أوراق عمل الوحدة السادسة Electrochemistry منهج انسباير

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Chemistry Worksheets 12 ADV

Unit 6 Electrochemistry

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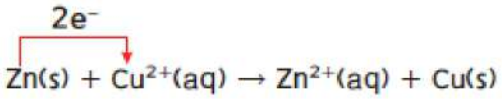
CHAPTER 6 / Electrochemistry

Section 1 : Votaic Cells

Redox in electrochemistry

- Electrochemistry : is the study of the redox processes by which chemical energy is converted to electrical energy and vise versa.
- study the following redox reaction and ansure the questions:



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

- Write the oxidation half reaction: --

- Write the reduction half reaction : ---

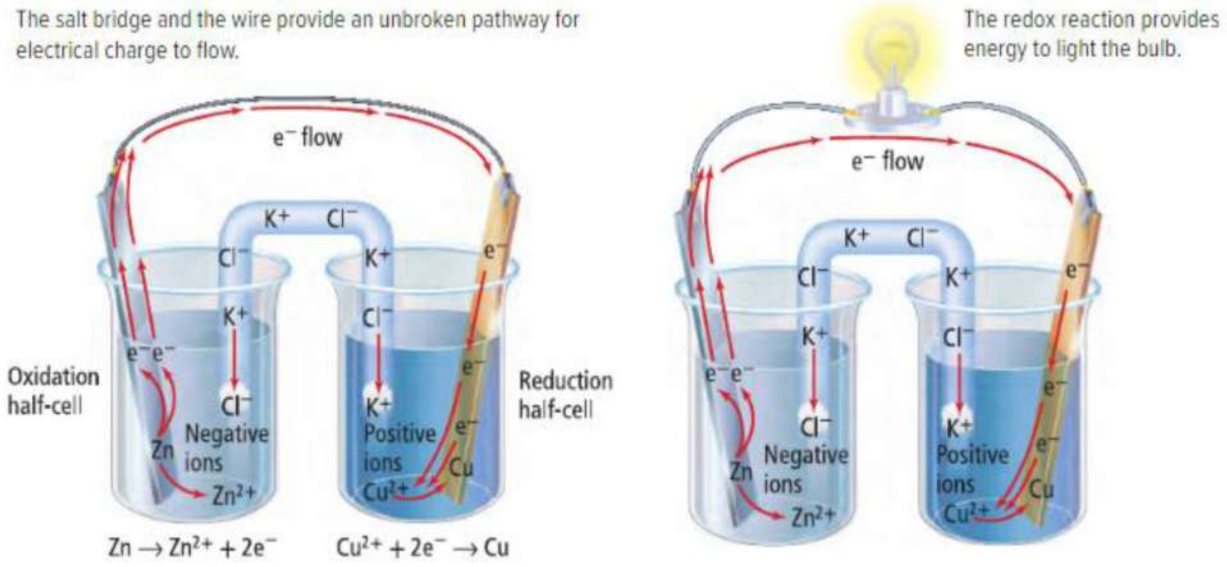
- What happened when we put a strip of zinc in $CuSO_4$ solution ?

- What is the type of energy which is produced from this reaction ?

- How can you generate electrical energy instead of heat energy ?

- What is a salt bridge? --

Electrochemical cells:



The salt bridge and the wire provide an unbroken pathway for electrical charge to flow.

The redox reaction provides energy to light the bulb.

Figure 2 The addition of a salt bridge completes the pathway. Negative ions move through the salt bridge to the zinc side. Positive ions move through the bridge to the copper side.

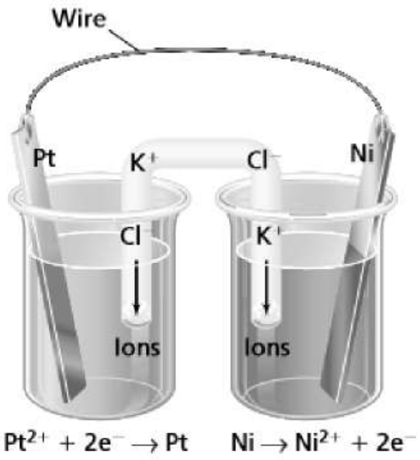
- **Electrochemical cell :** is an apparatus that uses a redox reaction to produce electrical energy or uses electrical energy to cause a chemical reaction **.**
- **Voltaic cell :** ....
  ....

Chemistry of voltaic cells :

- An electrochemical cell consists of tow parts called ....
- Each half – cell contains .... and ....
- **An electrode**: is an .... , usually a metallic strip or graphite.
- **The anode :** ....
- **The cathode** : ....

voltaic	electrochemical cell	electric current	salt bridge	galvanic

Oxidation and reduction reactions can occur in separate solutions, as long as there are two connections between the solutions. One connection is a(n) **(1)**____________________ through which ions can flow. The other connection is a metal wire through which electrons can flow. The flow of ions or electrons is known as a(n) **(2)**____________________. The complete setup, called a(n) **(3)**____________________, can convert chemical energy into electrical energy or electrical energy into chemical energy. These cells are also known as **(4)**____________________ cells or **(5)**____________________ cells.



Use the diagram of an electrochemical cell to answer the following questions.

6. The equation at the bottom of each beaker shows the half-reaction that is occurring in that beaker. What kind of reaction (oxidation or reduction) is occurring in each beaker?

Left beaker ____________________

Right beaker ____________________

7. Write the net ionic equation for this electrochemical cell.

____

8. In which direction do electrons move through the wire?

____

9. What kind of ions (positive or negative) move from the ∩-shaped tube into each beaker?

Left beaker ____________________ Right beaker ____________________

Voltaic cells and energy:

- Electric charge can flow between tow points when a difference in electric potential energy exists between the tow points.
- Electromotive force (EMF): is the force that pushes electrons generated at the anode toward the cathode.
- A volt : is a unit used to measure cell potential.
- The voltage of a cell is determined by comparing the difference in the tendency of the tow electrode materials to accept electrons .
- The greater the difference , the greater the potential energy difference between the tow electrodes and the larger the voltage of the cell will be.

Calculating electrochemical cell potential :

- **Reduction potential:**the tendency of a substance to gain electrons .
- The reduction potential of an electrode cannot be determined directly because the reduction half – reaction must be coupled with an oxidation half – reaction.

The standered hydrogen electrode:

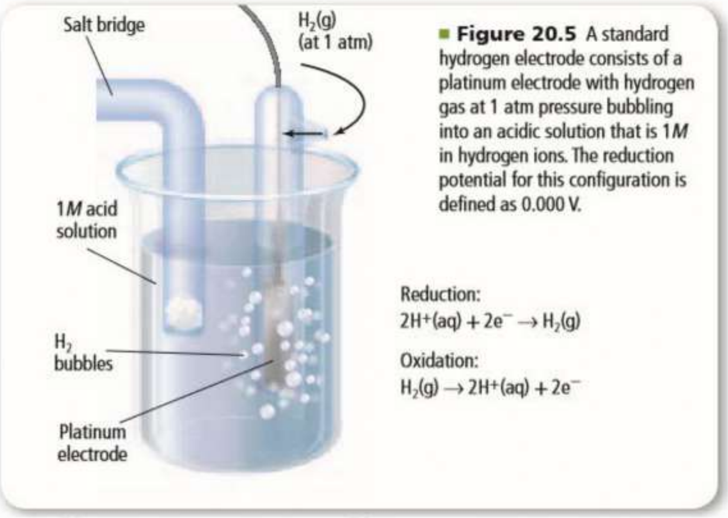


Figure 20.5 A standard hydrogen electrode consists of a platinum electrode with hydrogen gas at 1 atm pressure bubbling into an acidic solution that is 1*M* in hydrogen ions. The reduction potential for this configuration is defined as 0.000 V.

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

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

- It consists of :

  ……………………………………

  ……………………………………

  ……………………………………

  ……………………………………

  ……………………………………

  ……………………………………

- The potetial also called , the standered reduction potential (E°) , of this standered hydrogen electrode is defined as 0.00 V .

$$2H^+(aq) + 2e^- \underset{\text{Oxidation}}{\overset{\text{Reduction}}{\rightleftharpoons}} H_2(g) \qquad E^0 = 0.000\ V$$

Half – cell potential:

- Over the years chemists have measured and recorded the standered reduction potetials of many different half – cells .
- The following table lists some common half – cell reactions in order of increasing reduction potential.
- The values in the table were obtained by measuring the potential when each half – cell was connected to a standered hydrogen half – cell .
- All of the half – reactions are written as reductions .
- The half – reaction that is more positive will proceed as a reduction , and the half – reaction that is more negative will proceed as an oxidation .
- Standered conditions : 1M solution of its ions , at 25°C and 1 atm.

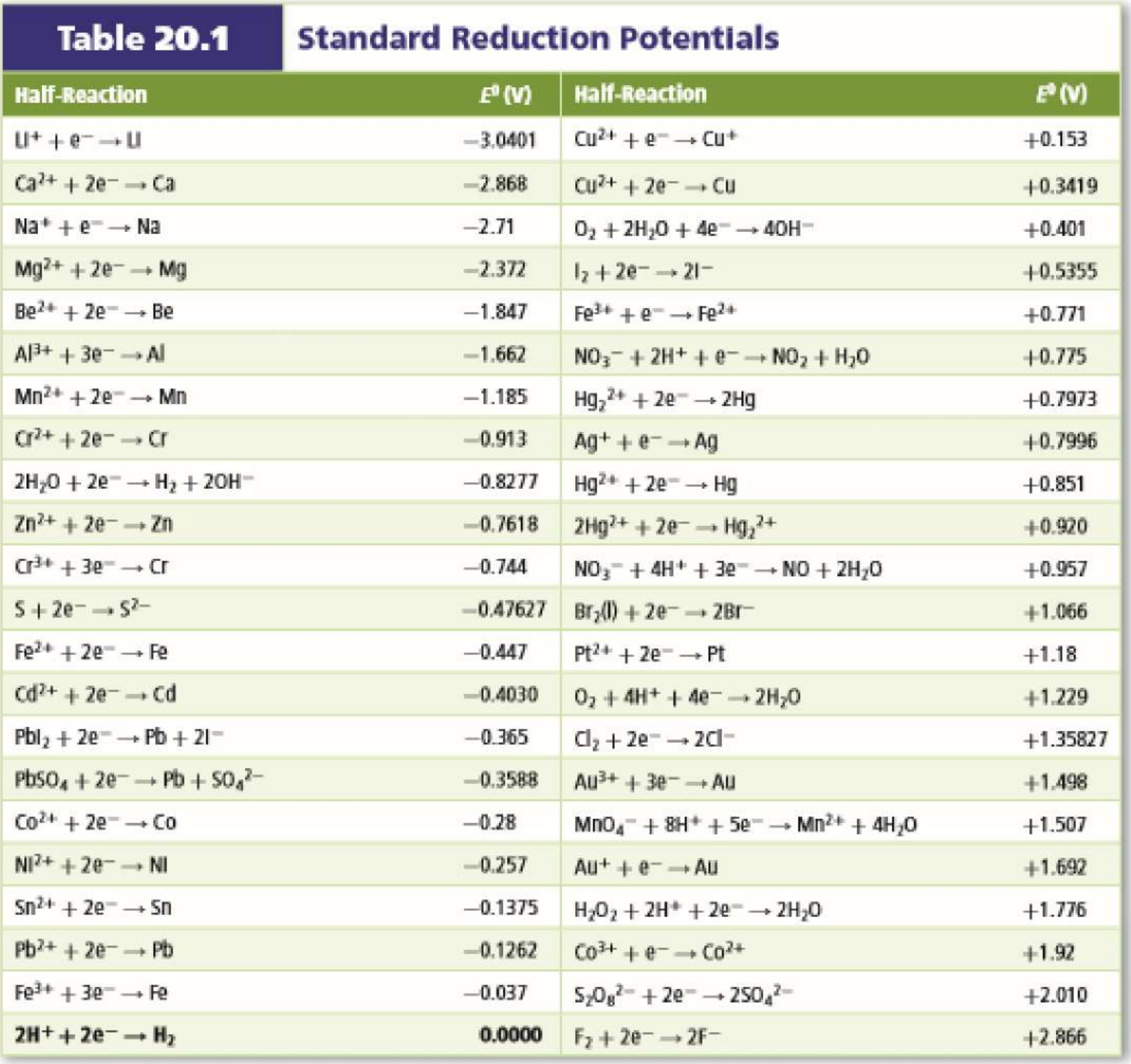


Table 20.1 Standard Reduction Potentials

Half-Reaction	E^0 (V)	Half-Reaction	E^0 (V)
$Li^+ + e^- \rightarrow Li$	−3.0401	$Cu^{2+} + e^- \rightarrow Cu^+$	+0.153
$Ca^{2+} + 2e^- \rightarrow Ca$	−2.868	$Cu^{2+} + 2e^- \rightarrow Cu$	+0.3419
$Na^+ + e^- \rightarrow Na$	−2.71	$O_2 + 2H_2O + 4e^- \rightarrow 4OH^-$	+0.401
$Mg^{2+} + 2e^- \rightarrow Mg$	−2.372	$I_2 + 2e^- \rightarrow 2I^-$	+0.5355
$Be^{2+} + 2e^- \rightarrow Be$	−1.847	$Fe^{3+} + e^- \rightarrow Fe^{2+}$	+0.771
$Al^{3+} + 3e^- \rightarrow Al$	−1.662	$NO_3^- + 2H^+ + e^- \rightarrow NO_2 + H_2O$	+0.775
$Mn^{2+} + 2e^- \rightarrow Mn$	−1.185	$Hg_2^{2+} + 2e^- \rightarrow 2Hg$	+0.7973
$Cr^{2+} + 2e^- \rightarrow Cr$	−0.913	$Ag^+ + e^- \rightarrow Ag$	+0.7996
$2H_2O + 2e^- \rightarrow H_2 + 2OH^-$	−0.8277	$Hg^{2+} + 2e^- \rightarrow Hg$	+0.851
$Zn^{2+} + 2e^- \rightarrow Zn$	−0.7618	$2Hg^{2+} + 2e^- \rightarrow Hg_2^{2+}$	+0.920
$Cr^{3+} + 3e^- \rightarrow Cr$	−0.744	$NO_3^- + 4H^+ + 3e^- \rightarrow NO + 2H_2O$	+0.957
$S + 2e^- \rightarrow S^{2-}$	−0.47627	$Br_2(l) + 2e^- \rightarrow 2Br^-$	+1.066
$Fe^{2+} + 2e^- \rightarrow Fe$	−0.447	$Pt^{2+} + 2e^- \rightarrow Pt$	+1.18
$Cd^{2+} + 2e^- \rightarrow Cd$	−0.4030	$O_2 + 4H^+ + 4e^- \rightarrow 2H_2O$	+1.229
$PbI_2 + 2e^- \rightarrow Pb + 2I^-$	−0.365	$Cl_2 + 2e^- \rightarrow 2Cl^-$	+1.35827
$PbSO_4 + 2e^- \rightarrow Pb + SO_4^{2-}$	−0.3588	$Au^{3+} + 3e^- \rightarrow Au$	+1.498
$Co^{2+} + 2e^- \rightarrow Co$	−0.28	$MnO_4^- + 8H^+ + 5e^- \rightarrow Mn^{2+} + 4H_2O$	+1.507
$Ni^{2+} + 2e^- \rightarrow Ni$	−0.257	$Au^+ + e^- \rightarrow Au$	+1.692
$Sn^{2+} + 2e^- \rightarrow Sn$	−0.1375	$H_2O_2 + 2H^+ + 2e^- \rightarrow 2H_2O$	+1.776
$Pb^{2+} + 2e^- \rightarrow Pb$	−0.1262	$Co^{3+} + e^- \rightarrow Co^{2+}$	+1.92
$Fe^{3+} + 3e^- \rightarrow Fe$	−0.037	$S_2O_8^{2-} + 2e^- \rightarrow 2SO_4^{2-}$	+2.010
$2H^+ + 2e^- \rightarrow H_2$	**0.0000**	$F_2 + 2e^- \rightarrow 2F^-$	+2.866

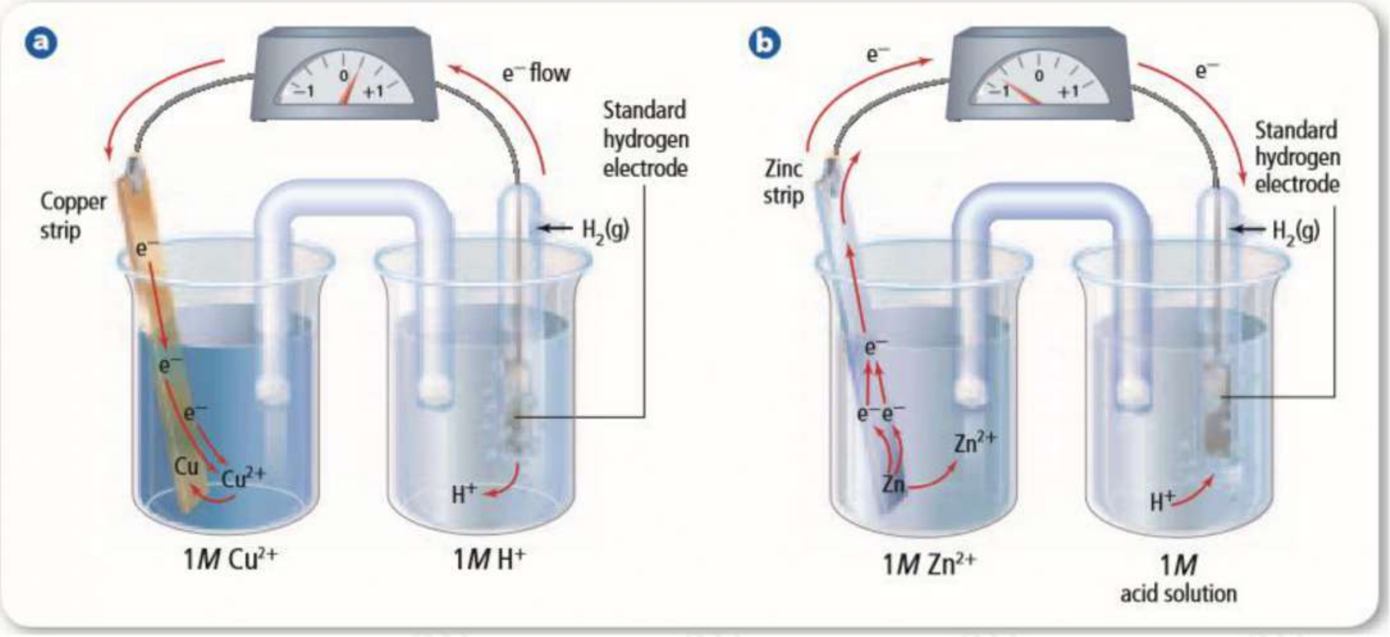


Figure 20.6 a. When a Cu | Cu^{2+} electrode is connected to the hydrogen electrode, electrons flow toward the copper strip and reduce Cu^{2+} ions to Cu atoms. The voltage of this reaction is +0.342 V. **b.** When a Zn | Zn^{2+} electrode is connected to the hydrogen electrode, electrons flow away from the zinc strip and zinc atoms are oxidized to Zn^{2+} ions. The voltage of this reaction is −0.762 V.

Determining electrochemical cell potentials:

- To calculate the electric potential of a voltic cell ($Zn/Zn^{+2}//Cu^{+2}/Cu$) :
- The first step is to determine (E°_{Cu}) :

$$H_2(g) \rightarrow 2H^+(aq) + 2e^- \text{ (oxidation half-cell reaction)}$$
$$Cu^{2+}(aq) + 2e^- \rightarrow Cu(s) \text{ (reduction half-cell reaction)}$$

$$H_2(g) + Cu^{2+}(aq) \rightarrow 2H^+(aq) + Cu(s) \text{ (overall redox reaction)}$$

This reaction can be written in a form called cell notation.

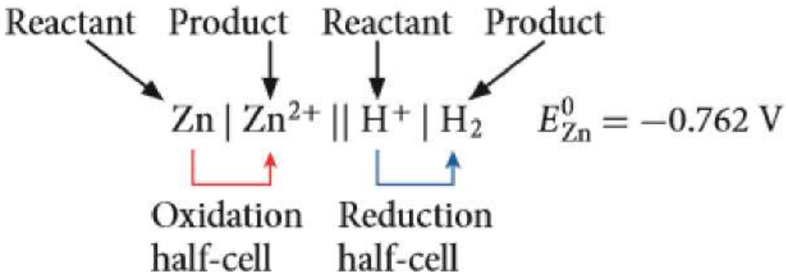
Reactant Product Reactant Product

$$\underbrace{H_2 \mid H^+}_{\text{Oxidation half-cell}} \parallel \underbrace{Cu^{2+} \mid Cu}_{\text{Reduction half-cell}} \quad E^0_{Cu} = +0.342 \text{ V}$$

- The next step is to determine (E°_{Zn}):

$$\begin{aligned} &Zn(s) \rightarrow Zn^{2+}(aq) + 2e^- \text{ (oxidation half-cell reaction)} \\ &2H^+(aq) + 2e^- \rightarrow H_2(g) \text{ (reduction half-cell reaction)} \\ \hline &Zn(s) + 2H^+(aq) \rightarrow Zn^{2+}(aq) + H_2(g) \text{ (overall redox cell reaction)} \end{aligned}$$

This reaction can be written in the following cell notation.



- The final step is to calculate E°_{cell} :

Formula for Cell Potential

$$E^0_{cell} = E^0_{reduction} - E^0_{oxidation}$$

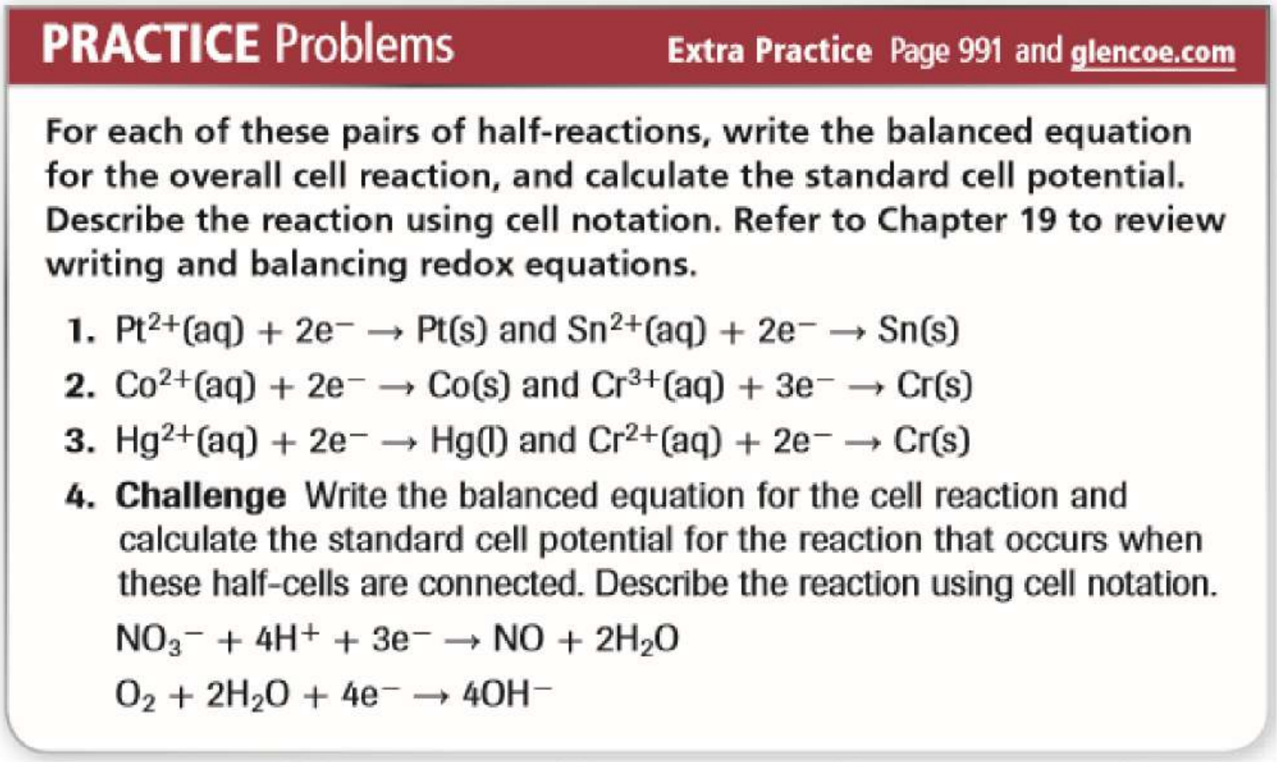
E^0_{cell} represents the overall standard cell potential.

$E^0_{reduction}$ represents the standard half-cell potential for the reduction.

$E^0_{oxidation}$ represents the standard half-cell potential for the oxidation.

The standard potential of a cell is the standard potential of the reduction half-cell minus the standard potential of the oxidation half-cell.

$$\begin{aligned} E^0_{cell} &= E^0_{Cu^{2+}|\,Cu} - E^0_{Zn^{2+}|\,Zn} \\ &= +\,0.342\text{ V} - (-0.762\text{ V}) \\ &= +1.104\text{ V} \end{aligned}$$



PRACTICE Problems Extra Practice Page 991 and glencoe.com

For each of these pairs of half-reactions, write the balanced equation for the overall cell reaction, and calculate the standard cell potential. Describe the reaction using cell notation. Refer to Chapter 19 to review writing and balancing redox equations.

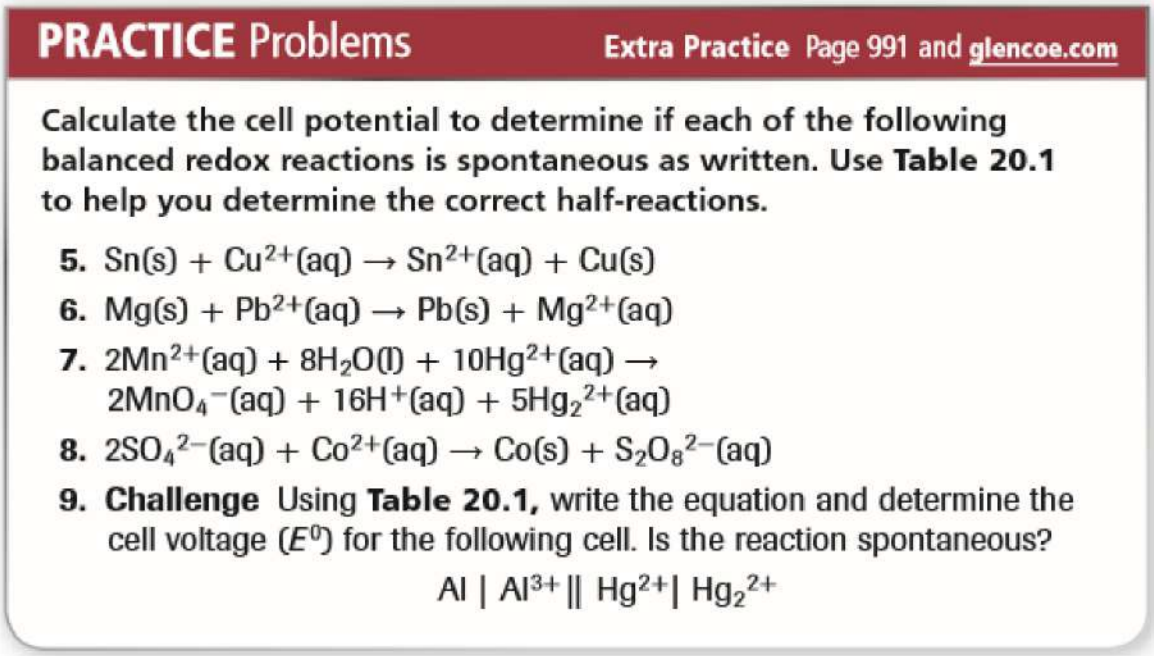
1. $Pt^{2+}(aq) + 2e^- \rightarrow Pt(s)$ and $Sn^{2+}(aq) + 2e^- \rightarrow Sn(s)$
2. $Co^{2+}(aq) + 2e^- \rightarrow Co(s)$ and $Cr^{3+}(aq) + 3e^- \rightarrow Cr(s)$
3. $Hg^{2+}(aq) + 2e^- \rightarrow Hg(l)$ and $Cr^{2+}(aq) + 2e^- \rightarrow Cr(s)$
4. **Challenge** Write the balanced equation for the cell reaction and calculate the standard cell potential for the reaction that occurs when these half-cells are connected. Describe the reaction using cell notation.

   $NO_3^- + 4H^+ + 3e^- \rightarrow NO + 2H_2O$

   $O_2 + 2H_2O + 4e^- \rightarrow 4OH^-$

Using standard reduction potentials:

- To calculate the standard potential of a voltaic cells .
- To determine if a proposed reaction will be spontaneous , if the E°_{cell} is positive the reaction is spontaneous .



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Calculate the cell potential to determine if each of the following balanced redox reactions is spontaneous as written. Use Table 20.1 to help you determine the correct half-reactions.

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

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

7. $2Mn^{2+}(aq) + 8H_2O(l) + 10Hg^{2+}(aq) \rightarrow 2MnO_4^{-}(aq) + 16H^{+}(aq) + 5Hg_2^{2+}(aq)$

8. $2SO_4^{2-}(aq) + Co^{2+}(aq) \rightarrow Co(s) + S_2O_8^{2-}(aq)$

9. **Challenge** Using **Table 20.1,** write the equation and determine the cell voltage (E^0) for the following cell. Is the reaction spontaneous?

$$Al \mid Al^{3+} \parallel Hg^{2+} \mid Hg_2^{2+}$$

SECTION 1 REVIEW

10. An electrochemical cell in which an oxidation half-reaction and a reduction half-reaction are connected by a salt bridge results in a flow of electrons (electric current) through a conducting wire.
11. A voltaic cell consists of an anode, a cathode, a salt bridge, and a connecting wire between the two electrodes. Oxidation takes place at the anode, reduction takes place at the cathode, the salt bridge allows movement of ions from one solution to the other, and the wire allows the passage of electrons from the anode to the cathode.
12. a. $2Ag^+ + Ni \rightarrow 2Ag + Ni^{2+}$
    b. $Mg + 2H^+ \rightarrow Mg^{2+} + H_2$
    c. $2Fe^{3+}(aq) + 3Sn(s) \rightarrow 2Fe(s) + 3Sn^{2+}(aq)$
    d. $Pb(s) + 2I^-(aq) + Pt^{2+}(aq) \rightarrow PbI_2(s) + Pt(s)$
13. a. E^0_{cell} = −2.004 V, nonspontaneous
    b. E^0_{cell} = +0.698 V, spontaneous
    c. E^0_{cell} = +1.178 V, spontaneous
14. Concept Maps will vary. Refer to the Solutions Manual.

H.W : Solve questions 30 – 42 page 264 .

For each item in Column A, write the letter of the matching item in Column B.

Column A

_______ **10.** One of the two parts of an electrochemical cell, where either oxidation or reduction takes place

_______ **11.** An electrode where oxidation takes place

_______ **12.** An electrode where reduction takes place

_______ **13.** One or more electrochemical cells in a single package that generates electrical current

_______ **14.** A measure of the amount of current that can be generated from an electrochemical cell to do work

Column B

a. battery

b. electrical potential

c. half-cell

d. cathode

e. anode

In your textbook, read about calculating cell electrochemical potential.

Circle the letter of the choice that best completes the statement or answers the question.

15. The tendency of an electrode to gain electrons is called

a. electron potential.
c. reduction potential.
b. gravitational potential.
d. oxidation potential.

16. A sheet of platinum covered with finely divided platinum particles is immersed in a 1*M* HCl solution containing hydrogen gas at a pressure of 1 atm and a temperature of 25°C. The platinum sheet is known as a

a. standard platinum electrode.
c. hydrogen chloride electrode.
b. standard hydrogen electrode.
d. platinum chloride electrode.

17. The standard reduction potential of a half-cell is a measure of

a. concentration.
c. temperature.
b. pressure.
d. voltage.

18. Which of the following is the correct way to represent the equation, $H_2(g) + Cu^{2+}(aq) \longrightarrow 2H^+(aq) + Cu(s)$?

a. $H_2|H^+||Cu^{2+}|Cu$

b. $H^+|H_2||Cu|Cu^{2+}$

c. $Cu^{2+}|Cu||H_2|H^+$

d. $Cu|Cu^{2+}||H^+|H_2$

19. When connected to a hydrogen electrode, an electrode with a negative standard reduction potential will carry out

a. reduction.

b. oxidation.

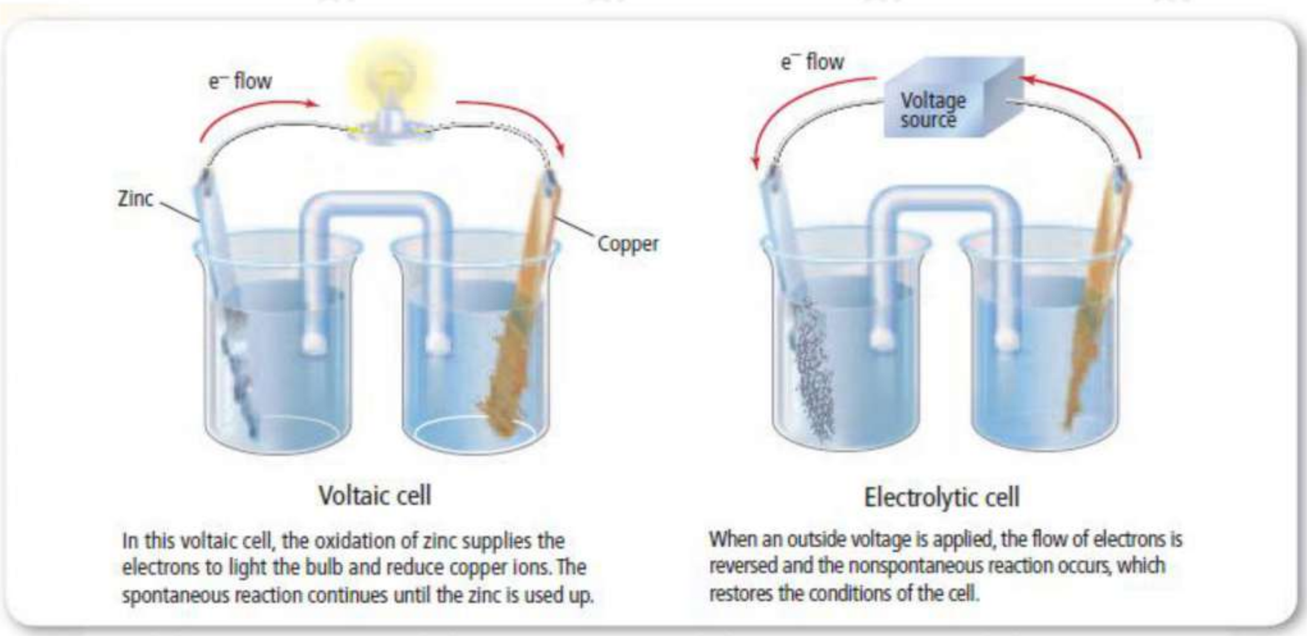
c. both oxidation and reduction.

d. neither oxidation nor reduction.

Section 3 : Electrolytes

Reversing redox reactions:

- **Electrolysis:** the use of electrical energy to bring about a chemical reaction .
- **Electrolytic cell:** is an electrochemical cell in which electrolysis occurs , like the recharging of a secondary battery



Voltaic cell

In this voltaic cell, the oxidation of zinc supplies the electrons to light the bulb and reduce copper ions. The spontaneous reaction continues until the zinc is used up.

Electrolytic cell

When an outside voltage is applied, the flow of electrons is reversed and the nonspontaneous reaction occurs, which restores the conditions of the cell.

	Voltaic cell	**Electrolytic cell**
The anode : The reaction:		
The cathode : The reaction:		
The charge of the anode :		
The charge of the cathode:		
The energy change:		
Is the reaction spontaneous?		

Applications of Electrolysis:

1- Electrolysis of water: to generate hydrogen gas for commercial use.

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

2- Electrolysis of molten NaCl :

- The name of the cell : **down's cell**
- The electrolyte : the molten NaCl
- The anode (+) made of carbon : ……………………………………………………………………
- The cathode (-) made of iron : ……………………………………………………………………
- The net cell reaction is : ……………………………………………………………………………

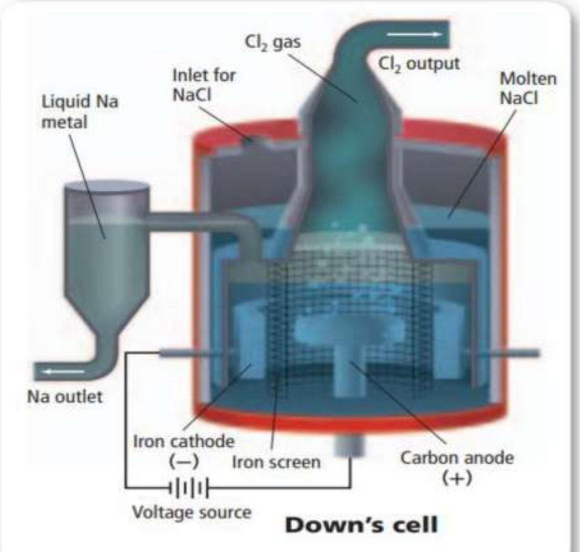


Figure 20.20 In a Down's cell, electrons supplied by a generator are used to reduce sodium ions. As electrons are removed from the anode, chloride ions are oxidized to chlorine gas.

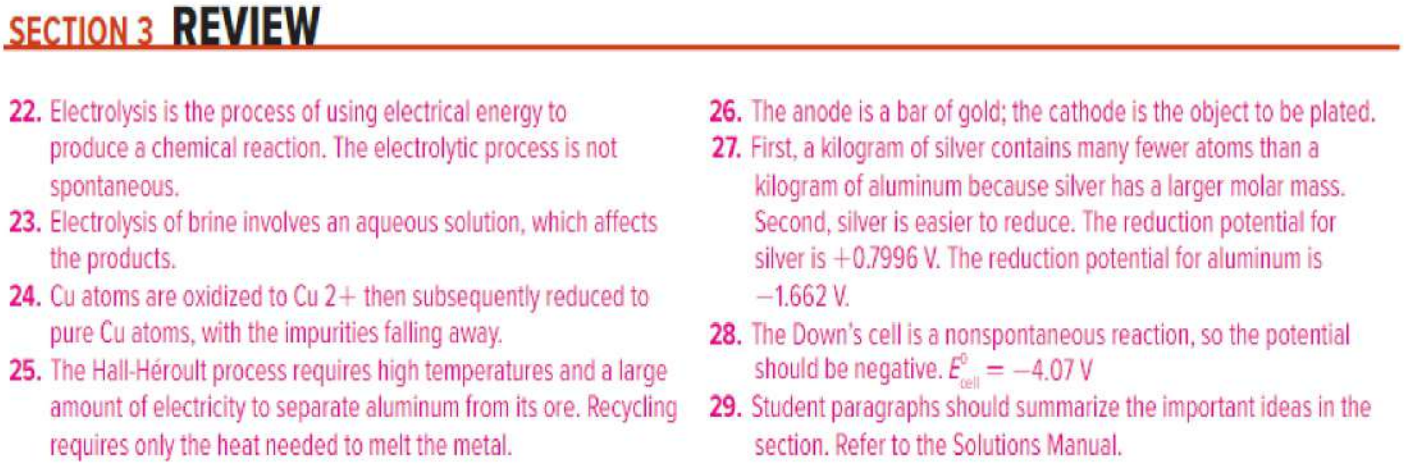
- The importance of chlorine :
  - to purify water for drinking and swimming
  - in cleaning products
  - in papers , plastics, insectisides, textiles , dyes and paints.
- The importance of sodium :
  - As a coolant in nuclear reactors
  - in sodium vapor lamps for outdoor lighting
  - to form ionic compounds (sodium salts) to use it in our foods .

Electrolytic cell	**The product at the Anode (+)**	**The product at the Cathode (-)**
1- Electrolysis of water		
2- **Electrolysis of molten NaCl (down's cell)**		

3- Electroplating :

- **Objects are** electroplated when a uniform coating is deposited usually as a protective or decorative layer.
- **The** cathode : is the object to be silver- plated for example, and silver ions are reduced to silver : ……………………………………………………………………………………………
- **The** anode is the silver bar or sheet , silver is oxeidized to silver ions: ……
- **The** electrolyte is a solution of silver ions .
- **Current** passing through the cell must be carefully controlled in order to get a smooth , even metal coating.

Figure 23 Power is needed to oxidize silver at the anode and reduce silver at the cathode. In an electrolytic cell used for silver plating, the object to be plated is the cathode where silver ions in the electrolyte solution are reduced to silver metal and deposited on the object.



SECTION 3 REVIEW

22. Electrolysis is the process of using electrical energy to produce a chemical reaction. The electrolytic process is not spontaneous.

23. Electrolysis of brine involves an aqueous solution, which affects the products.

24. Cu atoms are oxidized to Cu 2+ then subsequently reduced to pure Cu atoms, with the impurities falling away.

25. The Hall-Héroult process requires high temperatures and a large amount of electricity to separate aluminum from its ore. Recycling requires only the heat needed to melt the metal.

26. The anode is a bar of gold; the cathode is the object to be plated.

27. First, a kilogram of silver contains many fewer atoms than a kilogram of aluminum because silver has a larger molar mass. Second, silver is easier to reduce. The reduction potential for silver is +0.7996 V. The reduction potential for aluminum is −1.662 V.

28. The Down's cell is a nonspontaneous reaction, so the potential should be negative. $E^0_{cell} = -4.07$ V

29. Student paragraphs should summarize the important ideas in the section. Refer to the Solutions Manual.

H.W : Solve questions 55 – 63 page 265 .