Chemical Reactions

Nature of Chemical Reactions

Types of Chemical Reactions

Synthesis Reactions

Decomposition Reactions

Single-Replacement Reactions

Double-Replacement Reactions

Combustion Reactions

Identifying Oxidation-Reduction Reactions
Nature of Chemical Reactions

Objectives

1. Define chemical Reactions.
2. Describe four indications that a chemical reaction has occurred.
3. State the law of conservation of mass and describe how it relates to a chemical reaction.
4. Identify and use the common symbols used in writing chemical equations.
5. Translate chemical equations into word equations.
6. Translate word equations into chemical equations.
A chemical reaction is a process in which the chemical and physical properties of the original substance changes as new substances with different physical and chemical properties are formed.
In any chemical reaction, there are always two kinds of substances; the substances that are present before the change, called the reactants, and the substances that are formed by the change, called the products.

Reactants → Products
Identify the reactants and products in the following equation:

$$2\text{Al}(s) + 3\text{Cl}_2(g) \rightarrow 2\text{AlCl}_3(s)$$

**reactant(s)**  \( \text{Al, Cl}_2 \)

**products(s)**  \( \text{AlCl}_3 \)
Indications of a Chemical Reaction

What are the four observations that generally indicate that a chemical reaction has occurred?

1. Production of a gas.
2. Formation of a precipitate. (an insoluble solid that forms in an aqueous reaction)
3. Change in energy.
   - Endothermic reaction – energy is absorbed
   - Exothermic Reaction – energy is released
4. Change in color or odor.
Chemical equations are balanced to satisfy the Law of Conservation of Mass.

The Law of Conservation of Mass states that mass cannot be created or destroyed by ordinary physical or chemical means.

This means that the total mass of the reactions must equal the total mass of the products.
Remember that atoms don’t change in a chemical reaction; they just rearrange. For a chemical equation to accurately represent a reaction, the same number of each kind of atom must be on the left side of the arrow as are on the right side.
Symbols Used in Chemical Equations

+  Used to separate two reactants or two products

→  “yields”, separates reactants from products

⇌  Used in place of a → for reversible reactions

(s)  Designates a reactant or product in the solid state

(l)  Designates a reactant or product in the liquid state

(aq) Designates an aqueous solution, the substance is dissolved in water.
Symbols Used in Chemical Equations

\( (g) \) Designates a reactant or product in the gaseous state

\( \Delta \) or \( \text{heat} \) Indicates that heat is supplied to the reaction.

\( \text{Pt} \) A formula written above or below the yield sign indicates its use as a catalyst (in this example, platinum).

A catalyst is a substance that speeds up a reaction without being used up itself or permanently changed.

In biology, a catalyst is commonly called an enzyme.
Writing Sentences to Describe Balanced Equations

Sentences can be written to describe (interpret) what is happening in a balanced equation.

Example 1.

2NaOH(aq) + CO₂(g) → Na₂CO₃(s) + H₂O(l)

Aqueous sodium hydroxide reacts with carbon dioxide gas to produce solid sodium carbonate and water.
Writing Sentences to Describe Balanced Equations

Example 2.

\[ 2\text{Al}(s) + 3\text{Cl}_2(g) \rightarrow 2\text{AlCl}_3(s) \]

Aluminum metal reacts with chlorine gas to produce solid aluminum chloride.
You Try It

Write sentences to describe each of the following equations. (Do not worry about the coefficients at this time.)

i. $2H_2(g) + O_2(g) \rightarrow 2H_2O(l)$

Hydrogen gas reacts with oxygen gas to produce water.

ii. $Zn(s) + CuCl_2(aq) \rightarrow ZnCl_2(aq) + Cu(s)$

Zinc metal reacts with aqueous copper(II) chloride to produce aqueous zinc chloride and copper metal.
Before we can balance an equation, we must first make sure that everyone can count the atoms present in a compound. Here are some examples.

- **KClO$_3$**: \( \underline{1} \text{K}, \underline{1} \text{Cl}, \underline{3} \text{O} \)
- **Mg(OH)$_2$**: \( \underline{1} \text{Mg}, \underline{2} \text{O}, \underline{2} \text{H} \)
- **Al$_2$(SO$_4$)$_3$**: \( \underline{2} \text{Al}, \underline{3} \text{S}, \underline{12} \text{O} \)
- **2Al(NO$_3$)$_3$**: \( \underline{2} \text{Al}, \underline{6} \text{N}, \underline{18} \text{O} \)
You Try It

Determine the number of each atom present in each of the following.

\[\text{AlPO}_3\]  \[1\] \text{Al}, [1] \text{P}, [3] \text{O}

\[\text{Ba(NO}_3\text{)}_2\]  \[1\] \text{Ba}, [2] \text{N}, [6] \text{O}


\[3\text{Fe}_3\text{(PO}_4\text{)}_2\]  [9] \text{Fe}, [6] \text{P}, [24] \text{O}

Back to main menu
An equation in which the number of atoms of each element is the same on both sides of the equation is called a balanced chemical equation.

\[
Pb(NO_3)_2 + 2KI \rightarrow PbI_2 + 2KNO_3
\]

<table>
<thead>
<tr>
<th>Reactant Side</th>
<th>Element</th>
<th>Product Side</th>
</tr>
</thead>
<tbody>
<tr>
<td>1</td>
<td>Pb</td>
<td>1</td>
</tr>
<tr>
<td>2</td>
<td>N</td>
<td>2</td>
</tr>
<tr>
<td>6</td>
<td>O</td>
<td>6</td>
</tr>
<tr>
<td>2</td>
<td>K</td>
<td>2</td>
</tr>
<tr>
<td>2</td>
<td>I</td>
<td>2</td>
</tr>
</tbody>
</table>
You Try It

Classify each of the following equations as balanced or unbalanced.

a. $2\text{H}_2\text{O}_2(aq) \rightarrow 2\text{H}_2\text{O}(l) + \text{O}_2(g)$
   balanced

b. $\text{Mg}(s) + \text{HCl}(aq) \rightarrow \text{MgCl}_2(aq) + \text{H}_2(g)$
   unbalanced

c. $\text{Al}_4\text{C}_3(s) + \text{H}_2\text{O}(l) \rightarrow \text{CH}_4(g) + \text{Al(OH)}_3(s)$
   unbalanced

d. $\text{CH}_4(g) + 2\text{O}_2(g) \rightarrow \text{CO}_2(g) + 2\text{H}_2\text{O}(g)$
   balanced
Balancing Chemical Equations

The easiest equations to balance are the ones in which the skeleton equation is already written.

Example 1. \( \text{Cl}_2 + 2 \text{Na} \rightarrow 2 \text{NaCl} \)

Add coefficients to balance the equation. NEVER change subscripts.

What coefficient must be added in front of \( \text{NaCl} \) in order to balance the chlorine?

What coefficient must be added in front of \( \text{Na} \) in order to balance the sodium?
Example 2.

\[ \text{Mg} + 2\text{HCl} \rightarrow \text{MgCl}_2 + \text{H}_2 \]

What coefficient must be added in front of HCl in order to balance the chlorine and the hydrogen?
Example 3.

\[4 \text{Fe} + 3\text{O}_2 \rightarrow 2\text{Fe}_2\text{O}_3\]

Although your first instinct might be to balance the Fe first, you should actually balance the O first.

We have 2 O’s on the reactant side and 3 O’s on the product side. What is the least common multiple of 2 and 3? 6

What number can we put in front of \(\text{O}_2\) to obtain 6?
What number can we put in front of \(\text{Fe}_2\text{O}_3\) to obtain 6?

What coefficient must be added in front of Fe to balance the Fe?
Balance the following chemical equations.

i. \[ \text{Cl}_2 + 2\text{KI} \rightarrow 2\text{KCl} + \text{I}_2 \]

ii. \[ 4\text{Li} + \text{O}_2 \rightarrow 2\text{Li}_2\text{O} \]

iii. \[ 2\text{HC}_2\text{H}_3\text{O}_2 + \text{CaCO}_3 \rightarrow \text{Ca}((\text{C}_2\text{H}_3\text{O}_2)_2 + \text{CO}_2 + \text{H}_2\text{O} \]

iv. \[ \text{C}_2\text{H}_6\text{O}(l) + 3\text{O}_2(g) \rightarrow 2\text{CO}_2(g) + 3\text{H}_2\text{O}(g) \]

v. \[ 2\text{Al}(s) + 6\text{HCl}(aq) \rightarrow 2\text{AlCl}_3(aq) + 3\text{H}_2(g) \]

vi. \[ \text{Pb(NO}_3)_2 + \text{K}_2\text{CrO}_4 \rightarrow \text{PbCrO}_4 + 2\text{KNO}_3 \]

vii. \[ 2\text{C}_3\text{H}_6 + 9\text{O}_2 \rightarrow 6\text{CO}_2 + 6\text{H}_2\text{O} \]

viii. \[ 2\text{Al(OH)}_3 \rightarrow \text{Al}_2\text{O}_3 + 3\text{H}_2\text{O} \]

ix. \[ 6\text{HCl} + \text{Fe}_2\text{O}_3 \rightarrow 2\text{FeCl}_3 + 3\text{H}_2\text{O} \]
Balancing Chemical Equations

If the skeleton equation is not written for you, you must write your own.

Example 1. When an electric current is passed through water, the water molecules break down to produce hydrogen and oxygen. Bubbles of each gas are evidence of the reaction.

\[ 2\text{H}_2\text{O(aq)} \rightarrow 2\text{H}_2(g) + \text{O}_2(g) \]
Example 2. Solutions of aluminum sulfate and barium chloride react to produce solid barium sulfate and aqueous aluminum chloride.

aluminum sulfate = Al$^{3+}$, SO$_4^{2-}$ = Al$_2$(SO$_4$)$_3$
barium chloride = Ba$^{2+}$, Cl$^-$ = BaCl$_2$
barium sulfate = Ba$^{2+}$, SO$_4^{2-}$ = BaSO$_4$
aluminum chloride = Al$^{3+}$, Cl$^-$ = AlCl$_3$

$$\text{Al}_2\text{(SO}_4\text{)}_3(aq) + 3\text{BaCl}_2(aq) \rightarrow 3\text{BaSO}_4(s) + 2\text{AlCl}_3(aq)$$
You Try It

Write a balanced equation for each of the following reactions.

a. When magnesium and oxygen react, the product is solid magnesium oxide.

\[ 2\text{Mg}(s) \quad + \quad \text{O}_2(g) \quad \rightarrow \quad 2\text{MgO}(s) \]

b. When nitrogen and hydrogen react, the product is ammonia.

\[ \text{N}_2(g) \quad + \quad 3\text{H}_2(g) \quad \rightarrow \quad 2\text{NH}_3(g) \]
You Try It

Write a balanced equation for each of the following reactions.

c. Aluminum reacts with oxygen to produce aluminum oxide (rust).

\[ 4 \text{Al}(s) + 3 \text{O}_2(g) \rightarrow 2 \text{Al}_2\text{O}_3(s) \]

d. Solutions of calcium chloride and sodium sulfate react to produce aqueous sodium chloride and solid calcium sulfate.

\[ \text{CaCl}_2(aq) + \text{Na}_2\text{SO}_4(aq) \rightarrow 2 \text{NaCl}(aq) + \text{CaSO}_4(s) \]
In a synthesis reaction, two substances – either elements or compounds – combine to form a single compound.

The general equation for a synthesis reaction is

\[ A + B \rightarrow C \]
Synthesis Reactions – General Rules

a. Metal + Nonmetal → *Binary Ionic Compound

*You must remember to balance the charges when writing the formula for the binary ionic compound.

\[4 \text{Na} + \text{O}_2 \rightarrow 2\text{Na}_2\text{O}\]

What product is going to be formed?

Sodium oxide \(\text{Na}^+, \text{O}^{2-}\) \(\text{Na}_2\text{O}\)
b. Nonmetal + Nonmetal → Binary Covalent Compound

\[ S + O_2 \rightarrow SO_2 \]
c. Metal oxide + Water $\rightarrow$ *Base* (metal hydroxide)

*You must remember to balance the charges when writing the formula for the base formed.

$$\text{CaO} + \text{H}_2\text{O} \rightarrow \text{Ca(OH)}_2$$

What product is going to be formed?

Calcium hydroxide $\quad \text{Ca}^{2+}, \text{OH}^- \quad \text{Ca(OH)}_2$
d. Nonmetal oxide + Water → Acid

\[ \text{CO}_2 + \text{H}_2\text{O} \rightarrow \text{H}_2\text{CO}_3 \]
Synthesis Reactions – General Rules

e. Metallic oxide + Nonmetallic oxide $\rightarrow$ Ternary Ionic Compound (salt)

\[
\text{Na}_2\text{O} + \text{CO}_2 \rightarrow \text{Na}_2\text{CO}_3
\]
f. Here are two synthesis reactions that must be memorized.

\[ \text{N}_2(\text{g}) + 3\text{H}_2(\text{g}) \rightarrow 2\text{NH}_3(\text{g}) \]

\[ \text{NH}_3(\text{g}) + \text{H}_2\text{O}(\text{l}) \rightarrow \text{NH}_4\text{OH(aq)} \]
You Try It

Balance the following equations. You may need to complete the equation before balancing it.

a. Solid lithium metal reacts with oxygen gas.

\[ 4 \text{Li} + \text{O}_2 \rightarrow 2\text{Li}_2\text{O} \]

b. Magnesium metal burns in oxygen gas.

\[ 2\text{Mg} + \text{O}_2 \rightarrow 2\text{MgO} \]
You Try It

c. Magnesium metal reacts with chlorine gas
   \[ \text{Mg} + \text{Cl}_2 \rightarrow \text{MgCl}_2 \]

d. Solid calcium oxide is heated in the presence of sulfur trioxide.
   \[ \text{CaO} + \text{SO}_3 \rightarrow \text{CaSO}_4 \]

e. Sulfur dioxide gas is bubbled into distilled water.
   \[ \text{SO}_2 + \text{H}_2\text{O} \rightarrow \text{H}_2\text{SO}_3 \]
You Try It

f. Solid sodium oxide is added to water.

\[ \text{Na}_2\text{O} + \text{H}_2\text{O} \rightarrow 2\text{NaOH} \]

g. Calcium metal is heated strongly in nitrogen gas.

\[ 3\text{Ca} + \text{N}_2 \rightarrow \text{Ca}_3\text{N}_2 \]

h. Solid tetraphosphorus decoxide reacts with water to produce phosphoric acid.

\[ \text{P}_4\text{O}_{10} + 6\text{H}_2\text{O} \rightarrow 4\text{H}_3\text{PO}_4 \]
You Try It

i. Iodine crystals react with chlorine gas to form solid iodine trichloride.

\[ \text{I}_2 + 3 \text{Cl}_2 \rightarrow 2 \text{ICl}_3 \]

j. Solid strontium oxide reacts with water.

\[ \text{SrO} + \text{H}_2\text{O} \rightarrow \text{H}_2\text{SrO}_2 \]
Decomposition Reactions

In a decomposition reaction, a compound breaks down into two or more simpler substances.

The compound may break down into individual elements, a compound and an element, or into simpler compounds.

Many decomposition reaction take place only when energy is added in the form of heat or light.

Electrolysis is the decomposition of a substance by an electric current.

The general equation for a decomposition reaction is

\[ C \rightarrow A + B \]
a. Binary compounds break into their *elements.

*You must remember to watch out for the diatomic elements (H₂, Br₂, O₂, N₂, Cl₂, I₂, F₂).

\[ 2K_2O \rightarrow 4K + O_2 \]

What products are going to be formed?

Potassium

Oxygen
b. Metal carbonates generally break down to produce a metal oxide and carbon dioxide.

\[ \text{K}_2\text{CO}_3 \rightarrow \text{K}_2\text{O} + \text{CO}_2 \]

What products are going to be formed?

Potassium oxide \( \text{K}^+, \text{O}^{2-} \) \( \text{K}_2\text{O} \)

Carbon dioxide
c. Metal hydroxides generally break down to produce a metal oxide and water.

\[
\text{Sr(OH)}_2 \rightarrow \text{SrO} + \text{H}_2\text{O}
\]

What products are going to be formed?

Strontium oxide \(\text{Sr}^{2+}, \text{O}^{2-}\) \(\text{SrO}\)

Water
d. Metal chlorates generally break down to produce a metal chloride and oxygen.

\[
2\text{KClO}_3 \rightarrow 2\text{KCl} + 3\text{O}_2
\]

What products are going to be formed?

Potassium chloride, K\(^+\), Cl\(^-\), KCl

Oxygen
e. Many acids decompose to produce water and a nonmetal oxide.

$$\text{H}_2\text{CO}_3 \rightarrow \text{H}_2\text{O} \ + \ \text{CO}_2$$

What products are going to be formed?

Water

Nonmetal oxide
Decomposition Reactions – General Rules

f. Ternary Ionic compounds (salts) decompose to produce a **metal oxide** and a **nonmetal oxide**.

\[ \text{CaSO}_3 \rightarrow \text{CaO} + \text{SO}_2 \]

**What products are going to be formed?**

**Metal oxide**

**Nonmetal oxide**
Here are four decomposition reactions that must be memorized.

\[ 2\text{H}_2\text{O}_2 \rightarrow 2\text{H}_2\text{O} + \text{O}_2 \]

\[ \text{NH}_4\text{OH} \rightarrow \text{NH}_3 + \text{H}_2\text{O} \]

\[ (\text{NH}_4)_2\text{CO}_3 \rightarrow 2\text{NH}_3 + \text{H}_2\text{O} + \text{CO}_2 \]

\[ \text{NH}_4\text{NO}_3 \rightarrow \text{N}_2\text{O} + 2\text{H}_2\text{O} \]
Balance the following decomposition reactions. Use the examples above to help you predict the products when necessary. You do not have to indicate the states of the reactants and products.

a. Solid barium carbonate is heated.

\[ \text{BaCO}_3 \rightarrow \text{BaO} + \text{CO}_2 \]

b. Solid silver oxide is heated.

\[ 2\text{Ag}_2\text{O} \rightarrow 4 \text{Ag} + \text{O}_2 \]
c. Solid sodium chlorate is heated.

\[ 2\text{NaClO}_3 \rightarrow 2 \text{NaCl} + 3 \text{O}_2 \]

d. An aqueous solution of sulfuric acid is heated.

\[ \text{H}_2\text{SO}_4 \rightarrow \text{H}_2\text{O}(l) + \text{SO}_3(g) \]

e. Aluminum can be obtained from aluminum oxide with the addition of a large amount of electrical energy.

\[ 2\text{Al}_2\text{O}_3 \rightarrow 4 \text{Al} + 3\text{O}_2 \]
You Try It!

f. Heating tin(IV) hydroxide gives tin(IV) oxide and steam.

\[
\text{Sn(OH)}_4 \rightarrow \text{SnO}_2 + 2\text{H}_2\text{O}
\]

g. Solid sodium carbonate is heated.

\[
\text{Na}_2\text{CO}_3 \rightarrow \text{Na}_2\text{O} + \text{CO}_2
\]

h. Solid magnesium sulfite is heated.

\[
\text{MgSO}_3 \rightarrow \text{MgO} + \text{SO}_2
\]
You Try It!

**i.** Powdered magnesium carbonate is heated strongly.

\[ \text{MgCO}_3 \rightarrow \text{MgO}(s) + \text{CO}_2(g) \]

**j.** Solid calcium hydroxide decomposes.

\[ \text{Ca(OH)}_2(s) \rightarrow \text{CaO}(s) + \text{H}_2\text{O}(g) \]
In a single-replacement reaction, an uncombined element replaces an element in a compound.

There are three general equations for a single replacement reaction.

\[
A + BC \rightarrow AC + B \quad \text{(If } A \text{ is a metal)}
\]

\[
A + BC \rightarrow BA + C \quad \text{(If } A \text{ is a halogen)}
\]

\[
A + BC \rightarrow \text{Base} + H_2 \quad \text{(If } A \text{ is an active metal and } BC \text{ is } H_2O)
\]
Single-Replacement Reactions

Examples

\[ \text{Zn} + 2\text{HCl} \rightarrow \text{ZnCl}_2 + \text{H}_2 \]

\[ \text{Cl}_2 + 2\text{LiBr} \rightarrow 2\text{LiCl} + \text{Br}_2 \]

\[ 2\text{K} + 2\text{H}_2\text{O} \rightarrow 2\text{KOH} + \text{H}_2 \]

Back to main menu
Whether one metal will replace another from a compound can be determined by using the activity series of metals. The activity series of metals lists metals in order of chemical reactivity. A reactive metal will replace any metal found below it in the activity series.
The halogens can also take place in single-replacement reactions.

The order of reactivity for the halogens from increasing to decreasing reactivity is fluorine, chlorine, bromine, iodine.
Single-Replacement Reactions

Examples

a. \( \text{Li} + \text{KCl} \rightarrow \text{LiCl} + \text{K} \)

b. \( \text{Mg} + \text{K}_2\text{SO}_4 \rightarrow \text{no reaction} \)

c. \( \text{Cl}_2 + 2\text{KBr} \rightarrow 2\text{KCl} + \text{Br}_2 \)

d. \( 2\text{Li} + 2\text{H}_2\text{O} \rightarrow 2\text{LiOH} + \text{H}_2 \)
You Try It

Complete and balance the following equations.

a. \(\text{Zn} + \text{Pb(NO}_3\text{)}_2 \rightarrow \text{Zn(NO}_3\text{)}_2 + \text{Pb}\)

b. \(2\text{K} + \text{Ba(C}_2\text{H}_3\text{O}_2\text{)}_2 \rightarrow 2\text{KC}_2\text{H}_3\text{O}_2 + \text{Ba}\)

c. \(2\text{Al} + 3\text{NiSO}_4 \rightarrow \text{Al}_2(\text{SO}_4)_3 + 3\text{Ni}\)

d. \(2\text{Na} + 2\text{H}_2\text{O} \rightarrow 2\text{NaOH} + \text{H}_2\)

e. \(\text{Zn} + \text{BaCl}_2 \rightarrow \text{no reaction}\)

f. \(3\text{F}_2 + 2\text{AlCl}_3 \rightarrow 2\text{AlF}_3 + 3\text{Cl}_2\)
In a double-replacement reaction, the negative ions of two compounds exchange places in an aqueous solution to form two new compounds.

The general equation for a double replacement reaction is

\[ AB + CD \rightarrow AD + CB \]
Double-Replacement Reactions

For a double-replacement reaction to occur, one of the following statements is usually true concerning at least one of the products of the reaction.

a. It is a gas that bubbles out of the mixture.

\[ \text{FeS(s)} + 2\text{HCl(aq)} \rightarrow \text{FeCl}_2(\text{aq}) + \text{H}_2\text{S(g)} \]

Other gases commonly formed include:
\[ \text{H}_2\text{S} \]
\[ \text{CO}_2 \text{ (formed from the decomposition of H}_2\text{CO}_3 \text{)} \]
\[ \text{SO}_2 \text{ (formed from the decomposition of H}_2\text{SO}_3 \text{)} \]
\[ \text{NH}_3 \text{ (formed from the decomposition of NH}_4\text{OH)} \]
Double-Replacement Reactions

b. It is a molecular compound such as water. This is common in acid-base neutralization reactions.

\[ \text{HCl}(aq) + \text{NaOH}(aq) \rightarrow \text{H}_2\text{O}(l) + \text{NaCl}(aq) \]
c. It is only slightly soluble and precipitates from solution.

\[ 2 \text{KI}(aq) + \text{Pb(NO}_3\text{)}_2(aq) \rightarrow 2\text{KNO}_3(aq) + \text{PbI}_2(s) \]
A solubility chart can be used to identify insoluble solids (precipitates).

<table>
<thead>
<tr>
<th>Solubility of Common Ionic Compounds in Water</th>
</tr>
</thead>
<tbody>
<tr>
<td><strong>Soluble</strong></td>
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<tr>
<td>Compounds containing</td>
</tr>
<tr>
<td>( \text{NH}_4^+ )</td>
</tr>
<tr>
<td>( \text{C}_2\text{H}_3\text{O}_2^- )</td>
</tr>
<tr>
<td>( \text{Br}^- )</td>
</tr>
<tr>
<td>( \text{Cl}^- )</td>
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<tr>
<td>( \text{NO}_3^- )</td>
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<tr>
<td>( \text{SO}_4^{2-} )</td>
</tr>
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<thead>
<tr>
<th><strong>Common Exceptions</strong></th>
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<tr>
<td>Compounds of ( \text{Ag}^+ ), ( \text{Hg}_2^{2+} ), and ( \text{Pb}^{2+} )</td>
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<td>Compounds of ( \text{Ag}^+ ), ( \text{Hg}_2^{2+} ), and ( \text{Pb}^{2+} )</td>
</tr>
<tr>
<td>None</td>
</tr>
<tr>
<td>Compounds of ( \text{Sr}^{2+} ), ( \text{Ba}^{2+} ), ( \text{Hg}^{2+} ), and ( \text{Pb}^{2+} )</td>
</tr>
</tbody>
</table>

| **Insoluble**                                   |
| Compounds containing                           |
| \( \text{CO}_3^{2-} \)                         |
| \( \text{OH}^- \)                              |
| \( \text{PO}_4^{3-} \)                         |
| \( \text{S}^{2-} \)                            |
| \( \text{CrO}_4^{2-} \)                        |

<table>
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main menu
Example. Label the precipitate formed in the following reaction.

$$\text{Na}_2\text{CO}_3(aq) + \text{Ca(NO}_3)_2(aq) \rightarrow 2\text{NaNO}_3(aq) + \text{CaCO}_3(s)$$
You Try It

Label the precipitate formed in each of the following reactions.

i. \( \text{NH}_4\text{Cl}(aq) + \text{AgNO}_3(aq) \rightarrow \text{NH}_4\text{NO}_3(aq) + \text{AgCl}(s) \)

ii. \( \text{BaCl}_2(aq) + \text{Na}_2\text{SO}_4(aq) \rightarrow \text{BaSO}_4(s) + 2\text{NaCl}(aq) \)
You Try It

Complete and balance the following equations.

a. \( \text{AgNO}_3 + \text{NaCl} \rightarrow \text{AgCl} + \text{NaNO}_3 \)
b. \( \text{Mg(NO}_3)_2 + 2\text{KOH} \rightarrow \text{Mg(OH)}_2 + 2\text{KNO}_3 \)
c. \( 3\text{LiOH} + \text{Fe(NO}_3)_3 \rightarrow 3\text{LiNO}_3 + \text{Fe(OH)}_3 \)
d. \( \text{Pb(NO}_3)_2 + 2\text{NaOH} \rightarrow \text{Pb(OH)}_2 + 2\text{NaNO}_3 \)
e. \( \text{NH}_4\text{Cl(aq)} + \text{NaOH(aq)} \rightarrow \text{NH}_4\text{OH} + \text{H}_2\text{O} + \text{NaCl} \)
f. \( \text{K}_2\text{SO}_3(aq) + 2\text{HBr(aq)} \rightarrow 2\text{KBr} + \text{SO}_2 + \text{H}_2\text{O} \)
In a combustion reaction, a substance combines with oxygen, releasing a large amount of energy in the form of heat and light.

For our purpose, we will talk about hydrocarbons burning in the presence of oxygen.

The general equation for a complete combustion reaction is

$$C_xH_y + O_2 \rightarrow CO_2 + H_2O$$
Types of Combustion Reactions

Complete Combustion – the products are $\text{CO}_2$ and $\text{H}_2\text{O}$.

$$\text{CH}_4 + 2\text{O}_2 \rightarrow \text{CO}_2 + 2\text{H}_2\text{O}$$

Incomplete Combustion – the products are $\text{C}$ and/or $\text{CO}$ and $\text{H}_2\text{O}$.

$$2\text{C}_3\text{H}_8 + 7\text{O}_2 \rightarrow 6\text{CO} + 8\text{H}_2\text{O}$$
Example Problem. Write a balanced equation for the complete combustion of ethane, \( \text{C}_2\text{H}_6 \).

\[
2\text{C}_2\text{H}_6 + 7\text{O}_2 \rightarrow 4\text{CO}_2 + 6\text{H}_2\text{O}
\]
Write balanced equations for the complete combustion of the following substances.

a. \( \text{C}_3\text{H}_8 + 5\text{O}_2 \rightarrow 3\text{CO}_2 + 4\text{H}_2\text{O} \)

b. \( 2\text{C}_5\text{H}_{10} + 15\text{O}_2 \rightarrow 10\text{CO}_2 + 10\text{H}_2\text{O} \)

c. \( 2\text{C}_6\text{H}_{14} + 19\text{O}_2 \rightarrow 12\text{CO}_2 + 14\text{H}_2\text{O} \)
Oxidation-Reduction reactions are the chemical changes that occur when electrons are transferred between reactions. Examples include the burning of gasoline and the rusting of a nail.
Oxidation

Oxidation originally meant the combination of an element with oxygen to give oxides. However, today it is also defined as the loss of electrons. (Oxygen does not have to be present for oxidation to occur.)

Example: \(4\text{Fe} + 3\text{O}_2 \rightarrow 2\text{Fe}_2\text{O}_3\)
Reduction

Reduction originally meant the loss of oxygen from a compound. Today it is also defined as the gain of electrons.

Example: $2\text{Fe}_2\text{O}_3 + 3\text{O}_2 \rightarrow 4\text{Fe} + 3\text{CO}_2$
Oxidizing Agent

The oxidizing agent in a redox reaction gains electrons.

It is the substance being reduced.
Reducing Agent

The oxidizing agent loses electrons.

It is the substance being oxidized.
Helpful Mnemonics

Leo the Lion says Ger

L – Loss
e – of electrons
o – is oxidation

G – Gain
e – of electrons
r – is reduction

Oil Rig

O – Oxidation
i – is

R – Reduction
i – is

l – loss of electrons

g – gain of electrons
Rules for Assigning Oxidation Numbers

1. The oxidation number of a monatomic ion is equal to its charge.

Examples:

Br^- equals -1
Fe^{3+} equals +3
2. For a polyatomic ion, the sum of the oxidation numbers must equal the ionic charge of the ion.

Examples:

SO$_4^{2-}$ equals -2

NO$_3^-$ equals -1
3. The oxidation number of a metal cation is the same as its ionic charge.

Examples:
- Sodium is +1
- Calcium is +2
4. The oxidation number of hydrogen in a compound is $+1$ except in metal hydrides, for example, NaH, where it is $-1$. 
Rules for Assigning Oxidation Numbers

5. The oxidation number of oxygen in a compound is -2 except in peroxides, for example, $\text{H}_2\text{O}_2$, where it is -1.
6. The oxidation number of an uncombined element is 0.
For example, the oxidation number of the potassium atoms in potassium metal, K, and of the nitrogen atoms in nitrogen gas, N\textsubscript{2}, is zero.
7. For any neutral compound, the sum of the oxidation numbers of the atoms in the compound must equal 0.
Rules for Assigning Oxidation Numbers

8. Assign the oxidation numbers in the following order.

a. Metals (determined by looking at their group number)

b. Hydrogen (usually +1)

c. Oxygen (usually -2)

d. Transition Metals and Everything Else
What is the oxidation number of each element in the following?

\[ \text{SO}_2 \]

Start with oxygen. According to rule 5, the oxidation number of oxygen in a compound is usually -2. What number does sulfur need to be in order for the overall compound to have a sum of zero?

\[ S = +4 \]

\[ O = -2 \]
What is the oxidation number of each element in the following?

\[ \text{KClO}_3 \]

\[ +1 \ +5 \ -2 \]

Start with potassium. According to rule 3, the oxidation number of potassium is the same as its ionic charge, which is +1.

Next do oxygen. According to rule 5, the oxidation number of oxygen in a compound is usually -2.

What number does chlorine need to be in order for the overall compound to have a sum of zero?

K = +1

O = -2

Cl = +5
What is the oxidation number of each element in the following?

KClO\(_2\)

\[\begin{align*}
+1 & \quad +3 & \quad -2 \\
K & \quad O & \quad Cl \quad O
\end{align*}\]

Start with potassium. According to rule 3, the oxidation number of potassium is the same as its ionic charge, which is +1.

Next do oxygen. According to rule 5, the oxidation number of oxygen in a compound is usually -2.

What number does chlorine need to be in order for the overall compound to have a sum of zero?
What is the oxidation number of each element in the following?

$$\text{CO}_3^{2-}$$

Start with oxygen. According to rule 5, the oxidation number of oxygen in a compound is usually -2.

What number does carbon need to be in order for the polyatomic ion to have an overall charge of -2?

$$\text{C} = +4$$

$$\text{O} = -2$$
You Try It.  Determine the oxidation number of each element in each of the following.

1. \( \text{Na}_2\text{Cr}_2\text{O}_7 \)  \( \text{Na} = +1 \)  \( \text{O} = -2 \)  \( \text{Cr} = +6 \)

2. \( \text{BaH}_2 \)  \( \text{Ba} = +2 \)  \( \text{H} = -1 \)
   (Hint:  \( \text{BaH}_2 \) is a metal hydride.)

3. \( \text{Li}_2\text{O}_2 \)  \( \text{Li} = +1 \)  \( \text{O} = -1 \)
   (Hint:  \( \text{Li}_2\text{O}_2 \) is a peroxide.)

4. \( \text{ClO}_3^- \)  \( \text{Cl} = +5 \)  \( \text{O} = -2 \)
Oxidation Number Changes

The changes in oxidation number can be used to determine which elements are oxidized and which elements are reduced.

Remember – an increase in the oxidation number of an atom signifies oxidation and a decrease in the oxidation number signifies reduction.
Oxidation Number Changes

Use the change in oxidation number to identify which elements are oxidized and reduced in each of these reactions. Also identify the oxidizing (OA) and reducing (RA) agents.

a. \( \text{F}_2(g) + 2\text{HBr}(aq) \rightarrow 2\text{HF}(aq) + \text{Br}_2(l) \)

Fluorine: 0 to -1; reduced, OA

Bromine: -1 to 0; oxidized, RA
Oxidation Number Changes

Use the change in oxidation number to identify which elements are oxidized and reduced in each of these reactions. Also identify the oxidizing (OA) and reducing (RA) agents.

b. \(2 \text{KClO}_3(s) \rightarrow 2 \text{KCl}(s) + 3 \text{O}_2(g)\)

\[
\begin{array}{c}
+1 
\end{array} + \begin{array}{c}
5 
\end{array} - \begin{array}{c}
2 
\end{array} = 0
\]

\[
\begin{array}{c}
+1 
\end{array} + \begin{array}{c}
5 
\end{array} - \begin{array}{c}
6 
\end{array} = 0
\]

Chlorine: +5 to -1; reduced, OA

Oxygen: -2 to 0; oxidized, RA
Oxidation Number Changes

Use the change in oxidation number to identify which elements are oxidized and reduced in each of these reactions. Also identify the oxidizing (OA) and reducing (RA) agents.

c. \[ \text{SO}_3(g) + \text{H}_2\text{O}(l) \rightarrow \text{H}_2\text{SO}_4(aq) \]

\[
\begin{align*}
+6 &-2 =0 \\
+6 &-6 =0 \\
+2 &-2 =0 \\
+2 &+6 -8 =0
\end{align*}
\]

Since there is no change in oxidation number, this is not a redox reaction.
You Try It

Use the change in oxidation number to identify which elements are oxidized and reduced in each of these reactions. Also identify the oxidizing (OA) and reducing (RA) agents.

a. \[2\text{H}_2(g) + \text{O}_2(g) \rightarrow 2\text{H}_2\text{O}(g)\]

- hydrogen: 0 to +1; oxidized, RA
- oxygen: 0 to -2; reduced, OA
You Try It

Use the change in oxidation number to identify which elements are oxidized and reduced in each of these reactions. Also identify the oxidizing (OA) and reducing (RA) agents.

b. \[2\text{Al}(s) + 6\text{HCl}(aq) \rightarrow 2\text{AlCl}_3(s) + 3\text{H}_2(g)\]

\[
\begin{align*}
\text{aluminum: } &0 \text{ to } +3; \text{ oxidized, RA} \\
\text{hydrogen: } &+1 \text{ to } 0; \text{ reduced, OA}
\end{align*}
\]
You Try It

Use the change in oxidation number to identify which elements are oxidized and reduced in each of these reactions. Also identify the oxidizing (OA) and reducing (RA) agents.

c. \(2\text{NaCl}(aq) + \text{Li}_2\text{S}(aq) \rightarrow \text{Na}_2\text{S}(s) + \text{LiCl}(aq)\)

\[
\begin{align*}
+1 & \quad -1 & =0 \\
+1 & \quad -1 & =0 \\
+2 & \quad -2 & =0 \\
+1 & \quad -1 & =0
\end{align*}
\]

Since there is no change in oxidation number, this is not a redox reaction.