

Chapter 10 : Electrochemistry

Chapter 4: Types of chemical reactions and solution stoichiometry

4.10 Oxidation–Reduction Reactions

4.11 Balancing Oxidation–Reduction Equations

Chapter 11: Electrochemistry

11.1 Galvanic Cells

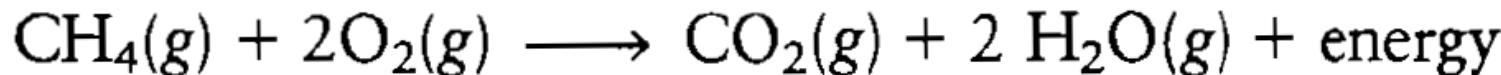
11.2 Standard Reduction Potentials

Electrochemistry is best defined as *the study of the interchange of chemical and electrical energy*. It is primarily concerned with two processes that involve oxidation-reduction reactions: the generation of an electric current from a chemical reaction and the opposite process, the use of a current to produce chemical change.

Oxidation-reduction reactions or redox reactions are reactions in which one or more electrons are transferred.

Most reactions used for energy production are redox reactions.

In humans the oxidation of sugars, fats, and proteins provides the energy necessary for life. Combustion reactions also involve oxidation and reduction.



Even though none of the reactants or products in the above reaction is ionic, the reaction is still assumed to involve a transfer of electrons from carbon to oxygen. To explain this, we introduce the concept of oxidation states.

Oxidation States

Oxidation states (also called oxidation numbers) provide a way to keep track of electrons in oxidation-reduction reactions.

Oxidation states are defined by a set of rules describing how to divide up the shared electrons in compounds containing covalent bonds.

Table 4.3Rules for Assigning Oxidation States

1. The oxidation state of an atom in an element is 0. For example, the oxidation state of each atom in the substances $\text{Na}(s)$, $\text{O}_2(g)$, $\text{O}_3(g)$, and $\text{Hg}(l)$ is 0.
 2. The oxidation state of a monatomic ion is the same as its charge. For example, the oxidation state of the Na^+ ion is +1.
 3. In its covalent compounds with nonmetals, hydrogen is assigned an oxidation state of +1. For example, in the compounds HCl , NH_3 , H_2O , and CH_4 , hydrogen is assigned an oxidation state of +1.
 4. Oxygen is assigned an oxidation state of -2 in its covalent compounds, such as CO , CO_2 , SO_2 , and SO_3 . The exception to this rule occurs in peroxides (compounds containing the O_2^{2-} group), where each oxygen is assigned an oxidation state of -1. The best-known example of a peroxide is hydrogen peroxide (H_2O_2).
 5. In binary compounds, the element with the greater attraction for the electrons in the bond is assigned a negative oxidation state equal to its charge in its ionic compounds. For example, fluorine is always assigned an oxidation state of -1. That is, for purposes of counting electrons, fluorine is assumed to be F^- . Nitrogen is usually assigned -3. For example, in NH_3 , nitrogen is assigned an oxidation state of -3; in H_2S , sulfur is assigned an oxidation state of -2; in HI , iodine is assigned an oxidation state of -1; and so on.
 6. The sum of the oxidation states must be zero for an electrically neutral compound and must be equal to the overall charge for an ionic species. For example, the sum of the oxidation states for the hydrogen and oxygen atoms in water is 0; the sum of the oxidation states for the carbon and oxygen atoms in CO_3^{2-} is -2; and the sum of oxidation states for the nitrogen and hydrogen atoms in NH_4^+ is +1.
-

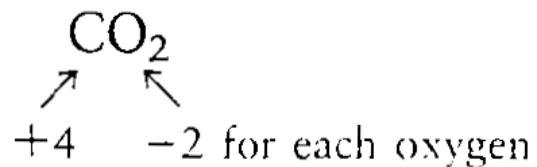
Example 4.13

Assign oxidation states to all atoms in the following:

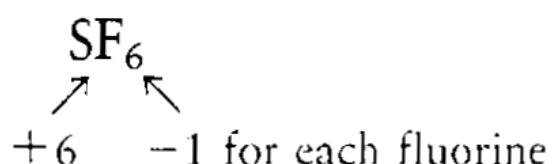
- a. CO_2
- b. SF_6
- c. NO_3^-

Solution

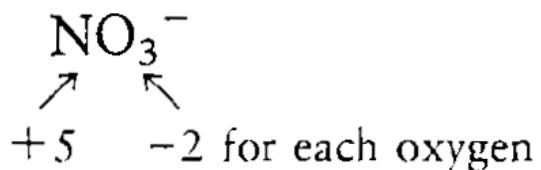
- a. The rule that takes precedence here is that oxygen is assigned an oxidation state of -2 . The oxidation state for carbon can be determined by recognizing that since CO_2 has no charge, the sum of the oxidation states for oxygen and carbon must be 0 . Since each oxygen is -2 and there are two oxygen atoms, the carbon atom must be assigned an oxidation state of $+4$:



- b. Since fluorine has the greater attraction for the shared electrons, we assign its oxidation state first. Because its charge in ionic compounds is 1^- , we assign -1 as the oxidation state of each fluorine atom. The sulfur must then be assigned an oxidation state of $+6$ to balance the total of -6 from the fluorine atoms:

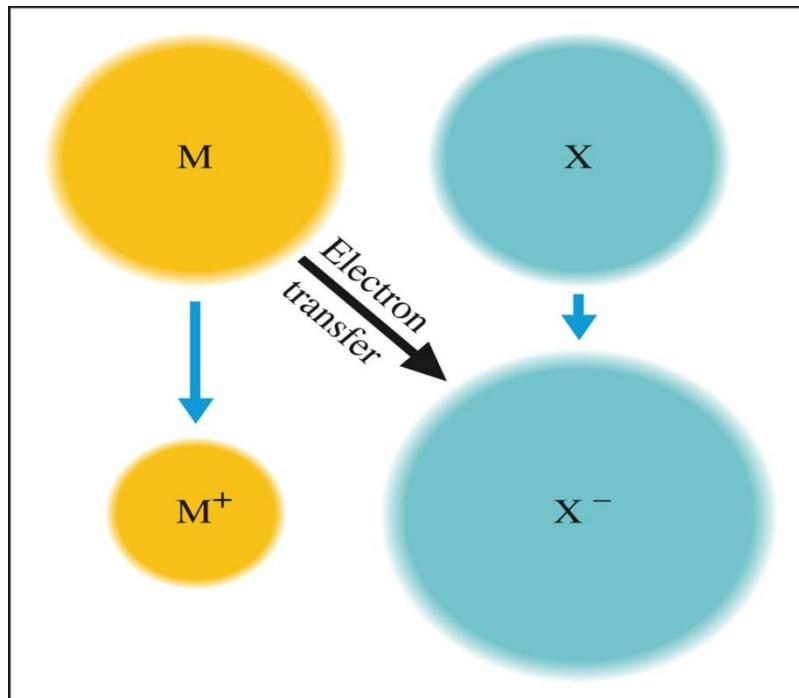


- c. Since oxygen has a greater attraction than does nitrogen for the shared electrons, we assign its oxidation state of -2 first. Because the sum of the oxidation states of the three oxygens is -6 and the net charge on the NO_3^- ion is 1^- , the nitrogen must have an oxidation state of $+5$:



The Characteristics of Oxidation-Reduction Reactions

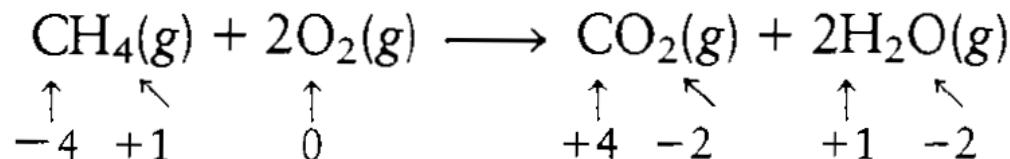
Oxidation is an increase in oxidation state (a loss of electrons).
Reduction is a decrease in oxidation state (a gain of electrons)



An oxidation-reduction (redox) reaction involves a transfer of electrons from the reducing agent to the oxidizing agent.

Oxidized <i>Loses electrons</i> <i>Oxidation state increases</i>	Reduced <i>Gains electrons</i> <i>Oxidation state decreases</i>
<i>Reducing agent</i>	<i>Oxidizing agent</i>

Concerning the reaction



we can say the following:

Carbon is oxidized because there is an increase in its oxidation state (carbon has formally lost electrons).

Oxygen is reduced as shown by the decrease in its oxidation state (oxygen has formally gained electrons).

CH_4 is the reducing agent.

O_2 is the oxidizing agent.

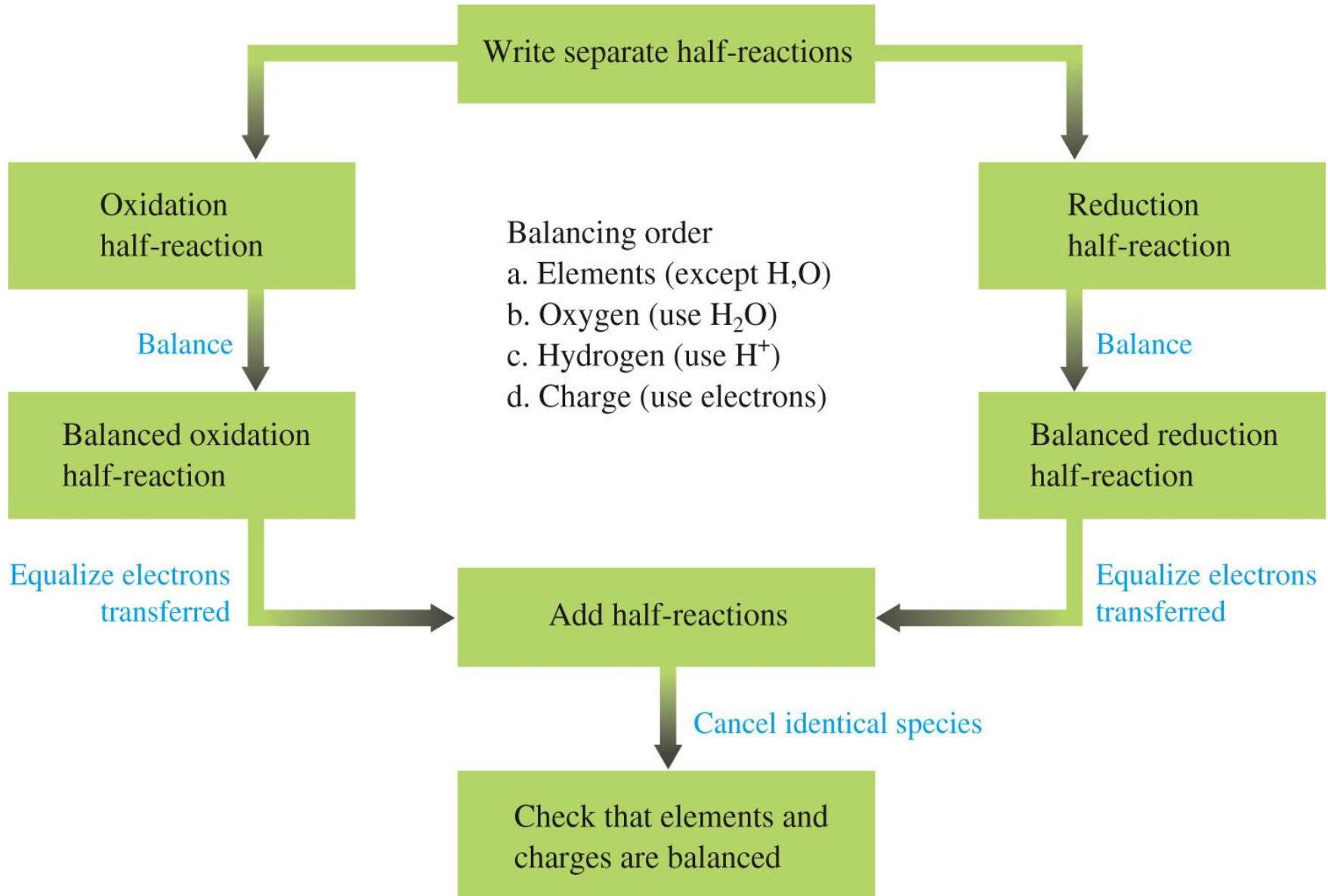
Note that when the oxidizing or reducing agent is named, the *whole compound* is specified, not just the element that undergoes the change in oxidation state.

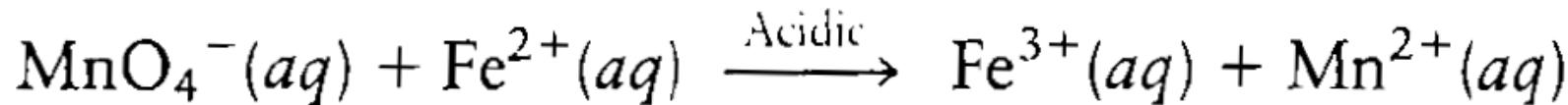
4.11 | Balancing Oxidation–Reduction Equations

The Half-Reaction Method

Balancing Oxidation–Reduction Equations Occurring in Acidic Solution by the Half-Reaction Method

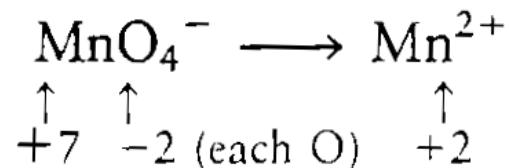
- Write the equations for the oxidation and reduction half-reactions.
- For each half-reaction:
 - a. Balance all the elements except hydrogen and oxygen.
 - b. Balance oxygen using H_2O .
 - c. Balance hydrogen using H^+ .
 - d. Balance the charge using electrons.
- If necessary, multiply one or both balanced half-reactions by integers to equalize the number of electrons transferred in the two half-reactions.
- Add the half-reactions, and cancel identical species.
- Check to be sure that the elements and charges balance.



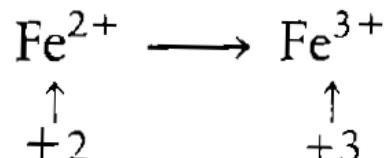


- Identify and write equations for the half-reactions.

The oxidation states for the half-reaction involving the permanganate ion show that manganese is reduced:



This is the *reduction half-reaction*. The other half-reaction involves the oxidation of iron(II) to iron(III) ion and is the *oxidation half-reaction*:



- Balance each half-reaction.

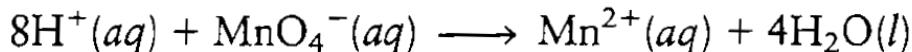
For the reduction reaction, we have



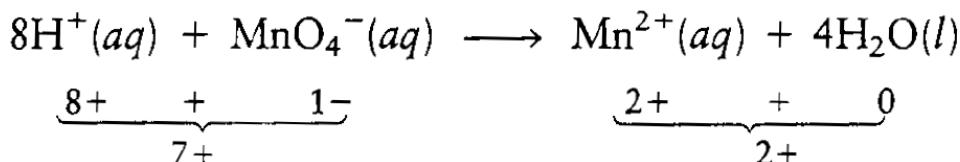
- The manganese is balanced.
- We balance oxygen by adding $4\text{H}_2\text{O}$ to the right side of the equation:



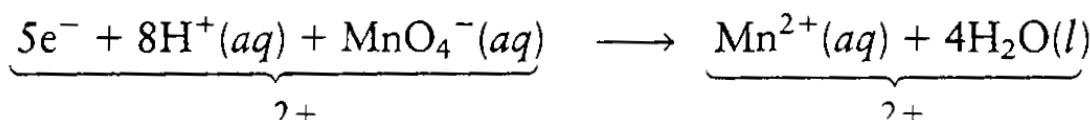
- Next, we balance hydrogen by adding 8H^+ to the left side:



- All the elements have been balanced, but we need to balance the charge by using electrons. At this point we have the following charges for reactants and products in the reduction half-reaction:



We can equalize the charges by adding five electrons to the left side:

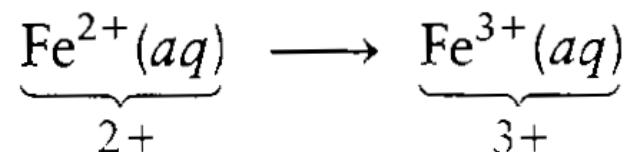


Both the *elements* and the *charges* are now balanced, so this represents the balanced reduction half-reaction. The fact that five electrons appear on the reactant side of the equation makes sense, since five electrons are required to reduce MnO_4^- (in which Mn has an oxidation state of +7) to Mn^{2+} (in which Mn has an oxidation state of +2).

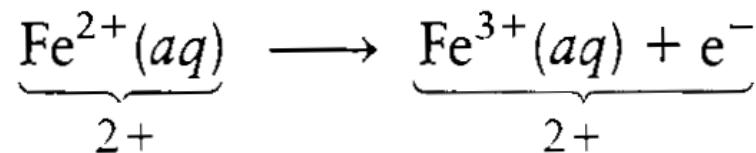
For the oxidation reaction,



the elements are balanced, and we must simply balance the charge:

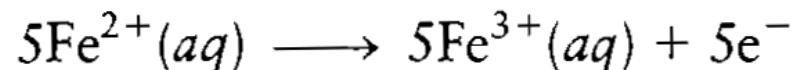


One electron is needed on the right side to give a net 2+ charge on both sides:



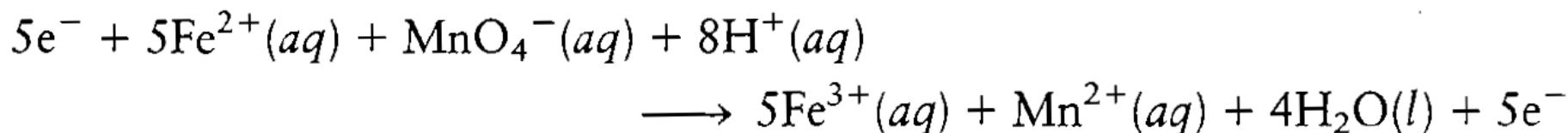
- Equalize the electron transfer in the two half-reactions.

Since the reduction half-reaction involves a transfer of five electrons and the oxidation half-reaction involves a transfer of only one electron, the oxidation half-reaction must be multiplied by 5:

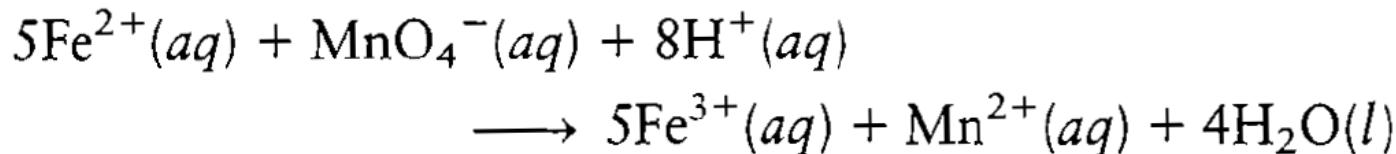


- Add the half-reactions.

The half-reactions are added to give



Note that the electrons cancel (as they must) to give the final balanced equation:



- Check that the elements and charges balance.

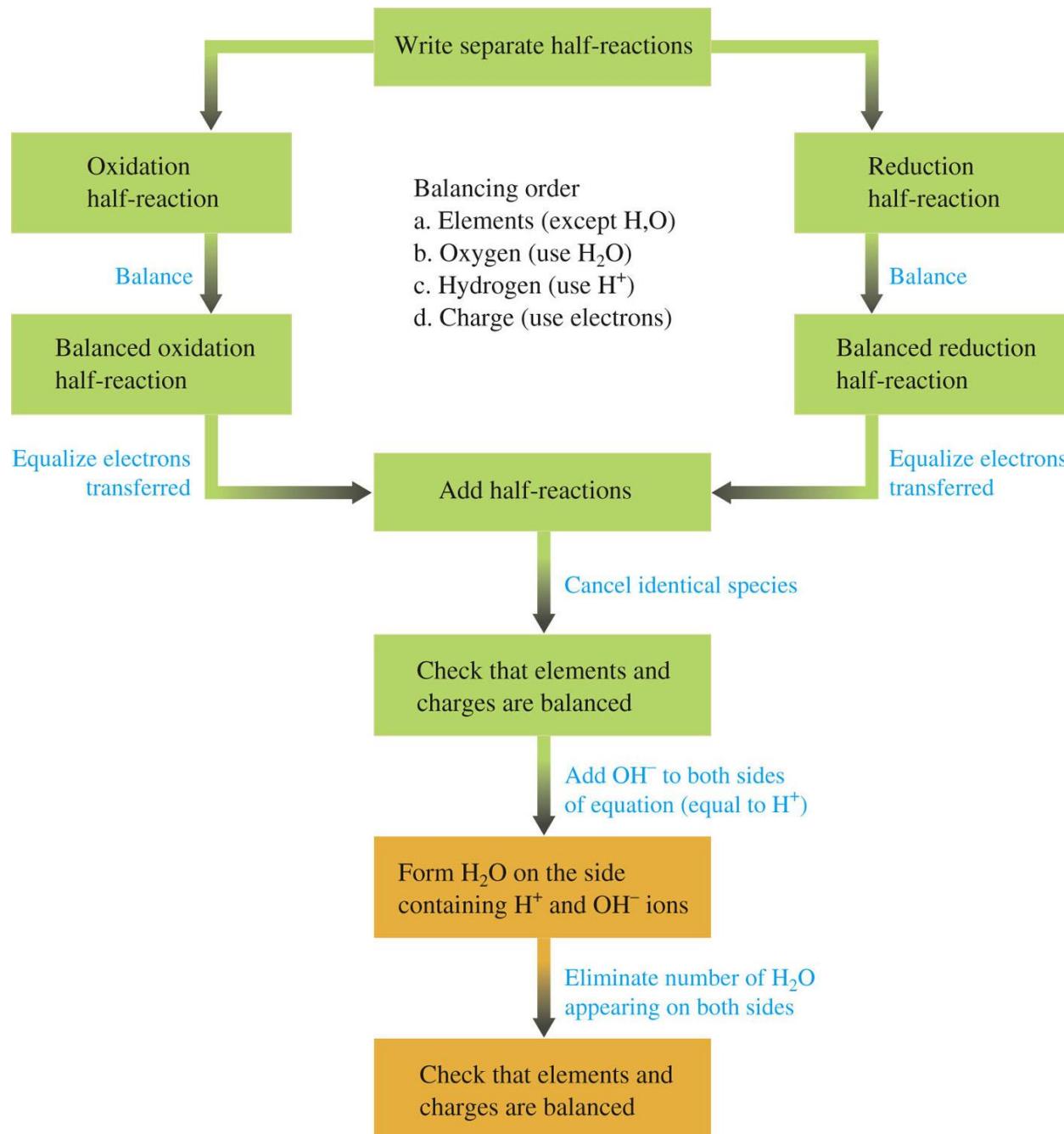
Elements balance: 5 Fe, 1 Mn, 4 O, 8 H \longrightarrow 5 Fe, 1 Mn, 4 O, 8 H

Charges balance: $5(2+) + (1-) + 8(1+) = 17+$
 $\longrightarrow 5(3+) + (2+) + 0 = 17+$

The equation is balanced.

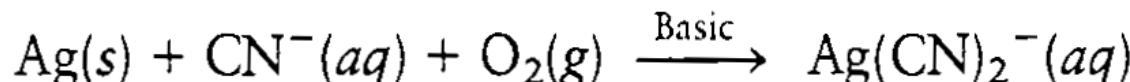
Balancing Oxidation–Reduction Equations Occurring in Basic Solution by the Half-Reaction Method

- Use the half-reaction method as specified for acidic solutions to obtain the final balanced equation as *if H⁺ ions were present*.
- To both sides of the equation obtained by the procedure above, add the number of OH[−] ions that is equal to the number of H⁺ ions. (We want to eliminate H⁺ by forming H₂O.)
- Form H₂O on the side containing both H⁺ and OH[−] ions, and eliminate the number of H₂O molecules that appear on both sides of the equation.
- Check that the elements and charges balance.



Example 4.15

Silver is sometimes found in nature as large nuggets; more often it is found mixed with other metals and their ores. Cyanide ion is often used to extract the silver by the following reaction that occurs in basic solution:

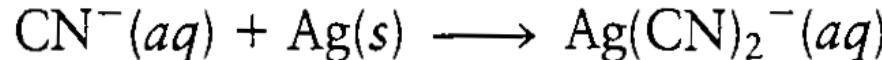


Balance this equation by using the half-reaction method.

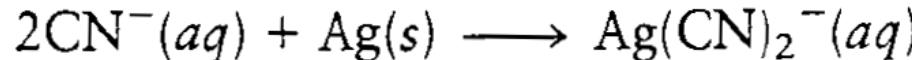
Solution

- Balance the equation as if H^+ ions were present.

Balance the oxidation half-reaction:



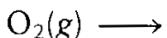
Balance carbon and nitrogen:



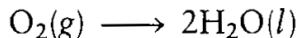
Balance the charge:



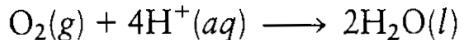
Balance the reduction half-reaction:



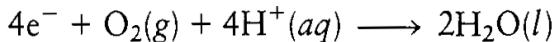
Balance oxygen:



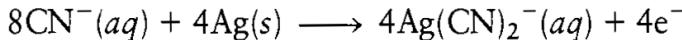
Balance hydrogen:



Balance the charge:

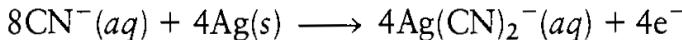


Multiply the balanced oxidation half-reaction by 4:

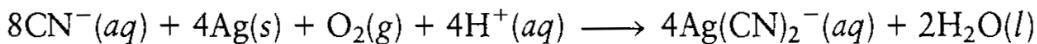
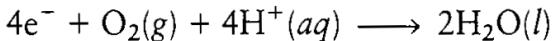


Add the half-reactions, and cancel identical species:

Oxidation half-reaction:

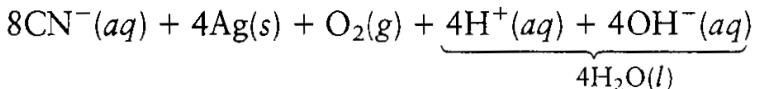


Reduction half-reaction:

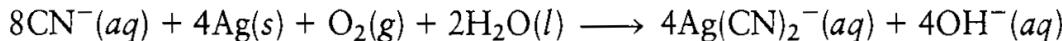


- Add OH^- ions to both sides of the balanced equation.

We need to add 4OH^- to each side:



- Eliminate as many H_2O molecules as possible.

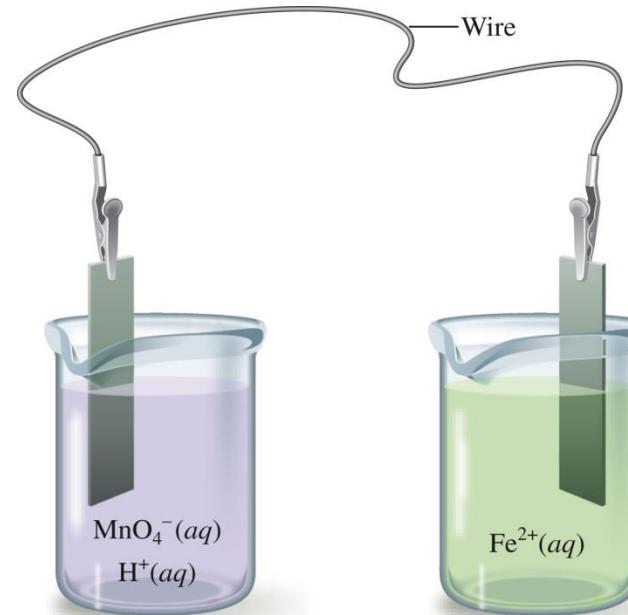


- Check that elements and charges balance.

11.1 | Galvanic Cells

When the oxidizing agent is separated from the reducing agent, the electron transfer occurs through a wire.

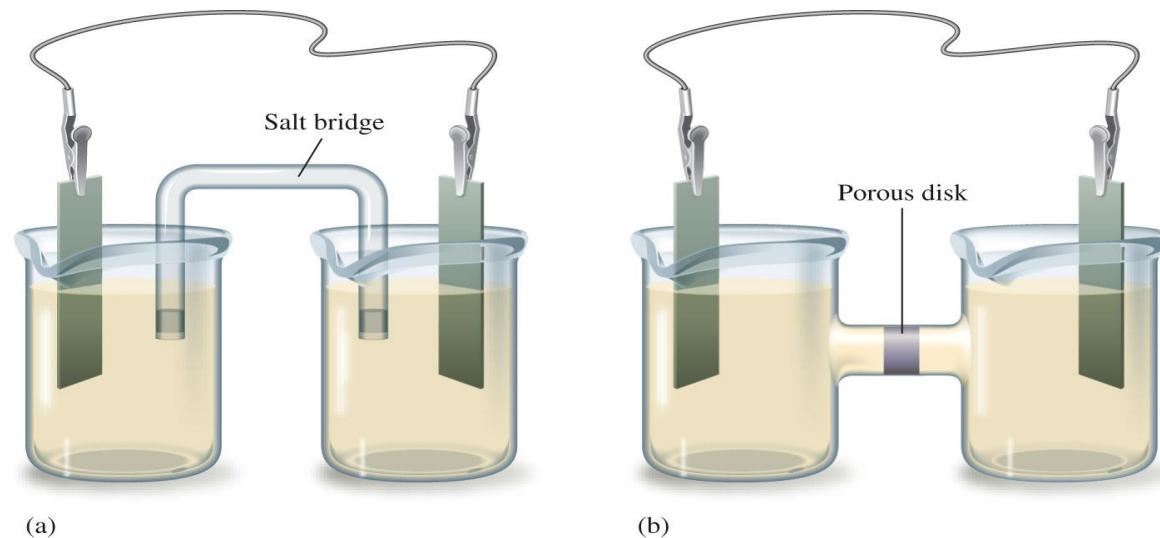
The current produced in the wire by the electron flow can then be directed through a device, such as an electric motor, to provide useful work.



The net charge in each compartment must be maintained at zero (otherwise the charge buildups and the current stops flowing). This is achieved by using a **salt bridge** or a **porous disk** connecting the two solutions.

A **salt bridge** or a **porous disk** allow ion flow without mixing of the solutions.

Electrons flow through the wire from reducing agent to oxidizing agent, and ions flow between the compartments to keep the net charge zero in each.

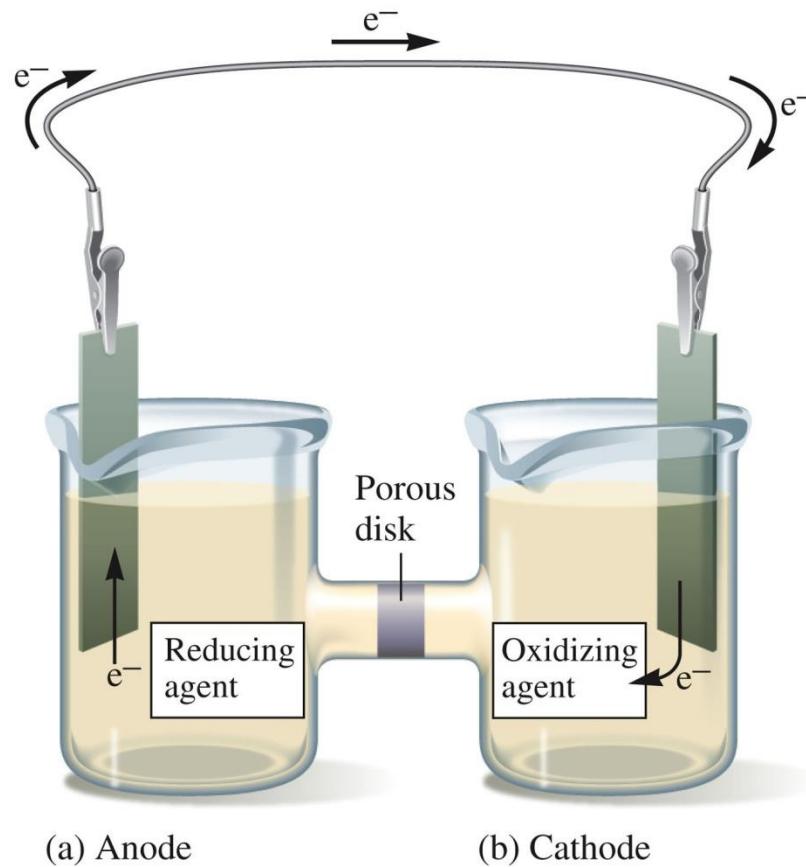


Complete circuit

The device in which chemical energy is changed to electrical energy is called a **Galvanic Cell**.

The electrode at which *oxidation* occurs is called the **anode**.

The electrode at which *reduction* occurs is called the **cathode**.

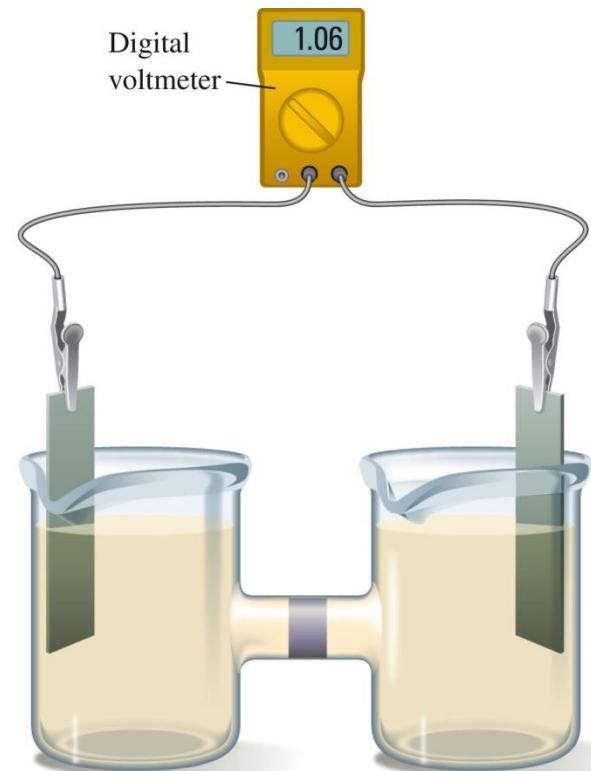


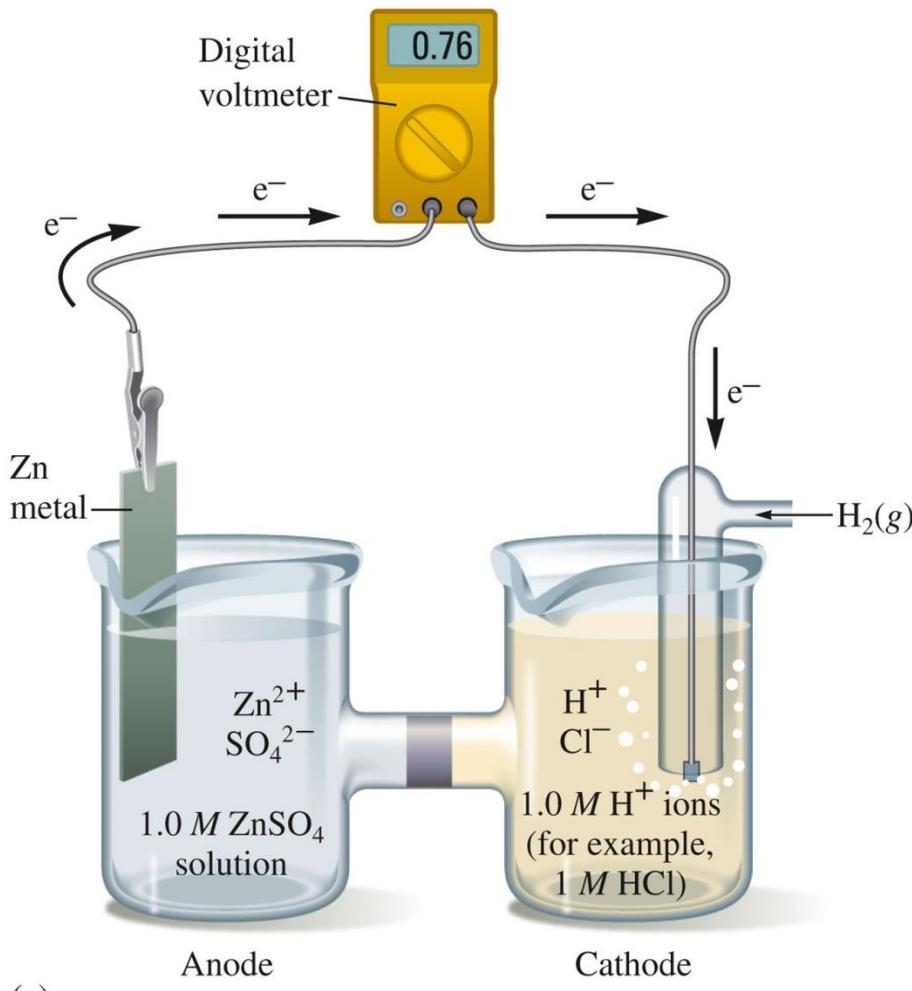
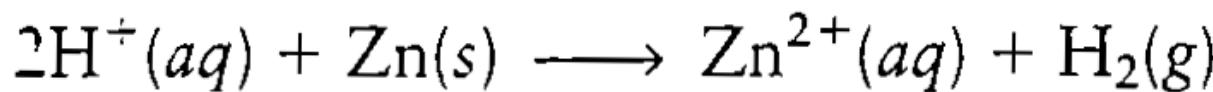
Cell Potential

The driving force that pulls electrons through a wire is called the **cell potential** ($\mathcal{E}_{\text{cell}}$) or the **electromotive force** (emf) of the cell.

Unit: Volt (V) = joule of work per coulomb of charge transferred (J/C)

Measured by a **voltmeter**.





Anode (oxidation half reaction):



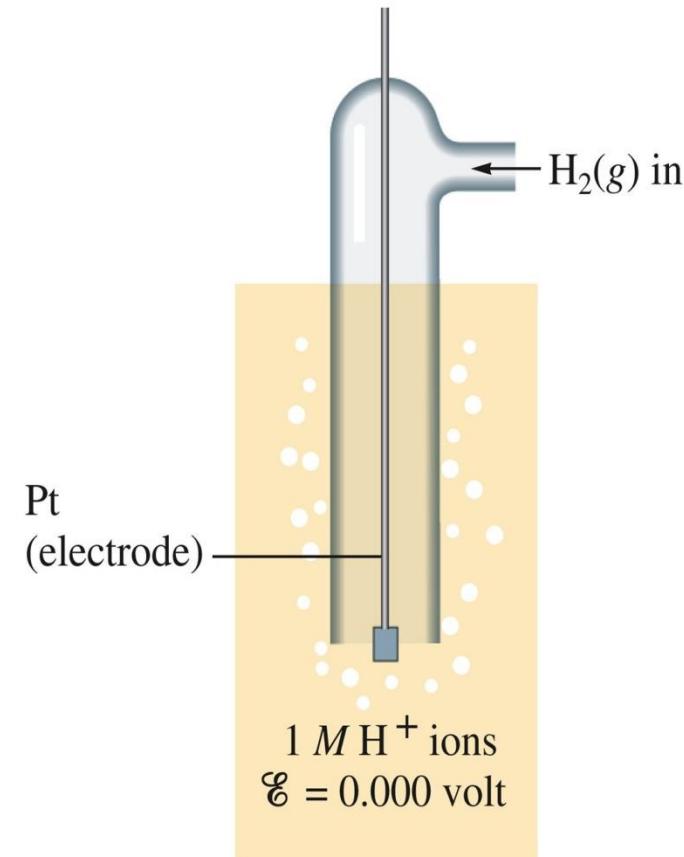
Cathode (reduction half reaction):



The **standard hydrogen electrode** consists of a platinum electrode in contact with 1M H^+ ions and bathed by hydrogen gas at 1 atm.

The standard hydrogen electrode has universally given a potential of exactly 0 volts.

By setting the standard potential for the standard hydrogen electrode equal to zero, we can assign values to all other half reactions.



(b)

The reaction: $\text{Zn} \rightarrow \text{Zn}^{2+} + 2\text{e}^-$
will have a potential of 0.76
volt .

The superscript 0 indicates
that standard states are used
(1 M concentration and 1 atm
pressure).

The accepted convention is to
give the potentials of half-
reactions as *reduction*
processes.

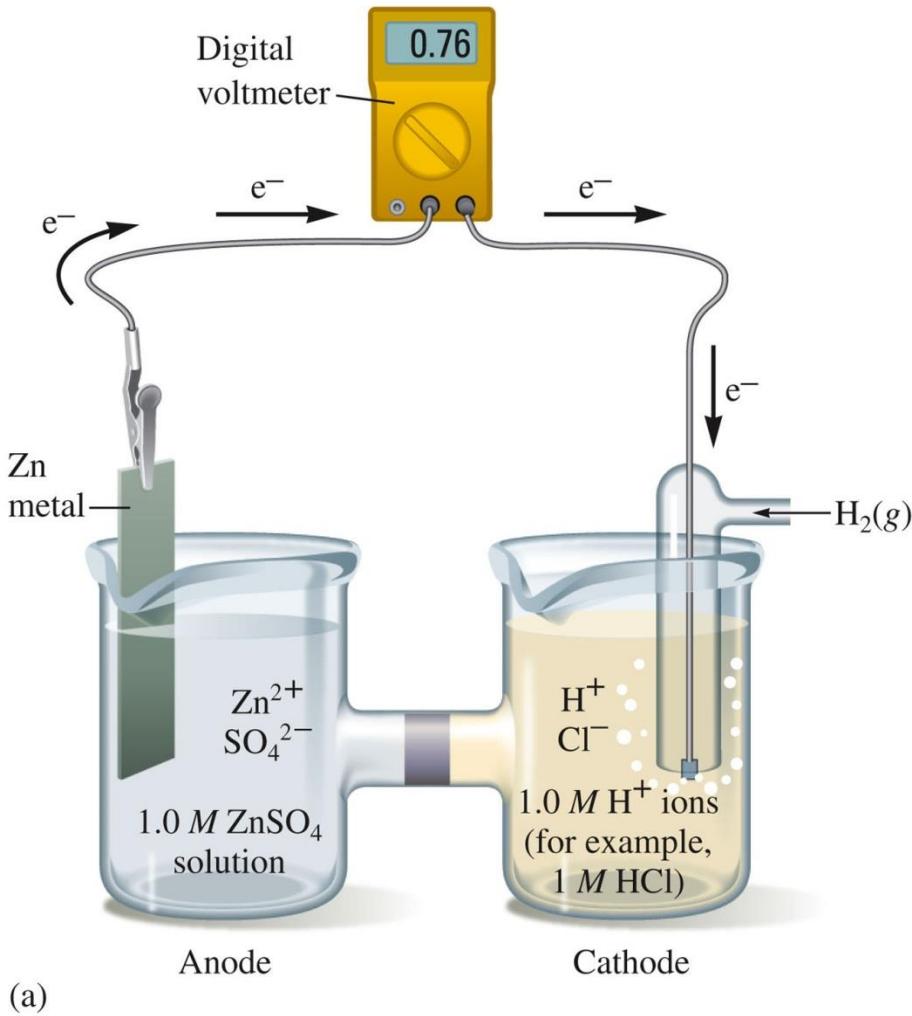


Table 11.1

Standard Reduction Potentials at 25°C (298 K) for Many Common Half-reactions

Half-reaction	\mathcal{E}° (V)	Half-reaction	\mathcal{E}° (V)
$\text{F}_2 + 2\text{e}^- \rightarrow 2\text{F}^-$	2.87	$\text{O}_2 + 2\text{H}_2\text{O} + 4\text{e}^- \rightarrow 4\text{OH}^-$	0.40
$\text{Ag}^{2+} + \text{e}^- \rightarrow \text{Ag}^+$	1.99	$\text{Cu}^{2+} + 2\text{e}^- \rightarrow \text{Cu}$	0.34
$\text{Co}^{3+} + \text{e}^- \rightarrow \text{Co}^{2+}$	1.82	$\text{Hg}_2\text{Cl}_2 + 2\text{e}^- \rightarrow 2\text{Hg} + 2\text{Cl}^-$	0.27
$\text{H}_2\text{O}_2 + 2\text{H}^+ + 2\text{e}^- \rightarrow 2\text{H}_2\text{O}$	1.78	$\text{AgCl} + \text{e}^- \rightarrow \text{Ag} + \text{Cl}^-$	0.22
$\text{Ce}^{4+} + \text{e}^- \rightarrow \text{Ce}^{3+}$	1.70	$\text{SO}_4^{2-} + 4\text{H}^+ + 2\text{e}^- \rightarrow \text{H}_2\text{SO}_3 + \text{H}_2\text{O}$	0.20
$\text{PbO}_2 + 4\text{H}^+ + \text{SO}_4^{2-} + 2\text{e}^- \rightarrow \text{PbSO}_4 + 2\text{H}_2\text{O}$	1.69	$\text{Cu}^{2+} + \text{e}^- \rightarrow \text{Cu}^+$	0.16
$\text{MnO}_4^- + 4\text{H}^+ + 3\text{e}^- \rightarrow \text{MnO}_2 + 2\text{H}_2\text{O}$	1.68	$2\text{H}^+ + 2\text{e}^- \rightarrow \text{H}_2$	0.00
$\text{IO}_4^- + 2\text{H}^+ + 2\text{e}^- \rightarrow \text{IO}_3^- + \text{H}_2\text{O}$	1.60	$\text{Fe}^{3+} + 3\text{e}^- \rightarrow \text{Fe}$	-0.036
$\text{MnO}_4^- + 8\text{H}^+ + 5\text{e}^- \rightarrow \text{Mn}^{2+} + 4\text{H}_2\text{O}$	1.51	$\text{Pb}^{2+} + 2\text{e}^- \rightarrow \text{Pb}$	-0.13
$\text{Au}^{3+} + 3\text{e}^- \rightarrow \text{Au}$	1.50	$\text{Sn}^{2+} + 2\text{e}^- \rightarrow \text{Sn}$	-0.14
$\text{PbO}_2 + 4\text{H}^+ + 2\text{e}^- \rightarrow \text{Pb}^{2+} + 2\text{H}_2\text{O}$	1.46	$\text{Ni}^{2+} + 2\text{e}^- \rightarrow \text{Ni}$	-0.23
$\text{Cl}_2 + 2\text{e}^- \rightarrow 2\text{Cl}^-$	1.36	$\text{PbSO}_4 + 2\text{e}^- \rightarrow \text{Pb} + \text{SO}_4^{2-}$	-0.35
$\text{Cr}_2\text{O}_7^{2-} + 14\text{H}^+ + 6\text{e}^- \rightarrow 2\text{Cr}^{3+} + 7\text{H}_2\text{O}$	1.33	$\text{Cd}^{2+} + 2\text{e}^- \rightarrow \text{Cd}$	-0.40
$\text{O}_2 + 4\text{H}^+ + 4\text{e}^- \rightarrow 2\text{H}_2\text{O}$	1.23	$\text{Fe}^{2+} + 2\text{e}^- \rightarrow \text{Fe}$	-0.44
$\text{MnO}_2 + 4\text{H}^+ + 2\text{e}^- \rightarrow \text{Mn}^{2+} + 2\text{H}_2\text{O}$	1.21	$\text{Cr}^{3+} + \text{e}^- \rightarrow \text{Cr}^{2+}$	-0.50
$\text{IO}_3^- + 6\text{H}^+ + 5\text{e}^- \rightarrow \frac{1}{2}\text{I}_2 + 3\text{H}_2\text{O}$	1.20	$\text{Cr}^{3+} + 3\text{e}^- \rightarrow \text{Cr}$	-0.73
$\text{Br}_2 + 2\text{e}^- \rightarrow 2\text{Br}^-$	1.09	$\text{Zn}^{2+} + 2\text{e}^- \rightarrow \text{Zn}$	-0.76
$\text{VO}_2^+ + 2\text{H}^+ + \text{e}^- \rightarrow \text{VO}^{2+} + \text{H}_2\text{O}$	1.00	$2\text{H}_2\text{O} + 2\text{e}^- \rightarrow \text{H}_2 + 2\text{OH}^-$	-0.83
$\text{AuCl}_4^- + 3\text{e}^- \rightarrow \text{Au} + 4\text{Cl}^-$	0.99	$\text{Mn}^{2+} + 2\text{e}^- \rightarrow \text{Mn}$	-1.18
$\text{NO}_3^- + 4\text{H}^+ + 3\text{e}^- \rightarrow \text{NO} + 2\text{H}_2\text{O}$	0.96	$\text{Al}^{3+} + 3\text{e}^- \rightarrow \text{Al}$	-1.66
$\text{ClO}_2 + \text{e}^- \rightarrow \text{ClO}_2^-$	0.954	$\text{H}_2 + 2\text{e}^- \rightarrow 2\text{H}^-$	-2.23
$2\text{Hg}^{2+} + 2\text{e}^- \rightarrow \text{Hg}_2^{2+}$	0.91	$\text{Mg}^{2+} + 2\text{e}^- \rightarrow \text{Mg}$	-2.37
$\text{Ag}^+ + \text{e}^- \rightarrow \text{Ag}$	0.80	$\text{La}^{3+} + 3\text{e}^- \rightarrow \text{La}$	-2.37
$\text{Hg}_2^{2+} + 2\text{e}^- \rightarrow 2\text{Hg}$	0.80	$\text{Na}^+ + \text{e}^- \rightarrow \text{Na}$	-2.71
$\text{Fe}^{3+} + \text{e}^- \rightarrow \text{Fe}^{2+}$	0.77	$\text{Ca}^{2+} + 2\text{e}^- \rightarrow \text{Ca}$	-2.76
$\text{O}_2 + 2\text{H}^+ + 2\text{e}^- \rightarrow \text{H}_2\text{O}_2$	0.68	$\text{Ba}^{2+} + 2\text{e}^- \rightarrow \text{Ba}$	-2.90
$\text{MnO}_4^- + \text{e}^- \rightarrow \text{MnO}_4^{2-}$	0.56	$\text{K}^+ + \text{e}^- \rightarrow \text{K}$	-2.92
$\text{I}_2 + 2\text{e}^- \rightarrow 2\text{I}^-$	0.54	$\text{Li}^+ + \text{e}^- \rightarrow \text{Li}$	-3.05
$\text{Cu}^+ + \text{e}^- \rightarrow \text{Cu}$	0.52		

$$\mathcal{E}_{cell}^0 = \mathcal{E}^0(\text{cathode}) - \mathcal{E}^0(\text{anode})$$

When the reduction half reaction is reversed, the sign of \mathcal{E}^0 will change.

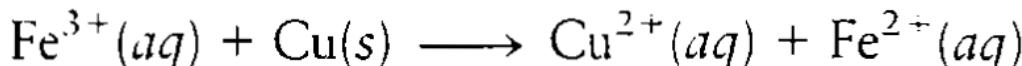
The value of \mathcal{E}^0 is not changed when a half-reaction is multiplied by an integer.

A cell will always run spontaneously in the direction that produces a positive cell potential.

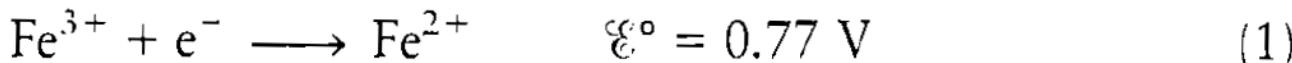
If $\mathcal{E}^0(\text{Ox}/\text{Red}) < 0$, then the substance red is a strong enough reducing agent to reduce H^+ ions.

If $\mathcal{E}^0(\text{Ox}/\text{Red}) > 0$, then the substance ox is strongly oxidizing.

Consider a galvanic cell based on the redox reaction



The pertinent half-reactions are

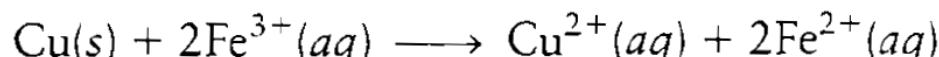
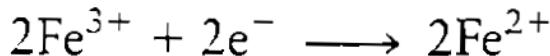


Since the Cu^{2+}/Cu half-reaction has the lower positive \mathcal{E}° value, it will be forced to run in reverse:



It is the anode reaction.

Then, since each Cu atom produces two electrons but each Fe^{3+} ion accepts only one electron, reaction (1) must be multiplied by 2:

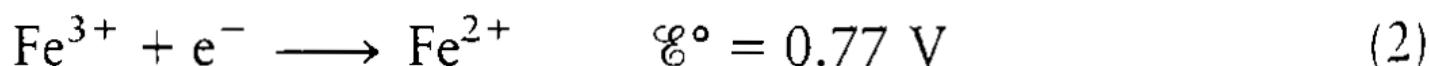
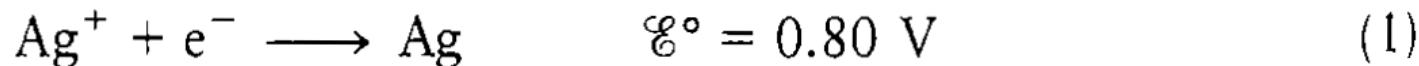


$$\mathcal{E}_{\text{cell}}^\circ = \mathcal{E}^\circ \text{ (cathode)} - \mathcal{E}^\circ \text{ (anode)}$$

$$= 0.77 \text{ V} - 0.34 \text{ V} = 0.43 \text{ V}$$

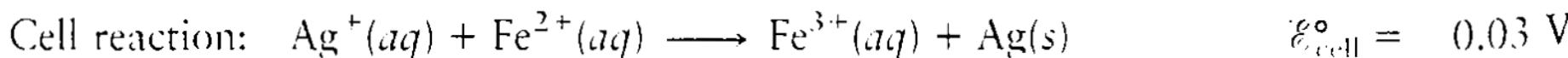
Example 11.1

Describe completely the galvanic cell based on the following half-reactions under standard conditions:



Solution

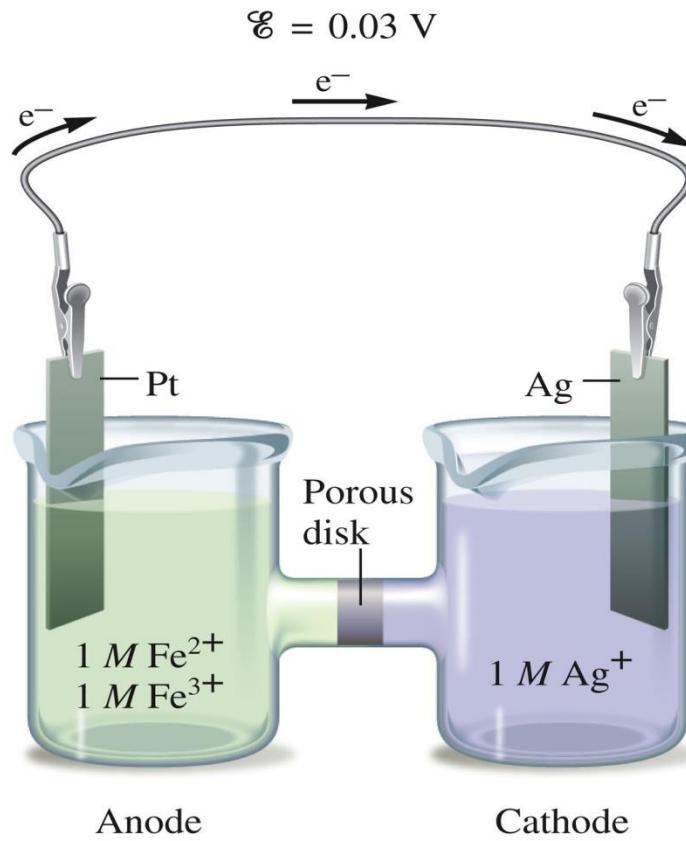
Since a positive $\mathcal{E}_{\text{cell}}^\circ$ value is required, reaction (2) must run in reverse:



Since Ag^+ receives electrons and Fe^{2+} loses electrons in the cell reaction, the electrons flow from the compartment containing Fe^{2+} to the compartment containing Ag^+ .

Oxidation occurs in the compartment containing Fe^{2+} . Hence this compartment functions as the anode. Reduction occurs in the compartment containing Ag^+ , so this compartment is the cathode.

The electrode in the Ag/Ag^+ compartment is silver metal; and an inert conductor, such as platinum, must be used in the $\text{Fe}^{2+}/\text{Fe}^{3+}$ compartment.



HOMEWORK

Chap.4: 64, 72

Chap.11: 15, 19, 22, 25, 27, 28, 31, 81

64

Assign oxidation states to all atoms in the following.

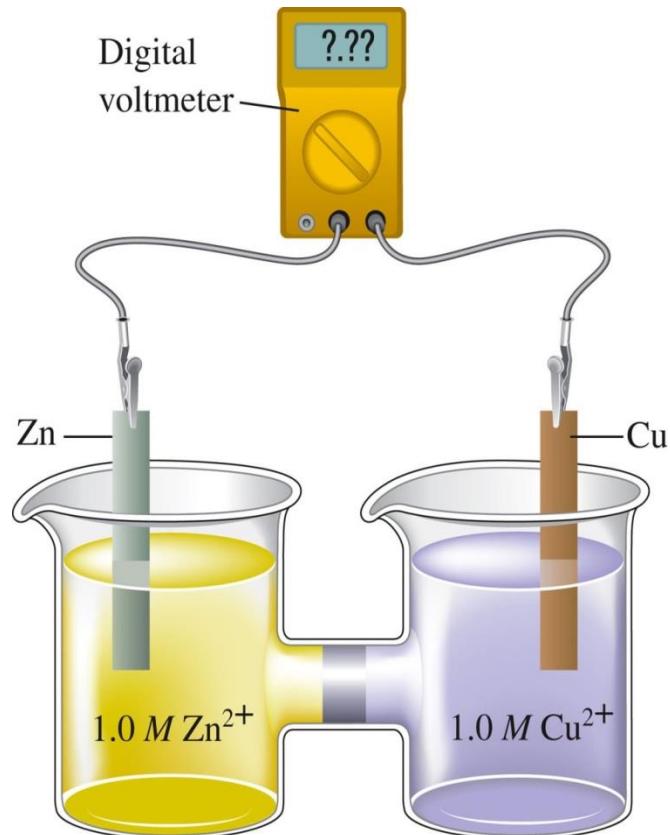
- | | | |
|----------------------------|--------------------------------------|--------------------------------------|
| a. UO_2^{2+} | d. As_4 | g. $\text{Na}_2\text{S}_2\text{O}_3$ |
| b. As_2O_3 | e. HAsO_2 | h. Hg_2Cl_2 |
| c. NaBiO_3 | f. $\text{Mg}_2\text{P}_2\text{O}_7$ | i. $\text{Ca}(\text{NO}_3)_2$ |

Balance the following oxidation-reduction reactions, which occur in basic solution, using the half-reaction method.

- $\text{Al}(s) + \text{MnO}_4^-(aq) \rightarrow \text{MnO}_2(s) + \text{Al(OH)}_4^-(aq)$
- $\text{Cl}_2(g) \rightarrow \text{Cl}^-(aq) + \text{ClO}^-(aq)$
- $\text{NO}_2^-(aq) + \text{Al}(s) \rightarrow \text{NH}_3(g) + \text{AlO}_2^-(aq)$
- $\text{MnO}_4^-(aq) + \text{S}^{2-}(aq) \rightarrow \text{MnS}(s) + \text{S}(s)$
- $\text{CN}^-(aq) + \text{MnO}_4^-(aq) \rightarrow \text{CNO}^-(aq) + \text{MnO}_2(s)$

15

Consider the following galvanic cell:

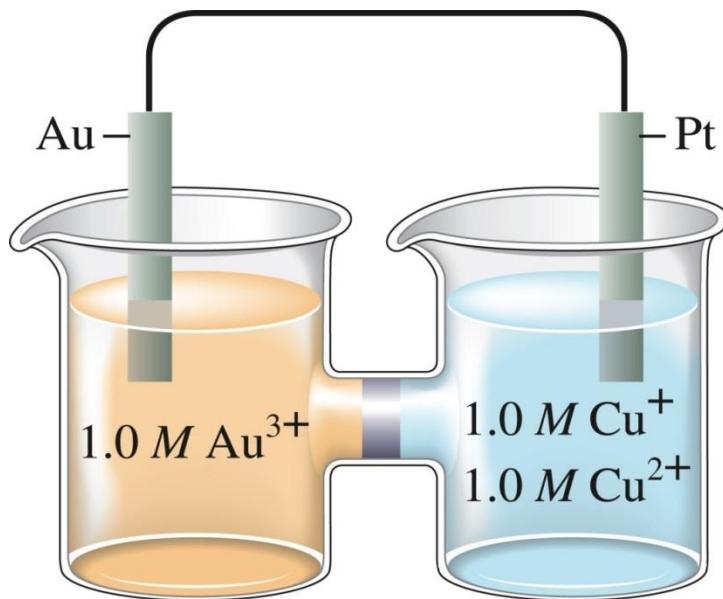


- Label the reducing agent and the oxidizing agent, and describe the direction of the electron flow.
- Determine the standard cell potential.
- Which electrode increases in mass as the reaction proceeds and which electrode decreases in mass?

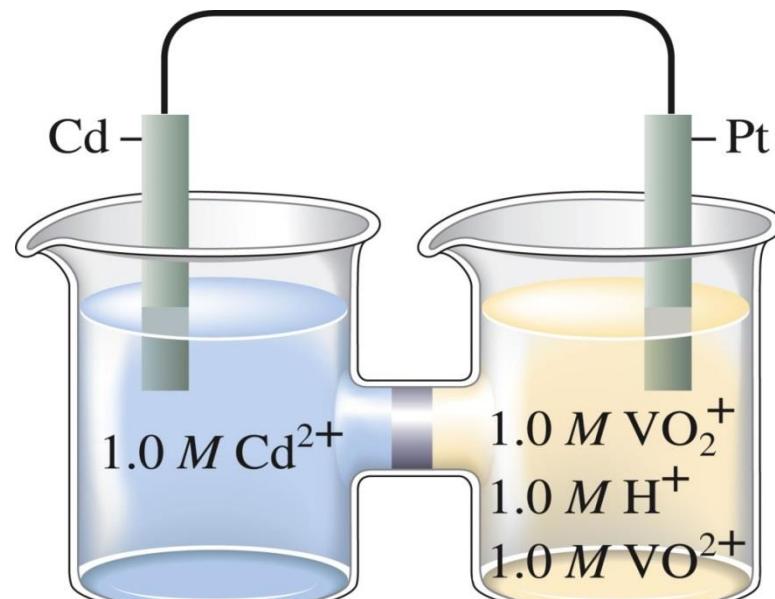
Sketch the galvanic cells based on the following overall reactions. Calculate \mathcal{E}° ; show the direction of electron flow and the direction of ion migration through the salt bridge; identify the cathode and anode; and give the overall balanced reaction. Assume that all concentrations are 1.0 M and that all partial pressures are 1.0 atm. Standard reduction potentials are found in Table 11.1.

- $\text{Cr}^{3+}(aq) + \text{Cl}_2(g) \rightleftharpoons \text{Cr}_2\text{O}_7^{2-}(aq) + \text{Cl}^-(aq)$
- $\text{Cu}^{2+}(aq) + \text{Mg}(s) \rightleftharpoons \text{Mg}^{2+}(aq) + \text{Cu}(s)$
- $\text{IO}_3^-(aq) + \text{Fe}^{2+}(aq) \rightleftharpoons \text{Fe}^{3+}(aq) + \text{I}_2(s)$
- $\text{Zn}(s) + \text{Ag}^+(aq) \rightleftharpoons \text{Zn}^{2+}(aq) + \text{Ag}(s)$

Consider the following galvanic cell:



a.



b.

For each galvanic cell, give the balanced cell reaction and determine \mathcal{E}° . Standard reduction potentials are found in Table 11.1.

25

Using data from Table 11.1, place the following in order of increasing strength as reducing agents (all under standard conditions):

Cu^+ F^- H^- H_2O I_2 and K

Use the table of standard reduction potentials (Table 11.1) to pick a reagent that is capable of each of the following oxidations (under standard conditions in acidic solution).

- oxidizes Br^- to Br_2 but does not oxidize Cl^- to Cl_2
- oxidizes Mn to Mn^{2+} but does not oxidize Ni to Ni^{2+}

Use the table of standard reduction potentials (Table 11.1) to pick a reagent that is capable of each of the following reductions (under standard conditions in acidic solution).

- a. reduces Cu^{2+} to Cu but does not reduce Cu^{2+} to Cu^+
- b. reduces Br_2 to Br^- but does not reduce I_2 to I^-

Answer the following questions using data from Table 11.1 (all under standard conditions).

- a. Is $\text{H}^+(\text{aq})$ capable of oxidizing $\text{Cu}(\text{s})$ to $\text{Cu}^{2+}(\text{aq})$?
- b. Is $\text{Fe}^{3+}(\text{aq})$ capable of oxidizing $\text{I}^-(\text{aq})$?
- c. Is $\text{H}_2(\text{g})$ capable of reducing $\text{Ag}^+(\text{aq})$?
- d. Is $\text{Fe}^{2+}(\text{aq})$ capable of reducing $\text{Cr}^{3+}(\text{aq})$ to $\text{Cr}^{2+}(\text{aq})$?

Copper can be plated onto a spoon by placing the spoon in an acidic solution of $\text{CuSO}_4 \text{ (aq)}$ and connecting it to a copper strip via a power source as illustrated in the figure.

- a. Label the anode and cathode, and describe the direction of the electron flow.
- b. Write out the chemical equations for the reactions that occur at each electrode.

