9.2 Balancing Redox Reaction Equations

Figure 1 Potassium permanganate is widely used to analyze chemicals that undergo redox reactions, including determining the concentration of iron(II) ions in a sample. When all the iron(II) ions are used up, the solution in the flask turns purple.

Geologists sometimes use potassium permanganate, $KMnO_4(aq)$, to analyze the iron content in a mineral ore sample. It is a common oxidizing agent used in quantitative analysis. A typical analysis involves titrating a prepared ore sample with a potassium permanganate solution of known concentration until an endpoint is reached (**Figure 1**). The analyst must know the balanced chemical equation for the titration reaction before analyzing the titration data. Most simple redox reactions can be balanced by inspection. For more complex reactions, we can use either of two strategies: the oxidation numbers method or the half-reactions method. CAREER LINK

The Oxidation Numbers Method

As shown in Section 9.1, the oxidation numbers of atoms or ions change during a redox reaction. One way to determine if a reaction is a redox reaction is to assign oxidation numbers to each element in the reaction and then see if oxidation numbers have changed between the reactant side and product side of the equation. A change in the oxidation number of an element indicates that electron transfer and, therefore, a redox reaction has occurred.

Redox reactions often take place in aqueous solutions. In these cases, water molecules, $H_2O(l)$, may participate in the redox reaction. An aqueous solution may be acidic, basic, or neutral. If a redox reaction occurs in an acidic solution, hydrogen ions, $H^+(aq)$, may be participants in the reaction. If a redox reaction occurs in a basic solution, hydroxide ions, $OH^-(aq)$, may be participants. In these cases, the redox reactions are balanced in the same way as other redox reactions. However, when you balance redox reactions that occur in acidic or basic solutions, you will need to include $H_2O(l)$, $H^+(aq)$, and/or $OH^-(aq)$ during the balancing procedure. CAREER LINK

Tutorial 1 Balancing Equations Using Oxidation Numbers

In this tutorial, you will balance redox reaction equations using the oxidation numbers method. For Sample Problem 1, you will use the oxidation numbers method to balance a straightforward equation for a single displacement reaction. In Sample Problems 2 and 3, this procedure is adjusted to balance equations for reactions that occur in acidic and basic solutions.

The following steps are an effective problem-solving approach:

1. Write the unbalanced chemical equation from the given information. Determine the oxidation numbers for each element in the equation and identify the elements for which the oxidation numbers change.
2. Adjust the values of the coefficients to balance the electrons transferred.
3. Balance the rest of the equation by inspection. If necessary, balance oxygen by adding water.
4. If necessary, balance hydrogen by adding $H^+(aq)$ and/or $OH^-(aq)$.
5. Check your answer.
6. Write the balanced equation.

Sample Problem 1: A Redox Reaction Involving Compounds

Concentrated nitric acid, $HNO_3(aq)$, is very reactive: it oxidizes copper metal to produce toxic nitrogen dioxide gas, dissolved copper(II) nitrate, and water (**Figure 2**). Use the oxidation numbers method to write a balanced chemical equation for this reaction.

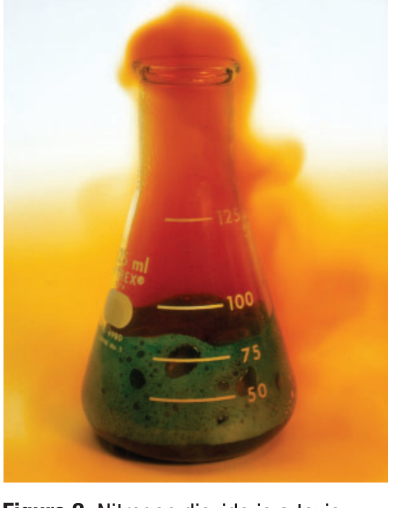

Figure 2 Nitrogen dioxide is a toxic reddish-brown gas produced when concentrated nitric acid oxidizes copper metal.

Solution

Step 1. Write the unbalanced chemical equation. Determine the oxidation numbers for each element in the equation and identify the elements for which the oxidation numbers change.

$2\,e^-$/Cu lost

$$\overset{+1+5-2}{HNO_3}(aq) + \overset{0}{Cu}(s) \longrightarrow \overset{+4-2}{NO_2}(g) + \overset{+2+5-2}{Cu(NO_3)_2}(aq) + \overset{+1-2}{H_2O}(l)$$

$1\,e^-$/N gained

The oxidation number of Cu increases from 0 to +2. Cu is oxidized by losing 2 electrons (**Figure 3**). The oxidation number of N decreases from +5 to +4. N is reduced by gaining 1 electron. Note that the oxidation number of nitrogen in $Cu(NO_3)_2$(aq) is unchanged.

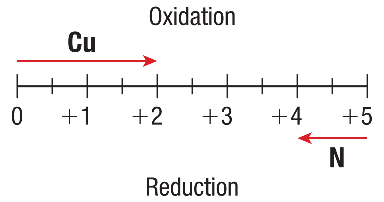

Figure 3 A redox number line shows the oxidation number of the copper atom increasing and the oxidation number of the nitrogen atom decreasing.

Step 2. Adjust the values of the coefficients to balance the electrons transferred.

During a redox reaction, electrons are transferred from one element to another. Therefore, the total number of electrons transferred must be equal. To balance the electrons, determine the simplest whole numbers that will equalize the electrons gained or lost in the reaction. Then, use those numbers as coefficients.

$2\,e^-$/Cu lost

$$2\ \overset{+1+5-2}{HNO_3}(aq) + \overset{0}{Cu}(s) \longrightarrow 2\ \overset{+4-2}{NO_2}(g) + \overset{+2+5-2}{Cu(NO_3)_2}(aq) + \overset{+1-2}{H_2O}(l)$$

$2\,(1\,e^-$/N gained$) = 2\,e^-$ gained

Since you multiplied the number of electrons gained by 2, the coefficients for HNO_3(aq) and NO_2(g) must be 2. The coefficient of copper remains unchanged.

Step 3. Balance the rest of the equation by inspection, assigning coefficients for the reactants and products as necessary.

Since there are 2 nitrate ions in $Cu(NO_3)_2$(aq), we need 2 more nitrate ions on the left side of the equation. Therefore, the coefficient of HNO_3(aq) changes to 4.

$$4\ HNO_3(aq) + Cu(s) \rightarrow 2\ NO_2(g) + Cu(NO_3)_2(aq) + 2\ H_2O(l)$$

Step 4. Since hydrogen is balanced, proceed to Step 5.

Step 5. Check your answer. Check the number of atoms (**Table 1**).

The equation is balanced if

(a) the total number of atoms of each element is equal on each side of the equation, and

(b) the sums of the charges on each side of the equation are equal.

Table 1 Number of Atoms of All Elements

$4\ HNO_3(aq) + Cu(s) \rightarrow 2\ NO_2(g) + Cu(NO_3)_2(aq) + 2\ H_2O(l)$		
Element	**Number of atoms in reactants**	**Number of atoms in products**
N	4 from 4 HNO_3	2 from 2 NO_2 + 2 from $Cu(NO_3)_2$ = 4
H	4 from 4 HNO_3	4 from 2 H_2O
O	12 from 4 HNO_3	4 from 2 NO_2 + 6 from $Cu(NO_3)_2$ + 2 from 2 H_2O = 12
Cu	1 from 1 Cu	1 from 1 $Cu(NO_3)_2$

The numbers of entities of each element are equal on both sides of the equation.

The reactants and products consist of only molecules. Since any molecule has a charge of 0, the sum of the charges of the reactants and products equals 0. This confirms that the number of electrons lost equals the number of electrons gained.

Step 6. Write the balanced equation:

$$4\ HNO_3(aq) + Cu(s) \rightarrow 2\ NO_2(g) + Cu(NO_3)_2(aq) + 2\ H_2O(l)$$

Sample Problem 2: A Redox Reaction in an Acidic Solution

To determine the iron content of an ore sample, you can perform a titration using an acidified solution of potassium permanganate, $KMnO_4(aq)$. Before the titration begins, all of the iron in the sample is converted to $Fe^{2+}(aq)$ ions. The net ionic equation that occurs during the titration is

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

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Balance the equation for this reaction.

Solution

Step 1. Write the unbalanced chemical equation. Determine the oxidation numbers for each element in the equation and identify the elements for which the oxidation numbers change.

1 e⁻/Fe lost

$$\overset{+7-2}{MnO_4^-}(aq) + \overset{+2}{Fe^{2+}}(aq) \longrightarrow \overset{+2}{Mn^{2+}}(aq) + \overset{+3}{Fe^{3+}}(aq)$$

5 e⁻/Mn gained

The oxidation number of Fe increases from +2 to +3. Fe is oxidized by losing 1 electron. The oxidation number of Mn decreases from +7 to +2. Mn is reduced by gaining 5 electrons.

Step 2. Adjust the values of the coefficients to balance the electrons transferred.

To do this, determine the simplest whole numbers that will equalize the electrons gained or lost in the reaction. Then, use those numbers as coefficients in the equation.

5 (1 e⁻/Fe lost) = 5 e⁻ lost

$$\overset{+7-2}{MnO_4^-}(aq) + 5\ \overset{+2}{Fe^{2+}}(aq) \longrightarrow \overset{+2}{Mn^{2+}}(aq) + 5\ \overset{+3}{Fe^{3+}}(aq)$$

5 e⁻/Mn gained

Since you multiplied the number of electrons lost by 5, the coefficient of $Fe^{2+}(aq)$ is 5. The coefficient of $MnO_4^-(aq)$ remains unchanged. The equation now becomes

$$1\ MnO_4^-(aq) + 5\ Fe^{2+}(aq) \rightarrow Mn^{2+}(aq) + 5\ Fe^{3+}(aq)$$

There are 5 Fe^{2+} ions on the left and 5 Fe^{3+} ions on the right. Therefore, the elements involved in the electron transfer are balanced. Note that this equation represents only the elements for which oxidation numbers change during the chemical reaction. However, since the reaction occurs in acidic solution, $H_2O(l)$ and $H^+(aq)$ also participate in the redox reaction.

Step 3. Balance the rest of the equation by inspection. If necessary, balance oxygen by adding water.

Although manganese and iron are now balanced, oxygen is not: there are 4 oxygen atoms on the left and none on the right. Since this reaction occurs in an aqueous solution, we may add 4 oxygen atoms (from 4 $H_2O(l)$) to equal the 4 oxygen atoms on the product side.

$$1\ MnO_4^-(aq) + 5\ Fe^{2+}(aq) \rightarrow Mn^{2+}(aq) + 5\ Fe^{3+}(aq) + 4\ H_2O(l)$$

Step 4. If necessary, balance hydrogen by adding $H^+(aq)$.

Since this reaction occurred in an acidic solution, we may add $H^+(aq)$ to balance the hydrogen atoms in the 4 water molecules. The water molecules in this example are products, so add 8 $H^+(aq)$ ions to the reactant side of the equation to balance the 8 hydrogen atoms on the product side:

$$1\ MnO_4^-(aq) + 5\ Fe^{2+}(aq) + 8\ H^+(aq) \rightarrow Mn^{2+}(aq) + 5\ Fe^{3+}(aq) + 4\ H_2O(l)$$

Step 5. Check your answer.

(a) Check the number of entities of each element in the balanced equation (**Table 2**).

Table 2 Number of Entities of All Elements

$MnO_4^-(aq) + 5\ Fe^{2+}(aq) + 8\ H^+(aq) \rightarrow Mn^{2+}(aq) + 5\ Fe^{3+}(aq) + 4\ H_2O(l)$		
Element	**Number in reactants**	**Number in products**
Mn	1 from MnO_4^-	1 from Mn^{2+}
Fe	5 from 5 Fe^{2+}	5 from 5 Fe^{3+}
O	4 from MnO_4^-	4 from 4 H_2O
H	8 from 8 H^+	8 from 4 H_2O

Therefore, all elements are balanced.

(b) To confirm that all the electrons involved in the reaction are accounted for, check that the sums of the charges on both sides of the equation are equal.

$$\underbrace{\underset{1\times1^-}{MnO_4^-(aq)} + \underset{5\times2^+}{5\ Fe^{2+}(aq)} + \underset{8\times1^+}{8\ H^+(aq)}}_{17+} \longrightarrow \underbrace{\underset{1\times2^+}{Mn^{2+}(aq)} + \underset{5\times3^+}{5\ Fe^{3+}(aq)} + \underset{4\times0}{4\ H_2O(l)}}_{17+}$$

The blue numbers are the charges for each side of the equation and not oxidation numbers. Since the sums of the charges on both sides are the same, the electrons transferred during the reaction are accounted for. Therefore, the equation is balanced.

Step 6. Write the balanced equation.

$$MnO_4^-(aq) + 5\ Fe^{2+}(aq) + 8\ H^+(aq) \rightarrow Mn^{2+}(aq) + 5\ Fe^{3+}(aq) + 4\ H_2O(l)$$

Sample Problem 3: A Redox Reaction in a Basic Solution

The process for balancing equations in basic solutions is similar to balancing equations in acidic solutions. The only difference is that you add hydroxide ions at the end to account for the reaction taking place in a basic solution. For example, iodate ions, $IO_3^-(aq)$, react with oxalate ions, $C_2O_4^{2-}(aq)$, in a basic solution to produce carbon dioxide gas and aqueous iodide ions.

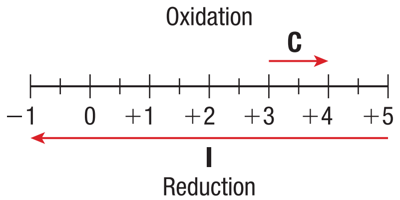

Figure 4 A redox number line shows the change in oxidation numbers of iodine and carbon.

Solution

Step 1. Write the unbalanced chemical equation. Identify the elements for which the oxidation numbers change.

1 e$^-$/C lost
2 e$^-$/$C_2O_4^{2-}$ lost

$$\overset{+5\,-2}{IO_3^-}(aq) + \overset{+3\,-2}{C_2O_4^{2-}}(aq) \longrightarrow \overset{-1}{I^-}(aq) + 2\,\overset{+4\,-2}{CO_2}(g)$$

6 e$^-$/I gained

The oxidation number of iodine decreases from +5 to −1 (**Figure 4**). Iodine is reduced by gaining 6 electrons. The oxidation number of C increases from +3 to +4. C is oxidized by losing 1 electron. However, there are 2 carbon atoms in $C_2O_4^{2-}$. Therefore, 2 electrons in total are lost. Note that a coefficient 2 is included on the right side to account for the second carbon atom.

Step 2. Adjust the values of the coefficients to balance the electrons transferred.

Determine the simplest whole numbers that will equalize the electrons gained or lost in the reaction. Then, use those numbers as coefficients in the equation.

1 e$^-$/C lost
3 (2 e$^-$/$C_2O_4^{2-}$ lost) = 6 e$^-$ lost

$$\overset{+5\,-2}{IO_3^-}(aq) + 3\,\overset{+3\,-2}{C_2O_4^{2-}}(aq) \longrightarrow \overset{-1}{I^-}(aq) + 6\,\overset{+4\,-2}{CO_2}(g)$$

6 e$^-$/I gained

Since you multiplied the number of electrons lost by 3, the coefficient of $C_2O_4^{2-}$(aq) is 3. As a result, the coefficient of CO_2 increases from 2 to 6. The coefficient of IO_3^-(aq) remains unchanged. The equation now becomes

$$IO_3^-(aq) + 3\,C_2O_4^{2-}(aq) \rightarrow I^-(aq) + 6\,CO_2(g)$$

The entities involved in the electron transfer are now balanced. Note that this equation represents only the entities for which oxidation numbers change during the chemical reaction. However, since the reaction occurs in a basic solution, H_2O(l) and OH^-(aq) may also participate in the redox reaction.

Step 3. Balance the rest of the equation by inspection. If necessary, balance oxygen by adding water.

Notice that oxygen is not balanced: there are 15 oxygen atoms on the left and 12 on the right. Since this reaction occurs in an aqueous solution, we may add H_2O molecules on the product side to balance the oxygen atoms on the reactant side. The product side now requires 3 oxygen atoms (from 3 H_2O(l)) to balance.

$$IO_3^-(aq) + 3\,C_2O_4^{2-}(aq) \rightarrow I^-(aq) + 6\,CO_2(g) + 3\,H_2O(l)$$

Step 4. If necessary, balance hydrogen by adding H^+(aq) and OH^-(aq).

Although the reaction occurs in a basic solution, it is simpler to add H^+(aq) to balance the hydrogen atoms rather than adding hydroxide. We will correct for this H^+(aq) in the next step. Since the water molecules in this example are products, add 6 H^+(aq) ions to the reactant side of the equation:

$$IO_3^-(aq) + 3\,C_2O_4^{2-}(aq) + 6\,H^+(aq) \rightarrow I^-(aq) + 6\,CO_2(g) + 3\,H_2O(l)$$

Since the reaction occurred in basic solution, the solution contains OH^-(aq) ions. Therefore, it is not likely that H^+(aq) ions took part in the reaction. You can eliminate H^+(aq) ions by adding OH^-(aq) ions to both sides of the equation. By adding OH^-(aq) ions to both sides, you are not changing the equation.

$$IO_3^-(aq) + 3\,C_2O_4^{2-}(aq) + 6\,H^+(aq) + \ 6\,OH^-(aq) \rightarrow$$
$$I^-(aq) + 6\,CO_2(g) + 3\,H_2O(l) + 6\,OH^-(aq)$$

Since $6\,H^+(aq) + 6\,OH^-(aq) \rightarrow 6\,H_2O(l)$, the H^+ and OH^- ions on the reactant side are equivalent to 6 water molecules. Subtracting 3 water molecules from each side eliminates the redundant water molecules from the equation:

$$IO_3^-(aq) + 3\ \ C_2O_4^{2-}(aq) + \cancel{6}\,3\,H_2O(l) \rightarrow$$
$$I^-(aq) + 6\,CO_2(g) + \cancel{3\,H_2O(l)} + 6\,OH^-(aq)$$

Step 5. Check your answer.

(a) Check the number of entities of each element in the balanced equation (**Table 3**).

Table 3 Number of Entities of All Elements

$IO_3^-(aq) + 3\ \ C_2O_4^{2-}(aq) + 3\,H_2O(l) \rightarrow I^-(aq) + 6\,CO_2(g) + 6\,OH^-(aq)$		
Element	**Number in reactants**	**Number in products**
I	1 from IO_3^-	1 from I^-
C	6 from 3 $C_2O_4^{2-}$	6 from 6 CO_2
O	3 from IO_3^- + 12 from $C_2O_4^{2-}$ + 3 from 3 H_2O = 18	12 from 6 CO_2 + 6 from 6 OH^- = 18
H	6 from 3 H_2O	6 from 6 OH^-

Therefore, all elements are balanced.

(b) Check that the sums of the charges on both sides of the equation are equal.

$$\underbrace{\underset{1 \times 1^-}{IO_3^-(aq)} \;+\; \underset{3 \times 2^-}{3\,C_2O_4^{2-}(aq)} \;+\; \underset{3 \times 0}{3\,H_2O(l)}}_{7-} \;\longrightarrow\; \underbrace{\underset{1 \times 1^-}{I^-(aq)} \;+\; \underset{6 \times 0}{6\,CO_2(g)} \;+\; \underset{6 \times 1^-}{6\,OH^-(aq)}}_{7-}$$

Since the sums of the charges on both sides are the same, the electrons transferred during the reaction are accounted for. Therefore, the equation is balanced.

Step 6. Write the balanced equation.

$$IO_3^-(aq) + 3\ \ C_2O_4^{2-}(aq) \ \ + 3\,H_2O(l) \rightarrow I^-(aq) + \ 6\,CO_2(g) \ \ + \ 6\,OH^-(aq)$$

Practice

1. Use the oxidation numbers method to write a balanced chemical equation for the reaction between lead(II) oxide, $PbO_2(s)$, and ammonia, $NH_3(g)$, to produce nitrogen gas, liquid water, and solid lead. T/I
2. Hydrogen sulfide, $H_2S(g)$, is a poisonous gas with an odour like rotten eggs. Natural gas in well water can be removed by reacting it with pressurized oxygen. The unbalanced equation for this reaction is

   $H_2S(g) + O_2(g) \rightarrow S(s) + H_2O(l)$

   Write a balanced chemical equation for this reaction using the oxidation numbers method. T/I
3. Balance the following chemical equations representing reactions that take place in acidic solutions: T/I

   (a) $MnO_4^-(aq) + Br^-(aq) \rightarrow MnO_2(s) + BrO_3^-(aq)$

   (b) $I_2(s) + OCl^-(aq) \rightarrow IO_3^-(aq) + Cl^-(aq)$

4. Balance the following chemical equations representing reactions that take place in basic solutions: T/I
   (a) $MnO_4^-(aq) + SO_3^{2-}(aq) \rightarrow SO_4^{2-}(aq) + MnO_2(s)$
   (b) $S^{2-}(aq) + I_2(s) \rightarrow SO_4^{2-}(aq) + I^-(aq)$
5. Iodate ions react in solution with hydrogen sulfite ions to produce sulfate ions and a precipitate of solid iodine. The unbalanced equation for this reaction is $IO_3^-(aq) + HSO_3^-(aq) \rightarrow SO_4^{2-}(aq) + I_2(s)$. T/I
   (a) Determine the entities that change oxidation number, and indicate the numbers of electrons that are lost or gained.
   (b) Balance the equation, assuming that the reaction takes place in an acidic solution.
6. Chromium hydroxide is a solid compound that reacts in a basic solution with chlorate ions. The products are aqueous chromate ions and chloride ions. The unbalanced chemical equation for this reaction is
   $Cr(OH)_3(s) + ClO_3^-(aq) \rightarrow CrO_4^{2-}(aq) + Cl^-(aq)$. T/I
   (a) Determine the entities that change oxidation number and indicate the numbers of electrons that are lost or gained.
   (b) Balance the equation. Assume that the reaction takes place in a basic solution.

The Half-Reactions Method

The second strategy to balance redox equations is to separate the reaction into two half-reaction equations: one involving oxidation and the other involving reduction. These equations are balanced separately, then added together to give the balanced chemical equation for the complete redox reaction. Although the half-reactions method differs from the oxidation numbers method in approach, the result is the same: a balanced redox reaction.

Tutorial 2 Balancing Equations Using Half-Reactions

To balance the equation of a redox reaction by this method, you will first identify and write the half-reaction equations for redox reactions. Then, you will balance the equations for the half-reactions separately. Finally, you will add the half-reaction equations to arrive at the balanced redox reaction equation. The following steps are an effective problem-solving approach:

1. Write the unbalanced equation and assign oxidation numbers to all entities.
2. Write unbalanced equations for the oxidation and reduction half-reactions.
3. Balance each half-reaction equation.
4. Use electrons to balance the charge in each half-reaction equation.
5. Equalize the electron transfer in the two half-reaction equations.
6. Add the half-reaction equations.
7. Check your answer.
8. Write the balanced equation.

Sample Problem 1: A Redox Reaction in an Acidic Solution

Consider again the permanganate reaction from Tutorial 1, Sample Problem 2. The net ionic equation for the reaction is

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

Using the half-reactions method, balance the equation for this reaction in an acidic solution. (Note that this will be a different way of solving the same problem as in Tutorial 1, Sample Problem 2.)

Solution

Step 1. Write the unbalanced equation and assign oxidation numbers to all entities.

$$\overset{+7\ -2}{MnO_4^-}(aq) + \overset{+2}{Fe^{2+}}(aq) \rightarrow \overset{+2}{Mn^{2+}}(aq) + \overset{+3}{Fe^{3+}}(aq)$$

Step 2. Write unbalanced equations for the oxidation and reduction half-reactions.

Look at the change in oxidation number of the manganese atom in the permanganate ion during the redox reaction. Since the oxidation number of Mn decreases, we know that Mn is reduced. The equation for the half-reaction involving the permanganate ion is

$$\overset{+7\ -2}{MnO_4^-}(aq) \rightarrow \overset{+2}{Mn^{2+}}(aq) \quad \text{(reduction)}$$

The other half-reaction involves the oxidation of iron(II) to iron(III):

$$\overset{+2}{Fe^{2+}}(aq) \rightarrow \overset{+3}{Fe^{3+}}(aq) \quad \text{(oxidation)}$$

Step 3. Balance each half-reaction equation. As you did in the oxidation numbers method, use $H_2O(l)$ to balance oxygen and $H^+(aq)$ ions to balance hydrogen.

Consider the reduction half-reaction:

$$MnO_4^-(aq) \rightarrow Mn^{2+}(aq)$$

Balancing the oxygen and hydrogen gives

$$8\ H^+(aq) + MnO_4^-(aq) \rightarrow Mn^{2+}(aq) + 4\ H_2O(l)$$

Iron in the oxidation half-reaction is already balanced:

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

Step 4. Use electrons to balance the charge in each half-reaction equation.

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

$$\underbrace{\underset{8+}{8\ H^+(aq)} \underset{+}{+} \underset{1-}{MnO_4^-(aq)}}_{7+} \longrightarrow \underbrace{\underset{2+}{Mn^{2+}(aq)} \underset{+}{+} \underset{0}{4\ H_2O(l)}}_{2+}$$

You may now equalize the charges by adding 5 electrons to the left side of the equation. Remember that electrons have a negative charge, so you are effectively adding a negative number to the total charge.

$$\underbrace{5e^- + 8\ H^+(aq) + MnO_4^-(aq)}_{2+} \longrightarrow \underbrace{Mn^{2+}(aq) + 4\ H_2O(l)}_{2+}$$

Now, both the elements and charges for the reduction half-reaction are balanced. The fact that 5 electrons appear on the reactant side of the equation makes sense when you consider the change in oxidation numbers that occurs. Manganese has an oxidation number of +7 in MnO_4^- and +2 in Mn^{2+}. Five electrons must be added in order for this change to occur.

In the oxidation half-reaction equation, the overall charges for the reactants and products are

$$\underbrace{Fe^{2+}(aq)}_{2+} \longrightarrow \underbrace{Fe^{3+}(aq)}_{3+}$$

You must add 1 electron to the right side of the equation to give a net charge of 2+ on both sides of the equation.

$$\underbrace{Fe^{2+}(aq)}_{2+} \longrightarrow \underbrace{Fe^{3+}(aq) + e^-}_{2+}$$

Step 5. Equalize the electron transfer in the two half-reaction equations.

The elements and charges for the oxidation half-reaction are now balanced. However, at this point, the reduction half-reaction involves a transfer of 5 electrons and the oxidation half-reaction involves a transfer of only 1 electron. To equalize the number of electrons transferred, multiply the oxidation half-reaction by 5:

$$5\,Fe^{2+}(aq) \rightarrow 5\,Fe^{3+}(aq) + 5\,e^-$$

Step 6. Add the half-reaction equations.

$$5\,e^- + 5\,Fe^{2+}(aq) + MnO_4^-(aq) + 8\,H^+(aq) \rightarrow 5\,Fe^{3+}(aq) + Mn^{2+}(aq) + 4\,H_2O(l) + 5\,e^-$$

Any identical entities on both sides of the arrow cancel out. Here, you can cross out the electrons on both sides, giving the final balanced redox equation:

$$5\,Fe^{2+}(aq) + MnO_4^-(aq) + 8\,H^+(aq) \rightarrow 5\,Fe^{3+}(aq) + Mn^{2+}(aq) + 4\,H_2O(l)$$

Step 7. Check your answer.

(a) Check the number of atoms (**Table 4**).

Table 4 Number of Entities of All Elements

$5\,Fe^{2+}(aq) + MnO_4^-(aq) + 8\,H^+(aq) \rightarrow 5\,Fe^{3+}(aq) + Mn^{2+}(aq) + 4\,H_2O(l)$		
Element	**Number in reactants**	**Number in products**
Fe	5 from 5 Fe^{2+}	5 from 5 Fe^{3+}
Mn	1 from MnO_4^-	1 from Mn^{2+}
O	4 from MnO_4^-	4 from 4 H_2O
H	8 from 8 H^+	8 from 4 H_2O

Therefore, all elements are balanced.

(b) To confirm that all the electrons involved in the reaction are accounted for, check that the sums of the charges on both sides of the equation are equal.

$$\underbrace{\underset{1-}{MnO_4^-(aq)} + \underset{10+}{5\,Fe^{2+}(aq)} + \underset{8+}{8\,H^+(aq)}}_{17+} \longrightarrow \underbrace{\underset{2+}{Mn^{2+}(aq)} + \underset{15+}{5\,Fe^{3+}(aq)} + \underset{0}{4\,H_2O(l)}}_{17+}$$

Since the sums of the charges on both sides are the same, the electrons transferred during the reaction are accounted for.

Step 8. Write the balanced equation.

$$MnO_4^-(aq) + 5\,Fe^{2+}(aq) + 8\,H^+(aq) \rightarrow Mn^{2+}(aq) + 5\,Fe^{3+}(aq) + 4\,H_2O(l)$$

Practice

1. Using the half-reaction method, balance the following equations: T/I
   (a) $Zn(s) + H^+(aq) \rightarrow Zn^{2+}(aq) + H_2(g)$
   (b) $HNO_3(aq) + Cu(s) \rightarrow NO_2(g) + Cu^{2+}(aq)$ in acid
2. Methanol reacts with permanganate in a basic solution:
   $CH_3OH(aq) + MnO_4^-(aq) \rightarrow CO_3^{2-}(aq) + MnO_4^{2-}(aq)$ T/I
   (a) Write the balanced half-reaction equations for this reaction.
   (b) Write the overall balanced equation for the reaction.

9.2 Review

Summary

- In a balanced redox equation, the total numbers of each type of atom or ion on either side of the equation are equal and the numbers of electrons transferred are equal.
- Redox equations may be balanced by using the oxidation numbers method or the half-reactions method.
- Oxidation–reduction reactions can occur in acidic or basic solutions. In these cases, it may be necessary to add water molecules, hydrogen ions, and/or hydroxide ions to balance the equation.

Questions

1. When magnesium metal is added to a beaker that contains hydrochloric acid, HCl(aq), a gas forms. T/I
   (a) If the magnesium is oxidized and the hydrogen is reduced, write the balanced half-reaction equations for the reaction.
   (b) Write the balanced equation for the reaction.
   (c) How many electrons transfer in the balanced equation?
2. Write balanced half-reactions for the following equations: T/I
   (a) $2\ AgNO_3(aq) + Cu(s) \rightarrow Cu(NO_3)_2(aq) + 2\ Ag(s)$
   (b) $Cr^{3+}(aq) + Cl_2(g) \rightarrow Cr_2O_7^{2-}(aq) + Cl^-(aq)$
   (c) $H_2SO_4(aq) + Ca(s) \rightarrow CaSO_4(aq) + H_2(g)$
3. Balance the following equations for reactions occurring in acidic conditions: T/I
   (a) $ClO_3^-(aq) + I_2(aq) \rightarrow Cl^-(aq) + IO_3^-(aq)$
   (b) $Cr_2O_7^{2-}(aq) + Cl^-(aq) \rightarrow Cr^{3+}(aq) + Cl_2(aq)$
4. Balance the following equations for reactions occurring in basic conditions: T/I
   (a) $Pb(OH)_4^{2-}(aq) + ClO^-(aq) \rightarrow PbO_2(s) + Cl^-(aq)$
   (b) $NO_2^-(aq) + Al(s) \rightarrow NH_3(aq) + Al(OH)_4^-(aq)$
5. Permanganate ions, $MnO_4^-(aq)$, and oxalate ions, $C_2O_4^{2-}(aq)$, react as follows: T/I
   $MnO_4^-(aq) + C_2O_4^{2-}(aq) \rightarrow Mn^{2+}(aq) + CO_2(aq)$
   (a) How many electrons are transferred for each $MnO_4^-(aq)$ ion?
   (b) How many electrons are transferred for each $C_2O_4^{2-}(aq)$ ion?
   (c) Write the balanced chemical equation using the oxidation numbers method.
6. Silver metal can be found in nature as large nuggets. Usually, it is mixed with other metals and ores. An aqueous solution of toxic cyanide ions can be used to extract the silver using the following reaction that occurs in basic solution: T/I A
   $Ag(s) + CN^-(aq) + O_2(g) \rightarrow Ag(CN)_2^-(aq)$
   (a) Balance the above equation using the half-reactions method.
   (b) If the cyanide ions are used up during the reaction, why is there still an environmental concern about the use of cyanide in this process? CAREER LINK
7. Potassium dichromate, $K_2Cr_2O_7(aq)$, is a bright orange compound that can be reduced to a blue-violet solution of $Cr^{3+}(aq)$ ions. Under certain conditions, $K_2Cr_2O_7(aq)$ reacts with ethanol, $C_2H_5OH(aq)$, as follows: T/I A
   $Cr_2O_7^{2-}(aq) + C_2H_5OH(aq) \rightarrow Cr^{3+}(aq) + CO_2(g)$
   (a) Balance the above equation in acidic conditions using either the oxidation numbers method or the half-reactions method.
   (b) Which method do you prefer to use? Explain.
8. Potassium permanganate is used to determine the amount of hydrogen peroxide in a sample:
   $MnO_4^-(aq) + H_2O_2(aq) \rightarrow Mn^{2+}(aq) + O_2(g)$ T/I
   (a) Write the balanced half-reaction equations for this reaction.
   (b) Balance the equation for the redox reaction involving these substances in acidic solution.
9. Write a balanced chemical equation for the oxidation of a sulfite solution by a nitrate solution. Assume the reaction products include dissolved sulfate and nitrogen dioxide gas. T/I
10. Chlorine bubbled into a basic solution produces the chloride and chlorate ions:
    $Cl_2(g) \rightarrow Cl^-(aq) + ClO_3^-(aq)$ K/U T/I
    (a) Balance this equation using the half-reactions method.
    (b) Assign oxidation numbers to each element in this reaction.
    (c) What is unusual about the role of the element chlorine in this reaction?