

3.5: Types of Chemical Reactions

Learning Objectives

- To identify fundamental types of chemical reactions.
- To predict the types of reactions substances will undergo.

The chemical reactions we have described are only a tiny sampling of the infinite number of chemical reactions possible. How do chemists cope with this overwhelming diversity? How do they predict which compounds will react with one another and what products will be formed? The key to success is to find useful ways to categorize reactions. Familiarity with a few basic types of reactions will help you to predict the products that form when certain kinds of compounds or elements come in contact.

Most chemical reactions can be classified into one or more of five basic types: acid–base reactionsA reaction of the general form acid + base → salt., exchange reactionsA chemical reaction that has the general form $AB + C \rightarrow AC + B$ or $AB + CD \rightarrow AD + CB$., condensation reactionsA chemical reaction that has the general form $A + B \rightarrow AB$. Condensation reactions are the reverse of cleavage reactions. Some, but not all, condensation reactions are also oxidation–reduction reactions. (and the reverse, cleavage reactionsA chemical reaction that has the general form $AB \rightarrow A + B$. Cleavage reactions are the reverse of condensation reactions.), and oxidation–reduction reactionsA chemical reaction that exhibits a change in the oxidation states of one or more elements in the reactants that has the general form oxidant + reductant → reduced oxidant + oxidized reductant.. The general forms of these five kinds of reactions are summarized in Table 3.5.1, along with examples of each. It is important to note, however, that many reactions can be assigned to more than one classification, as you will see in our discussion. The classification scheme is only for convenience; the same reaction can be classified in different ways, depending on which of its characteristics is most important. Oxidation–reduction reactions, in which there is a net transfer of electrons from one atom to another, and condensation reactions are discussed in this section. Acid–base reactions and one kind of exchange reaction—the formation of an insoluble salt such as barium sulfate when solutions of two soluble salts are mixed together—will be discussed in [Chapter 8](#).

Table 3.5.1 Basic Types of Chemical Reactions

Name of Reaction	General Form	Example(s)
oxidation–reduction (redox)	oxidant + reductant → reduced oxidant + oxidized reductant	$C_7H_{16}(l) + 11O_2(g) \rightarrow 7CO_2(g) + 8H_2O(g)$
acid–base	acid + base → salt	$NH_3(aq) + HNO_3(aq) \rightarrow NH_4^+(aq) + NO_3^-(aq)$
exchange	$AB + C \rightarrow AC + B$	$CH_3Cl + OH^- \rightarrow CH_3OH + Cl^-$
	$AB + CD \rightarrow AD + CB$	$BaCl_2(aq) + Na_2SO_4(aq) \rightarrow BaSO_4(s) + 2NaCl(aq)$
condensation	$A + B \rightarrow AB$	$CO_2(g) + H_2O(l) \rightarrow H_2CO_3(aq)$
		$HBr + H_2C=CH_2 \rightarrow CH_3CH_2Br^*$
cleavage	$AB \rightarrow A + B$	$CaCO_3(s) \rightarrow CaO(s) + CO_2(g)$
		$CH_3CH_2Cl \rightarrow H_2C=CH_2 + HCl^{**}$

* In more advanced chemistry courses you will learn that this reaction is also called an *addition* reaction.

** In more advanced chemistry courses you will learn that this reaction is also called an *elimination* reaction.

Oxidation–Reduction Reactions

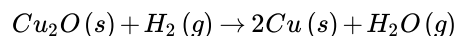
The term oxidation refers to the loss of one or more electrons in a chemical reaction. The substance that loses electrons is said to be oxidized. was first used to describe reactions in which metals react with oxygen in air to produce metal oxides. When iron is exposed to air in the presence of water, for example, the iron turns to rust—an iron oxide. When exposed to air, aluminum metal develops a continuous, coherent, transparent layer of aluminum oxide on its surface. In both cases, the metal acquires a positive charge by transferring electrons to the neutral oxygen atoms of an oxygen molecule. As a result, the oxygen atoms acquire a

negative charge and form oxide ions (O^{2-}). Because the metals have lost electrons to oxygen, they have been *oxidized*; oxidation is therefore the loss of electrons. Conversely, because the oxygen atoms have gained electrons, they have been *reduced*, so reduction is the gain of electrons. For every oxidation, there must be an associated reduction.

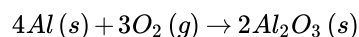
Note the Pattern

Any oxidation *must* be accompanied by a reduction and vice versa.

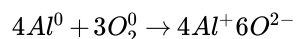
The term reduction refers to the gain of one or more electrons in a chemical reaction. The substance that gains electrons is said to be reduced. referred to the decrease in mass observed when a metal oxide was heated with carbon monoxide, a reaction that was widely used to extract metals from their ores. When solid copper(I) oxide is heated with hydrogen, for example, its mass decreases because the formation of pure copper is accompanied by the loss of oxygen atoms as a volatile product (water). The reaction is as follows:



Oxidation and reduction reactions are now defined as reactions that exhibit a change in the oxidation states of one or more elements in the reactants, which follows the mnemonic *oxidation is loss reduction is gain*, or *oil rig*. The oxidation state: The charge that each atom in a compound would have if all its bonding electrons were transferred to the atom with the greater attraction for electrons. of each atom in a compound is the charge an atom would have if all its bonding electrons were transferred to the atom with the greater attraction for electrons. Atoms in their elemental form, such as O_2 or H_2 , are assigned an oxidation state of zero. For example, the reaction of aluminum with oxygen to produce aluminum oxide is



Each neutral oxygen atom gains two electrons and becomes negatively charged, forming an oxide ion; thus, oxygen has an oxidation state of -2 in the product and has been reduced. Each neutral aluminum atom loses three electrons to produce an aluminum ion with an oxidation state of $+3$ in the product, so aluminum has been oxidized. In the formation of Al_2O_3 , electrons are transferred as follows (the superscript 0 emphasizes the oxidation state of the elements):



Equation 3.5.1 and Equation 3.5.2 are examples of oxidation–reduction (redox) reactions. In redox reactions, there is a net transfer of electrons from one reactant to another. In any redox reaction, the total number of electrons lost must equal the total of electrons gained to preserve electrical neutrality. In Equation 3.5.3, for example, the total number of electrons lost by aluminum is equal to the total number gained by oxygen:

$$\begin{aligned} \text{electrons lost} &= 4 \text{ Al } \cancel{\text{atoms}} \times \frac{3 \text{ e}^- \text{ lost}}{\cancel{\text{Al atoms}}} = 12 \text{ e}^- \text{ lost} \\ \text{electrons gained} &= 6 \text{ O } \cancel{\text{atoms}} \times \frac{2 \text{ e}^- \text{ gained}}{\cancel{\text{O atoms}}} = 12 \text{ e}^- \text{ gained} \end{aligned} \quad (3.5.1)$$

The same pattern is seen in all oxidation–reduction reactions: the number of electrons lost must equal the number of electrons gained.

Note the Pattern

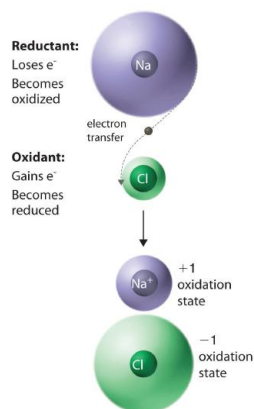
In all oxidation–reduction (redox) reactions, the number of electrons lost equals the number of electrons gained.

Assigning Oxidation States

Assigning oxidation states to the elements in binary ionic compounds is straightforward: the oxidation states of the elements are identical to the charges on the monatomic ions. In [Chapter 5](#), you learned how to predict the formulas of simple ionic compounds based on the sign and magnitude of the charge on monatomic ions formed by the neutral elements. Examples of such compounds are sodium chloride ($NaCl$; Figure 3.5.1, magnesium oxide (MgO), and calcium chloride ($CaCl_2$). In covalent compounds, in contrast, atoms share electrons. Oxidation states in covalent compounds can be understood in terms of polar covalent bonds and electronegativity. Oxidation states assume that electrons are completely transferred. While this is not true, it is almost true, and the oxidation state model can help you understand and predict many reactions. For example, in water, the oxidation state of the oxygen atom is -2 and that of each hydrogen is $+1$. That is the oxidation state model assumes that the hydrogen atoms transfer an electron

completely to the oxygen atom. We know this is not the case, but that the electrons from the hydrogen atoms are found closer to the oxygen atom. A more extreme case is carbon dioxide, where the oxidation states of the more electronegative oxygen atoms are -2 and that of the carbon atom is $+4$. Again, the idea is that the oxidation state model assumes complete transfer, and the atomic bond model tell us that the transfer is only partial. Still, the oxidation state model is useful as we shall see

Figure 3.5.1 The Reaction of a Neutral Sodium Atom with a Neutral Chlorine Atom



The result is the transfer of one electron from sodium to chlorine, forming the ionic compound NaCl.

A set of rules for assigning oxidation states to atoms in chemical compounds follows.

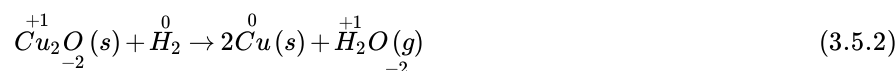
Rules for Assigning Oxidation States:

1. The oxidation state of an atom in any pure element, whether monatomic, diatomic, or polyatomic, is zero.
2. The oxidation state of a monatomic ion is the same as its charge—for example, $\text{Na}^+ = +1$, $\text{Cl}^- = -1$.
3. The oxidation state of fluorine in chemical compounds is always -1 . Other halogens usually have oxidation states of -1 as well, except when combined with oxygen or other halogens.
4. Hydrogen is assigned an oxidation state of $+1$ in its compounds with nonmetals and -1 in its compounds with metals.
5. Oxygen is normally assigned an oxidation state of -2 in compounds, with two exceptions: in compounds that contain oxygen–fluorine or oxygen–oxygen bonds, the oxidation state of oxygen is determined by the oxidation states of the other elements present.
6. The sum of the oxidation states of all the atoms in a neutral molecule or ion must equal the charge on the molecule or ion.

In any chemical reaction, the net charge must be conserved; that is, in a chemical reaction, the total number of electrons is constant, just like the total number of atoms. Consistent with this, rule 1 states that the sum of the individual oxidation states of the atoms in a molecule or ion must equal the net charge on that molecule or ion. In NaCl, for example, Na has an oxidation state of $+1$ and Cl is -1 . The net charge is zero, as it must be for any compound.

Rule 3 is required because fluorine attracts electrons more strongly than any other element, for reasons you discovered in [Chapter 7](#). Hence fluorine provides a reference for calculating the oxidation states of other atoms in chemical compounds. Rule 4 reflects the difference in chemistry observed for compounds of hydrogen with nonmetals (such as chlorine) as opposed to compounds of hydrogen with metals (such as sodium). For example, NaH contains the H^- ion, whereas HCl forms H^+ and Cl^- ions when dissolved in water. Rule 5 is necessary because fluorine has a greater attraction for electrons than oxygen does; this rule also prevents violations of rule 2. So the oxidation state of oxygen is $+2$ in OF_2 but $-\frac{1}{2}$ in KO_2 . Note that an oxidation state of $-\frac{1}{2}$ for O in KO_2 is perfectly acceptable. Nonintegral oxidation states are encountered occasionally. They are usually due to the presence of two or more atoms of the same element with different oxidation states, so that the "average" oxidation state is a fraction. As with resonance structures, this is a failure of the model.

The reduction of copper(I) oxide shown in Equation 3.5.5 demonstrates how to apply these rules. Rule 1 states that atoms in their elemental form have an oxidation state of zero, which applies to H_2 and Cu. From rule 4, hydrogen in H_2O has an oxidation state of $+1$, and from rule 5, oxygen in both Cu_2O and H_2O has an oxidation state of -2 . Rule 6 states that the sum of the oxidation states in a molecule or formula unit must equal the net charge on that compound. This means that each Cu atom in Cu_2O must have a charge of $+1$: $2(+1) + (-2) = 0$. So the oxidation states are as follows:



Assigning oxidation states allows us to see that there has been a net transfer of electrons from hydrogen ($0 \rightarrow +1$) to copper ($+1 \rightarrow 0$). So this is a redox reaction. Once again, the number of electrons lost equals the number of electrons gained, and there is a net conservation of charge:

$$\begin{aligned} \text{electrons lost} &= 2 \text{ H } \cancel{\text{atoms}} \times \frac{1 \text{ e}^- \text{ lost}}{\cancel{\text{H atoms}}} = 2 \text{ e}^- \text{ lost} \\ \text{electrons gained} &= 2 \text{ Cu } \cancel{\text{atoms}} \times \frac{1 \text{ e}^- \text{ gained}}{\cancel{\text{Cu atoms}}} = 2 \text{ e}^- \text{ gained} \end{aligned} \quad (3.5.3)$$

Remember that oxidation states are useful for visualizing the transfer of electrons in oxidation–reduction reactions, but the oxidation state of an atom and its actual charge are the same only for simple ionic compounds. Oxidation states are a convenient way of assigning electrons to atoms, and they are useful for predicting the types of reactions that substances undergo.

✓ Example 3.5.1

Assign oxidation states to all atoms in each compound.

1. sulfur hexafluoride (SF_6)
2. methanol (CH_3OH)
3. ammonium sulfate $[(\text{NH}_4)_2\text{SO}_4]$
4. magnetite (Fe_3O_4)
5. ethanoic (acetic) acid ($\text{CH}_3\text{CO}_2\text{H}$)

Given: molecular or empirical formula

Asked for: oxidation states

Strategy:

Begin with atoms whose oxidation states can be determined unambiguously from the rules presented (such as fluorine, other halogens, oxygen, and monatomic ions). Then determine the oxidation states of other atoms present according to rule 1.

Solution:

1. We know from rule 3 that fluorine always has an oxidation state of -1 in its compounds. The six fluorine atoms in sulfur hexafluoride give a total negative charge of -6 . Because rule 1 requires that the sum of the oxidation states of all atoms be zero in a neutral molecule (here SF_6), the oxidation state of sulfur must be $+6$: $[(6 \text{ F atoms})(-1)] + [(1 \text{ S atom})(+6)] = 0$
2. According to rules 4 and 5, hydrogen and oxygen have oxidation states of $+1$ and -2 , respectively. Because methanol has no net charge, carbon must have an oxidation state of -2 : $[(4 \text{ H atoms})(+1)] + [(1 \text{ O atom})(-2)] + [(1 \text{ C atom})(-2)] = 0$
3. Note that $(\text{NH}_4)_2\text{SO}_4$ is an ionic compound that consists of both a polyatomic cation (NH_4^+) and a polyatomic anion (SO_4^{2-}) (see Table 6.2.1). We assign oxidation states to the atoms in each polyatomic ion separately. For NH_4^+ , hydrogen has an oxidation state of $+1$ (rule 4), so nitrogen must have an oxidation state of -3 : $[(4 \text{ H atoms})(+1)] + [(1 \text{ N atom})(-3)] = +1$, the charge on the NH_4^+ ion

For SO_4^{2-} , oxygen has an oxidation state of -2 (rule 5), so sulfur must have an oxidation state of $+6$:

$[(4 \text{ O atoms})(-2)] + [(1 \text{ S atom})(+6)] = -2$, the charge on the sulfate ion

4. Oxygen has an oxidation state of -2 (rule 5), giving an overall charge of -8 per formula unit. This must be balanced by the positive charge on three iron atoms, giving an oxidation state of $+8/3$ for iron:

Fractional oxidation states are allowed because oxidation states are a somewhat arbitrary way of keeping track of electrons. In fact, Fe_3O_4 can be viewed as having two Fe^{3+} ions and one Fe^{2+} ion per formula unit, giving a net positive charge of $+8$ per formula unit. Fe_3O_4 is a magnetic iron ore commonly called magnetite. In ancient times, magnetite was known as lodestone because it could be used to make primitive compasses that pointed toward Polaris (the North Star), which was called the “lodestar.”

5. Initially, we assign oxidation states to the components of $\text{CH}_3\text{CO}_2\text{H}$ in the same way as any other compound. Hydrogen and oxygen have oxidation states of +1 and -2 (rules 4 and 5, respectively), resulting in a total charge for hydrogen and oxygen of [

$$(4H_{atoms})(+1) + [(2O_{atoms})(-2)] = 0 \quad (3.5.4)$$

So the oxidation state of carbon must also be zero (rule 6). This is, however, an *average* oxidation state for the two carbon atoms present. Because each carbon atom has a different set of atoms bonded to it, they are likely to have different oxidation states. To determine the oxidation states of the individual carbon atoms, we use the same rules as before but with the additional assumption that bonds between atoms of the same element do not affect the oxidation states of those atoms. The carbon atom of the methyl group ($-\text{CH}_3$) is bonded to three hydrogen atoms and one carbon atom. We know from rule 4 that hydrogen has an oxidation state of +1, and we have just said that the carbon-carbon bond can be ignored in calculating the oxidation state of the carbon atom. For the methyl group to be electrically neutral, its carbon atom must have an oxidation state of -3. Similarly, the carbon atom of the carboxylic acid group ($-\text{CO}_2\text{H}$) is bonded to one carbon atom and two oxygen atoms. Again ignoring the bonded carbon atom, we assign oxidation states of -2 and +1 to the oxygen and hydrogen atoms, respectively, leading to a net charge of

$$[(2 \text{ O atoms})(-2)] + [(1 \text{ H atom})(+1)] = -3$$

To obtain an electrically neutral carboxylic acid group, the charge on this carbon must be +3. The oxidation states of the individual atoms in acetic acid are thus



Thus the sum of the oxidation states of the two carbon atoms is indeed zero.

? Exercise 3.5.1

Assign oxidation states to all atoms in each compound.

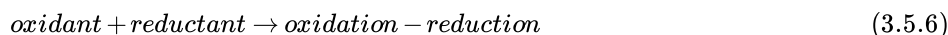
1. barium fluoride (BaF_2)
2. formaldehyde (CH_2O)
3. potassium dichromate ($\text{K}_2\text{Cr}_2\text{O}_7$)
4. cesium oxide (CsO_2)
5. ethanol ($\text{CH}_3\text{CH}_2\text{OH}$)

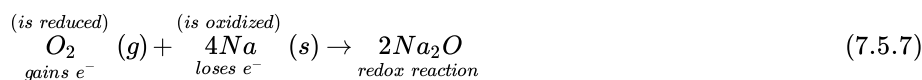
Answer

1. Ba, +2; F, -1
2. C, 0; H, +1; O, -2
3. K, +1; Cr, +6; O, -2
4. Cs, +1; O, -1/2
5. C, -3; H, +1; C, -1; H, +1; O, -2; H, +1

Oxidants and Reductants

Compounds that are capable of accepting electrons, such as O_2 or F_2 , are called oxidants (or oxidizing agents). A compound that is capable of accepting electrons; thus it is reduced, because they can oxidize other compounds. *In the process of accepting electrons, an oxidant is reduced.* Compounds that are capable of donating electrons, such as sodium metal or cyclohexane (C_6H_{12}), are called reductants (or reducing agents). A compound that is capable of donating electrons; thus it is oxidized, because they can cause the reduction of another compound. *In the process of donating electrons, a reductant is oxidized.* These relationships are summarized in [Equation 7.5.7](#):





Some oxidants have a greater ability than others to remove electrons from other compounds. Oxidants can range from very powerful, capable of oxidizing most compounds with which they come in contact, to rather weak. Both F_2 and Cl_2 are powerful oxidants: for example, F_2 will oxidize H_2O in a vigorous, potentially explosive reaction. In contrast, S_8 is a rather weak oxidant, and O_2 falls somewhere in between. Conversely, reductants vary in their tendency to donate electrons to other compounds. Reductants can also range from very powerful, capable of giving up electrons to almost anything, to weak. The alkali metals are powerful reductants, so they must be kept away from atmospheric oxygen to avoid a potentially hazardous redox reaction.

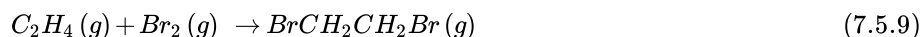
A combustion reactionAn oxidation–reduction reaction in which the oxidant is O_2 , first introduced in [Section 7.2](#) is an oxidation–reduction reaction in which the oxidant is O_2 . One example of a combustion reaction is the burning of a candle, shown in [Figure 7.3.3](#). Consider, for example, the combustion of cyclohexane, a typical hydrocarbon, in excess oxygen. The balanced chemical equation for the reaction, with the oxidation state shown for each atom, is as follows:



If we compare the oxidation state of each element in the products and the reactants, we see that hydrogen is the only element whose oxidation state does not change; it remains +1. Carbon, however, has an oxidation state of –2 in cyclohexane and +4 in CO_2 ; that is, each carbon atom changes its oxidation state by six electrons during the reaction. Oxygen has an oxidation state of 0 in the reactants, but it gains electrons to have an oxidation state of –2 in CO_2 and H_2O . Because carbon has been oxidized, cyclohexane is the reductant; because oxygen has been reduced, it is the oxidant. All combustion reactions are therefore oxidation–reduction reactions.

Condensation Reactions

The reaction of bromine with ethylene to give 1,2-dibromoethane, which is used in agriculture to kill nematodes in soil, is as follows:

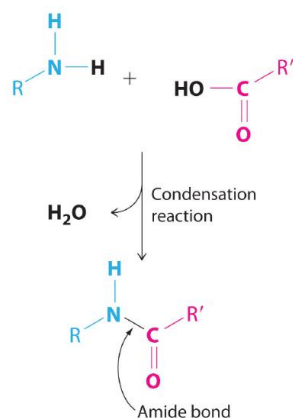


According to [Table 7.1.1](#), this is a condensation reaction because it has the general form $A + B \rightarrow AB$. This reaction, however, can also be viewed as an oxidation–reduction reaction, in which electrons are transferred from carbon ($-2 \rightarrow -1$) to bromine ($0 \rightarrow -1$). Another example of a condensation reaction is the one used for the industrial synthesis of ammonia:



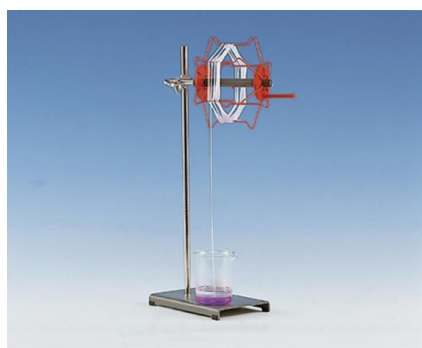
Although this reaction also has the general form of a condensation reaction, hydrogen has been oxidized ($0 \rightarrow +1$) and nitrogen has been reduced ($0 \rightarrow -3$), so it can also be classified as an oxidation–reduction reaction.

Not all condensation reactions are redox reactions. The reaction of an amine with a carboxylic acid, for example, is a variant of a condensation reaction ($A + B \rightarrow A'B' + C$): two large fragments condense to form a single molecule, and a much smaller molecule, such as H_2O , is eliminated. In this reaction, the $-OH$ from the carboxylic acid group and $-H$ from the amine group are eliminated as H_2O , and the reaction forms an *amide bond* (also called a *peptide bond*) that links the two fragments. Amide bonds are the essential structural unit linking the building blocks of proteins and many polymers together. Nylon, for example, is produced from a condensation reaction ([Figure 7.5.2](#)).



Amide bonds. The reaction of an amine with a carboxylic acid proceeds by eliminating water and forms a new C–N (amide) bond.

Figure 7.5.2 The Production of Nylon



✓ Example 3.5.2

The following reactions have important industrial applications. Using Table 3.5.1, classify each reaction as an oxidation–reduction reaction, an acid–base reaction, an exchange reaction, a condensation reaction, or a cleavage reaction. For each redox reaction, identify the oxidant and reductant and specify which atoms are oxidized or reduced. (Don’t forget that some reactions can be placed into more than one category.)

1. $\text{C}_2\text{H}_4(\text{g}) + \text{Cl}_2(\text{g}) \rightarrow \text{ClCH}_2\text{CH}_2\text{Cl}(\text{g})$
2. $\text{AgNO}_3(\text{aq}) + \text{NaCl}(\text{aq}) \rightarrow \text{AgCl}(\text{s}) + \text{NaNO}_3(\text{aq})$
3. $\text{CaCO}_3(\text{s}) \rightarrow \text{CaO}(\text{s}) + \text{CO}_2(\text{g})$
4. $\text{Ca}_5(\text{PO}_4)_3(\text{OH})(\text{s}) + 7\text{H}_3\text{PO}_4(\text{aq}) + 4\text{H}_2\text{O}(\text{l}) \rightarrow 5\text{Ca}(\text{H}_2\text{PO}_4)_2 \cdot \text{H}_2\text{O}(\text{s})$
5. $\text{Pb}(\text{s}) + \text{PbO}_2(\text{s}) + 2\text{H}_2\text{SO}_4(\text{aq}) \rightarrow 2\text{PbSO}_4(\text{s}) + 2\text{H}_2\text{O}(\text{l})$

Given: balanced chemical equation

Asked for: classification of chemical reaction

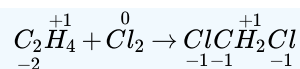
Strategy:

A Determine the general form of the equation by referring to Table 3.5.1 and then classify the reaction.

B For redox reactions, assign oxidation states to each atom present in the reactants and the products. If the oxidation state of one or more atoms changes, then the reaction is a redox reaction. If not, the reaction must be one of the other types of reaction listed in Table 3.5.1

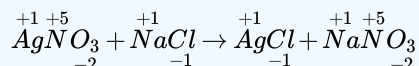
Solution

1. **A** This reaction is used to prepare 1,2-dichloroethane, one of the top 25 industrial chemicals. It has the general form $\text{A} + \text{B} \rightarrow \text{AB}$, which is typical of a condensation reaction. **B** Because reactions may fit into more than one category, we need to look at the oxidation states of the atoms:



The oxidation states show that chlorine is reduced from 0 to -1 and carbon is oxidized from -2 to -1, so this is a redox reaction as well as a condensation reaction. Ethylene is the reductant, and chlorine is the oxidant.

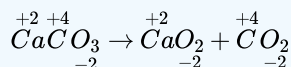
2. **A** This reaction is used to prepare silver chloride for making photographic film. The chemical equation has the general form $AB + CD \rightarrow AD + CB$, so it is classified as an exchange reaction. **B** The oxidation states of the atoms are as follows



There is no change in the oxidation states, so this is not a redox reaction.

AgCl(s) precipitates when solutions of AgNO₃(aq) and NaCl(aq) are mixed. NaNO₃ (aq) is in solution as Na⁺ and NO₃⁻ ions.

3. **A** This reaction is used to prepare lime (CaO) from limestone (CaCO₃) and has the general form $AB \rightarrow A + B$. The chemical equation's general form indicates that it can be classified as a cleavage reaction, the reverse of a condensation reaction. **B** The oxidation states of the atoms are as follows:

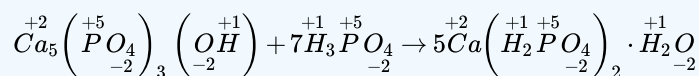


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Because the oxidation states of all the atoms are the same in the products and the reactant, this is *not* a redox reaction.

4. **A** This reaction is used to prepare "super triple phosphate" in fertilizer. One of the reactants is phosphoric acid, which transfers a proton (H⁺) to the phosphate and hydroxide ions of hydroxyapatite [Ca₅(PO₄)₃(OH)] to form H₂PO₄⁻ and H₂O, respectively. This is an acid-base reaction, in which H₃PO₄ is the acid (H⁺ donor) and Ca₅(PO₄)₃(OH) is the base (H⁺ acceptor).

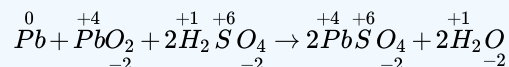
B To determine whether it is also a redox reaction, we assign oxidation states to the atoms:



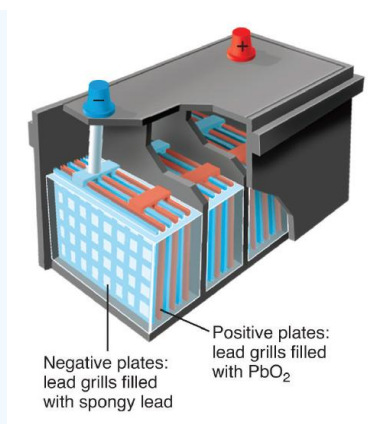
Because there is no change in oxidation state, this is not a redox reaction.

5. **A** This reaction occurs in a conventional car battery every time the engine is started. An acid (H₂SO₄) is present and transfers protons to oxygen in PbO₂ to form water during the reaction. The reaction can therefore be described as an acid-base reaction.

B The oxidation states are as follows:



The oxidation state of lead changes from 0 in Pb and +4 in PbO₂ (both reactants) to +2 in PbSO₄. This is also a redox reaction, in which elemental lead is the reductant, and PbO₂ is the oxidant. Which description is correct? Both.



Schematic drawing of a 12-volt car battery. The locations of the reactants (lead metal in a spongy form with large surface area) and PbO_2 are shown. The product (PbSO_4) forms as a white solid between the plates.

? Exercise 3.5.2

Using Table 7.5.1, classify each reaction as an oxidation–reduction reaction, an acid–base reaction, an exchange reaction, a condensation reaction, or a cleavage reaction. For each redox reaction, identify the oxidant and the reductant and specify which atoms are oxidized or reduced.

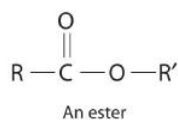
- $\text{Al(s)} + \text{OH}^{\text{--}}(\text{aq}) + 3\text{H}_2\text{O(l)} \rightarrow 3/2\text{H}_2(\text{g}) + [\text{Al}(\text{OH})_4]^{\text{--}}(\text{aq})$
- $\text{TiCl}_4(\text{l}) + 2\text{Mg(l)} \rightarrow \text{Ti(s)} + 2\text{MgCl}_2(\text{l})$
- $\text{MgCl}_2(\text{aq}) + \text{Na}_2\text{CO}_3(\text{aq}) \rightarrow \text{MgCO}_3(\text{s}) + 2\text{NaCl}(\text{aq})$
- $\text{CO(g)} + \text{Cl}_2(\text{g}) \rightarrow \text{Cl}_2\text{CO(l)}$
- $\text{H}_2\text{SO}_4(\text{l}) + 2\text{NH}_3(\text{g}) \rightarrow (\text{NH}_4)_2\text{SO}_4(\text{s})$

Answer

- Redox reaction; reductant is Al, oxidant is H_2O ; Al is oxidized, H is reduced. This is the reaction that occurs when Drano is used to clear a clogged drain.
- Redox reaction; reductant is Mg, oxidant is TiCl_4 ; Mg is oxidized, Ti is reduced.
- Exchange reaction. This reaction is responsible for the scale that develops in coffee makers in areas that have hard water.
- Both a condensation reaction and a redox reaction; reductant is CO, oxidant is Cl_2 ; C is oxidized, Cl is reduced. The product of this reaction is phosgene, a highly toxic gas used as a chemical weapon in World War I. Phosgene is now used to prepare polyurethanes, which are used in foams for bedding and furniture and in a variety of coatings.
- Acid–base reaction.

Catalysts

Many chemical reactions, including some of those discussed previously, occur more rapidly in the presence of a catalyst. A substance that increases the rate of a chemical reaction without undergoing a net chemical change itself, which is a substance that participates in a reaction and causes it to occur more rapidly but can be recovered unchanged at the end of a reaction and reused. Because catalysts are not involved in the stoichiometry of a reaction, they are usually shown above the arrow in a net chemical equation. Chemical processes in industry rely heavily on the use of catalysts, which are usually added to a reaction mixture in trace amounts, and most biological reactions do not take place without a biological catalyst or enzyme. Catalysts that occur naturally in living organisms and catalyze biological reactions. Examples of catalyzed reactions in industry are the use of platinum in petroleum cracking and reforming, the reaction of SO_2 and O_2 in the presence of V_2O_5 to produce SO_3 in the industrial synthesis of sulfuric acid, and the use of sulfuric acid in the synthesis of compounds such as ethyl acetate and procaine. Not only do catalysts greatly increase the rates of reactions, but in some cases such as in petroleum refining, they also control which products are formed. The acceleration of a reaction by a catalyst is called catalysis. The acceleration of a chemical reaction by a catalyst.



A heterogeneous catalyst. This large circular gauze, woven from rhodium-platinum wire, is a heterogeneous catalyst in the commercial production of nitric acid by the oxidation of ammonia.

Catalysts may be classified as either homogeneous or heterogeneous. A homogeneous catalyst is uniformly dispersed throughout the reactant mixture to form a solution. Sulfuric acid, for example, is a homogeneous catalyst used in the synthesis of esters such as procaine (Example 13). An ester has a structure similar to that of a carboxylic acid, in which the hydrogen atom attached to oxygen has been replaced by an R group. They are responsible for the fragrances of many fruits, flowers, and perfumes. Other examples of homogeneous catalysts are the enzymes that allow our bodies to function. In contrast, a heterogeneous catalyst is in a different physical state than the reactants. For economic reasons, most industrial processes use heterogeneous catalysts in the form of solids that are added to solutions of the reactants. Because such catalysts often contain expensive precious metals such as platinum or palladium, it makes sense to formulate them as solids that can be easily separated from the liquid or gaseous reactant-product mixture and recovered. Examples of heterogeneous catalysts are the iron oxides used in the industrial synthesis of ammonia and the catalytic converters found in virtually all modern automobiles, which contain precious metals like palladium and rhodium. Catalysis will be discussed in more detail when we discuss reaction rates, in the second semester but you will encounter the term frequently throughout the text.

Summary

Chemical reactions may be classified as an **acid–base reaction**, an **exchange reaction**, a **condensation reaction** and its reverse, a **cleavage reaction**, and an **oxidation–reduction (or redox) reaction**. To keep track of electrons in chemical reactions, oxidation states are assigned to atoms in compounds. The **oxidation state** is the charge an atom would have if all its bonding electrons were transferred completely to the atom that has the greater attraction for electrons. In an **oxidation–reduction reaction**, one atom must lose electrons and another must gain electrons. **Oxidation** is the loss of electrons, and an element whose oxidation state increases is said to be *oxidized*. **Reduction** is the gain of electrons, and an element whose oxidation state decreases is said to be *reduced*. **Oxidants** are compounds that are capable of accepting electrons from other compounds, so they are reduced during an oxidation–reduction reaction. In contrast, **reductants** are compounds that are capable of donating electrons to other compounds, so they are oxidized during an oxidation–reduction reaction. A **combustion reaction** is a redox reaction in which the oxidant is $\text{O}_2(\text{g})$. An **amide bond** is formed from the condensation reaction between a carboxylic acid and an amine; it is the essential structural unit of proteins and many polymers. A **catalyst** is a substance that increases the rate of a chemical reaction without undergoing a net chemical change itself. A biological catalyst is called an **enzyme**. **Catalysis** is an acceleration in the rate of a reaction caused by the presence of a substance that does not appear in the chemical equation. A **homogeneous catalyst** is uniformly dispersed in a solution of the reactants, whereas a **heterogeneous catalyst** is present as a different phase, usually a solid.

Key Takeaway

- Chemical reactions may be classified as acid–base, exchange, condensation, cleavage, and oxidation–reduction (redox).

Conceptual Problems

- What is a combustion reaction? How can it be distinguished from an exchange reaction?
- What two products are formed in the combustion of an organic compound containing only carbon, hydrogen, and oxygen? Is it possible to form only these two products from a reaction that is not a combustion reaction? Explain your answer.

3. What factors determine whether a reaction can be classified as a redox reaction?
4. Name three characteristics of a balanced redox reaction.
5. Does an oxidant accept electrons or donate them?
6. Does the oxidation state of a reductant become more positive or more negative during a redox reaction?
7. Nitrogen, hydrogen, and ammonia are known to have existed on primordial earth, yet mixtures of nitrogen and hydrogen do not usually react to give ammonia. What natural phenomenon would have enough energy to initiate a reaction between these two primordial gases?
8. Catalysts are not added to reactions in stoichiometric quantities. Why?
9. State whether each of the following uses a homogeneous catalyst or a heterogeneous catalyst.
 1. Platinum metal is used in the catalytic converter of an automobile.
 2. Nitrogen is biologically converted to ammonia by an enzyme.
 3. Carbon monoxide and hydrogen combine to form methane and water with a nickel catalyst.
 4. A dissolved rhodium compound is used as a catalyst for the conversion of an alkene to an alkane.
10. State whether each of the following uses a homogeneous catalyst or a heterogeneous catalyst.
 1. Pellets of ZSM-5, an aluminum- and silicon-containing mineral, are used to catalyze the conversion of methanol to gasoline.
 2. The conversion of glucose to a carboxylic acid occurs with catalysis by the enzyme glucose oxidase.
 3. Metallic rhodium is used to the conversion of carbon monoxide and water to carbon dioxide and hydrogen.
11. Complete the following table to describe some key differences between homogeneous and heterogeneous catalysis.

	Homogeneous	Heterogeneous
number of phases		
ease of separation from product		
ease of recovery of catalyst		

12. To increase the rate of a reaction, a scientist decided to use a catalyst. Unexpectedly, the scientist discovered that the catalyst decreased the yield of the desired product, rather than increasing it. What might have happened?

Answer

- 1.
- 2.
- 3.
- 4.
- 5.
- 6.
- 7.
- 8.
- 9.
- 10.

	Homogeneous	Heterogeneous
number of phases	single phase	at least two phases
ease of separation from product	difficult	easy
ease of recovery of catalyst	difficult	easy

- 12.

Numerical Problems

Please be sure you are familiar with the topics discussed in Essential Skills 2 (Section 7.7) before proceeding to the Numerical Problems.

1. Classify each chemical reaction according to the types listed in Table 7.5.1

1. $12\text{FeCl}_2(\text{s}) + 3\text{O}_2(\text{g}) \rightarrow 8\text{FeCl}_3(\text{s}) + 2\text{Fe}_2\text{O}_3(\text{s})$
2. $\text{CaCl}_2(\text{aq}) + \text{K}_2\text{SO}_4(\text{aq}) \rightarrow \text{CaSO}_4(\text{s}) + 2\text{KCl}(\text{aq})$
3. $\text{HCl}(\text{aq}) + \text{NaOH}(\text{aq}) \rightarrow \text{NaCl}(\text{aq}) + \text{H}_2\text{O}(\text{l})$
4. $\text{Br}_2(\text{l}) + \text{C}_2\text{H}_4(\text{g}) \rightarrow \text{BrCH}_2\text{CH}_2\text{Br}(\text{l})$

2. Classify each chemical reaction according to the types listed in Table 7.5.1

1. $4\text{FeO}(\text{s}) + \text{O}_2(\text{g}) \rightarrow 2\text{Fe}_2\text{O}_3(\text{s})$
2. $\text{Ca}_3(\text{PO}_4)_2(\text{s}) + 3\text{H}_2\text{SO}_4(\text{aq}) \rightarrow 3\text{CaSO}_4(\text{s}) + 2\text{H}_3\text{PO}_4(\text{aq})$
3. $\text{HNO}_3(\text{aq}) + \text{KOH}(\text{aq}) \rightarrow \text{KNO}_3(\text{aq}) + \text{H}_2\text{O}(\text{l})$
4. ethane(g) + oxygen(g) → carbon dioxide(g) + water(g)

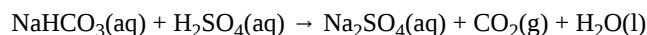
3. Assign oxidation states to the atoms in each compound or ion.

1. $(\text{NH}_4)_2\text{S}$
2. the phosphate ion
3. $[\text{AlF}_6]^{3-}$
4. CuS
5. HCO_3^-
6. NH_4^+
7. H_2SO_4
8. formic acid
9. *n*-butanol

4. Assign oxidation states to the atoms in each compound or ion.

1. ClO_2
2. HO_2^-
3. sodium bicarbonate
4. MnO_2
5. PCl_5
6. $[\text{Mg}(\text{H}_2\text{O})_6]^{2+}$
7. N_2O_4
8. butanoic acid
9. methanol

5. Balance this chemical equation:



What type of reaction is this? Justify your answer.

6. Assign oxidation states to the atoms in each compound.

1. iron(III) nitrate
2. Al_2O_3
3. potassium sulfate
4. Cr_2O_3
5. sodium perchlorate
6. Cu_2S
7. hydrazine (N_2H_4)
8. NO_2
9. *n*-pentanol

7. Assign oxidation states to the atoms in each compound.

1. calcium carbonate
2. NaCl
3. CO₂
4. potassium dichromate
5. KMnO₄
6. ferric oxide
7. Cu(OH)₂
8. Na₂SO₄
9. *n*-hexanol

8. For each redox reaction, determine the identities of the oxidant, the reductant, the species oxidized, and the species reduced.

1. $\text{H}_2(\text{g}) + \text{I}_2(\text{s}) \rightarrow 2\text{HI}(\text{g})$
2. $2\text{Na}(\text{s}) + 2\text{H}_2\text{O}(\text{l}) \rightarrow 2\text{NaOH}(\text{aq}) + \text{H}_2(\text{g})$
3. $2\text{F}_2(\text{g}) + 2\text{NaOH}(\text{aq}) \rightarrow \text{OF}_2(\text{g}) + 2\text{NaF}(\text{aq}) + \text{H}_2\text{O}(\text{l})$

9. For each redox reaction, determine the identities of the oxidant, the reductant, the species oxidized, and the species reduced.

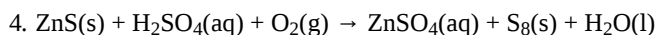
1. $2\text{Na}(\text{s}) + \text{Cl}_2(\text{g}) \rightarrow 2\text{NaCl}(\text{s})$
2. $\text{SiCl}_4(\text{l}) + 2\text{Mg}(\text{s}) \rightarrow 2\text{MgCl}_2(\text{s}) + \text{Si}(\text{s})$
3. $2\text{H}_2\text{O}_2(\text{aq}) \rightarrow 2\text{H}_2\text{O}(\text{l}) + \text{O}_2(\text{g})$

10. Balance each chemical equation. Then identify the oxidant, the reductant, the species oxidized, and the species reduced. (Δ indicates that the reaction requires heating.)

1. $\text{H}_2\text{O}(\text{g}) + \text{CO}(\text{g}) \rightarrow \text{CO}_2(\text{g}) + \text{H}_2(\text{g})$
2. the reaction of aluminum oxide, carbon, and chlorine gas at 900°C to produce aluminum chloride and carbon monoxide

11. Balance each chemical equation. Then identify the oxidant, the reductant, the species oxidized, and the species reduced. (Δ indicates that the reaction requires heating.)

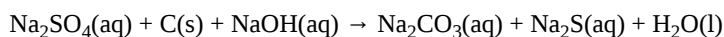
1. the reaction of water and carbon at 800°C to produce hydrogen and carbon monoxide
2. $\text{Mn}(\text{s}) + \text{S}_8(\text{s}) + \text{CaO}(\text{s}) \rightarrow \text{CaS}(\text{s}) + \text{MnO}(\text{s})$
3. the reaction of ethylene and oxygen at elevated temperature in the presence of a silver catalyst to produce ethylene oxide



12. Silver is tarnished by hydrogen sulfide, an atmospheric contaminant, to form a thin layer of dark silver sulfide (Ag₂S) along with hydrogen gas.

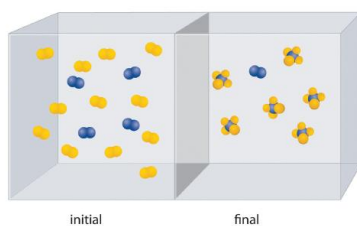
1. Write a balanced chemical equation for this reaction.
2. Which species has been oxidized and which has been reduced?
3. Assuming 2.2 g of Ag has been converted to silver sulfide, construct a table showing the reaction in terms of the number of atoms in the reactants and products, the moles of reactants and products, the grams of reactants and products, and the molecules of reactants and products.

13. The following reaction is used in the paper and pulp industry:

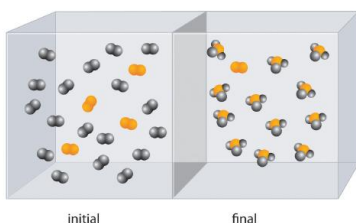


1. Balance the chemical equation.
2. Identify the oxidant and the reductant.
3. How much carbon is needed to convert 2.8 kg of sodium sulfate to sodium sulfide?
4. If the yield of the reaction were only 78%, how many kilograms of sodium carbonate would be produced from 2.80 kg of sodium sulfate?
5. If 240 g of carbon and 2.80 kg of sodium sulfate were used in the reaction, what would be the limiting reactant (assuming an excess of sodium hydroxide)?

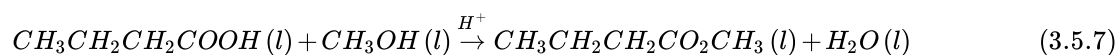
14. The reaction of A₂ (blue) with B₂ (yellow) is shown below. The initial reaction mixture is shown on the left and the mixture after the reaction has gone to completion is shown on the right.



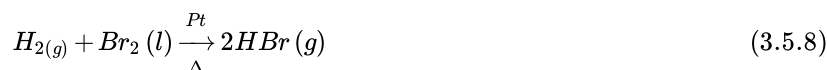
1. Write a balanced chemical equation for the reaction.
 2. Which is the limiting reactant in the initial reaction mixture?
 3. How many moles of the product AB_4 could you obtain from a mixture of 0.020 mol A_2 and 0.060 mol B_2 ?
15. The reaction of X_4 (orange) with Y_2 (black) is shown below. The initial reaction mixture is shown on the left and the mixture after the reaction has gone to completion is shown on the right.



1. Write a balanced chemical equation for the reaction.
 2. Which is the limiting reactant in the initial reaction mixture?
 3. How many moles of the product XY_3 could you obtain from a mixture of 0.100 mol X_4 and 0.300 mol Y_2 ?
16. Methyl butyrate, an artificial apple flavor used in the food industry, is produced by the reaction of butanoic acid with methanol in the presence of an acid catalyst (H^+):



1. Given 7.8 g of butanoic acid, how many grams of methyl butyrate would be synthesized, assuming 100% yield?
 2. The reaction produced 5.5 g of methyl butyrate. What was the percent yield?
 3. Is the catalyst used in this reaction heterogeneous or homogeneous?
17. In the presence of a platinum catalyst, hydrogen and bromine react at elevated temperatures ($300^\circ C$) to form hydrogen bromide (heat is indicated by Δ):



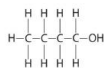
Given the following, calculate the mass of hydrogen bromide produced:

1. 8.23×10^{22} molecules of H_2
2. 6.1×10^3 mol of H_2
3. 1.3×10^5 g of H_2
4. Is the catalyst used in this reaction heterogeneous or homogeneous?

Answers

1. redox reaction
 2. exchange
 3. acid-base
 4. condensation
- 2.
3. 1. S, -2; N, -3; H, +1

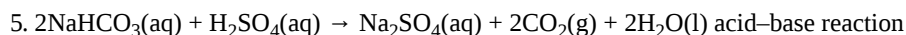
2. P, +5; O, -2
3. F, -1; Al, +3
4. S, -2; Cu, +2
5. H, +1; O, -2; C, +4
6. H, +1; N, -3
7. H, +1; O, -2; S, +6
8. H, +1, O, -2; C, +2
9. butanol:



O, -2; H, +1

From left to right: C, -3-2-2-1

4.



6.

7. 1. Ca, +2; O, -2; C, +4
2. Na, +1; Cl, -1
3. O, -2; C, +4
4. K, +1; O, -2; Cr, +6
5. K, +1; O, -2; Mn, +7
6. O, -2; Fe, +3
7. O, -2; H, +1; Cu, +2
8. O, -2; S, +6
9. Hexanol

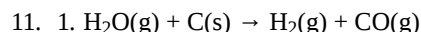
O, -2; H, +1

From left to right: C: -3, -2, -2, -2, -2, -1

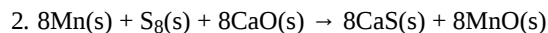
8.

9. 1. Na is the reductant and is oxidized. Cl_2 is the oxidant and is reduced.
2. Mg is the reductant and is oxidized. Si is the oxidant and is reduced.
3. H_2O_2 is both the oxidant and reductant. One molecule is oxidized, and one molecule is reduced.

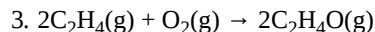
10.



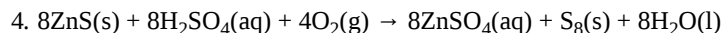
C is the reductant and is oxidized. H_2O is the oxidant and is reduced.



Mn is the reductant and is oxidized. The S_8 is the oxidant and is reduced.



Ethylene is the reductant and is oxidized. O_2 is the oxidant and is reduced.



Sulfide in ZnS is the reductant and is oxidized. O_2 is the oxidant and is reduced.

12.

13. 1. $\text{Na}_2\text{SO}_4 + 2\text{C} + 4\text{NaOH} \rightarrow 2\text{Na}_2\text{CO}_3 + \text{Na}_2\text{S} + 2\text{H}_2\text{O}$
2. The sulfate ion is the oxidant, and the reductant is carbon.
3. 470 g
4. 3300 g
5. carbon

- 14.
- 15.
- 16.
17.
 1. 22.1 g
 2. 9.9×10^5 g
 3. 1.0×10^7 g
 4. heterogeneous

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