9.1 Electron Transfer Reactions

Copper wire is shiny and slightly pinkish. **Figure 1(a)** shows a piece of copper wire when it is first placed in a solution of silver nitrate, $AgNO_3(aq)$. Over time, a fuzz of silver metal, Ag(s), forms on the copper wire (**Figure 1(b)**). The solution turns blue as copper(II) ions are released.

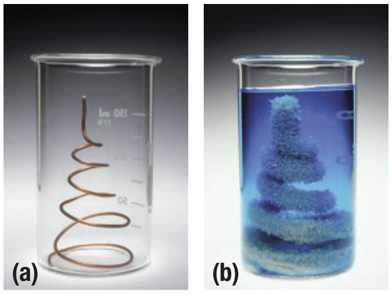


Figure 1 Copper wire placed in a solution of silver nitrate. Over time, copper atoms in the wire are displaced by silver ions in solution. The blue tint of the solution is due to copper ions, and the fuzzy coating on the wire is silver metal.

$$Cu(s) + 2\,AgNO_3(aq) \rightarrow 2\,Ag(s) + Cu(NO_3)_2(aq)$$

Remember that the copper(II) nitrate produced is completely dissociated into ions in the solution: $Cu^{2+}(aq)$ and $NO_3^-(aq)$. The equation above may also be written with its ionic compounds dissociated:

$$Cu(s) + 2\,Ag^+(aq) + 2\,NO_3^-(aq) \rightarrow 2\,Ag(s) + Cu^{2+}(aq) + 2\,NO_3^-(aq)$$

This form of the equation is sometimes called the total ionic equation. Notice that nitrate ions appear on both sides of the equation. This means that they remain unchanged. Ions that do not participate in a chemical reaction are called spectator ions. These ions can be eliminated from the total ionic equation to give the net ionic equation: WEB LINK

$$Cu(s) + 2\,Ag^+(aq) \rightarrow 2\,Ag(s) + Cu^{2+}(aq)$$

During this reaction, copper atoms, Cu(s), each lose 2 electrons to form $Cu^{2+}(aq)$ ions. Each $Ag^+(aq)$ ion gains 1 of these electrons to become a neutral silver atom, Ag(s). Since there are 2 silver ions in the equation, a total of 2 electrons are transferred for each atom of copper that reacts.

oxidation the process in which one or more electrons is lost by a chemical entity

reduction the process in which one or more electrons is gained by a chemical entity

oxidation–reduction (redox) reaction the reaction in which one or more electrons are transferred between chemical entities

In chemistry, the loss of electrons is called **oxidation**, and the gain of electrons is called **reduction**. Thus, a reaction in which electrons are transferred from one entity to another is called an **oxidation–reduction reaction**, or **redox** reaction. We can summarize the oxidation and reduction occurring in the copper–silver reaction as follows:

2 e⁻ lost (oxidation)

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

2 e⁻ gained (reduction)

Half-Reaction Equations

half-reaction equation the part of an oxidation–reduction reaction equation representing either the oxidation reaction or the reduction reaction

You have seen that a redox reaction involves the transfer of electrons. One element in the reaction gains electrons and another element loses electrons. Not all reactions are redox reactions. To clarify the behaviour of the electrons, we can break the chemical equation for this reaction into two separate parts called half-reaction equations. A **half-reaction equation** is a chemical equation that represents one of the two parts of a redox reaction. For the reaction in Figure 1, copper atoms lose 2 electrons to form $Cu^{2+}(aq)$ ions. The half-reaction equation for this oxidation is

$$Cu^0(s) \rightarrow Cu^{2+}(aq) + 2\,e^-$$

The oxidation of copper supplies the electrons needed for the reduction of silver ions. The half-reaction equation for this reduction is

$$Ag^+(aq) + e^- \rightarrow Ag^0(s)$$

Note that when the silver ion, with a charge of +1, gains an electron, the atom of silver metal that forms has a charge of 0. Since the silver ion gains electrons, it becomes reduced in the chemical reaction. It may seem odd to say that the silver is reduced, when in fact it gained an electron. However, its charge is reduced by going from +1 to 0. WEB LINK

LEARNING **TIP**

A Redox Mnemonic

Use the mnemonic "LEO says GER" to help remember the difference between reduction and oxidation.

LEO: Losing Electrons is Oxidation

GER: Gaining Electrons is Reduction

If we compare the number of electrons in the two half-reaction equations, we see that the number of electrons is not the same. In a redox reaction, the number of electrons lost by oxidation must always be equal to the number of electrons gained by reduction. To make the number of electrons equal in our example, we must multiply all of the entities in the reduction half-reaction equation by 2 (**Figure 2**).

$$2\,Ag^{+}(aq) + 2\,e^{-} \rightarrow 2\,Ag(s)$$

The overall balanced redox reaction equation can be shown as

$$Cu(s) + 2\,Ag^{+}(aq) \rightarrow 2\,Ag(s) + Cu^{2+}(aq)$$

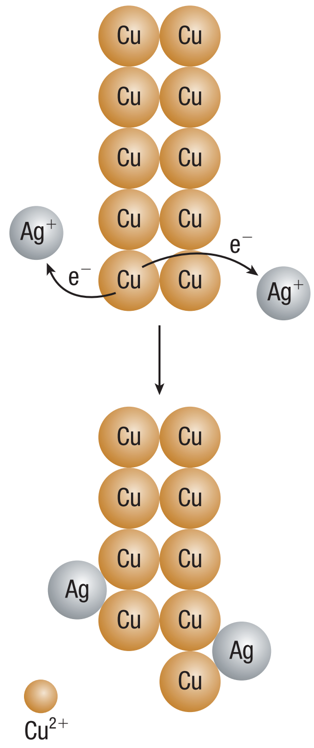


Figure 2 In the reaction of copper metal with silver ions, the copper atom transfers 2 electrons to 2 silver ions. The added electrons reduce the silver ions to uncharged silver atoms, while the electron loss oxidizes the uncharged copper atom to a copper ion.

Tutorial 1 Writing Half-Reaction Equations for Redox Reactions

In this tutorial, you will write the oxidation and reduction half-reaction equations from the net ionic equation of a redox reaction.

Sample Problem 1: Writing a Half-Reaction Equation

Write the oxidation and reduction half-reaction equations for the reaction of aluminum metal in an aqueous solution containing silver ions:

$$Al(s) + 3\,Ag^{+}(aq) \rightarrow Al^{3+}(aq) + 3\,Ag(s)$$

Solution

Step 1. Separate the equation into two half-reactions one for each element.

$$Al(s) \rightarrow Al^{3+}(aq)$$

$$3\,Ag^{+}(aq) \rightarrow 3\,Ag(s)$$

Step 2. If necessary, divide each equation by a whole number so that the coefficients in the equation are in the simplest whole-number ratio.

In this case, simplify the silver half-reaction by dividing both sides by 3:

$$Al(s) \rightarrow Al^{3+}(aq)$$

$$Ag^{+}(aq) \rightarrow Ag(s)$$

Step 3. Add electrons to both equations so that the net charge on both sides of each equation is equal. If you add electrons on one side of the arrow in one half-reaction equation, then you will add electrons on the opposite side of the arrow in the other half-reaction equation. Note, though, that the *number* of electrons added may not be the same in both half-reaction equations.

$$Al(s) \rightarrow Al^{3+}(aq) + 3\,e^{-}$$

$$Ag^{+}(aq) + e^{-} \rightarrow Ag(s)$$

Step 4. Determine whether each half-reaction equation represents an oxidation or a reduction.

The first equation represents an oxidation half-reaction since the aluminum atom loses electrons. The second equation represents a reduction half-reaction since the silver ion gains an electron.

Oxidation: $Al(s) \rightarrow Al^{3+}(aq) + 3\,e^{-}$

Reduction: $Ag^{+}(aq) + e^{-} \rightarrow Ag(s)$

Practice

1. Write the oxidation and reduction half-reaction equations for the following redox reactions: T/I
   (a) $Sn^{2+}(aq) + Co(s) \rightarrow Sn(s) + Co^{2+}(aq)$
   (b) $2\,Ag^{+}(aq) + Pb(s) \rightarrow 2\,Ag(s) + Pb^{2+}(aq)$
   (c) $2\,Fe^{2+}(aq) + I_2(s) \rightarrow 2\,I^{-}(aq) + 2\,Fe^{3+}(aq)$

2. Identify the reactant oxidized and the reactant reduced in Question 1. T/I
3. Write the oxidation and reduction half-reaction equations for each of the following redox reactions. Identify the spectator ion(s) in each reaction. T/I
   (a) $Ni(s) + CuCl_2(aq) \rightarrow NiCl_2(aq) + Cu(s)$
   (b) A solution of tin(II) nitrate reacts with a solution of chromium(II) nitrate to produce tin metal and a solution of chromium(III) nitrate.
4. Chlorine gas reacts with potassium iodide solution to produce solid iodine in a solution of potassium chloride. T/I
   (a) Write the balanced chemical equation for this reaction.
   (b) Write the net ionic equation for the reaction.
   (c) Write the oxidation and reduction half-reaction equations for the reaction.
   (d) Identify the reactant oxidized and the reactant reduced.

Oxidation Numbers

oxidation number a number used to keep track of electrons in oxidation–reduction reactions according to certain rules; also known as oxidation state

Chemists use a type of "electron bookkeeping" to keep track of which atoms are losing electrons and which atoms are gaining electrons in a redox reaction. In this system, an atom's **oxidation number**, also known as its oxidation state, is defined as the apparent net electric charge that the atom would have if electron pairs in covalent bonds belonged entirely to the more electronegative atom. The oxidation number system is a useful way to keep track of electrons, but it does not usually represent an actual charge on an atom. An oxidation number can be a positive or a negative number.

Consider the oxidation numbers for atoms in a covalently bonded molecule. Recall that atoms share electrons in covalent bonds. You can determine the oxidation numbers of atoms in covalent compounds by assigning the shared electrons to particular atoms. In covalent bonds between two identical atoms, the electrons are shared equally between the two atoms. However, in cases where a covalent bond exists between two different atoms, the atoms share the electrons unequally. For the purposes of assigning oxidation numbers, the shared electrons go to the atom that has the stronger attraction for electrons (the more electronegative atom).

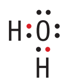


Figure 3 In a molecule of water, oxygen and hydrogen atoms share electrons. The oxygen atom is more electronegative, so it has a greater attraction for the electrons than the hydrogen atoms do.

For example, in a molecule of water, oxygen is more electronegative than hydrogen (**Figure 3**). When you assign oxidation numbers to oxygen and hydrogen in water, you assume that the oxygen atom actually possesses all of the electrons, even the electrons it shares with hydrogen. Recall that a hydrogen atom has 1 electron. Therefore, in water, you assume that the oxygen atom has "taken" the electrons from the 2 hydrogen atoms. This gives the oxygen atom an excess of 2 electrons. Thus, the oxidation number of oxygen is -2. Each hydrogen atom has no electrons and is given an oxidation number of $+1$.

Therefore, we define the oxidation numbers of the atoms in a covalent compound as the imaginary charges the atoms would have if (a) the shared electrons were divided equally between identical atoms bonded to each other, and (b) the shared electrons were assigned to the atom that has the greater attraction for electrons if the atoms are different. CAREER LINK

Table 1 (next page) provides a set of rules for assigning oxidation numbers to atoms. Applying these rules allows you to assign oxidation numbers to the atoms or ions of most compounds.

There are two additional rules for assigning oxidation numbers:

1. The sum of the oxidation numbers of all atoms in an electrically neutral compound must equal zero. This rule is often referred to as the zero-sum rule.
2. The sum of the oxidation numbers of all atoms in ions containing 2 or more atoms must equal the overall charge of the ion.

Table 1 Rules for Assigning Oxidation Numbers

The oxidation number of . . .	Summary	Examples
an atom in an element is 0.	element: 0	Na(s), O_2(g), O_3(g), Hg(l)
a monatomic ion is the same as its charge.	monatomic ion: charge of ion	Na^+, Cl^-
fluorine is −1 in its compounds.	fluorine: −1	HF, PF_3
oxygen is usually −2 in its compounds. (Exception: peroxides, containing O_2^{2-}, in which oxygen is −1)	oxygen: −2	H_2O, CO_2 (Exception: H_2O_2)
hydrogen is +1 in its covalent compounds. (Exception: metal hydrides, in which hydrogen is −1)	hydrogen: +1	H_2O, HCl, NH_3 (Exception: CaH_2)

The convention is to write actual charges on ions as $^{n+}$ or $^{n-}$, placing the number before the plus or minus sign as a superscript. However, you write oxidation numbers (not actual charges) as $+n$ or $-n$, placing the number after the plus or minus sign. Thus, you would write a magnesium ion as Mg^{2+}, but the oxidation number of the magnesium ion as +2.

Tutorial 2 Assigning Oxidation Numbers

In this tutorial, you will learn how to correctly assign an oxidation number to an element within a compound or an ion. The following steps are a useful guide:

1. Assign oxidation numbers to elements as listed in Table 1.
2. Identify any elements not mentioned in Table 1, and use the zero-sum rule or ion-charge rule to assign their oxidation numbers.
3. Check that the sum of the oxidation numbers is equal to zero (for a neutral compound) or the charge (for a polyatomic ion).

Sample Problem 1: Assigning Oxidation Numbers for a Molecular Compound

Assign the oxidation numbers to all atoms in a molecule of carbon dioxide.

Solution

Step 1. Assign oxidation numbers to elements as listed in Table 1.

The oxidation number of oxygen is usually –2.

Step 2. Identify any elements not mentioned in Table 1, and use the zero-sum rule to assign their oxidation numbers.

There is no specific oxidation number for carbon atoms. Since CO_2 is an electrically neutral compound, the sum of the oxidation numbers of oxygen and carbon must be zero. Since each oxygen atom has an oxidation number of −2, and there are 2 oxygen atoms in carbon dioxide, the carbon atom in a molecule of carbon dioxide must have an oxidation number of +4 to balance the −4 of the oxygen atoms.

$$\overset{+4}{C}\overset{-2}{O_2} \quad \text{for each atom}$$

Step 3. Check that the sum of the oxidation numbers is equal to zero. Make sure that you account for the number of each atom in the carbon dioxide molecule.

$$\underset{\substack{\uparrow \\ \text{number of C} \\ \text{atoms}}}{1}(+4) + \underset{\substack{\uparrow \\ \text{number of O} \\ \text{atoms}}}{2}(-2) = 0$$

In a CO_2 molecule, the oxidation number of the carbon atom is +4 and the oxidation number of each oxygen atom is −2.

Sample Problem 2: Assigning Oxidation Numbers for a Polyatomic Ion

Assign oxidation numbers to all atoms in the nitrate ion, NO_3^-.

Solution

Step 1. Assign oxidation numbers to elements as listed in Table 1.

The oxidation number of oxygen is −2.

Step 2. Identify any elements not mentioned in Table 1, and use the ion-charge rule to assign their oxidation numbers.

Nitrogen is not mentioned in Table 1. The nitrate ion has a net charge of −1. Therefore, the sum of the oxidation numbers of all the atoms in the nitrate ion must equal −1. Each nitrate ion contains 3 oxygen atoms, each of which has an oxidation number of −2. Therefore, the total charge due to oxygen is −6.

Since the net charge of the ion is −1, the oxidation number of the nitrogen atom must be +5.

$$\overset{+5}{N}\overset{-2}{O_3^-} \quad \text{for each atom}$$

Step 3. Check that the sum of the oxidation numbers is equal to the charge on the polyatomic ion.

$$\underset{\text{number of N atoms}}{\underset{\uparrow}{1}}(+5) + \underset{\text{number of O atoms}}{\underset{\uparrow}{3}}(-2) = -1$$

In the nitrate ion, NO_3^-, the oxidation number of the nitrogen atom is +5 and the oxidation number of each oxygen atom is –2.

Sample Problem 3: Assigning Oxidation Numbers for an Ionic Compound

Assign the oxidation numbers to all atoms or ions in sodium thiosulfate, $Na_2S_2O_3$.

Solution

Step 1. Assign oxidation numbers to elements as listed in Table 1.

The oxidation number of the sodium ion, a monatomic ion, is +1. The oxidation number of oxygen is –2.

Step 2. Identify any elements not mentioned in Table 1, and use the zero-sum rule to assign their oxidation numbers.

The sum of the oxidation numbers of the atoms in a compound must equal zero. Since each sodium ion has an oxidation number of +1 and there are 2 sodium ions, the total contribution of sodium is +2. The polyatomic thiosulfate ion must therefore have an overall charge of −2.

Since each oxygen atom has an oxidation number of −2, and there are 3 oxygen atoms, the total contribution from oxygen is −6. Since the total charge of the ion is −2, the total contribution from the 2 sulfur atoms is +4. Therefore, the oxidation number for each sulfur atom is +2.

$$\overset{+1}{Na_2}\overset{+2}{S_2}\overset{-2}{O_3} \quad \text{for each atom}$$

Step 3. Check that the sum of the oxidation numbers is equal to zero.

$$\underset{\text{number of Na atoms}}{\underset{\uparrow}{2}}(+1) + \underset{\text{number of S atoms}}{\underset{\uparrow}{2}}(+2) + \underset{\text{number of O atoms}}{\underset{\uparrow}{3}}(-2) = 0$$

In sodium thiosulfate, $Na_2S_2O_3$, the oxidation number of each sodium ion is +1; the oxidation number of each sulfur atom is +2; and the oxidation number of each oxygen atom is −2.

Practice

1. Determine the oxidation number of nitrogen in each of the following substances: T/I
   (a) N_2
   (b) NO_2
   (c) N_2O
   (d) $NaNO_3$
   (e) NH_3
2. Determine the oxidation number of carbon in each of the following compounds: T/I
   (a) CO
   (b) CH_4
   (c) Na_2CO_3
   (d) $C_6H_{12}O_6$
3. Determine the oxidation number of sulfur in each of the following substances: T/I
   (a) SO_2
   (b) SO_3^{2-}
   (c) SO_4^{2-}
   (d) $S_2O_8^{2-}$
4. Assign the oxidation number to each element in each of the following compounds: T/I
   (a) Na_2CO_3
   (b) $K_2Cr_2O_7$
   (c) $HClO_4$
   (d) $Cu_3(PO_4)_2$

Oxidation Numbers in Redox Reactions

We can use oxidation numbers to identify the reactant that is oxidized and the reactant that is reduced in a redox reaction. The reaction of nickel(II) oxide with carbon is a useful example. The mining industry uses this reaction during the extraction of nickel from its ore. CAREER LINK

Assigning oxidation numbers to each atom/ion in the chemical equation gives

$$\overset{+2-2}{\mathrm{NiO(s)}} + \overset{0}{\mathrm{C(s)}} \longrightarrow \overset{0}{\mathrm{Ni(s)}} + \overset{+2\,-2}{\mathrm{CO(g)}}$$

Note that the oxidation number of carbon changes from 0 to +2. This means that the carbon atom loses 2 electrons and becomes oxidized. The oxidation number of nickel changes from +2 to 0. This means that the Ni^{2+} ion in NiO gains 2 electrons and is reduced to nickel metal, Ni (**Figure 4**).

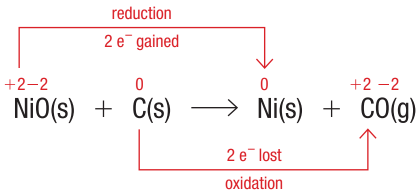


Figure 4 Note how the oxidation numbers change during the reaction of nickel oxide with carbon.

The reaction of nickel(II) oxide with carbon involves the direct transfer of electrons from one reactant to another. However, many redox reactions involve the partial transfer of electrons from one reactant to another. This occurs in the redox reactions involving molecular substances. Consider the synthesis of water from its elements:

$$2\,H_2(g) + O_2(g) \rightarrow 2\,H_2O(g)$$

Using the rules from Table 1, we can assign the following oxidation numbers:

$$2\,\overset{0}{H_2}(g) + \overset{0}{O_2}(g) \longrightarrow 2\,\overset{+1\,-2}{H_2O}(g)$$

Note that the oxidation number of each atom in H_2 and O_2 is 0 because both hydrogen and oxygen are in their elemental form. Also, the oxidation number of hydrogen changes from 0 in H_2 to +1 in H_2O. This change indicates that each hydrogen atom effectively loses 1 electron in the reaction. However, the oxidation number of each oxygen atom changes from 0 in O_2 to −2 in H_2O. This means that each oxygen atom effectively gains 2 electrons. Since water is a molecular compound rather than an ionic compound, we know that the electrons are not completely transferred from the hydrogen atoms to the oxygen atom. Instead, the electrons in the covalent bond between oxygen and hydrogen shift toward the more electronegative element: oxygen. Hence, only a partial transfer of electrons occurs during the synthesis of water.

oxidizing agent the reactant that is reduced (gains electrons from another substance) during an oxidation–reduction reaction

reducing agent the reactant that is oxidized (loses electrons to another substance) during an oxidation–reduction reaction

Oxidizing Agents and Reducing Agents

Oxidation–reduction reactions involve the transfer of electrons (**Figure 5**). The substance that gains electrons is referred to as an **oxidizing agent**. The oxidizing agent is reduced in a redox reaction. In the previous example, the oxidizing agent is oxygen gas, since it gained electrons from hydrogen. The corresponding substance that loses electrons is called the **reducing agent**. The reducing agent in the example above is hydrogen, since it lost electrons to oxygen. The reducing agent is oxidized in a redox reaction. In all oxidation–reduction reactions, electrons transfer from the reducing agent to the oxidizing agent (**Table 2**).

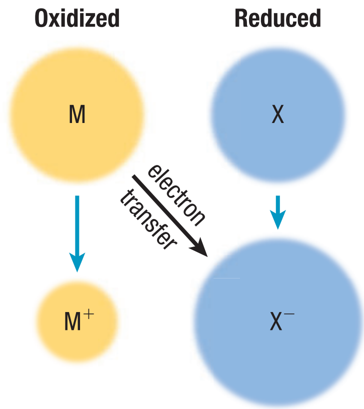


Figure 5 A summary of the oxidation–reduction process in which M is oxidized and X is reduced.

Table 2 Characteristics of Oxidizing and Reducing Agents

Oxidizing agent	Reducing agent
causes oxidation	causes reduction
gains electrons	loses electrons
is reduced	is oxidized

As an example, consider the synthesis of water:

$$2\,\overset{0}{H_2}(g) + \overset{0}{O_2}(g) \longrightarrow 2\,\overset{+1\,-2}{H_2O}(g)$$

For this reaction, we can say the following:

- Hydrogen is oxidized because its oxidation number increases. Each hydrogen atom has partially lost its electron.
- Oxygen is reduced because its oxidation number decreases. The oxygen atom has partially gained electrons.
- H_2 is the reducing agent.
- O_2 is the oxidizing agent.

Note that when you name the reducing or oxidizing agent, you specify the whole substance and not just the element that undergoes the change in oxidation number. This becomes more relevant when we consider redox reactions involving compounds: the compound is the oxidizing (or reducing) agent.

Investigation 9.1.1

Single Displacement Reactions (page 624)
In this investigation, you will design and perform tests on metals and metal ions, and then rank the strength of the ions as oxidizing agents.

Tutorial 3 Analyzing Redox Reactions

In this tutorial, you will identify atoms that become oxidized and reduced during an oxidation–reduction reaction. It will also help you identify the oxidizing and reducing agents in a redox reaction. Note that you always start by assigning an oxidation number to each element, just as you did in Tutorial 2.

Sample Problem 1: Analyzing a Redox Reaction

Metallurgy, the process of producing a metal from its ore, always involves oxidation–reduction reactions. The valuable component of lead ore is the compound lead(II) sulfide, PbS(s), also called galena. The first step in extracting lead from its ore is the conversion of lead(II) sulfide to its oxide by a process called roasting:

(a) $2\ PbS(s) + 3\ O_2(g) \rightarrow 2\ PbO(s) + 2\ SO_2(g)$

The oxide then reacts with carbon monoxide to produce the metallic element:

(b) $PbO(s) + CO(g) \rightarrow Pb(s) + CO_2(g)$

For each reaction, identify the entities that are oxidized and reduced, and specify the oxidizing and reducing agents.

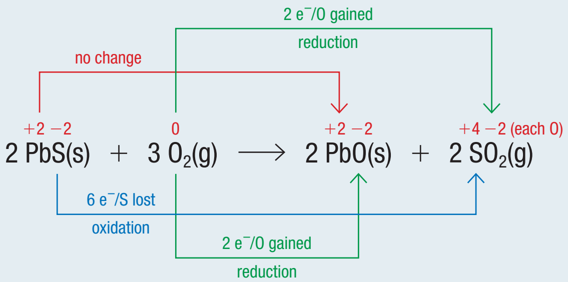
Solution (a)

Step 1. Assign oxidation numbers to elements as listed in Table 1.

Lead(II) sulfide and lead(II) oxide are simple ionic compounds, so the oxidation numbers of their elements are the same as the charges on those elements.

$$\overset{+2\ -2}{2\ PbS(s)} + \overset{0}{3\ O_2(g)} \rightarrow \overset{+2\ -2}{2\ PbO(s)} + \overset{+4\ -2\ (\text{each O})}{2\ SO_2(g)}$$

Step 2. Determine how the oxidation number on each element changes, and identify these changes as either oxidation or reduction.



The oxidation number for sulfur increases from −2 to +4. Sulfur is oxidized (**Figure 6**). The oxidation number for oxygen decreases from 0 to −2 (in both products). Oxygen is reduced. The oxidation number for lead does not change. The oxidizing agent (the substance that gains electrons) is O_2. The reducing agent (the substance that loses electrons) is PbS.

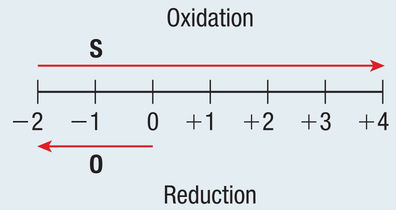


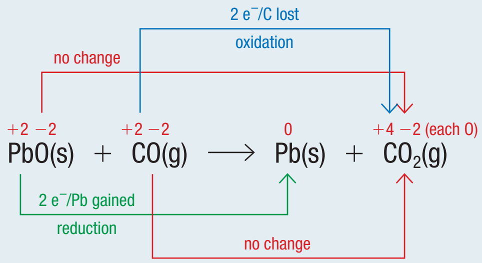
Figure 6 This number line shows the changes in oxidation number of sulfur and oxygen.

Solution (b)

Step 1. Assign oxidation numbers to elements as listed in Table 1.

$$\overset{+2\ -2}{PbO(s)} + \overset{+2\ -2}{CO(g)} \rightarrow \overset{0}{Pb(s)} + \overset{+4\ -2\ (\text{each O})}{CO_2(g)}$$

Step 2. Determine how the oxidation number on each element changes, and identify these changes as either oxidation or reduction.



The oxidation number for lead decreases from +2 to 0. Lead is reduced. The oxidation number for carbon increases from +2 to +4. Carbon is oxidized. The oxidation number for oxygen does not change (in both products). The oxidizing agent (the substance that gains electrons) is PbO. The reducing agent (the substance that loses electrons) is CO.

Practice

1. Identify the entities that are oxidized and reduced in the following reactions: T/I
   (a) $Cu(s) + 2\ Ag^+(aq) \rightarrow 2\ Ag(s) + Cu^{2+}(aq)$ (b) $4\ Fe(s) + 3\ O_2(g) \rightarrow 2\ Fe_2O_3(s)$
2. Identify the oxidizing agent and the reducing agent in the following reactions: T/I
   (a) $Zn(s) + 2\ HCl(aq) \rightarrow ZnCl_2(aq) + H_2(g)$ (b) $SnO_2(s) + C(s) \rightarrow Sn(s) + CO_2(g)$
3. For the following reactions, assign the oxidation numbers to each atom, then indicate the oxidizing and reducing agents: T/I
   (a) $2\ H_2S(g) + 3\ O_2(g) \rightarrow SO_2(g) + H_2O(g)$ (b) $CH_4(g) + 2\ O_2(g) \rightarrow CO_2(g) + 2\ H_2O(g)$

9.1 Review

Summary

- Many reactions involve the transfer of electrons.
- Oxidation–reduction reactions, or redox reactions, are reactions in which electrons transfer from one entity to another. Neither oxidation nor reduction can occur without the other.
- Chemists assign oxidation numbers to atoms or ions to keep track of the electrons in an oxidation–reduction reaction.
- The rules in Table 1 (page 603) allow oxidation numbers to be assigned for common entities. Oxidation numbers can then be assigned to remaining atoms or ions.
- In an oxidation–reduction reaction, one element is oxidized (loses electrons). Another element is reduced (gains electrons).
- An oxidizing agent causes oxidation to occur (and is reduced in the process). A reducing agent causes reduction to occur (and is oxidized in the process).

UNIT TASK BOOKMARK

You can apply what you learned about redox reactions to the Unit Task described on page 684.

Questions

1. Explain, in your own words, the process that occurs in a redox reaction. K/U
2. Write the oxidation and reduction half-reaction equations for the reactions represented by each of the following equations: T/I
   (a) $Mg(s) + 2\,H^+(aq) \rightarrow Mg^{2+}(aq) + H_2(g)$
   (b) $2\,Al(s) + Fe_2O_3(s) \rightarrow 2\,Fe(l) + Al_2O_3(s)$
3. Identify the oxidation number of the specified element in each of the following entities: T/I
   (a) S in S_8
   (b) Cr in $Cr_2O_7^{\,2-}$
   (c) N in N_2H_4
   (d) I in MgI_2
   (e) C in CO
   (f) N in NH_3
   (g) P in P_4O_6
   (h) Mn in $MnO_4^{\,-}$
   (i) C in C_2H_5OH
   (j) S in $Al_2(SO_3)_3$
4. Identify the entity that is oxidized and the entity that is reduced in the chemical reactions represented by the following equations: T/I
   (a) $CH_4(g) + H_2O(g) \rightarrow CO(g) + 3\,H_2(g)$
   (b) $8\,H^+(aq) + MnO_4^-(aq) + Fe^{2+}(aq) \rightarrow Mn^{2+}(aq) + Fe^{3+}(aq) + 4\,H_2O(l)$
   (c) a single displacement reaction in which copper metal reacts with silver nitrate
5. (a) Describe the change in the oxidation number of an entity when it becomes oxidized.
   (b) Describe the change in the oxidation number of an entity when it is reduced.
   (c) Why do oxidation and reduction half-reactions both have to occur in the same reaction? K/U
6. Use oxidation numbers to determine which of the following chemical equations represent a redox reaction. For all that ARE redox reactions, indicate the oxidizing agent, the reducing agent, the element being oxidized, and the element being reduced. T/I
   (a) $HCl(g) + NH_3(g) \rightarrow NH_4Cl(s)$
   (b) $SiCl_4(l) + 2\,Mg(s) \rightarrow 2\,MgCl_2(s) + Si(s)$
   (c) $CO(g) + H_2O(g) \rightarrow CO_2(g) + H_2(g)$
7. Assign oxidation numbers to each element in the following equations: T/I
   (a) $4\,PH_3(g) + 8\,O_2(g) \rightarrow P_4O_{10}(s) + 6\,H_2O(l)$
   (b) $2\,KClO_3(s) \rightarrow 2\,KCl(s) + 3\,O_2(g)$
   (c) $Pb(s) + PbO_2(s) + 2\,H_2SO_4(aq) \rightarrow 2\,PbSO_4(s) + 2\,H_2O(g)$
8. Describe each of the reactions in Question 7 using the terms "oxidation," "reduction," "oxidizing agent," and "reducing agent." K/U
9. In the process of photosynthesis, carbon dioxide and water form glucose and oxygen: T/I A
   $6\,CO_2(g) + 6\,H_2O(l) \rightarrow C_6H_{12}O_6(aq) + 6\,O_2(g)$
   (a) What is the oxidation number of carbon in CO_2 and $C_6H_{12}O_6$?
   (b) Describe the process of photosynthesis using the terms "oxidation," "reduction," "oxidizing agent," and "reducing agent."
10. Carbon dioxide can be progressively reduced to methane through a series of reactions. What is the oxidation number of carbon in each of the following compounds? T/I
    $CO_2 \rightarrow CH_2O_2 \rightarrow CH_2O \rightarrow CH_4O \rightarrow CH_4$