Example Exercise 17.1 Calculating Oxidation Numbers for Carbon

Calculate the oxidation number for carbon in each of the following compounds:

(a) diamond, C
(b) dry ice, CO_2
(c) marble, $CaCO_3$
(d) baking soda, $NaHCO_3$

Solution

Let's begin by recalling that uncombined elements, as well as compounds, are electrically neutral. Thus, free elements and compounds have no charge.

(a) In diamond, the oxidation number of carbon is zero.

(b) In dry ice, we assign oxygen an oxidation number of –2. We can determine the oxidation number of carbon in CO_2 as follows:

$$\begin{aligned} \text{ox no C} + 2\,(\text{ox no O}) &= 0 \\ \text{ox no C} + 2\,(-2) &= 0 \\ \text{ox no C} &= +4 \end{aligned}$$

(c) In marble, we assign calcium ion an oxidation number of +2, and oxygen a value of –2. We can determine the value of carbon in $CaCO_3$ as follows:

$$\begin{aligned} \text{ox no Ca} + \text{ox no C} + 3\,(\text{ox no O}) &= 0 \\ +2 + \text{ox no C} + 3\,(-2) &= 0 \\ \text{ox no C} &= +4 \end{aligned}$$

(d) In baking soda, we assign sodium ion an oxidation number of +1, hydrogen a value of +1, and oxygen a value of –2. We can determine the oxidation number of carbon in $NaHCO_3$ as follows:

$$\begin{aligned} \text{ox no Na} + \text{ox no H} + \text{ox no C} + 3\,(\text{ox no O}) &= 0 \\ +1 + +1 + \text{ox no C} + 3\,(-2) &= 0 \\ +1 + +1 + \text{ox no C} + -6 &= 0 \\ \text{ox no C} &= +4 \end{aligned}$$

Example Exercise 17.1 Calculating Oxidation Numbers for Carbon

Continued

Practice Exercise

Calculate the oxidation number for iodine in each of the following compounds:

(a) iodine, I2
(b) potassium iodide, KI
(c) silver periodate, $AgIO_4$
(d) zinc iodate, $Zn(IO_3)_2$

Answers: (a) 0; (b) –1; (c) +7; (d) +5

Concept Exercise

Calculate the oxidation number for nonmetal X in each of the following compounds:

(a) X2
(b) X_2O
(c) CaX_2
(d) HXO_4

Answer: See Appendix G.

Example Exercise 17.2 Calculating Oxidation Numbers for Sulfur

Calculate the oxidation number for sulfur in each of the following ions.

(a) sulfide ion, S^{2-}

(b) sulfite ion, SO_3^{2-}

(c) sulfate ion, SO_4^{2-}

(d) thiosulfate ion, $S_2O_3^{2-}$

Solution

We can begin by recalling that the charge on an ion corresponds to the sum of the oxidation numbers.

(a) In S^{2-}, the oxidation number of sulfur is –2.

(b) In SO_3^{2-}, the polyatomic anion has a charge of 2–. We assign oxygen an oxidation number of –2 and write the equation

$$\begin{aligned} \text{ox no S} + 3\,(\text{ox no O}) &= -2 \\ \text{ox no S} + 3\,(-2) &= -2 \\ \text{ox no S} &= +4 \end{aligned}$$

(c) In SO_4^{2-}, the polyatomic anion has a charge of 2–. We assign oxygen an oxidation number of –2 and write the equation

$$\begin{aligned} \text{ox no S} + 4\,(\text{ox no O}) &= -2 \\ \text{ox no S} + 4\,(-2) &= -2 \\ \text{ox no S} &= +6 \end{aligned}$$

(d) In $S_2O_3^{2-}$, the polyatomic anion has a charge of 2–. We assign oxygen an oxidation number of –2 and write the equation

$$\begin{aligned} 2\,(\text{ox no S}) + 3\,(\text{ox no O}) &= -2 \\ 2\,(\text{ox no S}) + 3\,(-2) &= -2 \\ 2\,(\text{ox no S}) &= +4 \\ \text{ox no S} &= +2 \end{aligned}$$

Example Exercise 17.2 Calculating Oxidation Numbers for Sulfur

Continued

Practice Exercise

Calculate the oxidation number for chlorine in each of the following ions:

(a) hypochlorite ion, ClO^-

(b) chlorite ion, ClO_2^-

(c) chlorate ion, ClO_3^-

(d) perchlorate ion, ClO_4^-

Answers: (a) +1; (b) +3; (c) +5; (d) +7

Concept Exercise

Calculate the oxidation number for nonmetal X in each of the following ions:

(a) X^-

(b) XO^-

(c) XO_2^-

(d) XO_3^-

Answer: See Appendix G.

Example Exercise 17.3 Identifying Oxidizing and Reducing Agents

An oxidation–reduction reaction occurs when a stream of hydrogen gas is passed over hot copper(II) sulfide.

$$CuS(s) + H_2(g) \xrightarrow{\Delta} Cu(s) + H_2S(g)$$

Indicate each of the following for the above redox reaction:

(a) substance oxidized (b) substance reduced

(c) oxidizing agent (d) reducing agent

Solution

By definition, the substance oxidized loses electrons, and its oxidation number increases. The substance reduced gains electrons, and its oxidation number decreases. After assigning oxidation numbers to each atom, we have the following:

$$\overset{+2}{Cu}S(s) + \overset{0}{H_2}(g) \rightarrow \overset{0}{Cu}(s) + \overset{+1}{H_2}S(g)$$

The oxidation number of hydrogen increases from 0 to +1. Thus, H_2 is oxidized. The oxidation number of copper decreases from +2 to 0. Thus, the Cu in CuS is reduced. Note that the oxidation number of S remains constant (–2).

$$\underset{\text{oxidizing agent}}{CuS(s)} + \underset{\text{reducing agent}}{H_2(g)} \rightarrow Cu(s) + H_2S(g)$$

reduction

oxidation

The oxidizing agent is CuS because it causes hydrogen to be oxidized from 0 to +1. The reducing agent is H_2 because it causes copper to be reduced from +2 to 0.

Example Exercise 17.3 Identifying Oxidizing and Reducing Agents

Continued

Practice Exercise

A redox reaction occurs when molten aluminum reacts with iron(III) oxide.

$$Fe_2O_3(l) + 2\,Al(l) \xrightarrow{\Delta} 2\,Fe(l) + Al_2O_3(l)$$

Indicate each of the following for the preceding redox reaction:

(a) substance oxidized (b) substance reduced

(c) oxidizing agent (d) reducing agent

Answers: (a) Al; (b) Fe_2O_3; (c) Fe_2O_3; (d) Al

Concept Exercise

Identify the oxidizing agent and reducing agent in the following redox reaction:

$$M(s) + 2\,HX(aq) \longrightarrow MX_2(aq) + H_2(g)$$

Answer: See Appendix G.

Example Exercise 17.4 Identifying Oxidizing and Reducing Agents

The amount of iodine in a solution can be determined by a redox method using a sulfite solution:

$$I_2(s) + SO_3^{2-}(aq) + H_2O(l) \longrightarrow 2\,I^-(aq) + SO_4^{2-}(aq) + 2\,H^+(aq)$$

Indicate each of the following for the preceding reaction:

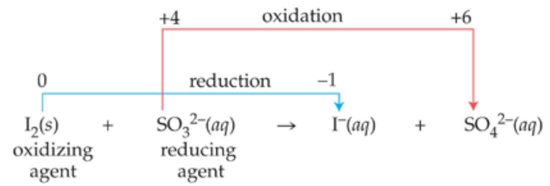
(a) substance oxidized (b) substance reduced

(c) oxidizing agent (d) reducing agent

Solution

Notice that iodine solid is converted to iodide ion in an aqueous solution. Since I_2 gains electrons, it is *reduced* and I_2 is the *oxidizing agent.*

The reducing agent is not as obvious. If we calculate the oxidation number for sulfur in SO_3^{2-} and SO_4^{2-}, we find sulfur changes from +4 to +6 and loses electrons. Thus, the sulfur in SO_3^{2-} is *oxidized*, and SO_3^{2-} is the *reducing agent*. We can illustrate the oxidation and reduction processes for the redox reaction as follows:



Example Exercise 17.4 Identifying Oxidizing and Reducing Agents

Continued

Practice Exercise

A redox reaction occurs when the tin(II) ion reacts with the iodate ion as follows:

$$6\,H^+(aq) + 3\,Sn^{2+}(aq) + IO_3^-(aq) \longrightarrow 3\,Sn^{4+}(aq) + I^-(aq) + 3\,H_2O(l)$$

Indicate each of the following for the preceding redox reaction:

(a) substance oxidized (b) substance reduced
(c) oxidizing agent (d) reducing agent

Answers: (a) Sn^{2+}; (b) IO_3^-; (c) IO_3^-; (d) Sn^{2+}

Concept Exercise

Identify the oxidizing agent and reducing agent in the following redox reaction:

$$Y_2(g) + 2\,X^-(aq) \longrightarrow 2\,Y^-(aq) + X_2(g)$$

Answer: See Appendix G.

Example Exercise 17.5 Balancing Redox Equations by Oxidation Number

An industrial blast furnace reduces iron ore, Fe_2O_3, to molten iron. Balance the following redox equation using the oxidation number method:

$$Fe_2O_3(l) + CO(g) \longrightarrow Fe(l) + CO_2(g)$$

Solution

In this reaction, the oxidation number of iron decreases from +3 in Fe_2O_3 to 0 in Fe. Simultaneously, the oxidation number of carbon increases from +2 to +4. Thus, each Fe gains three electrons, while each C loses two electrons. We can diagram the redox process as follows:

+2 C loses 2 e^- +4

+3 Fe gains 3 e^- 0

$$\overset{+3}{Fe_2O_3(l)} + \overset{+2}{CO(g)} \rightarrow \overset{0}{Fe(l)} + \overset{+4}{CO_2(g)}$$

Because the number of electrons gained and lost must be equal, we find the lowest common multiple. In this case, it is 6. Each Fe gains three electrons, so we place the coefficient 3 in front of CO and CO_2.

$$Fe_2O_3(l) + 3\,CO(g) \longrightarrow Fe(l) + 3CO_2(g)$$

Each carbon atom loses two electrons, so we place the coefficient 2 in front of each iron atom. Because Fe_2O_3 has two iron atoms, it does not require a coefficient.

$$Fe_2O_3(l) + 3\,CO(g) \longrightarrow 2\,Fe(l) + 3CO_2(g)$$

Finally, we verify that the equation is balanced. We check (√)each element in the equation:

$$\overset{\surd}{Fe_2}\overset{\surd}{O_3}(l) + 3\,\overset{\surd}{C}\overset{\surd}{O}(g) \longrightarrow 2\,\overset{\surd}{Fe}(l) + 3\overset{\surd}{C}\overset{\surd}{O_2}(g)$$

Because all the elements are balanced, we have a balanced redox equation.

Example Exercise 17.5 Balancing Redox Equations by Oxidation Number

Continued

Practice Exercise

Balance the following redox equation by the oxidation number method:

$$Cl_2O_5(g) + CO(g) \longrightarrow Cl_2(g) + CO_2(g)$$

Answer: $Cl_2O_5(g) + 5\,CO(g) \longrightarrow Cl_2(g) + 5CO_2(g)$

Concept Exercise

Balance the following redox reaction using the oxidation number method:

$$M_2O_3(s) + X_2(g) \longrightarrow MX_3(aq) + O_2(g)$$

Answer: See Appendix G.

Example Exercise 17.6 Balancing Redox Equations by Oxidation Number

Aqueous sodium iodide reacts with a potassium dichromate solution. Write a balanced equation for the following redox reaction:

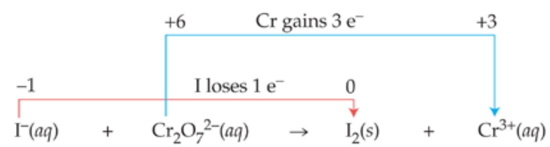
$$H^+(aq) + I^-(aq) + Cr_2O_7^{2-}(aq) \longrightarrow I_2(s) + Cr^{3+}(aq) + H_2O(l)$$

Solution

In this special example, $Cr_2O_7^{2-}$ and I_2 each contain a subscript that affects electron transfer. Note that there are two atoms of chromium in the reactant and two atoms of iodine in the product. Let's balance the chromium and iodine atoms first.

$$H^+(aq) + 2\,I^-(aq) + Cr_2O_7^{2-}(aq) \longrightarrow I_2(s) + 2\,Cr^{3+}(aq) + H_2O(l)$$

In this reaction, the oxidation number of each iodine atom increases from –1 to 0, and the oxidation number of each chromium atom decreases from +6 to +3. We can show the loss and gain of electrons as



There are two chromium atoms, and so the total electron gain is six electrons. Thus, the total electron loss must also be six electrons. Because there are two iodine atoms, and only one iodide ion, we place the coefficients as follows:

$$H^+(aq) + 6\,I^-(aq) + Cr_2O_7^{2-}(aq) \longrightarrow 3\,I_2(s) + 2\,Cr^{3+}(aq) + H_2O(l)$$

Next, we balance the oxygen and hydrogen atoms. Because there are 7 oxygen atoms as reactants, we place the coefficient 7 in front of H_2O. This gives 14 hydrogen atoms, and so we place the coefficient 14 in front of H^+.

$$14\,H^+(aq) + 6\,I^-(aq) + Cr_2O_7^{2-}(aq) \longrightarrow 3\,I_2(s) + 2\,Cr^{3+}(aq) + 7\,H_2O(l)$$

Example Exercise 17.6 Balancing Redox Equations by Oxidation Number

Continued

We can verify that the equation is balanced by checking

$$14\,H^+(aq) + 6\,I^-(aq) + Cr_2O_7^{2-}(aq) \longrightarrow 3I_2(s) + 2\,Cr^{3+}(aq) + 7\,H_2O(l)$$

Last, we verify that the ionic charges are balanced. On the left side of the equation, we have +14 – 6 – 2 = +6. On the right side of the equation, we have +6. Because the ionic charge on each side is +6, the equation is balanced.

Practice Exercise

Write a balanced equation for the following redox reaction:

$$H^+(aq) + MnO_4^-(aq) + NO_2^-(aq) \longrightarrow Mn^{2+}(aq) + NO_3^-(aq) + H_2O(l)$$

Answer: $6\,H^+(aq) + 2\,MnO_4^-(aq) + 5\,NO_2^-(aq) \longrightarrow 2\,Mn^{2+}(aq) + 5\,NO_3^-(aq) + 3\,H_2O(l)$

Concept Exercise

Balance the following redox reaction in an acidic solution using the oxidation number method:

$$MnO_4^-(aq) + SO_3^{2-}(aq) \longrightarrow Mn^{2+}(aq) + SO_4^{2-}(aq)$$

Answer: See Appendix G.

Example Exercise 17.7 Balancing Redox Equations by Half-Reaction

Write a balanced ionic equation for the reaction of iron(II) sulfate and potassium permanganate in acidic solution. The ionic equation is

$$MnO_4^-(aq) + Fe^{2+}(aq) \longrightarrow Fe^{3+}(aq) + Mn^{2+}(aq)$$

Solution

We can balance the redox reaction by the half-reaction method as follows:

Step 1: Because Fe^{2+} is oxidized from +2 to +3, MnO_4^- must be reduced. The two half-reactions are as follows:

$$\text{Oxidation:}\quad Fe^{2+} \longrightarrow Fe^{3+}$$

$$\text{Reduction:}\quad MnO_4^- \longrightarrow Mn^{2+}$$

Step 2: We can balance each half-reaction as follows:

$$Fe^{2+} \longrightarrow Fe^{3+} + e^-$$

$$5\,e^- + 8\,H^+ + MnO_4^- \longrightarrow Mn^{2+} + 4\,H_2O$$

Step 3: Because Fe^{2+} loses 1 e^- and MnO_4^- gains 5 e^-, we multiply the Fe^{2+} half-reaction by 5.

$$5\,Fe^{2+} \longrightarrow 5\,Fe^{3+} + 5\,e^-$$

$$5\,e^- + 8\,H^+ + MnO_4^- \longrightarrow Mn^{2+} + 4\,H_2O$$

Step 4: Then, we add the two half-reactions together and cancel the 5 e^-.

$$5\,Fe^{2+} \longrightarrow 5Fe^{3+} + \cancel{5\,e^-}$$

$$\cancel{5e^-} + 8\,H^+ + MnO_4^- \longrightarrow Mn^{2+} + 4\,H_2O$$

$$5\,Fe^{2+} + 8\,H^+ + MnO_4^- \longrightarrow 5\,Fe^{3+} + Mn^{2+} + 4\,H_2O$$

Step 5: Finally, let's check the atoms and ionic charges to verify that the equation is balanced. We have

Atoms: 5 Fe, 8 H, 1 Mn, 4 O = 5 Fe, 1 Mn, 8 H, 4 O

Charges: $5(+2) + 8(+1) + (-1) = 5(+3) + (+2)$

$$+17 = +17$$

Example Exercise 17.7 Balancing Redox Equations by Half-Reaction

Continued

Because the atoms and ionic charges are equal for reactants and products, the redox equation is balanced.

Practice Exercise

Write a balanced ionic equation for the reaction of sodium nitrite and potassium permanganate in an acidic solution. The ionic equation is

$$MnO_4^-(aq) + NO_2^-(aq) \longrightarrow Mn^{2+}(aq) + NO_3^-(aq)$$

Answer: $6\,H^+ + 2\,MnO_4^- + 5\,NO_2^- \longrightarrow 2\,Mn^{2+} + 5\,NO_3^- + 3\,H_2O$

Concept Exercise

Balance the following redox reaction in a basic solution using the half-reaction method:

$$MnO_4^-(aq) + SO_3^{2-}(aq) \longrightarrow MnO_2(s) + SO_4^{2-}(aq)$$

Answer: See Appendix G.

Example Exercise 17.8 Predicting Spontaneous Redox Reactions

Predict whether the following reaction is spontaneous or nonspontaneous:

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

Solution

Let's refer to the table of reduction potentials to predict whether or not the reaction is spontaneous. Table 17.3 lists $Ni^{2+}(aq)$ as a weaker oxidizing agent than $Sn^{2+}(aq)$. Moreover, $Sn(s)$ is a weaker reducing agent than $Ni(s)$.

$Ni^{2+}(aq)$	+	$Sn(s)$	$\longrightarrow$	$Ni(s)$	+	$Sn^{2+}(aq)$
weaker oxidizing agent		weaker reducing agent		stronger reducing agent		stronger oxidizing agent

Because the reactants are the weaker pair of oxidizing and reducing agents, the reaction is *nonspontaneous*. Conversely, the reverse reaction is spontaneous because the products are the stronger oxidizing and reducing agents.

TABLE 17.3 REDUCTION POTENTIALS OF SELECTED OXIDIZING AND REDUCING AGENTS

STRONGEST Oxidizing Agent — Weakest Reducing Agent

$F_2(g)$	$+ 2\,e^- \rightarrow$	$2\,F^-(aq)$
$Cl_2(g)$	$+ 2\,e^- \rightarrow$	$2\,Cl^-(aq)$
$Br_2(l)$	$+ 2\,e^- \rightarrow$	$2\,Br^-(aq)$
$Ag^+(aq)$	$+ e^- \rightarrow$	$Ag(s)$
$Fe^{3+}(aq)$	$+ e^- \rightarrow$	$Fe^{2+}(aq)$
$I_2(s)$	$+ 2\,e^- \rightarrow$	$2\,I^-(aq)$
$Cu^{2+}(aq)$	$+ 2\,e^- \rightarrow$	$Cu(s)$
$2\,H^+(aq)$	**$+ 2\,e^- \rightarrow$**	**$H_2(g)$**
$Pb^{2+}(aq)$	$+ 2\,e^- \rightarrow$	$Pb(s)$
$Sn^{2+}(aq)$	$+ 2\,e^- \rightarrow$	$Sn(s)$
$Ni^{2+}(aq)$	$+ 2\,e^- \rightarrow$	$Ni(s)$
$Fe^{2+}(aq)$	$+ 2\,e^- \rightarrow$	$Fe(s)$
$Cr^{3+}(aq)$	$+ 3\,e^- \rightarrow$	$Cr(s)$
$Zn^{2+}(aq)$	$+ 2\,e^- \rightarrow$	$Zn(s)$
$Mn^{2+}(aq)$	$+ 2\,e^- \rightarrow$	$Mn(s)$
$Al^{3+}(aq)$	$+ 3\,e^- \rightarrow$	$Al(s)$
$Mg^{2+}(aq)$	$+ 2\,e^- \rightarrow$	$Mg(s)$
$Na^+(aq)$	$+ e^- \rightarrow$	$Na(s)$
$Ca^{2+}(aq)$	$+ 2\,e^- \rightarrow$	$Ca(s)$
$K^+(aq)$	$+ e^- \rightarrow$	$K(s)$
$Li^+(aq)$	$+ e^- \rightarrow$	$Li(s)$

Weakest Oxidizing Agent — STRONGEST Reducing Agent

Example Exercise 17.8 Predicting Spontaneous Redox Reactions

Continued

Practice Exercise

Predict whether the following reaction is spontaneous or nonspontaneous:

$$Ni^{2+}(aq) + Al(s) \longrightarrow Ni(s) + Al^{3+}(aq)$$

Answer: spontaneous

Concept Exercise

Refer to Table 17.3 and predict which of the following metals reacts spontaneously in an aqueous $FeSO_4$ solution:

(a) Ag (b) Al

(c) Cr (d) Ni

Answer: See Appendix G.

TABLE 17.3 REDUCTION POTENTIALS OF SELECTED OXIDIZING AND REDUCING AGENTS

STRONGEST Oxidizing Agent — Weakest Reducing Agent

$F_2(g)$	$+ 2\,e^- \rightarrow$	$2\,F^-(aq)$
$Cl_2(g)$	$+ 2\,e^- \rightarrow$	$2\,Cl^-(aq)$
$Br_2(l)$	$+ 2\,e^- \rightarrow$	$2\,Br^-(aq)$
$Ag^+(aq)$	$+ e^- \rightarrow$	$Ag(s)$
$Fe^{3+}(aq)$	$+ e^- \rightarrow$	$Fe^{2+}(aq)$
$I_2(s)$	$+ 2\,e^- \rightarrow$	$2\,I^-(aq)$
$Cu^{2+}(aq)$	$+ 2\,e^- \rightarrow$	$Cu(s)$
$2\,H^+(aq)$	**$+ 2\,e^- \rightarrow$**	**$H_2(g)$**
$Pb^{2+}(aq)$	$+ 2\,e^- \rightarrow$	$Pb(s)$
$Sn^{2+}(aq)$	$+ 2\,e^- \rightarrow$	$Sn(s)$
$Ni^{2+}(aq)$	$+ 2\,e^- \rightarrow$	$Ni(s)$
$Fe^{2+}(aq)$	$+ 2\,e^- \rightarrow$	$Fe(s)$
$Cr^{3+}(aq)$	$+ 3\,e^- \rightarrow$	$Cr(s)$
$Zn^{2+}(aq)$	$+ 2\,e^- \rightarrow$	$Zn(s)$
$Mn^{2+}(aq)$	$+ 2\,e^- \rightarrow$	$Mn(s)$
$Al^{3+}(aq)$	$+ 3\,e^- \rightarrow$	$Al(s)$
$Mg^{2+}(aq)$	$+ 2\,e^- \rightarrow$	$Mg(s)$
$Na^+(aq)$	$+ e^- \rightarrow$	$Na(s)$
$Ca^{2+}(aq)$	$+ 2\,e^- \rightarrow$	$Ca(s)$
$K^+(aq)$	$+ e^- \rightarrow$	$K(s)$
$Li^+(aq)$	$+ e^- \rightarrow$	$Li(s)$

Weakest Oxidizing Agent — STRONGEST Reducing Agent

Example Exercise 17.9 Voltaic Cells—Spontaneous Processes

Nickel can react with an aqueous silver nitrate solution according to the following ionic equation:

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

Assume the half-reactions are separated into two compartments. A Ni electrode is placed in a compartment with 1.00 *M* $Ni(NO_3)_2$, and a Ag electrode is placed in a compartment with 1.00 *M* $AgNO_3$. Indicate each of the following:

(a) oxidation half-cell reaction
(b) reduction half-cell reaction
(c) anode and cathode
(d) direction of electron flow
(e) direction of NO_3^- in the salt bridge

Solution

Referring to Table 17.3, we see that Ag^+ has a higher reduction potential than Ni^{2+}. Therefore, the process is spontaneous and Ni is oxidized as Ag^+ is reduced. The two half-cell processes are

(a) Oxidation: $Ni \longrightarrow Ni^{2+} + 2\,e^-$

(b) Reduction: $Ag^+ + e^- \longrightarrow Ag$

(c) The anode is where oxidation occurs; thus, Ni is the anode. The cathode is where reduction occurs; thus, Ag is the cathode.

(d) Ni loses electrons while Ag^+ gains electrons, so the direction of electron flow is from Ni to Ag. (*Electrons flow from the anode to the cathode*.)

(e) As the Ni anode loses electrons, the Ni compartment acquires a net positive charge due to Ni^{2+}. As the Ag cathode gains electrons, the Ag compartment acquires a net negative charge due to excess NO_3^-. A salt bridge allows the cell to operate continuously as NO_3^- ions travel from the Ag compartment to the Ni compartment. (*Anions flow from the cathode to the anode*.)

TABLE 17.3 REDUCTION POTENTIALS OF SELECTED OXIDIZING AND REDUCING AGENTS

STRONGEST Oxidizing Agent		Weakest Reducing Agent
$F_2(g)$	$+ 2\,e^- \rightarrow$	$2\,F^-(aq)$
$Cl_2(g)$	$+ 2\,e^- \rightarrow$	$2\,Cl^-(aq)$
$Br_2(l)$	$+ 2\,e^- \rightarrow$	$2\,Br^-(aq)$
$Ag^+(aq)$	$+ e^- \rightarrow$	$Ag(s)$
$Fe^{3+}(aq)$	$+ e^- \rightarrow$	$Fe^{2+}(aq)$
$I_2(s)$	$+ 2\,e^- \rightarrow$	$2\,I^-(aq)$
$Cu^{2+}(aq)$	$+ 2\,e^- \rightarrow$	$Cu(s)$
$2\,H^+(aq)$	**$+ 2\,e^- \rightarrow$**	**$H_2(g)$**
$Pb^{2+}(aq)$	$+ 2\,e^- \rightarrow$	$Pb(s)$
$Sn^{2+}(aq)$	$+ 2\,e^- \rightarrow$	$Sn(s)$
$Ni^{2+}(aq)$	$+ 2\,e^- \rightarrow$	$Ni(s)$
$Fe^{2+}(aq)$	$+ 2\,e^- \rightarrow$	$Fe(s)$
$Cr^{3+}(aq)$	$+ 3\,e^- \rightarrow$	$Cr(s)$
$Zn^{2+}(aq)$	$+ 2\,e^- \rightarrow$	$Zn(s)$
$Mn^{2+}(aq)$	$+ 2\,e^- \rightarrow$	$Mn(s)$
$Al^{3+}(aq)$	$+ 3\,e^- \rightarrow$	$Al(s)$
$Mg^{2+}(aq)$	$+ 2\,e^- \rightarrow$	$Mg(s)$
$Na^+(aq)$	$+ e^- \rightarrow$	$Na(s)$
$Ca^{2+}(aq)$	$+ 2\,e^- \rightarrow$	$Ca(s)$
$K^+(aq)$	$+ e^- \rightarrow$	$K(s)$
$Li^+(aq)$	$+ e^- \rightarrow$	$Li(s)$
Weakest Oxidizing Agent		STRONGEST Reducing Agent

Example Exercise 17.9 Voltaic Cells—Spontaneous Processes

Continued

Practice Exercise

Iron can react with an aqueous tin(II) nitrate solution according to the following ionic equation:

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

An Fe electrode is placed in a compartment with 1.00 *M* $Fe(NO_3)_2$, and a Sn electrode is placed in another compartment with 1.00 *M* $Sn(NO_3)_2$. Indicate each of the following:

(a) oxidation half-cell reaction
(b) reduction half-cell reaction
(c) anode and cathode
(d) direction of electron flow
(e) direction of NO_3^- in the salt bridge

Answers:

(a) Oxidation $Fe \longrightarrow Fe^{2+} + 2\,e^-$
(b) Reduction $Sn^{2+} + 2\,e^- \longrightarrow Sn$
(c) Fe is the anode; Sn is the cathode.
(d) Electrons flow from the Fe anode to the Sn cathode.
(e) A salt bridge allows NO_3^- ions to travel from the Sn cathode compartment to the Fe anode compartment.

Concept Exercise

Draw and illustrate the voltaic cell described in the practice exercise. Refer to Figure 17.4 and label the anode and cathode; show the direction of electron and anion flow.

Answer: See Appendix G.

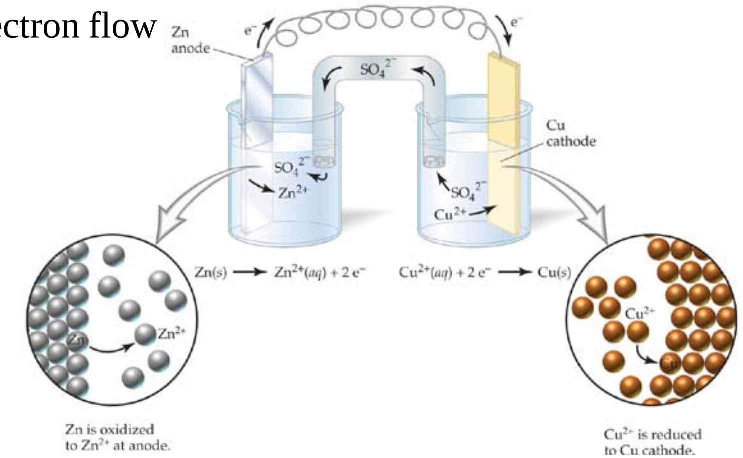


Figure 17.4 Voltaic Cell The two half-cells are connected by a salt bridge, and the negative sulfate ions travel from the right half-cell to the left half-cell. The salt bridge reduces the positive charge buildup in the left half-cell and the negative charge buildup in the right half-cell. The cell continues to operate spontaneously as electrons flow from the zinc anode to the copper cathode.

Example Exercise 17.10 Electrolytic Cells—Nonspontaneous Processes

Aluminum metal is produced by passing an electric current through bauxite, which contains Al_2O_3 dissolved in the molten mineral cryolite. Graphite rods serve as electrodes, and we can write the redox equation as follows:

$$3\,C(s) + 2\,Al_2O_3(l) \xrightarrow{\text{electricity}} 4\,Al(l) + 3\,CO_2(g)$$

Indicate each of the following for this nonspontaneous process:

(a) oxidation half-cell reaction
(b) reduction half-cell reaction
(c) anode and cathode
(d) direction of electron flow

Solution

In the equation, we see that C is oxidized to CO_2 and that Al_2O_3 is reduced to Al metal. The two half-cell processes are

(a) Oxidation: $C + 2\,O^{2-} \longrightarrow CO_2 + 4\,e^-$

(b) Reduction: $Al^{3+} + 3\,e^- \longrightarrow Al$

(c) Oxidation occurs at the graphite anode, where CO_2 is released. Reduction occurs at the graphite cathode, where Al metal is produced.

(d) The electrons flow from the anode, where CO_2 gas is released, to the cathode, where Al metal is produced.

Practice Exercise

Magnesium metal is produced by passing an electric current through molten $MgCl_2$ obtained from evaporated seawater. Carbon and platinum rods serve as electrodes, and we can write the redox equation as follows:

$$C(s) + 2\,MgCl_2(l) \xrightarrow{\text{electricity}} 2\,Mg(l) + CCl_4(g)$$

Example Exercise 17.10 Electrolytic Cells—Nonspontaneous Processes

Continued

Indicate each of the following for this nonspontaneous process:

(a) oxidation half-cell reaction (b) reduction half-cell reaction
(c) anode and cathode (d) direction of electron flow

Answers:

(a) Oxidation: $C + 4\,Cl^- \longrightarrow CCl_4 + 4\,e^-$
(b) Reduction: $Mg^{2+} + 2\,e^- \longrightarrow Mg$
(c) The C electrode is the anode; the Pt electrode is the cathode.
(d) The electrons flow from the C anode to the Pt cathode.

Concept Exercise

Draw and illustrate the electrolytic cell described in the practice exercise. Refer to Figure 17.6 and label the anode and cathode; show the direction of electron flow.

Answer: See Appendix G.

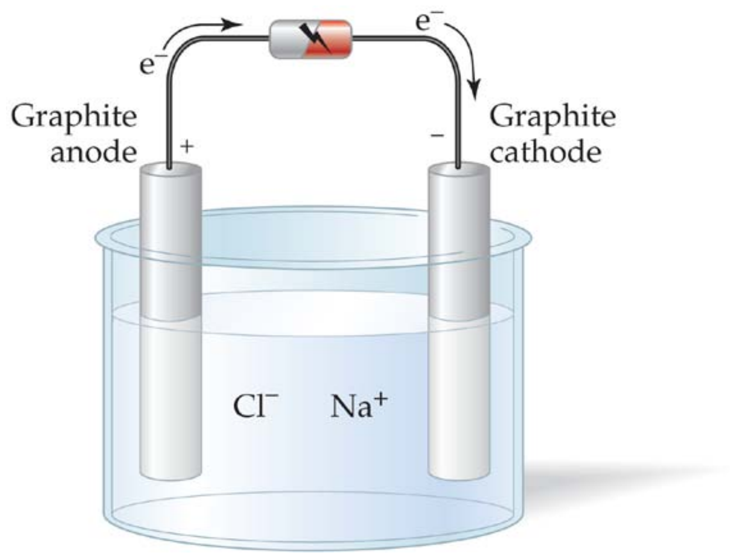


Figure 17.6 Electrolytic Cell An electrolytic cell can produce sodium metal and chlorine gas from molten NaCl. A source of direct current forces electrons to travel toward the cathode for the reduction of sodium ions, as electrons are released from the anode from the oxidation of chloride ions.