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:

$$CS_2(g) + O_2(g) \rightarrow CO_2(g) + SO_2(g)$$

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
Step 1 Assign an oxidation number to each atom in the equation and determine whether it is a redox reaction.	$CS_2(g) + O_2(g) \longrightarrow CO_2(g) + SO_2(g)$ +4 -2 0 +4 -2 +4 -2 The oxidation numbers of sulfur and oxygen change during the reaction. It is a redox reaction.
Step 2 Identify the atom or atoms that undergo an increase in oxidation number and the atom or atoms that undergo a decrease in oxidation number.	Each of two sulfur atoms lost 6 electrons and each of two oxygen atoms gained 2 electrons. So the O_2 lost 12 electrons and the CS_2 gained 4 electrons. $1(-12e^-) = -12e^-$ $CS_2(g) + O_2(g) \longrightarrow CO_2(g) + SO_2(g)$ +4 -2 0 +4 -2 +4 -2 $3(+4e^-) = +12e^-$

Step 3 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.	A coefficient of 3 is needed for the O_2. A coefficient of 1 is already implied for the CS_2. This ratio of 1:3 must be maintained for the remainder of the balancing. $CS_2(g) + 3O_2(g) \rightarrow CO_2(g) + SO_2(g)$
Step 4 Determine the coefficients of the products that are needed to complete the balancing of the equation.	There are 2 S atoms on the left and 1 on the right. Place a coefficient of 2 in front of the SO_2 in the products. $CS_2(g) + 3O_2(g) \rightarrow CO_2(g) + 2SO_2(g)$ 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.

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

$$B_2O_3(aq) + Mg(s) \rightarrow MgO(s) + Mg_3B_2(aq)$$

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: $B_2O_3(aq) + Mg(s) \rightarrow MgO(s) + Mg_3B_2(aq)$
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
Step 1 Assign an oxidation number to each atom in the equation and determine whether it is a redox reaction.	$\underset{+3\ -2}{B_2O_3(aq)} + \underset{0}{Mg(s)} \longrightarrow \underset{+2\ -2}{MgO(s)} + \underset{+2\ -3}{Mg_3B_2(aq)}$ The oxidation numbers of boron and magnesium change during the reaction. It is a redox reaction.
Step 2 Identify the atom or atoms that undergo an increase in oxidation number and the atom or atoms that undergo a decrease in oxidation number.	Each of two B atoms loses 6 electrons and the Mg atom gained 2 electrons. $1(+12e^-) = +12e^-$ $\underset{+3\ -2}{B_2O_3(aq)} + \underset{0}{Mg(s)} \longrightarrow \underset{+2\ -2}{MgO(s)} + \underset{+2\ -3}{Mg_3B_2(aq)}$ $6(-2e^-) = -12e^-$
Step 3 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.	Place a coefficient of 6 in front of the Mg reactant. A coefficient of 1 is already implied for the B_2O_3. This ratio of 1:6 must be maintained for the remainder of the balancing. $B_2O_3(aq) + 6Mg(s) \rightarrow MgO(s) + Mg_3B_2(aq)$
Step 4 Determine the coefficients of the products that are needed to complete the balancing of the equation.	Place a 3 in front of the magnesium oxide to balance the oxygen atoms $B_2O_3(aq) + 6Mg(s) \rightarrow 3MgO(s) + Mg_3B_2(aq)$ 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.

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

$$H_2S(g) + H_2O_2(aq) \rightarrow S_8(s) + H_2O(\ell)$$

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: $H_2S(g) + H_2O_2(aq) \rightarrow S_8(s) + H_2O(\ell)$
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
Step 1 Assign an oxidation number to each atom in the equation and determine whether it is a redox reaction.	$\underset{+1\ -2}{H_2S(g)} + \underset{+1\ -1}{H_2O_2(aq)} \longrightarrow \underset{0}{S_8(s)} + \underset{+1\ -2}{H_2O(\ell)}$ The oxidation numbers of sulfur and oxygen change during the reaction. It is a redox reaction.
Step 2 Identify the atom or atoms that undergo an increase in oxidation number and the atom or atoms that undergo a decrease in oxidation number.	One sulfur atom lost 2 electrons and each of two oxygen atoms gained 1 electron for a total of 2. $1(-2e^-) = -2e^-$ $\underset{+1\ -2}{H_2S(g)} + \underset{+1\ -1}{H_2O_2(aq)} \longrightarrow \underset{0}{S_8(s)} + \underset{+1\ -2}{H_2O(\ell)}$ $1(+2e^-) = +2e^-$
Step 3 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.	Coefficients of 1 are already implied for H_2S and H_2O_2. This ratio of 1:1 must be maintained while balancing the products. $H_2S(g) + H_2O_2(aq) \rightarrow S_8(s) + H_2O(\ell)$

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 H_2S on the left. If the H_2S has a coefficient of 8, then H_2O_2 must also have a coefficient of 8 in order to maintain the 1:1 ratio. $8H_2S(g) + 8H_2O_2(aq) \rightarrow S_8(s) + H_2O(\ell)$ Now there are 32 H atoms on the left and 2 on the right. Give the H_2O on the right a coefficient of 16. $8H_2S(g) + 8H_2O_2(aq) \rightarrow S_8(s) + 16H_2O(\ell)$ 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:

$$ClO_2(aq) + SbO_2^-(aq) \rightarrow ClO_2^-(aq) + Sb(OH)_6^-(aq)$$

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: $ClO_2(aq) + SbO_2^-(aq) \rightarrow ClO_2^-(aq) + Sb(OH)_6^-(aq)$
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
Step 1 Assign an oxidation number to each atom in the equation and determine whether it is a redox reaction.	$\underset{+4\ -2}{ClO_2}(aq) + \underset{+3\ -2}{SbO_2^-}(aq) \longrightarrow \underset{+3\ -2}{ClO_2^-}(aq) + \underset{+5\ -2\ +1}{Sb(OH)_6^-}(aq)$ The oxidation numbers of chlorine and antimony change during the reaction. It is a redox reaction.
Step 2 Identify the atom or atoms that undergo an increase in oxidation number and the atom or atoms that undergo a decrease in oxidation number.	The chlorine atom gained 1 electron and the antimony atom lost 2 electrons. $2(+1e^-) = +2e^-$ $\underset{+4\ -2}{ClO_2}(aq) + \underset{+3\ -2}{SbO_2^-}(aq) \longrightarrow \underset{+3\ -2}{ClO_2^-}(aq) + \underset{+5\ -2\ +1}{Sb(OH)_6^-}(aq)$ $1(-2e^-) = -2e^-$
Step 3 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.	The ClO_2 needs a coefficient of 2 and a coefficient of 1 is already implied for the SbO_2^-. This ratio of 2:1 must be maintained for the remainder of the balancing. $2ClO_2(aq) + SbO_2^-(aq) \rightarrow$ $ClO_2^-(aq) + Sb(OH)_6^-(aq)$
Step 4 The presence of the $Sb(OH)_6^-(aq)$ 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.	There are 2 Cl atoms on the left and one on the right. Give the ClO_2^- a coefficient of 2. $2ClO_2(aq) + SbO_2^-(aq) \rightarrow$ $2ClO_2^-(aq) + Sb(OH)_6^-(aq)$

Step 5 Add water molecules to balance oxygen atoms.	There are 6 O atoms on the left and 10 on the right. Add 4 water molecules to the left side of the equation. $2ClO_2(aq) + SbO_2^-(aq) + 4H_2O(\ell) \rightarrow$ $2ClO_2^-(aq) + Sb(OH)_6^-(aq)$
Step 6 Balance the hydrogen atoms by adding hydrogen ions.	There are 8 H atoms on the left and 6 on the right. Add 2 H ions to the right side. $2ClO_2(aq) + SbO_2^-(aq) + 4H_2O(\ell) \rightarrow$ $2ClO_2^-(aq) + Sb(OH)_6^-(aq) + 2H^+(aq)$
Step 7 Add as many OH^- ions to both sides as there are H^+ ions. On the right, combine the OH^- and H^+ ions to make water.	There are 2 H^+ ions on the right. Add 2 OH^- to both sides. $2ClO_2(aq) + SbO_2^-(aq) + 4H_2O(\ell) +$ $2(OH)^-(aq) \rightarrow$ $2ClO_2^-(aq) + Sb(OH)_6^-(aq) + 2H_2O(\ell)$
Step 8 Cancel the number of water molecules that are on both sides of the equation.	There are 2 water molecules on the right and 4 on the left. Delete 2 water molecules from each side of the equation. $2ClO_2(aq) + SbO_2^-(aq) + 2H_2O(\ell) +$ $2(OH)^-(aq) \rightarrow$ $2ClO_2^-(aq) + Sb(OH)_6^-(aq)$ 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 –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:

$$Cr_2O_7^{2-}(aq) + Fe^{2+}(aq) \rightarrow Cr^{3+}(aq) + Fe^{3+}(aq)$$

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: $Cr_2O_7^{2-}(aq) + Fe^{2+}(aq) \rightarrow Cr^{3+}(aq) + Fe^{3+}(aq)$
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
Step 1 Assign an oxidation number to each atom in the equation and determine whether it is a redox reaction.	$\underset{+6}{Cr_2}\underset{-2}{O_7^{2}}(aq) + \underset{+2}{Fe^{2+}}(aq) \longrightarrow \underset{+3}{2Cr^{3+}}(aq) + \underset{+3}{Fe^{3+}}(aq)$ The oxidation numbers of chromium and iron change during the reaction. It is a redox reaction.
Step 2 Identify the atom or atoms that undergo an increase in oxidation number and the atom or atoms that undergo a decrease in oxidation number.	Each of two chromium atoms gains 3 electrons for a total of 6. The iron ion loses 1 electron. 1(+6e⁻) = +6e⁻ $\underset{+6}{Cr_2}\underset{-2}{O_7^{2-}}(aq) + \underset{+2}{Fe^{2+}}(aq) \longrightarrow \underset{+3}{2Cr^{3+}}(aq) + \underset{+3}{Fe^{3+}}(aq)$ 6(-1e⁻) = -6e⁻
Step 3 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.	Place a coefficient of 6 in front of the iron ion and a coefficient of 1 is already implied for the $Cr_2O_7^{2-}$ ion. This ratio of 1:6 must be maintained for the remainder of the balancing. $Cr_2O_7^{2-}(aq) + 6Fe^{2+}(aq) \rightarrow Cr^{3+}(aq) + Fe^{3+}(aq)$
Step 4 Balance atoms of all elements except oxygen and hydrogen.	There are 2 Cr atoms on the left and 1 on the right. There are 6 Fe^{2+} ions on the left and 1 on the right. Place a coefficient of 2 in front of the Cr^{3+} ion and a 6 in front of the Fe^{3+} ion in the products. $Cr_2O_7^{2-}(aq) + 6Fe^{2+}(aq) \rightarrow 2Cr^{3+}(aq) + 6Fe^{3+}(aq)$

Step 5 Balance the O atoms by adding H_2O.	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. $Cr_2O_7^{2-}(aq) + 6Fe^{2+}(aq) \rightarrow$ $2Cr^{3+}(aq) + 6Fe^{3+}(aq) + 7H_2O(\ell)$
Step 6 Balance the hydrogen atoms by adding hydrogen ions.	There are 14 H atoms on the right and none on the left, so add 14 H^+ ions to the left side of the equation. $Cr_2O_7^{2-}(aq) + 6Fe^{2+}(aq) + 14H^+ \rightarrow$ $2Cr^{3+}(aq) + 6Fe^{3+}(aq) + 7H_2O(\ell)$ 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 +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:

$$I_2(g) + NO_3^-(aq) \rightarrow IO_3^-(aq) + NO_2(g)$$

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: $I_2(g) + NO_3^-(aq) \rightarrow IO_3^-(aq) + NO_2(g)$
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
Step 1 Assign an oxidation number to each atom in the equation and determine whether it is a redox reaction.	$I_2(g) + NO_3^-(aq) \longrightarrow IO_3^-(aq) + NO_2(g)$ 0 + 5 -2 +5 -2 +4 -2 The oxidation numbers of iodine and nitrogen change during the reaction. It is a redox reaction.
Step 2 Identify the atom or atoms that undergo an increase in oxidation number and the atom or atoms that undergo a decrease in oxidation number.	Each of 2 I atoms in I_2 loses 5 electrons for a total of 10 electrons. The nitrogen atom in NO_3^- gains one electron. $1(-10e^-) = -10e^-$ $I_2(g) + NO_3^-(aq) \longrightarrow IO_3^-(aq) + NO_2(g)$ 0 + 5 -2 +5 -2 +4 -2 $10(+1e^-) = +10e^-$
Step 3 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.	Give a coefficient of 10 to the NO_3^- and a coefficient of 1 is already implied for the I_2 molecule. $I_2(g) + 10NO_3^-(aq) \rightarrow IO_3^-(aq) + NO_2(g)$ This ratio of 1:10 must be maintained for the remainder of balancing.

Step 4 Balance atoms of all elements except oxygen and hydrogen.	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_2 molecule on the right of the equation. $I_2(g) + 10NO_3^-(aq) \rightarrow 2IO_3^-(aq) + 10NO_2(g)$
Step 5 Balance the O atoms by adding H_2O.	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. $I_2(g) + 10NO_3^-(aq) \rightarrow$ $2IO_3^-(aq) + 10NO_2(g) + 4H_2O(\ell)$
Step 6 Balance the hydrogen atoms by adding hydrogen ions	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. $8H^+(aq) + I_2(g) + 10NO_3^-(aq) \rightarrow$ $2IO_3^-(aq) + 10NO_2(g) + 4H_2O(\ell)$ 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 –2 and the total charge on the right side is also –2. The charge is balanced.

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

$$PbSO_4(aq) \rightarrow Pb(s) + PbO_2(aq) + SO_4^{2-}(aq)$$

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: $PbSO_4(aq) \rightarrow Pb(s) + PbO_2(aq) + SO_4^{2-}(aq)$
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
Step 1 Assign an oxidation number to each atom in the equation and determine whether it is a redox reaction.	$\underset{+2\ +6\ -2}{PbSO_4(aq)} \longrightarrow \underset{0}{Pb(s)} + \underset{+4\ -2}{PbO_2(aq)} + \underset{+6\ -2}{SO_4^{2-}(aq)}$ The oxidation numbers of lead changes during the reaction. It is a disproportionation reaction.
Step 2 Identify the atom or atoms that undergo an increase in oxidation number and the atom or atoms that undergo a decrease in oxidation number.	For every Pb atom that loses an electron, another atom of Pb gains an electron. $-2e^-$ $\underset{+2\ +6\ -2}{PbSO_4(aq)} \longrightarrow \underset{0}{Pb(s)} + \underset{+4\ -2}{PbO_2(aq)} + \underset{+6\ -2}{SO_4^{2-}(aq)}$ $+2e^-$
Step 3 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.	The same number of electrons that are gained by one Pb atom are lost by another. Therefore, there must be an even number of Pb atoms on the left so exactly half of them can lose electrons and half of them can gain electrons. Place a coefficient of 2 in front of the $PbSO_4$ on the left side of the equation. This coefficient must remain even for the remainder of the balancing. $2PbSO_4(aq) \rightarrow Pb(s) + PbO_2(aq) + SO_4^{2-}(aq)$

Step 4 Balance atoms of all elements except oxygen and hydrogen.	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 SO_4^{2-} on the right side of the equation. $2PbSO_4(aq) \rightarrow Pb(s) + PbO_2(aq) + 2SO_4^{2-}(aq)$
Step 5 Balance the oxygen atoms by adding water molecules.	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. $2H_2O(\ell) + 2PbSO_4(aq) \rightarrow$ $Pb(s) + PbO_2(aq) + 2SO_4^{2-}(aq)$
=**Step 6** Balance the hydrogen atoms by adding hydrogen ions	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. $2H_2O(\ell) + 2PbSO_4(aq) \rightarrow$ $Pb(s) + PbO_2(aq) + 2SO_4^{2-}(aq) + 4H^+(aq)$ 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.

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

$$Cl^-(aq) + CrO_4^{2-}(aq) \rightarrow ClO^-(aq) + CrO_2^-(aq)$$

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: $Cl^-(aq) + CrO_4^{2-}(aq) \rightarrow ClO^-(aq) + CrO_2^-(aq)$
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
Step 1 Assign an oxidation number to each atom in the equation and determine whether it is a redox reaction.	$\underset{-1}{Cl^-}(aq) + \underset{+6\ -2}{CrO_4^{2-}}(aq) \longrightarrow \underset{+1\ -2}{ClO^-}(aq) + \underset{+3\ -2}{CrO_2^-}(aq)$ The oxidation numbers of chromium and chlorine change during the reaction. It is a redox reaction.
Step 2 Identify the atom or atoms that undergo an increase in oxidation number and the atom or atoms that undergo a decrease in oxidation number.	The chlorine atom lost 2 electrons and the chromium atom gained 3 electrons. $3(-2e^-) = -6e^-$ $\underset{-1}{Cl^-}(aq) + \underset{+6\ -2}{CrO_4^{2-}}(aq) \longrightarrow \underset{+1\ -2}{ClO^-}(aq) + \underset{+3\ -2}{CrO_2^-}(aq)$ $2(+3e^-) = +6e^-$
Step 3 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.	The Cl^- needs a coefficient of 3 and the CrO_4^{2-} needs a coefficient of 2. This ratio of 3:2 must be maintained for the remainder of the balancing. $3Cl^-(aq) + 2CrO_4^{2-}(aq) \rightarrow ClO^-(aq) + CrO_2^-(aq)$
Step 4 Balance atoms of all elements except oxygen and hydrogen.	There are 3 Cl atoms on the left and 1 on the right. Place a coefficient of 3 in front of the ClO^- 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 CrO_2^- on the right side of the equation. $3Cl^-(aq) + 2CrO_4^{2-}(aq) \rightarrow 3ClO^-(aq) + 2CrO_2^-(aq)$

Step 5 Balance the oxygen atoms by adding water molecules.	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. $3Cl^-(aq) + 2CrO_4^{2-}(aq) \rightarrow$ $3ClO^-(aq) + 2CrO_2^-(aq) + H_2O(\ell)$
Step 6 Balance the hydrogen atoms by adding hydrogen ions	There are 2 hydrogen atoms on the right side and none on the left, so add 2 hydrogen ions to the left side. $2H^+(aq) + 3Cl^-(aq) + 2CrO_4^{2-}(aq) \rightarrow$ $3ClO^-(aq) + 2CrO_2^-(aq) + H_2O(\ell)$
Step 7 Add as many OH^- ions to both sides as there are H^+ ions. When H^+ and OH^- ions appear on the same side of the equation, combine them to make water.	There are 2 hydrogen ions on the left side, so add 2 hydroxide ions to each side. $2H_2O(\ell) + 3Cl^-(aq) + 2CrO_4^{2-}(aq) \rightarrow$ $3ClO^-(aq) + 2CrO_2^-(aq) + H_2O(\ell) + 2OH^-(aq)$
Step 8 Cancel the number of water molecules that are on both sides of the equation.	There are 2 water molecules on the left and 1 on the right. Delete 1 water molecule from each side. $H_2O(\ell) + 3Cl^-(aq) + 2CrO_4^{2-}(aq) \rightarrow$ $3ClO^-(aq) + 2CrO_2^-(aq) + 2OH^-(aq)$ 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 –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:

$$Ni(s) + MnO_4^-(aq) \rightarrow NiO(s) + MnO_2(s)$$

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: $Ni(s) + MnO_4^-(aq) \rightarrow NiO(s) + MnO_2(s)$
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
Step 1 Assign an oxidation number to each atom in the equation and determine whether it is a redox reaction.	$\underset{0}{Ni(s)} + \underset{+7\ -2}{MnO_4^-(aq)} \longrightarrow \underset{+2\ -2}{NiO(s)} + \underset{+4\ -2}{MnO_2(s)}$ The oxidation numbers of nickel and manganese change during the reaction. It is a redox reaction.
Step 2 Identify the atom or atoms that undergo an increase in oxidation number and the atom or atoms that undergo a decrease in oxidation number.	The nickel atom lost 2 electron and the manganese atom gained 3 electrons. $3(-2e^-) = -6e^-$ $\underset{0}{Ni(s)} + \underset{+7\ -2}{MnO_4^-(aq)} \longrightarrow \underset{+2\ -2}{NiO(s)} + \underset{+4\ -2}{MnO_2(s)}$ $2(+3e^-) = +6e^-$
Step 3 Place coefficients in front of the reactants that will balance the number of electrons that are lost by one reactant and gained by another.	The Ni needs a coefficient of 3 and the MnO_4^- needs a coefficient of 2. This ratio of 3:2 must be maintained for the remainder of the balancing. $3Ni(s) + 2MnO_4^-(aq) \rightarrow NiO(s) + MnO_2(s)$
Step 4 Balance atoms of all elements except oxygen and hydrogen.	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 MnO_2 on the right side of the equation. $3Ni(s) + 2MnO_4^-(aq) \rightarrow 3NiO(s) + 2MnO_2(s)$

Step 5 Balance the oxygen atoms by adding water molecules.	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. $3Ni(s) + 2MnO_4^-(aq) \rightarrow$ $3NiO(s) + 2MnO_2(s) + H_2O(\ell)$
Step 6 Balance the hydrogen atoms by adding hydrogen ions.	There are 2 hydrogen atoms on the right side and none on the left, so add 2 hydrogen ions to the left side. $2H^+(aq) + 3Ni(s) + 2MnO_4^-(aq) \rightarrow$ $3NiO(s) + 2MnO_2(s) + H_2O(\ell)$
Step 7 Add as many OH^- ions to both sides as there are H^+ ions. When H^+ and OH^- ions appear on the same side of the equation, combine them to make water.	There are 2 hydrogen ions on the left side, so add 2 hydroxide ions to each side. $3Ni(s) + 2MnO_4^-(aq) + 2H_2O(\ell) \rightarrow$ $3NiO(s) + 2MnO_2(s) + H_2O(\ell) + 2OH^-(aq)$
Step 8 Cancel the number of water molecules that are on both sides of the equation.	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. $3Ni(s) + 2MnO_4^-(aq) + H_2O(\ell) \rightarrow$ $3NiO(s) + 2MnO_2(s) + 2OH^-(aq)$ 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 –2 and the total charge on the right side is also –2. The charge is balanced.