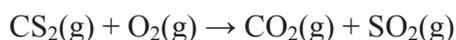


**Balancing a Redox Equation for an Acidic Solution  
Using the Oxidation Number Method  
(Student textbook page 615)**

35. Use the oxidation number method to balance the following equation for the combustion of carbon disulfide:



**What Is Required?**

You must balance an equation for the combustion of carbon disulfide, using the oxidation number method.

**What Is Given?**

You are given the reactants, carbon disulfide and oxygen, and the products, carbon dioxide and sulfur dioxide.

You are given rules for assigning oxidation numbers in **Table 9.3** on page 604 of the student textbook.

Plan Your Strategy	Act on Your Strategy
<p><b>Step 1</b> Assign an oxidation number to each atom in the equation and determine whether it is a redox reaction.</p>	$\overset{+4}{\text{C}}\overset{-2}{\text{S}}_2(\text{g}) + \overset{0}{\text{O}}_2(\text{g}) \rightarrow \overset{+4}{\text{C}}\overset{-2}{\text{O}}_2(\text{g}) + \overset{+4}{\text{S}}\overset{-2}{\text{O}}_2(\text{g})$ <p>The oxidation numbers of sulfur and oxygen change during the reaction. It is a redox reaction.</p>
<p><b>Step 2</b> Identify the atom or atoms that undergo an increase in oxidation number and the atom or atoms that undergo a decrease in oxidation number.</p>	<p>Each of two sulfur atoms lost 6 electrons and each of two oxygen atoms gained 2 electrons. So the <math>\text{O}_2</math> lost 12 electrons and the <math>\text{CS}_2</math> gained 4 electrons.</p> $\begin{array}{c} \text{1}(-12\text{e}^-) = -12\text{e}^- \\ \downarrow \\ \overset{+4}{\text{C}}\overset{-2}{\text{S}}_2(\text{g}) + \overset{0}{\text{O}}_2(\text{g}) \rightarrow \overset{+4}{\text{C}}\overset{-2}{\text{O}}_2(\text{g}) + \overset{+4}{\text{S}}\overset{-2}{\text{O}}_2(\text{g}) \\ \uparrow \qquad \qquad \qquad \uparrow \qquad \qquad \qquad \uparrow \\ \text{3}(+4\text{e}^-) = +12\text{e}^- \end{array}$

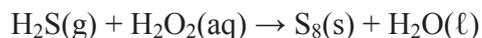
<p><b>Step 3</b> Balance the atoms that undergo a change in oxidation number. Place coefficients in front of the reactants that will balance the number of electrons that are lost by one reactant and gained by another.</p>	<p>A coefficient of 3 is needed for the O<sub>2</sub>. A coefficient of 1 is already implied for the CS<sub>2</sub>. This ratio of 1:3 must be maintained for the remainder of the balancing.</p> $\text{CS}_2(\text{g}) + 3\text{O}_2(\text{g}) \rightarrow \text{CO}_2(\text{g}) + \text{SO}_2(\text{g})$
<p><b>Step 4</b> Determine the coefficients of the products that are needed to complete the balancing of the equation.</p>	<p>There are 2 S atoms on the left and 1 on the right. Place a coefficient of 2 in front of the SO<sub>2</sub> in the products.</p> $\text{CS}_2(\text{g}) + 3\text{O}_2(\text{g}) \rightarrow \text{CO}_2(\text{g}) + 2\text{SO}_2(\text{g})$ <p>Atoms of all elements are now balanced.</p>

### Check Your Solution

Atoms of all of the elements are balanced. The total charge on the left side is 0 and the total charge on the right side is also 0. The charge is balanced.



37. Use the oxidation number method to balance the following equation:



**What Is Required?**

You must balance an equation for the given redox reaction using the oxidation number method.

**What Is Given?**

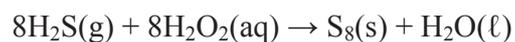
You are given the reaction:  $\text{H}_2\text{S}(\text{g}) + \text{H}_2\text{O}_2(\text{aq}) \rightarrow \text{S}_8(\text{s}) + \text{H}_2\text{O}(\ell)$

You are given rules for assigning oxidation numbers in **Table 9.3** on page 604 of the student textbook.

Plan Your Strategy	Act on Your Strategy
<p><b>Step 1</b> Assign an oxidation number to each atom in the equation and determine whether it is a redox reaction.</p>	$\overset{+1}{\text{H}}_2\overset{-2}{\text{S}}(\text{g}) + \overset{+1}{\text{H}}_2\overset{-1}{\text{O}}_2(\text{aq}) \longrightarrow \overset{0}{\text{S}}_8(\text{s}) + \overset{+1}{\text{H}}_2\overset{-2}{\text{O}}(\ell)$ <p>The oxidation numbers of sulfur and oxygen change during the reaction. It is a redox reaction.</p>
<p><b>Step 2</b> Identify the atom or atoms that undergo an increase in oxidation number and the atom or atoms that undergo a decrease in oxidation number.</p>	<p>One sulfur atom lost 2 electrons and each of two oxygen atoms gained 1 electron for a total of 2.</p> $\overset{+1}{\text{H}}_2\overset{-2}{\text{S}}(\text{g}) + \overset{+1}{\text{H}}_2\overset{-1}{\text{O}}_2(\text{aq}) \longrightarrow \overset{0}{\text{S}}_8(\text{s}) + \overset{+1}{\text{H}}_2\overset{-2}{\text{O}}(\ell)$
<p><b>Step 3</b> Balance the atoms that undergo a change in oxidation number. Place coefficients in front of the reactants that will balance the number of electrons that are lost by one reactant and gained by another.</p>	<p>Coefficients of 1 are already implied for <math>\text{H}_2\text{S}</math> and <math>\text{H}_2\text{O}_2</math>. This ratio of 1:1 must be maintained while balancing the products.</p> $\text{H}_2\text{S}(\text{g}) + \text{H}_2\text{O}_2(\text{aq}) \rightarrow \text{S}_8(\text{s}) + \text{H}_2\text{O}(\ell)$

**Step 4** Determine the coefficients of the products that are needed to complete the balancing of the equation.

There are 8 S atoms on the right so a coefficient of 8 is needed for the  $\text{H}_2\text{S}$  on the left. If the  $\text{H}_2\text{S}$  has a coefficient of 8, then  $\text{H}_2\text{O}_2$  must also have a coefficient of 8 in order to maintain the 1:1 ratio.



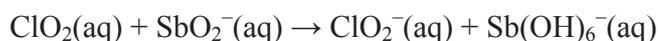
Now there are 32 H atoms on the left and 2 on the right. Give the  $\text{H}_2\text{O}$  on the right a coefficient of 16.  $8\text{H}_2\text{S}(\text{g}) + 8\text{H}_2\text{O}_2(\text{aq}) \rightarrow \text{S}_8(\text{s}) + 16\text{H}_2\text{O}(\ell)$

Atoms of all elements are now balanced.

**Check Your Solution**

Atoms of all of the elements are balanced. The total charge on the left side is 0 and the total charge on the right side is also 0. The charge is balanced.

38. Use the oxidation number method to balance the following equation:



### What Is Required?

You must balance an equation for the given redox reaction using the oxidation number method.

### What Is Given?

You are given the reaction:  $\text{ClO}_2(\text{aq}) + \text{SbO}_2^-(\text{aq}) \rightarrow \text{ClO}_2^-(\text{aq}) + \text{Sb}(\text{OH})_6^-(\text{aq})$

You are given rules for assigning oxidation numbers in **Table 9.3** on page 604 of the student textbook.

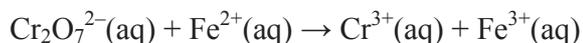
Plan Your Strategy	Act on Your Strategy
<p><b>Step 1</b> Assign an oxidation number to each atom in the equation and determine whether it is a redox reaction.</p>	$\overset{+4}{\text{Cl}}\overset{-2}{\text{O}_2}(\text{aq}) + \overset{+3}{\text{Sb}}\overset{-2}{\text{O}_2}(\text{aq}) \rightarrow \overset{+3}{\text{Cl}}\overset{-2}{\text{O}_2}(\text{aq}) + \overset{+5}{\text{Sb}}\overset{-2}{\text{O}}\overset{+1}{\text{H}}_6(\text{aq})$ <p>The oxidation numbers of chlorine and antimony change during the reaction. It is a redox reaction.</p>
<p><b>Step 2</b> Identify the atom or atoms that undergo an increase in oxidation number and the atom or atoms that undergo a decrease in oxidation number.</p>	<p>The chlorine atom gained 1 electron and the antimony atom lost 2 electrons.</p> $\overset{+4}{\text{Cl}}\overset{-2}{\text{O}_2}(\text{aq}) + \overset{+3}{\text{Sb}}\overset{-2}{\text{O}_2}(\text{aq}) \rightarrow \overset{+3}{\text{Cl}}\overset{-2}{\text{O}_2}(\text{aq}) + \overset{+5}{\text{Sb}}\overset{-2}{\text{O}}\overset{+1}{\text{H}}_6(\text{aq})$ <p style="text-align: center;"> <math>2(+1e^-) = +2e^-</math> (above the ClO<sub>2</sub> species)  <math>1(-2e^-) = -2e^-</math> (below the Sb(OH)<sub>6</sub> species)         </p>
<p><b>Step 3</b> Determine the numerical values of the total increase and decrease in oxidation numbers. Place coefficients in front of the reactants that will balance the number of electrons that are lost by one reactant and gained by another.</p>	<p>The ClO<sub>2</sub> needs a coefficient of 2 and a coefficient of 1 is already implied for the SbO<sub>2</sub><sup>-</sup>. This ratio of 2:1 must be maintained for the remainder of the balancing.</p> $2\text{ClO}_2(\text{aq}) + \text{SbO}_2^-(\text{aq}) \rightarrow \text{ClO}_2^-(\text{aq}) + \text{Sb}(\text{OH})_6^-(\text{aq})$
<p><b>Step 4</b> The presence of the Sb(OH)<sub>6</sub><sup>-</sup>(aq) on the right implies that the reaction was carried out in a basic solution. Continue to balance the equation according to basic conditions. Balance atoms of all elements except oxygen and hydrogen.</p>	<p>There are 2 Cl atoms on the left and one on the right. Give the ClO<sub>2</sub><sup>-</sup> a coefficient of 2.</p> $2\text{ClO}_2(\text{aq}) + \text{SbO}_2^-(\text{aq}) \rightarrow 2\text{ClO}_2^-(\text{aq}) + \text{Sb}(\text{OH})_6^-(\text{aq})$

<p><b>Step 5</b> Add water molecules to balance oxygen atoms.</p>	<p>There are 6 O atoms on the left and 10 on the right. Add 4 water molecules to the left side of the equation.</p> $2\text{ClO}_2(\text{aq}) + \text{SbO}_2^-(\text{aq}) + 4\text{H}_2\text{O}(\ell) \rightarrow 2\text{ClO}_2^-(\text{aq}) + \text{Sb}(\text{OH})_6^-(\text{aq})$
<p><b>Step 6</b> Balance the hydrogen atoms by adding hydrogen ions.</p>	<p>There are 8 H atoms on the left and 6 on the right. Add 2 H ions to the right side.</p> $2\text{ClO}_2(\text{aq}) + \text{SbO}_2^-(\text{aq}) + 4\text{H}_2\text{O}(\ell) \rightarrow 2\text{ClO}_2^-(\text{aq}) + \text{Sb}(\text{OH})_6^-(\text{aq}) + 2\text{H}^+(\text{aq})$
<p><b>Step 7</b> Add as many <math>\text{OH}^-</math> ions to both sides as there are <math>\text{H}^+</math> ions. On the right, combine the <math>\text{OH}^-</math> and <math>\text{H}^+</math> ions to make water.</p>	<p>There are 2 <math>\text{H}^+</math> ions on the right. Add 2 <math>\text{OH}^-</math> to both sides.</p> $2\text{ClO}_2(\text{aq}) + \text{SbO}_2^-(\text{aq}) + 4\text{H}_2\text{O}(\ell) + 2(\text{OH})^-(\text{aq}) \rightarrow 2\text{ClO}_2^-(\text{aq}) + \text{Sb}(\text{OH})_6^-(\text{aq}) + 2\text{H}_2\text{O}(\ell)$
<p><b>Step 8</b> Cancel the number of water molecules that are on both sides of the equation.</p>	<p>There are 2 water molecules on the right and 4 on the left. Delete 2 water molecules from each side of the equation.</p> $2\text{ClO}_2(\text{aq}) + \text{SbO}_2^-(\text{aq}) + 2\text{H}_2\text{O}(\ell) + 2(\text{OH})^-(\text{aq}) \rightarrow 2\text{ClO}_2^-(\text{aq}) + \text{Sb}(\text{OH})_6^-(\text{aq})$ <p>Atoms of all elements are now balanced.</p>

### Check Your Solution

Atoms of all of the elements are balanced. The total charge on the left side is  $-3$  and the total charge on the right side is also  $-3$ . The charge is balanced.

39. Use the oxidation number method to balance the ionic equation in an acidic solution:



### What Is Required?

You must balance an equation for the given redox reaction using the oxidation number method in an acidic solution.

### What Is Given?

You are given the reaction:  $\text{Cr}_2\text{O}_7^{2-}(\text{aq}) + \text{Fe}^{2+}(\text{aq}) \rightarrow \text{Cr}^{3+}(\text{aq}) + \text{Fe}^{3+}(\text{aq})$

You are given rules for assigning oxidation numbers in **Table 9.3** on page 604 of the student textbook.

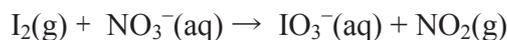
Plan Your Strategy	Act on Your Strategy
<p><b>Step 1</b> Assign an oxidation number to each atom in the equation and determine whether it is a redox reaction.</p>	$\overset{+6}{\text{Cr}_2}\overset{-2}{\text{O}_7}(\text{aq}) + \overset{+2}{\text{Fe}^{2+}}(\text{aq}) \rightarrow \overset{+3}{2\text{Cr}^{3+}}(\text{aq}) + \overset{+3}{\text{Fe}^{3+}}(\text{aq})$ <p>The oxidation numbers of chromium and iron change during the reaction. It is a redox reaction.</p>
<p><b>Step 2</b> Identify the atom or atoms that undergo an increase in oxidation number and the atom or atoms that undergo a decrease in oxidation number.</p>	<p>Each of two chromium atoms gains 3 electrons for a total of 6. The iron ion loses 1 electron.</p> $\begin{array}{c} \xrightarrow{1(+6e^-) = +6e^-} \\ \text{Cr}_2\text{O}_7^{2-}(\text{aq}) + \text{Fe}^{2+}(\text{aq}) \rightarrow 2\text{Cr}^{3+}(\text{aq}) + \text{Fe}^{3+}(\text{aq}) \\ \xleftarrow{6(-1e^-) = -6e^-} \end{array}$
<p><b>Step 3</b> Balance the atoms that undergo a change in oxidation number. Place coefficients in front of the reactants that will balance the number of electrons that are lost by one reactant and gained by another.</p>	<p>Place a coefficient of 6 in front of the iron ion and a coefficient of 1 is already implied for the <math>\text{Cr}_2\text{O}_7^{2-}</math> ion. This ratio of 1:6 must be maintained for the remainder of the balancing.</p> $\text{Cr}_2\text{O}_7^{2-}(\text{aq}) + 6\text{Fe}^{2+}(\text{aq}) \rightarrow \text{Cr}^{3+}(\text{aq}) + \text{Fe}^{3+}(\text{aq})$
<p><b>Step 4</b> Balance atoms of all elements except oxygen and hydrogen.</p>	<p>There are 2 Cr atoms on the left and 1 on the right. There are 6 <math>\text{Fe}^{2+}</math> ions on the left and 1 on the right. Place a coefficient of 2 in front of the <math>\text{Cr}^{3+}</math> ion and a 6 in front of the <math>\text{Fe}^{3+}</math> ion in the products.</p> $\text{Cr}_2\text{O}_7^{2-}(\text{aq}) + 6\text{Fe}^{2+}(\text{aq}) \rightarrow 2\text{Cr}^{3+}(\text{aq}) + 6\text{Fe}^{3+}(\text{aq})$

<p><b>Step 5</b> Balance the O atoms by adding H<sub>2</sub>O.</p>	<p>There are 7 O atoms on the left and none on the right so add 7 water molecules to the right side of the equation.</p> $\text{Cr}_2\text{O}_7^{2-}(\text{aq}) + 6\text{Fe}^{2+}(\text{aq}) \rightarrow 2\text{Cr}^{3+}(\text{aq}) + 6\text{Fe}^{3+}(\text{aq}) + 7\text{H}_2\text{O}(\ell)$
<p><b>Step 6</b> Balance the hydrogen atoms by adding hydrogen ions.</p>	<p>There are 14 H atoms on the right and none on the left, so add 14 H<sup>+</sup> ions to the left side of the equation.</p> $\text{Cr}_2\text{O}_7^{2-}(\text{aq}) + 6\text{Fe}^{2+}(\text{aq}) + 14\text{H}^+ \rightarrow 2\text{Cr}^{3+}(\text{aq}) + 6\text{Fe}^{3+}(\text{aq}) + 7\text{H}_2\text{O}(\ell)$ <p>Atoms of all elements are now balanced.</p>

**Check Your Solution**

Atoms of all of the elements are balanced. The total charge on the left side is +24 and the total charge on the right side is also +24. The charge is balanced.

40. Use the oxidation number method to balance the ionic equation in an acidic solution:



**What Is Required?**

You must balance an equation for the given redox reaction using the oxidation number method in an acidic solution.

**What Is Given?**

You are given the reaction:  $\text{I}_2(\text{g}) + \text{NO}_3^-(\text{aq}) \rightarrow \text{IO}_3^-(\text{aq}) + \text{NO}_2(\text{g})$

You are given rules for assigning oxidation numbers in **Table 9.3** on page 604 of the student textbook.

Plan Your Strategy	Act on Your Strategy
<p><b>Step 1</b> Assign an oxidation number to each atom in the equation and determine whether it is a redox reaction.</p>	$\text{I}_2(\text{g}) + \text{NO}_3^-(\text{aq}) \rightarrow \text{IO}_3^-(\text{aq}) + \text{NO}_2(\text{g})$ <p style="text-align: center;"> <math>\begin{matrix} 0 &amp; +5 &amp; -2 &amp; &amp; +5 &amp; -2 &amp; &amp; +4 &amp; -2 \end{matrix}</math> </p> <p>The oxidation numbers of iodine and nitrogen change during the reaction. It is a redox reaction.</p>
<p><b>Step 2</b> Identify the atom or atoms that undergo an increase in oxidation number and the atom or atoms that undergo a decrease in oxidation number.</p>	<p>Each of 2 I atoms in <math>\text{I}_2</math> loses 5 electrons for a total of 10 electrons. The nitrogen atom in <math>\text{NO}_3^-</math> gains one electron.</p> $1(-10e^-) = -10e^-$ $\text{I}_2(\text{g}) + \text{NO}_3^-(\text{aq}) \rightarrow \text{IO}_3^-(\text{aq}) + \text{NO}_2(\text{g})$ <p style="text-align: center;"> <math>\begin{matrix} 0 &amp; +5 &amp; -2 &amp; &amp; +5 &amp; -2 &amp; &amp; +4 &amp; -2 \end{matrix}</math> </p> $10(+1e^-) = +10e^-$
<p><b>Step 3</b> Balance the atoms that undergo a change in oxidation number. Place coefficients in front of the reactants that will balance the number of electrons that are lost by one reactant and gained by another.</p>	<p>Give a coefficient of 10 to the <math>\text{NO}_3^-</math> and a coefficient of 1 is already implied for the <math>\text{I}_2</math> molecule.</p> $\text{I}_2(\text{g}) + 10\text{NO}_3^-(\text{aq}) \rightarrow \text{IO}_3^-(\text{aq}) + \text{NO}_2(\text{g})$ <p>This ratio of 1:10 must be maintained for the remainder of balancing.</p>

<p><b>Step 4</b> Balance atoms of all elements except oxygen and hydrogen.</p>	<p>There are 2 iodine atoms on the left and 1 on the right so place a coefficient of 2 in front of the iodate ion on the right of the equation. There are 10 nitrogen atoms on the left and 1 on the right so place a coefficient of 10 in front of the NO<sub>2</sub> molecule on the right of the equation.</p> $\text{I}_2(\text{g}) + 10\text{NO}_3^-(\text{aq}) \rightarrow 2\text{IO}_3^-(\text{aq}) + 10\text{NO}_2(\text{g})$
<p><b>Step 5</b> Balance the O atoms by adding H<sub>2</sub>O.</p>	<p>There are 30 O atoms on the left and 26 on the right so add 4 water molecules to the right side of the equation.</p> $\text{I}_2(\text{g}) + 10\text{NO}_3^-(\text{aq}) \rightarrow 2\text{IO}_3^-(\text{aq}) + 10\text{NO}_2(\text{g}) + 4\text{H}_2\text{O}(\ell)$
<p><b>Step 6</b> Balance the hydrogen atoms by adding hydrogen ions</p>	<p>There are 8 hydrogen atoms on the right side and none on the left, so add 8 hydrogen ions to the left side of the equation.</p> $8\text{H}^+(\text{aq}) + \text{I}_2(\text{g}) + 10\text{NO}_3^-(\text{aq}) \rightarrow 2\text{IO}_3^-(\text{aq}) + 10\text{NO}_2(\text{g}) + 4\text{H}_2\text{O}(\ell)$ <p>Atoms of all elements are now balanced.</p>

### Check Your Solution

Atoms of all of the elements are balanced. The total charge on the left side is  $-2$  and the total charge on the right side is also  $-2$ . The charge is balanced.

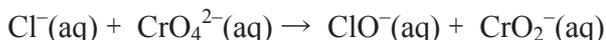


<p><b>Step 4</b> Balance atoms of all elements except oxygen and hydrogen.</p>	<p>The Pb atoms are balanced. There are 2 S atoms on the left and 1 on the right. Place a coefficient of 2 in front of the <math>\text{SO}_4^{2-}</math> on the right side of the equation.</p> $2\text{PbSO}_4(\text{aq}) \rightarrow \text{Pb}(\text{s}) + \text{PbO}_2(\text{aq}) + 2\text{SO}_4^{2-}(\text{aq})$
<p><b>Step 5</b> Balance the oxygen atoms by adding water molecules.</p>	<p>There are 8 atoms of oxygen on the left side and 10 on the right side, so add 2 water molecules to the left side of the equation.</p> $2\text{H}_2\text{O}(\ell) + 2\text{PbSO}_4(\text{aq}) \rightarrow \text{Pb}(\text{s}) + \text{PbO}_2(\text{aq}) + 2\text{SO}_4^{2-}(\text{aq})$
<p><b>Step 6</b> Balance the hydrogen atoms by adding hydrogen ions</p>	<p>There are 4 hydrogen atoms on the left side and none on the right, so add 4 hydrogen ions to the right side of the equation.</p> $2\text{H}_2\text{O}(\ell) + 2\text{PbSO}_4(\text{aq}) \rightarrow \text{Pb}(\text{s}) + \text{PbO}_2(\text{aq}) + 2\text{SO}_4^{2-}(\text{aq}) + 4\text{H}^+(\text{aq})$ <p>Atoms of all elements are now balanced.</p>

### Check Your Solution

Atoms of all of the elements are balanced. The total charge on the left side is 0 and the total charge on the right side is also 0. The charge is balanced.

42. Use the oxidation number method to balance the ionic equation in a basic solution:



### What Is Required?

You must balance an equation for the given redox reaction using the oxidation number method in a basic solution.

### What Is Given?

You are given the reaction:  $\text{Cl}^-(\text{aq}) + \text{CrO}_4^{2-}(\text{aq}) \rightarrow \text{ClO}^-(\text{aq}) + \text{CrO}_2^-(\text{aq})$

You are given rules for assigning oxidation numbers in **Table 9.3** on page 604 of the student textbook.

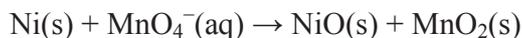
Plan Your Strategy	Act on Your Strategy
<p><b>Step 1</b> Assign an oxidation number to each atom in the equation and determine whether it is a redox reaction.</p>	$\overset{-1}{\text{Cl}}^-(\text{aq}) + \overset{+6}{\text{Cr}}\overset{-2}{\text{O}_4}^{2-}(\text{aq}) \rightarrow \overset{+1}{\text{Cl}}\overset{-2}{\text{O}}^-(\text{aq}) + \overset{+3}{\text{Cr}}\overset{-2}{\text{O}_2}^-(\text{aq})$ <p>The oxidation numbers of chromium and chlorine change during the reaction. It is a redox reaction.</p>
<p><b>Step 2</b> Identify the atom or atoms that undergo an increase in oxidation number and the atom or atoms that undergo a decrease in oxidation number.</p>	<p>The chlorine atom lost 2 electrons and the chromium atom gained 3 electrons.</p> $\begin{array}{c} 3(-2e^-) = -6e^- \\ \downarrow \\ \overset{-1}{\text{Cl}}^-(\text{aq}) + \overset{+6}{\text{Cr}}\overset{-2}{\text{O}_4}^{2-}(\text{aq}) \rightarrow \overset{+1}{\text{Cl}}\overset{-2}{\text{O}}^-(\text{aq}) + \overset{+3}{\text{Cr}}\overset{-2}{\text{O}_2}^-(\text{aq}) \\ \uparrow \\ 2(+3e^-) = +6e^- \end{array}$
<p><b>Step 3</b> Determine the numerical values of the total increase and decrease in oxidation numbers. Place coefficients in front of the reactants that will balance the number of electrons that are lost by one reactant and gained by another.</p>	<p>The <math>\text{Cl}^-</math> needs a coefficient of 3 and the <math>\text{CrO}_4^{2-}</math> needs a coefficient of 2. This ratio of 3:2 must be maintained for the remainder of the balancing.</p> $3\text{Cl}^-(\text{aq}) + 2\text{CrO}_4^{2-}(\text{aq}) \rightarrow \text{ClO}^-(\text{aq}) + \text{CrO}_2^-(\text{aq})$
<p><b>Step 4</b> Balance atoms of all elements except oxygen and hydrogen.</p>	<p>There are 3 Cl atoms on the left and 1 on the right. Place a coefficient of 3 in front of the <math>\text{ClO}^-</math> on the right side of the equation. There are 2 Cr atoms on the left and 1 on the right. Place a coefficient of 2 in front of the <math>\text{CrO}_2^-</math> on the right side of the equation.</p> $3\text{Cl}^-(\text{aq}) + 2\text{CrO}_4^{2-}(\text{aq}) \rightarrow 3\text{ClO}^-(\text{aq}) + 2\text{CrO}_2^-(\text{aq})$

<p><b>Step 5</b> Balance the oxygen atoms by adding water molecules.</p>	<p>There are 8 atoms of oxygen on the left side and 7 on the right side, so add 1 water molecule to the right side.</p> $3\text{Cl}^{-}(\text{aq}) + 2\text{CrO}_4^{2-}(\text{aq}) \rightarrow 3\text{ClO}^{-}(\text{aq}) + 2\text{CrO}_2^{-}(\text{aq}) + \text{H}_2\text{O}(\ell)$
<p><b>Step 6</b> Balance the hydrogen atoms by adding hydrogen ions</p>	<p>There are 2 hydrogen atoms on the right side and none on the left, so add 2 hydrogen ions to the left side.</p> $2\text{H}^{+}(\text{aq}) + 3\text{Cl}^{-}(\text{aq}) + 2\text{CrO}_4^{2-}(\text{aq}) \rightarrow 3\text{ClO}^{-}(\text{aq}) + 2\text{CrO}_2^{-}(\text{aq}) + \text{H}_2\text{O}(\ell)$
<p><b>Step 7</b> Add as many <math>\text{OH}^{-}</math> ions to both sides as there are <math>\text{H}^{+}</math> ions. When <math>\text{H}^{+}</math> and <math>\text{OH}^{-}</math> ions appear on the same side of the equation, combine them to make water.</p>	<p>There are 2 hydrogen ions on the left side, so add 2 hydroxide ions to each side.</p> $2\text{H}_2\text{O}(\ell) + 3\text{Cl}^{-}(\text{aq}) + 2\text{CrO}_4^{2-}(\text{aq}) \rightarrow 3\text{ClO}^{-}(\text{aq}) + 2\text{CrO}_2^{-}(\text{aq}) + \text{H}_2\text{O}(\ell) + 2\text{OH}^{-}(\text{aq})$
<p><b>Step 8</b> Cancel the number of water molecules that are on both sides of the equation.</p>	<p>There are 2 water molecules on the left and 1 on the right. Delete 1 water molecule from each side.</p> $\text{H}_2\text{O}(\ell) + 3\text{Cl}^{-}(\text{aq}) + 2\text{CrO}_4^{2-}(\text{aq}) \rightarrow 3\text{ClO}^{-}(\text{aq}) + 2\text{CrO}_2^{-}(\text{aq}) + 2\text{OH}^{-}(\text{aq})$ <p>Atoms of all elements are now balanced.</p>

### Check Your Solution

Atoms of all of the elements are balanced. The total charge on the left side is  $-7$  and the total charge on the right side is also  $-7$ . The charge is balanced.

43. Use the oxidation number method to balance the ionic equation in a basic solution:



**What Is Required?**

You must balance an equation for the given redox reaction using the oxidation number method in a basic solution.

**What Is Given?**

You are given the reaction:  $\text{Ni(s)} + \text{MnO}_4^-(\text{aq}) \rightarrow \text{NiO(s)} + \text{MnO}_2(\text{s})$

You are given rules for assigning oxidation numbers in **Table 9.3** on page 604 of the student textbook.

Plan Your Strategy	Act on Your Strategy
<p><b>Step 1</b> Assign an oxidation number to each atom in the equation and determine whether it is a redox reaction.</p>	$\overset{0}{\text{Ni}}(\text{s}) + \overset{+7}{\text{Mn}}\overset{-2}{\text{O}_4}(\text{aq}) \rightarrow \overset{+2}{\text{Ni}}\overset{-2}{\text{O}}(\text{s}) + \overset{+4}{\text{Mn}}\overset{-2}{\text{O}_2}(\text{s})$ <p>The oxidation numbers of nickel and manganese change during the reaction. It is a redox reaction.</p>
<p><b>Step 2</b> Identify the atom or atoms that undergo an increase in oxidation number and the atom or atoms that undergo a decrease in oxidation number.</p>	<p>The nickel atom lost 2 electron and the manganese atom gained 3 electrons.</p> $\overset{0}{\text{Ni}}(\text{s}) + \overset{+7}{\text{Mn}}\overset{-2}{\text{O}_4}(\text{aq}) \rightarrow \overset{+2}{\text{Ni}}\overset{-2}{\text{O}}(\text{s}) + \overset{+4}{\text{Mn}}\overset{-2}{\text{O}_2}(\text{s})$ <p style="text-align: center;"> <math>3(-2e^-) = -6e^-</math>    <math>2(+3e^-) = +6e^-</math> </p>
<p><b>Step 3</b> Place coefficients in front of the reactants that will balance the number of electrons that are lost by one reactant and gained by another.</p>	<p>The Ni needs a coefficient of 3 and the <math>\text{MnO}_4^-</math> needs a coefficient of 2. This ratio of 3:2 must be maintained for the remainder of the balancing.</p> $3\text{Ni(s)} + 2\text{MnO}_4^-(\text{aq}) \rightarrow \text{NiO(s)} + \text{MnO}_2(\text{s})$
<p><b>Step 4</b> Balance atoms of all elements except oxygen and hydrogen.</p>	<p>There are three Ni atoms on the left and 1 on the right. Place a coefficient of 3 in front of the NiO on the right side of the equation. There are 2 Mn atoms on the left and 1 on the right. Place a coefficient of 2 in front of the <math>\text{MnO}_2</math> on the right side of the equation.</p> $3\text{Ni(s)} + 2\text{MnO}_4^-(\text{aq}) \rightarrow 3\text{NiO(s)} + 2\text{MnO}_2(\text{s})$

<p><b>Step 5</b> Balance the oxygen atoms by adding water molecules.</p>	<p>There are 8 atoms of oxygen on the left side and 7 on the right side, so add 1 water molecule to the right side.</p> $3\text{Ni(s)} + 2\text{MnO}_4^-(\text{aq}) \rightarrow 3\text{NiO(s)} + 2\text{MnO}_2(\text{s}) + \text{H}_2\text{O}(\ell)$
<p><b>Step 6</b> Balance the hydrogen atoms by adding hydrogen ions.</p>	<p>There are 2 hydrogen atoms on the right side and none on the left, so add 2 hydrogen ions to the left side.</p> $2\text{H}^+(\text{aq}) + 3\text{Ni(s)} + 2\text{MnO}_4^-(\text{aq}) \rightarrow 3\text{NiO(s)} + 2\text{MnO}_2(\text{s}) + \text{H}_2\text{O}(\ell)$
<p><b>Step 7</b> Add as many <math>\text{OH}^-</math> ions to both sides as there are <math>\text{H}^+</math> ions. When <math>\text{H}^+</math> and <math>\text{OH}^-</math> ions appear on the same side of the equation, combine them to make water.</p>	<p>There are 2 hydrogen ions on the left side, so add 2 hydroxide ions to each side.</p> $3\text{Ni(s)} + 2\text{MnO}_4^-(\text{aq}) + 2\text{H}_2\text{O}(\ell) \rightarrow 3\text{NiO(s)} + 2\text{MnO}_2(\text{s}) + \text{H}_2\text{O}(\ell) + 2\text{OH}^-(\text{aq})$
<p><b>Step 8</b> Cancel the number of water molecules that are on both sides of the equation.</p>	<p>There are 2 water molecules of the left side and 1 on the right side of the equation. Delete 1 water molecule from each side.</p> $3\text{Ni(s)} + 2\text{MnO}_4^-(\text{aq}) + \text{H}_2\text{O}(\ell) \rightarrow 3\text{NiO(s)} + 2\text{MnO}_2(\text{s}) + 2\text{OH}^-(\text{aq})$ <p>Atoms of all elements are now balanced.</p>

### Check Your Solution

Atoms of all of the elements are balanced. The total charge on the left side is  $-2$  and the total charge on the right side is also  $-2$ . The charge is balanced.