.

21.3 Electrolytic Cells

7. In a(n) ________________ cell, electron flow comes from a spontaneous chemical reaction.

8. In a(n) ________________ cell, electron flow comes from an outside power source.

9. The process of ________________ is commonly used to plate, purify, and refine metals.

If You Have Trouble With...									
Question	1	2	3	4	5	6	7	8	9
See Page	728	730	732	737	738	741	746	746	746

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Review Key Equations

Use the following equation and Table 21.2 to determine whether each of the listed reactions is spontaneous or not. Show your work. In each blank, write *S* if the reaction is spontaneous and *N* if the reaction is nonspontaneous.

$$E^0_{cell} = E^0_{red} - E^0_{oxid}$$

N **1.** $2Cr^{3+}(aq) + 3Ni(s) \rightarrow 2Cr(s) + 3Ni^{2+}(aq)$

Reduction: $Cr^{3+}(aq) + 3e^- \rightarrow Cr(s)$; $E^0_{Cr^{3+}} = -0.74$ V
Oxidation: $Ni(s) \rightarrow Ni^{2+}(aq) + 2e^-$; $E^0_{Ni^{2+}} = +0.25$ V
$E^0_{cell} = E^0_{Cr^{3+}} - E^0_{Ni^{2+}}$
$E^0_{cell} = -0.74\text{ V} - (+0.25\text{ V}) = -0.99\text{ V}$
The E^0_{cell} is negative, so the reaction is not spontaneous.

N **2.** $2K(s) + Zn^{2+}(aq) \rightarrow 2K^{+}(aq) + Zn(s)$

Reduction: $Zn^{2+}(aq) + 2e^- \rightarrow Zn(s)$; $E^0_{Zn^{2+}} = -0.76$ V
Oxidation: $K(s) \rightarrow K^{+}(aq) + e^-$; $E^0_{K^{+}} = +2.93$ V
$E^0_{cell} = E^0_{Zn^{2+}} - E^0_{K^{+}}$
$E^0_{cell} = -0.76\text{ V} - (+2.93\text{ V}) = -3.69\text{ V}$
The E^0_{cell} is negative, so the reaction is not spontaneous.

S **3.** $Cu^{2+}(aq) + Zn(s) \rightarrow Cu(s) + Zn^{2+}(aq)$

Reduction: $Cu^{2+}(aq) + 2e^- \rightarrow Cu(s)$; $E^0_{Cu^{2+}} = +0.34$ V
Oxidation: $Zn(s) \rightarrow Zn^{2+}(aq) + 2e^-$; $E^0_{Zn^{+}} = -0.76$ V
$E^0_{cell} = E^0_{Cu^{2+}} - E^0_{Zn^{+}}$
$E^0_{cell} = +0.34\text{ V} - (-0.76\text{ V}) = +1.10\text{ V}$
The E^0_{cell} is positive, so the reaction is spontaneous.

Review Vocabulary

For each set of vocabulary terms, write a sentence that explains how the listed terms are related. ***Sample answers provided.***

1. electrochemical cell, voltaic cell, half-cell, salt bridge

A voltaic cell is an electrochemical cell in which two half-cells are connected by a salt bridge.

2. dry cell, fuel cell, battery

A battery is a series of voltaic cells, such as dry cells and fuel cells.

3. electrode, anode, cathode, voltaic cell, electrolytic cell

An electrolytic cell, like a voltaic cell, contains two electrodes—the anode, where oxidation occurs, and the cathode, where reduction occurs.

4. reduction potential, standard cell potential, cell potential

The cell potential is the difference in the reduction potentials of the two half-cells, and this number is the standard cell potential if the concentration, pressure, and temperature are standard.