Chapter 7

Chemical Reactions

Based on slides provided with Introductory Chemistry, Fifth Edition
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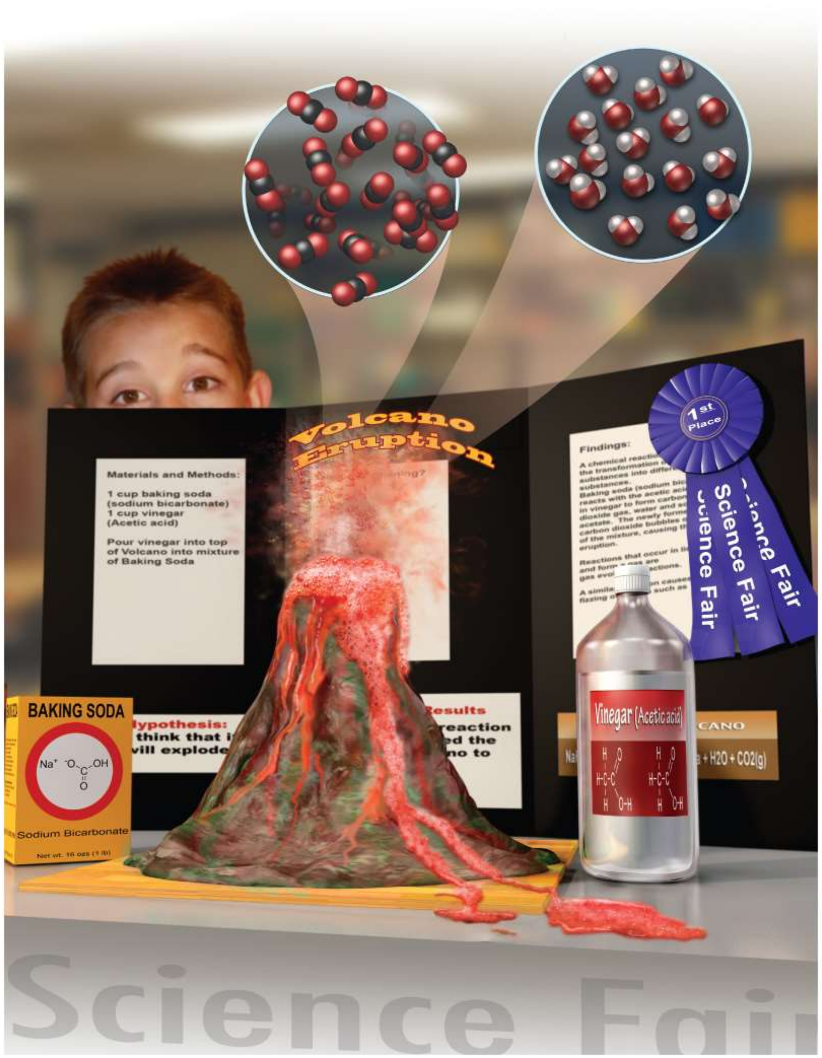
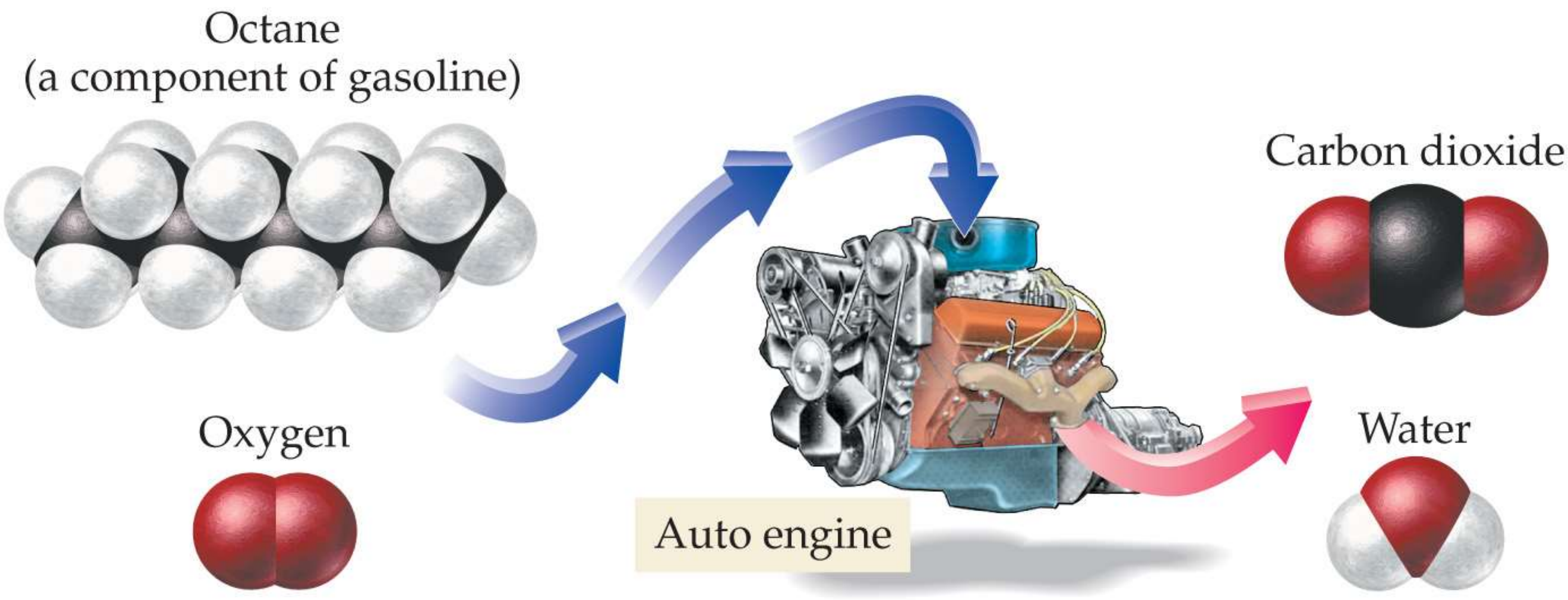
Evidence of a Chemical Reaction

- In the classic grade school volcano, the baking soda (which is sodium bicarbonate) reacts with acetic acid in the vinegar to form carbon dioxide gas, water, and sodium acetate.
- The newly formed carbon dioxide bubbles out of the mixture, causing the eruption.
- Reactions that occur in liquids and form a gas are gas evolution reactions.
- Red food coloring helps us see the bubbles.



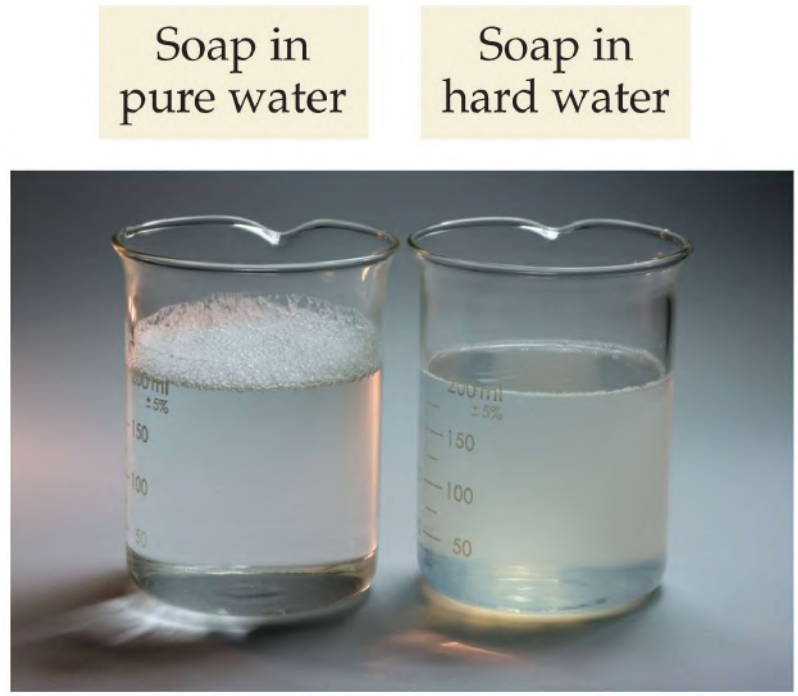
Chemical Reactions in Automobiles: Combustion



A combustion reaction In an automobile engine, hydrocarbons such as octane (C_8H_{18}) from gasoline combine with oxygen from the air and react to form carbon dioxide and water.
Combustion reactions are a subcategory of *oxidation–reduction reactions,* in which electrons are transferred from one substance to another.

Chemical Reactions in Laundry Detergents

- Laundry detergent works better than soap because it contains substances that soften hard water.
- Hard water contains dissolved calcium and magnesium ions.
- These ions interfere with the action of soap by reacting with it to form a gray, slimy substance called *curd* or *soap scum*.
- If you have ever washed your clothes in ordinary soap, you may have noticed gray soap scum residue on your clothes.



Soap forms suds with pure water (left), but reacts with the ions in hard water (right) to form a gray residue that adheres to clothes.

Chemical Reactions in Laundry Detergents

- Laundry detergents contain substances such as sodium carbonate (Na_2CO_3) that remove calcium and magnesium ions from the water.
- The dissolved carbonate ions react with calcium and magnesium ions in the hard water to form solid calcium carbonate ($CaCO_3$) and solid magnesium carbonate ($MgCO_3$).
- These solids settle to the bottom of the laundry mixture, resulting in the removal of the ions from the water.
- Laundry detergents contain substances that react with the ions in hard water to immobilize them.
- Reactions such as these—that form solid substances in water—are *precipitation reactions.*

Evidence of a Chemical Reaction

If we could see the atoms and molecules that compose matter, we could easily identify a chemical reaction:

- Atoms combine with other atoms to form compounds.
- New molecules form.
- The original molecules decompose.
- Atoms in one molecule change places with atoms in another.

Often, these molecular changes cause macroscopic changes that we can directly experience.

Evidence of a Chemical Reaction

color change

emission of light

formation of a solid

emission or absorption of heat

formation of a gas

Evidence of a Chemical Reaction

Gas formation

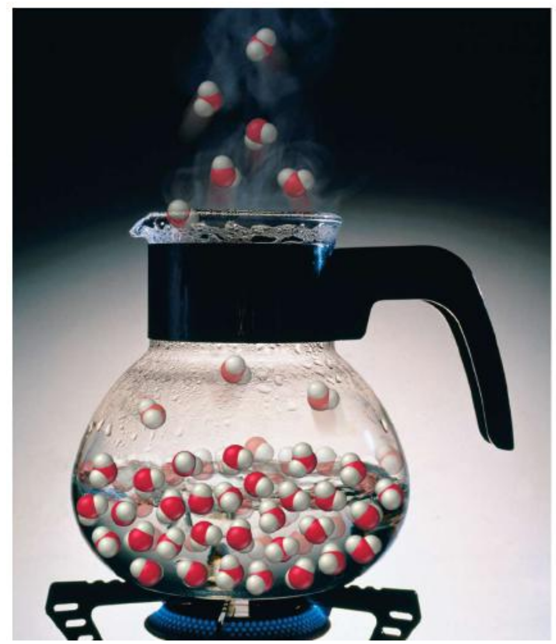
The formation of a gas is evidence of
a chemical reaction.

Solid formation

A precipitation reaction:
The formation of a solid in a previously clear solution is evidence of a chemical reaction.

NOT Evidence of a Chemical Reaction

- We can be fooled.
- When water boils, bubbles form and a gas is evolved, but no chemical reaction has occurred.
- Boiling water forms gaseous steam, but both water and steam are composed of water molecules—no chemical change has occurred.



Evidence of a Chemical Reaction: Changes Occurring at the Atomic and Molecular level Determine Whether a Chemical Reaction has Occurred

- ***Only chemical analysis*** that shows that the initial substances have changed into other substances conclusively proves that a chemical reaction has occurred.
- Chemical reactions may occur without any obvious signs, yet chemical analysis may show that a reaction has indeed occurred.
- The changes occurring at the atomic and molecular level determine whether a chemical reaction has taken place.

The Chemical Equation

- We represent chemical reactions with *chemical equations.*
- The substances on the left side of the equation are the *reactants,* and the substances on the right side are the *products.*
- We often specify the state of each reactant or product in parentheses next to the formula.

Use These Abbreviations to Add States to the Equation

TABLE 7.1 Abbreviations Indicating the States of Reactants and Products in Chemical Equations

Abbreviation	State
(*g*)	gas
(*l*)	liquid
(*s*)	solid
(*aq*)	aqueous (water solution)*

*The (*aq*) designation stands for *aqueous*, which indicates that a substance is dissolved in water. When a substance dissolves in water, the mixture is called a *solution* (see Section 7.5).

The Chemical Equation: Combustion of Methane

- The reaction occurring in a natural gas flame is methane (CH_4) reacting with oxygen (O_2) to form carbon dioxide (CO_2) and water (H_2O).
- We represent this reaction with the following equation:

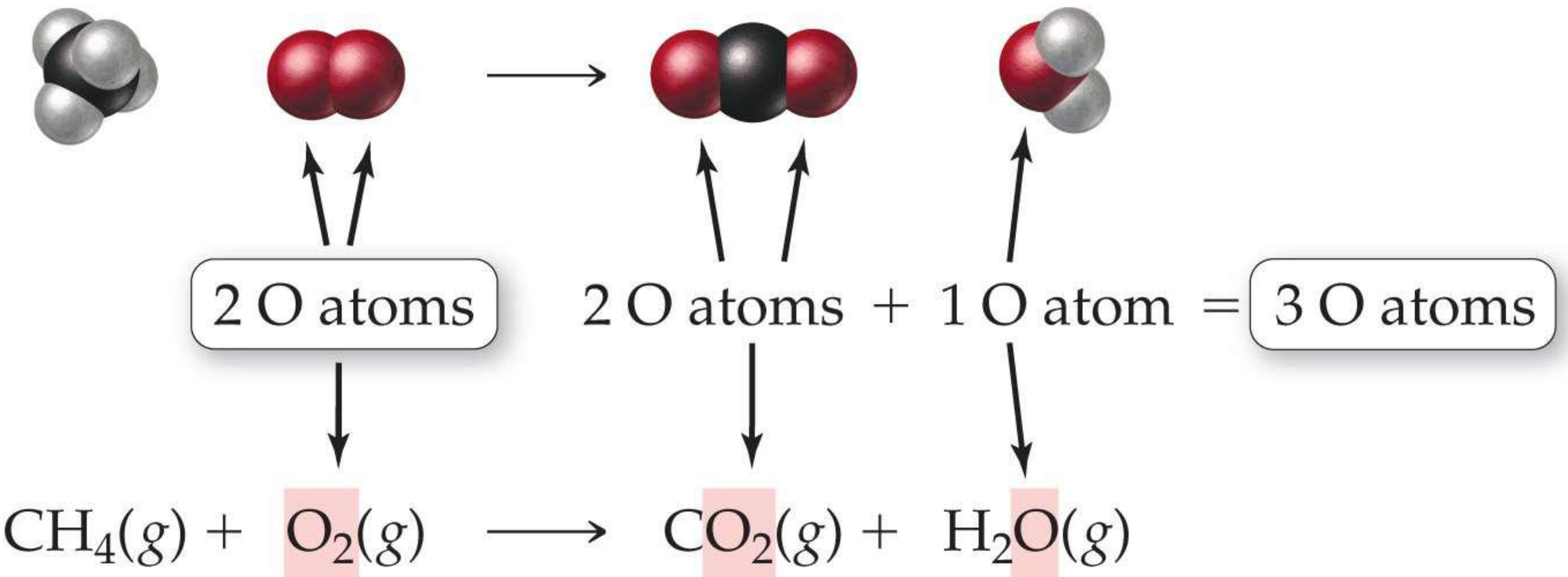
$$\underbrace{CH_4 + O_2}_{\text{reactants}} \longrightarrow \underbrace{CO_2 + H_2O}_{\text{products}}$$

- With states included, the equation becomes

$$CH_4(g) + O_2(g) \longrightarrow CO_2(g) + H_2O(g)$$

The Chemical Equation: Combustion of Methane

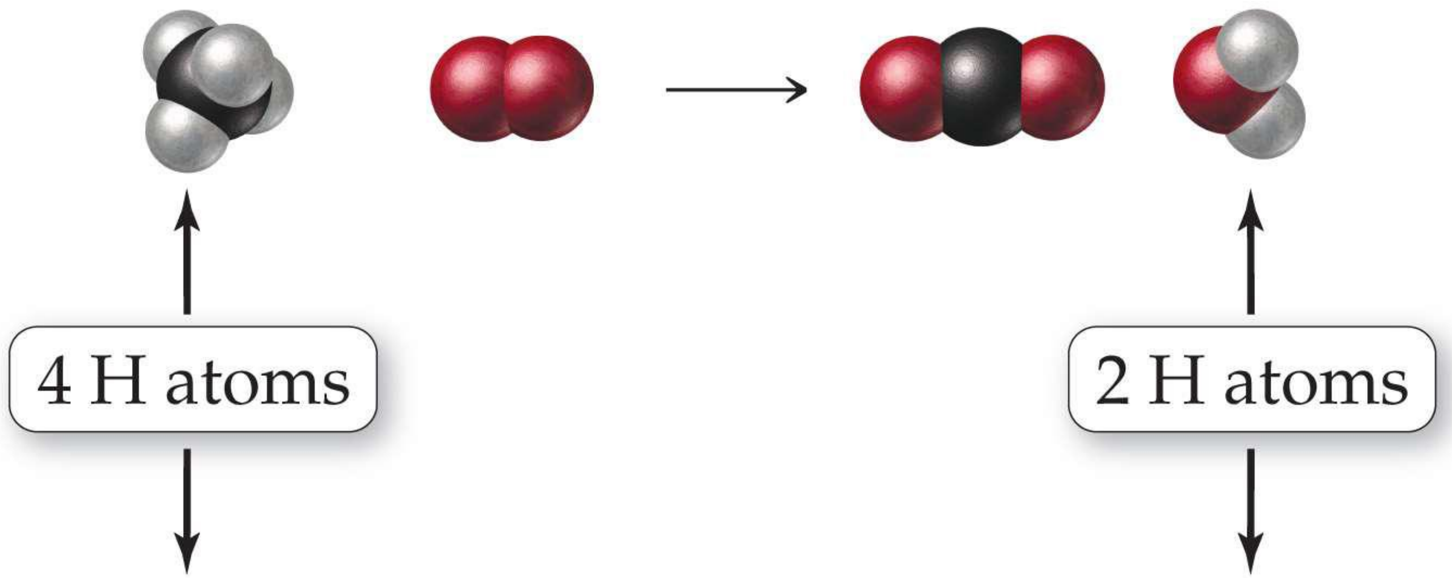
- How many oxygen atoms are on each side of the equation?



The left side of the equation has two oxygen atoms and the right side has three.
Atoms cannot simply appear or disappear in chemical equations. We must account for the atoms on both sides of the equation.

The Chemical Equation: Combustion of Methane

Notice also that the left side of the equation has four hydrogen atoms and the right side has only two.

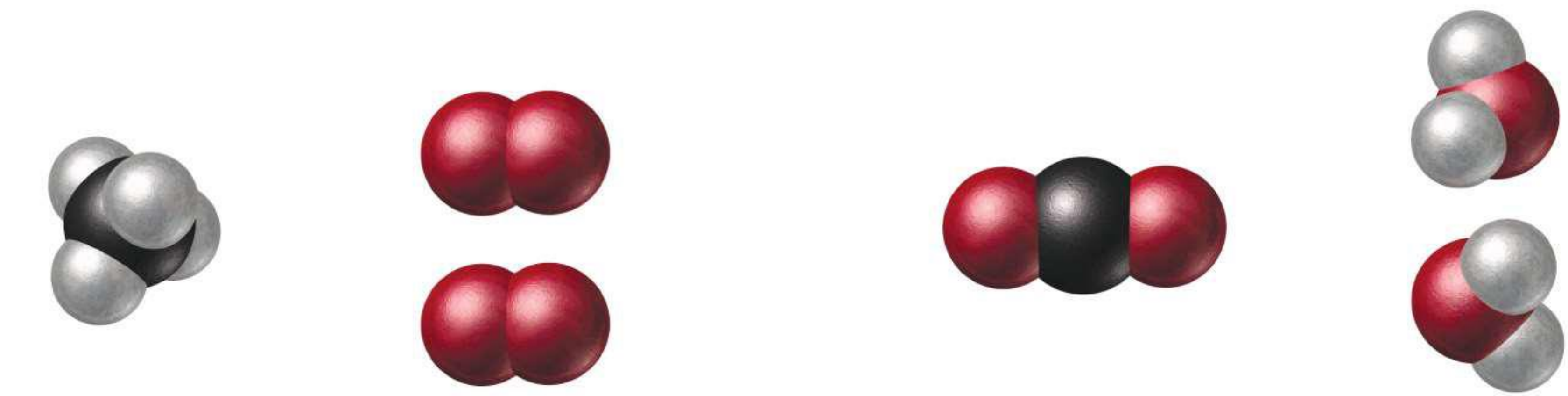


$$CH_4(g) + O_2(g) \longrightarrow CO_2(g) + H_2O(g)$$

To correct these problems, we must create a **balanced equation.**

The Chemical Equation: Combustion of Methane

- To balance an equation, we insert coefficients—not subscripts—in front of the chemical formulas as needed to make the number of each type of atom in the reactants equal to the number of each type of atom in the products.



$$CH_4(g) + \mathbf{2}\,O_2(g) \longrightarrow CO_2(g) + \mathbf{2}\,H_2O(g)$$

The equation is now balanced because the numbers of each type of atom on both sides of the equation are equal.

Checking the Balanced Equation

- The number of a particular type of atom within a chemical formula embedded in an equation is obtained by multiplying the subscript for the atom by the coefficient for the chemical formula.
- If there is no coefficient or subscript, a 1 is implied.
- The balanced equation for the combustion of natural gas is as follows:

$$CH_4(g) + 2\,O_2(g) \longrightarrow CO_2(g) + 2\,H_2O(g)$$

Reactants	Products
1 C atom (1 × $\underline{C}H_4$)	1 C atom (1 × $\underline{C}O_2$)
4 H atoms (1 × $C\underline{H}_4$)	4 H atoms (2 × $\underline{H}_2O$)
4 O atoms (2 × $\underline{O}_2$)	4 O atoms (1 × $C\underline{O}_2$ + 2 × $H_2\underline{O}$)

The Numbers of Each Type of Atom on Both Sides of the Equation are Equal—the Equation is Balanced

$$CH_4(g) + 2\,O_2(g) \longrightarrow CO_2(g) + 2\,H_2O(g)$$

A balanced chemical equation represents a chemical reaction. In this image, methane molecules combine with oxygen to form carbon dioxide and water.

- Balancing chemical equations is ultimately a mathematical process that can be completely automated.
- We can do it manually for relatively simple equations.
- Balancing manually is best done by following a set of rules and guidelines that work for many reactions.
- Avoid “trial and error” as much as possible.
- If you recognize the class of reaction, you can greatly simplify the balancing because you know where different pieces are moving.
- We will see a “general” set of guidelines for balancing, but those would be overkill for reactions we will become familiar with, such as combustion reactions, where there is a specific and simple theme.

Balancing combustion reactions

$$C_cH_hO_o + \; x\, O_2 \rightarrow \; c\, CO_2 + \; h/2\, H_2O$$

Combustion:

A "fuel" containing C, H, O reacting with O_2 to produce CO_2 and H_2O

- We don't need to use the general rules
- We "follow" and balance the carbon and hydrogen:
  - All the C in CO_2 comes from the C in the fuel
  - All the H in H_2O comes from the H in the fuel
- Oxygen is not so simple, but becomes simple to balance after we balance C and H

$x = [c + (h/2 - o)/2]$ **but there is no need memorize that**

$$1\,C_{2}H_{6}O + \quad O_2 \rightarrow 2\,CO_2 + 3\,H_2O$$

We start with **one** molecule of "fuel" (C_2H_6O here)

Balance carbon and hydrogen first:

- **Number of CO_2 molecules** equals the number of C atoms in the fuel molecule (each CO_2 molecule has one C atom)

$$(1)(2) = x\,(1) \Longrightarrow x = 2$$

- **Number of H_2O molecules** is half the number of H atoms in the fuel molecule (each H_2O molecule has 2 H atoms)

$$(1)(6) = x\,(2) \Longrightarrow x = 3$$

- Then balance oxygen

We now balance O:

$$C_2H_6O + 3\,O_2 \rightarrow 2\,CO_2 + 3\,H_2O$$

$$(1)(1) + \mathbf{x}\,(2) = (2)(2) + (3)(1) \Rightarrow \mathbf{x} = 3$$

Practice:
Check your solution on next page

Balance

$C_6H_{12}O_6 + O_2 \rightarrow CO_2 + H_2O$

Practice

Balance

$C_6H_{12}O_6 + 6O_2 \rightarrow 6CO_2 + 6H_2O$

Practice:
Check your solution on next page

Balance

$$C_8H_{18}(l) + O_2(g) \rightarrow CO_2(g) + H_2O(g)$$

Practice

$$2\ C_8H_{18}(l) + 25\ O_2(g) \rightarrow 16\ CO_2(g) + 18\ H_2O(g)$$

How to Write Balanced Chemical Equations

1. Write a skeletal equation by writing correct chemical formulas for each of the reactants and products.

   If a skeletal equation is provided, proceed to balance.

   – by putting appropriate coefficients in front of the species in which the elements occur

How to Write Balanced Chemical Equations

2. If an element occurs in only one compound on each side of the equation, balance it first.

   If there is more than one such element, balance metals before nonmetals

How to Write Balanced Chemical Equations

3. Look for an element that that occurs in only one coeficientless compound (one that hasn't received a coefficient yet) anywhere in the equation, and balance it.

How to Write Balanced Chemical Equations

4. If an element occurs as free (Cu(s), O_2(g), etc.) anywhere, balance it last.

Always balance free elements by adjusting the coefficient *on the free element.*

How to Write Balanced Chemical Equations

- At any point, if you obtain a fractional coefficient, change into a whole number by multiplying all the known coefficients by the denominator of that fraction.
  - Don’t treat “no coefficient” as an implicit 1; leave coefficientless compouds coeficientless

How to Write Balanced Chemical Equations

- During the balancing process, if a coefficient turns out to be 1, write it explicitly to mark that substance as “balanced” (since no coefficient in front of a species means we have not yet produced a coefficient for it).
  If the coefficient is still 1 at the end, delete it to make it an implicit 1.

How to Write Balanced Chemical Equations

- Check to make certain the equation is balanced by summing the total number of each type of atom on both sides of the equation.
- Remember, change only the <u>*coefficients*</u> to balance a chemical equation; <u>*never change the subscripts*</u>.

Example: Write a Balanced Equation

- Use your knowledge of chemical nomenclature from Chapter 5 to write a skeletal equation containing formulas for each of the reactants and products.
- The formulas for each compound MUST BE CORRECT before you begin to balance the equation.

For the reaction of solid aluminum with aqueous sulfuric acid to form aqueous aluminum sulfate and hydrogen gas:

$$Al(s) + H_2SO_4(aq) \longrightarrow Al_2(SO_4)_3(aq) + H_2(g)$$

Example: The Reaction of Solid Aluminum with Aqueous Sulfuric Acid to Form Aqueous Aluminum Sulfate and Hydrogen Gas

- Since both aluminum and hydrogen occur as pure elements, balance those last.
- Sulfur and oxygen occur in only one compound on each side of the equation, so balance these first.
- Sulfur and oxygen are part of a polyatomic ion that stays intact on both sides of the equation.
- *Balance polyatomic ions such as these as a unit*.
- There are 3 SO_4^{2-} ions on the right side of the equation, so put a 3 in front of H_2SO_4.

$$\mathrm{Al}(s) + 3\,\mathrm{H_2SO_4}(aq) \longrightarrow 1\,\mathrm{Al_2(SO_4)_3}(aq) + \mathrm{H_2}(g)$$

Example: The Reaction of Solid Aluminum with Aqueous Sulfuric Acid to Form Aqueous Aluminum Sulfate and Hydrogen Gas

Al and H both occur in elemental form (Al(s) and H_2(g)) so they are balanced last. We can proceed with either one.

- Balance Al next. Since there are 2 Al atoms on the right side of the equation, place a 2 in front of Al on the left side of the equation.

$$2\,\mathrm{Al}(s) + 3\,\mathrm{H_2SO_4}(aq) \longrightarrow 1\,\mathrm{Al_2(SO_4)_3}(aq) + \mathrm{H_2}(g)$$

- Balance H next. Since there are 6 H atoms on the left side, place a 3 in front of H_2 on the right side.

$$2\,\mathrm{Al}(s) + 3\,\mathrm{H_2SO_4}(aq) \longrightarrow 1\,\mathrm{Al_2(SO_4)_3}(aq) + 3\,\mathrm{H_2}(g)$$

Example: The Reaction of Solid Aluminum with Aqueous Sulfuric Acid to Form Aqueous Aluminum Sulfate and Hydrogen Gas

- Sum the number of atoms on each side to make sure that the equation is balanced.

$$2\,Al(s) + 3\,H_2SO_4(aq) \longrightarrow Al_2(SO_4)_3(aq) + 3\,H_2(g)$$

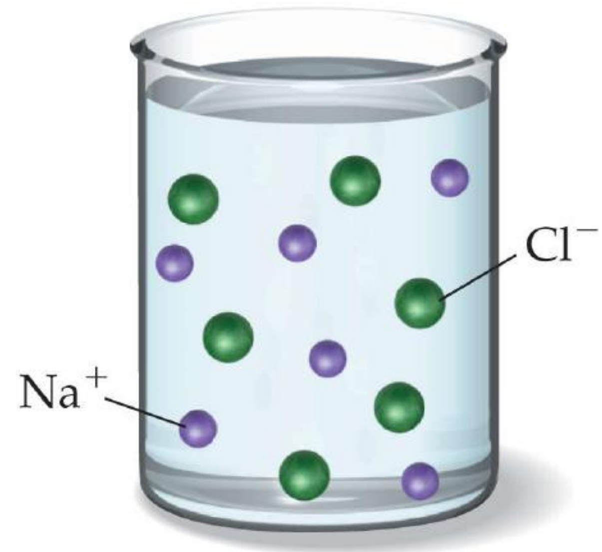
Reactants		Products
2 Al atoms	⟶	2 Al atoms
6 H atoms	⟶	6 H atoms
3 S atoms	⟶	3 S atoms
12 O atoms	⟶	12 O atoms

Aqueous Solutions and Solubility: Terminology for Compounds Dissolved in Water

- A compound is **soluble** in a particular liquid if it dissolves in that liquid.
- A compound is **insoluble** if it does not dissolve in the liquid.
- An **aqueous solution** is a homogeneous mixture of a substance with water.
- When ionic compounds dissolve in water, they dissociate into their component ions.

Aqueous Solutions: NaCl Dissolves in Water

- A sodium chloride solution, NaCl(*aq*), does not contain any NaCl units.
- Only dissolved Na^+ ions and Cl^- ions are present.
- Substances (such as NaCl) that completely dissociate into ions in solution are called *strong electrolytes.*



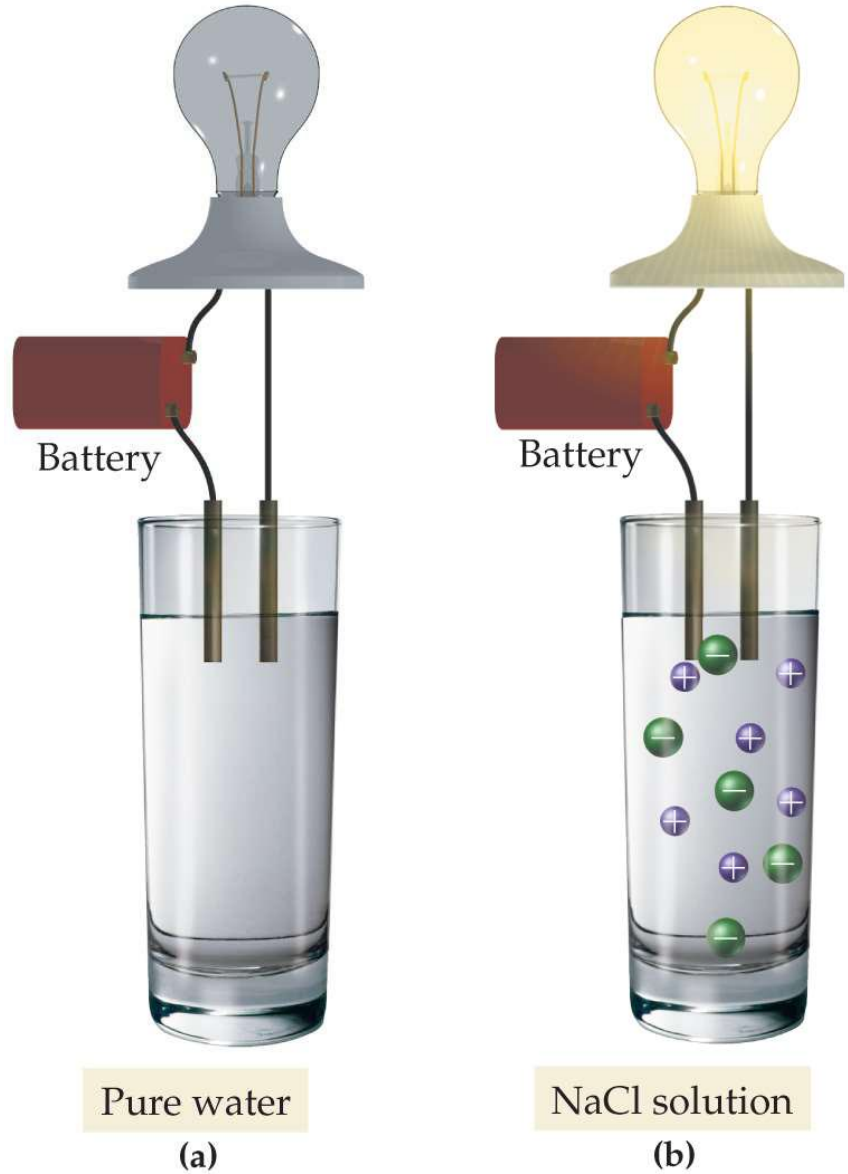
A sodium chloride solution contains independent **Na^+** and **Cl^-** ions.

Ions as Conductors: Strong Electrolyte Solutions

(a) Pure water does not conduct electricity.

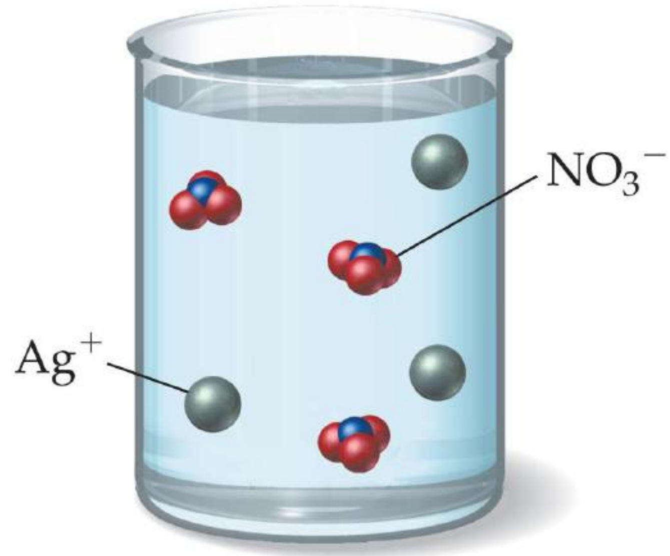
(b) Ions in a sodium chloride solution conduct electricity, causing the bulb to light.

- Solutions such as NaCl(*aq*) are called **strong electrolyte solutions.**



Aqueous Solutions: $AgNO_3$ Dissolves in Water

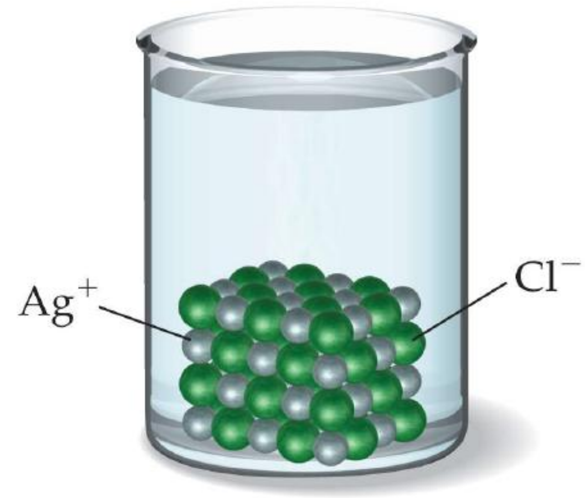
- A silver nitrate solution, $AgNO_3$(*aq*), does not contain any $AgNO_3$ units.
- Only dissolved Ag^+ ions and NO_3^- ions are present.
- $AgNO_3$(*aq*) is a strong electrolyte solution.
- When compounds containing polyatomic ions such as NO_3^- dissolve, the polyatomic ions dissolve as intact units.



A silver nitrate solution contains independent **Ag^+** and **NO_3^-** ions.

Aqueous Solutions: AgCl Does Not Dissolve in Water

- Not all ionic compounds dissolve in water.
- AgCl does ***not*** dissolve in water.
- AgCl remains as a solid, AgCl(*s*), separate from the liquid water.



When silver chloride is added to water, it remains as solid AgCl—it does not dissolve into independent ions.

Empirical Rules of Solubility

- A compound is soluble in a particular liquid if it dissolves in that liquid; a compound is insoluble if it does not dissolve in the liquid.

Empirical Rules of Solubility

- For ionic compounds, empirical rules of solubility have been deduced from observations of many compounds.
- As ionic compounds are made by pairing a cation and an anion, the rules are about whether a cation is soluble with other anions, or whether an anion is soluble with other cations

Solubility: Mostly Soluble

For example:

- The solubility rules indicate that compounds containing the potassium ion are *soluble.*
- Compounds such as KBr, KOH, and K_2CO_3 dissolve in water to form strong electrolyte solutions.
- The solubility rules state that compounds containing the NO_3^- ion are *soluble.*
- Compounds such as $AgNO_3$, $Pb(NO_3)_2$, and $Ca(NO_3)_2$ dissolve in water to form strong electrolyte solutions.

Solubility: Mostly Insoluble

For example:

- The solubility rules state that, with some exceptions, compounds containing the CO_3^{2-} ion are *insoluble.*
- Compounds such as $CaCO_3$, $FeCO_3$, $SrCO_3$, and $CuCO_3$ do not dissolve in water.
- Note that the solubility rules contain many *exceptions.*
- For example, compounds containing CO_3^{2-} *are soluble when paired with* Li^+, Na^+, K^+, or NH_4^+.
- Thus, Li_2CO_3, Na_2CO_3, K_2CO_3, and $(NH_4)_2CO_3$ are soluble.

Empirical Rules of Solubility

The observations about the solubilities are summarized as rules about whether a cation is soluble with other anions, or whether an anion is soluble with other cations.

That makes sense, since ionic compounds are made by pairing a cation and an anion.

So the rules are about which cation+anion combinations are soluble/insoluble.

Solubility Rules for cation-anion pairs

Before you get scared by a bunch of rules to memorize, know that the solubility rules are ordered in decreasing solubility:

Always soluble

Less soluble as we go down the list of rules

"Always" insoluble

Always **soluble** cations (*with any anion!*)
- They never precipitate!

Always **soluble** anions (*with any cation!*)
- They never precipitate!

Generally **soluble** anions
-- with exceptions

Generally **insoluble** anions
-- with exceptions

"Others": **insoluble** anions
-- except with "always soluble" cations

Solubility Rules for cation-anion pairs

1. Alkali metal (**Group 1**) cations and **NH_4^+** are <u>always soluble</u>

2. **Nitrate, perchlorate, chlorate, acetate** are <u>always soluble</u>.

3. **Cl^-**, **Br^-**, and **I^-** are generally soluble (except with Ag^+, Pb^{2+}, Hg_2^{2+}).
   halides

4. **SO_4^{2-}** is generally soluble (except with Sr^{2+}, Ba^{2+}, Ca^{2+}, Hg_2^{2+}, Pb^{2+}; and Ag_2SO_4 is only "marginally soluble").

5. **S^{2-}** and **OH^-** are generally insoluble.
   Exceptions: see Rule 1, <u>and</u>
   Soluble or somewhat soluble with Group 2 cations

6. **"Other"** anions we see in General Chemistry (CO_3^{2-}, CrO_4^{2-}, PO_4^{3-}, $C_2O_4^{2-}$ etc.) are insoluble (exceptions: see Rule 1)

Practice:
Check your solution on next page

Determine if each of the following is soluble in water

KOH

$AgBr$

$CaCl_2$

$Pb(NO_3)_2$

$PbSO_4$

Practice – Determine if each of the following is soluble in water

KOH	KOH is soluble because it contains K^+
AgBr	AgBr is insoluble; most bromides are soluble, but AgBr is an exception
$CaCl_2$	$CaCl_2$ is soluble; most chlorides are soluble, and $CaCl_2$ is not an exception
$Pb(NO_3)_2$	$Pb(NO_3)_2$ is soluble because it contains NO_3^-
$PbSO_4$	$PbSO_4$ is insoluble; most sulfates are soluble, but $PbSO_4$ is an exception

Precipitation Reactions: Reactions in Aqueous Solution that Form a Solid

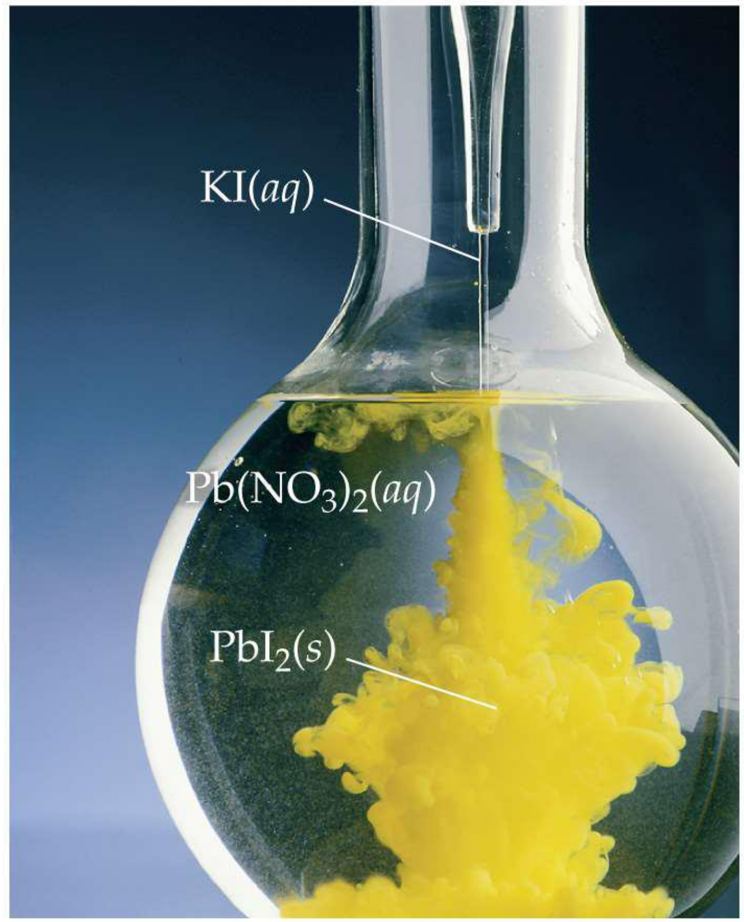
- Sodium carbonate in laundry detergent reacts with dissolved Mg^{2+} and Ca^{2+} ions to form solids that precipitate from solution.
- These reactions are examples of **precipitation reactions,** reactions that **form a solid (*s*),** called a **precipitate,** upon mixing two aqueous solutions.

$2\ KI(aq) + Pb(NO_3)_2(aq) \rightarrow PbI_2(s) + 2\ KNO_3(aq)$

$$2\ KI(aq) + Pb(NO_3)_2(aq) \longrightarrow PbI_2(s) + 2\ KNO_3(aq)$$

Precipitation

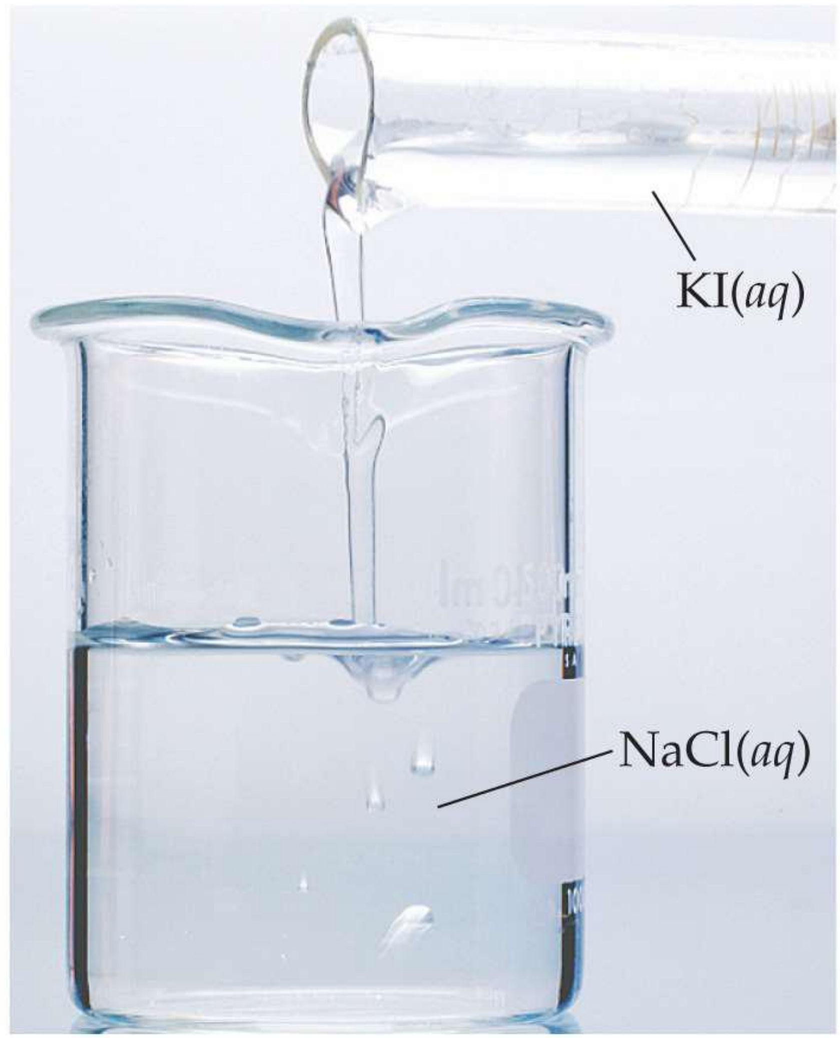
- When a potassium iodide solution is mixed with a lead(II) nitrate solution,
  a brilliant yellow precipitate of $PbI_2(s)$ forms.



KI(*aq*) + NaCl(*aq*) → NO REACTION

$$KI(aq) + NaCl(aq) \longrightarrow \text{NO REACTION}$$

- Precipitation reactions do not always occur when mixing two aqueous solutions.
- When a potassium iodide solution is mixed with a sodium chloride solution, no reaction occurs.



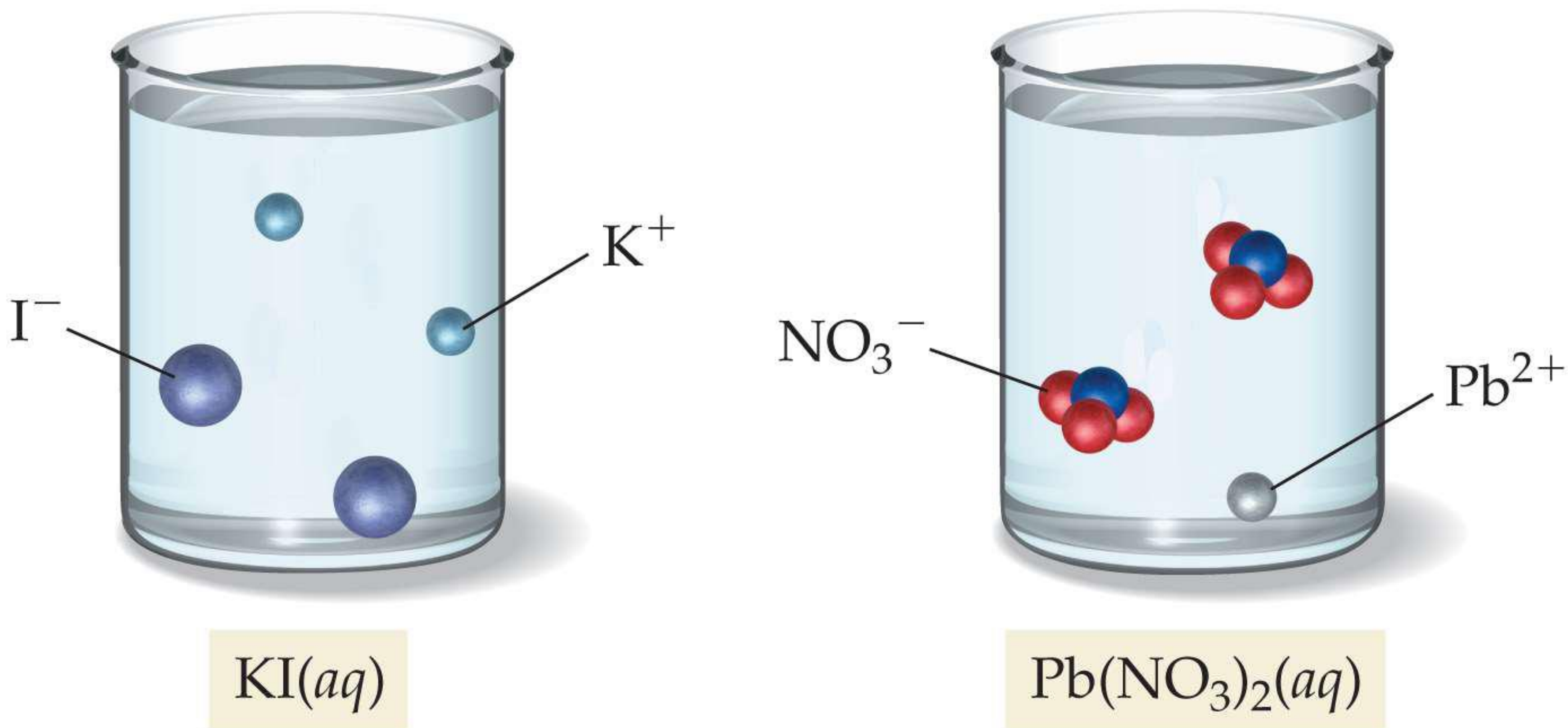
Predicting Precipitation Reactions

- The key to predicting precipitation reactions is understanding that *only insoluble compounds form precipitates.*
- In a precipitation reaction, two solutions containing soluble compounds combine and an insoluble compound precipitates.

If there is an insoluble cation-anion combination, it will precipitate

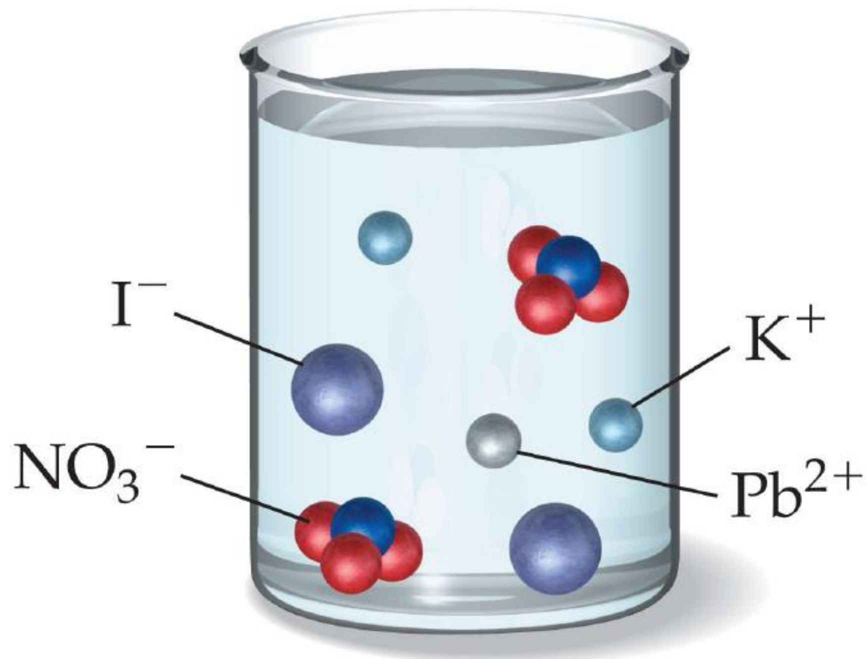
$2\ KI(aq) + Pb(NO_3)_2(aq) \rightarrow PbI_2(s) + 2\ KNO_3(aq)$

Before mixing, $KI(aq)$ and $Pb(NO_3)_2(aq)$ are both dissociated in their respective solutions.



$2\ KI(aq) + Pb(NO_3)_2(aq) \rightarrow PbI_2(s) + 2\ KNO_3(aq)$

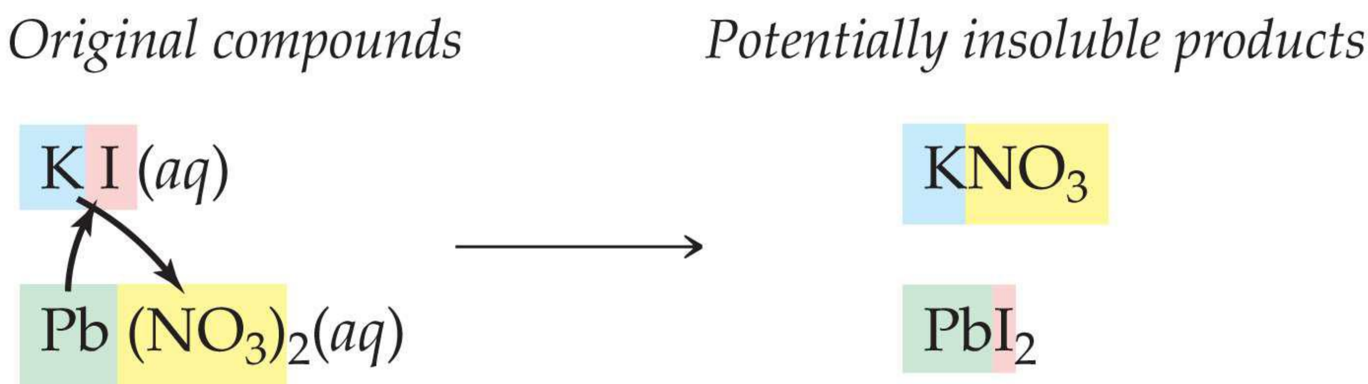
The instant that the solutions are mixed, all four ions are present.



$KI(aq)$ and $Pb(NO_3)_2(aq)$

2 KI(*aq*) + $Pb(NO_3)_2$(*aq*) → PbI_2(*s*) + 2 KNO_3(*aq*)

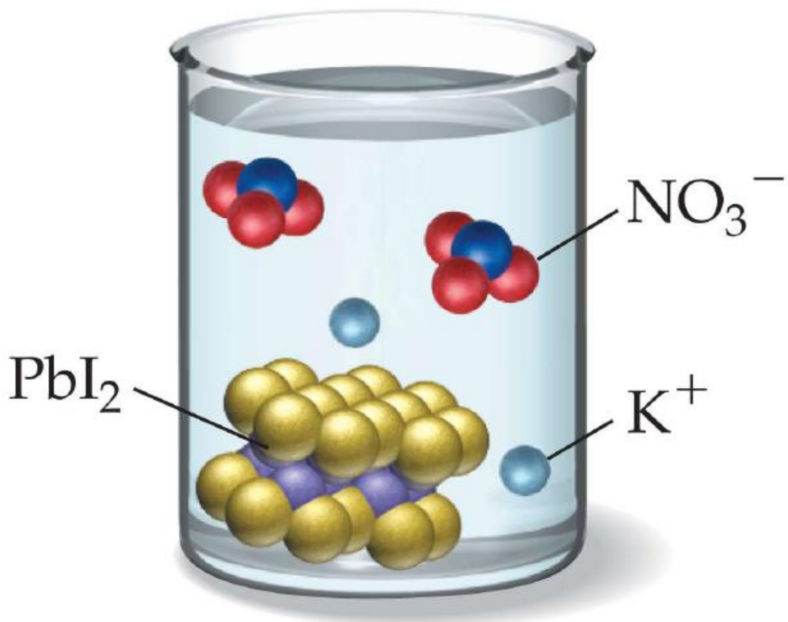
- The cation from one compound can now pair with the anion from the other compound to form new (and potentially insoluble) products.



If the *potentially insoluble* products are both *soluble,* then no reaction occurs.
If one or both of the potentially insoluble products are *insoluble,* a precipitation reaction occurs.

$2\ KI(aq) + Pb(NO_3)_2(aq) \rightarrow PbI_2(s) + 2\ KNO_3(aq)$

- In this case, KNO_3 is soluble, but PbI_2 is insoluble. Consequently, PbI_2 precipitates.

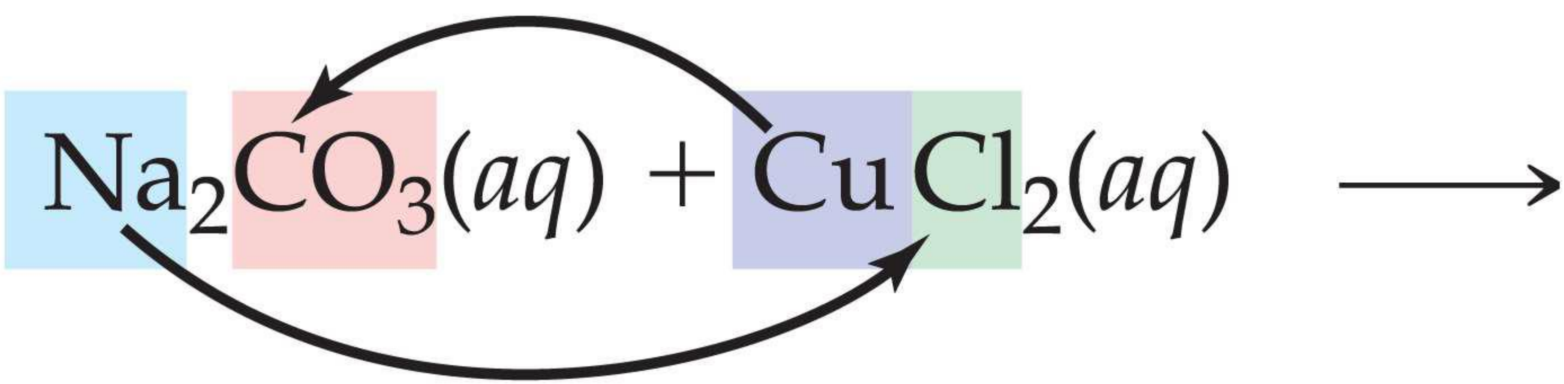


$PbI_2(s)$ and $KNO_3(aq)$

Writing Equations for Precipitation Reactions
Shown through an example:

Write an equation for the precipitation reaction that occurs (if any) when solutions of sodium carbonate $Na_2CO_3(aq)$ and copper(II) chloride $CuCl_2(aq)$ are mixed.

Combine the cation from one reactant with the anion from the other.



Writing Equations for Precipitation Reactions

Use the solubility rules to determine whether any of the potential new products are indeed insoluble.

Potentially Insoluble Products:

$NaCl$ $CuCO_3$

Make sure to write correct (charge-neutral) formulas for the new ionic compounds

NaCl is *soluble* (compounds containing Cl^- are usually soluble, and Na^+ is not an exception).

$CuCO_3$ is *insoluble* (compounds containing CO_3^{2-} are usually insoluble, and Cu^{2+} is not an exception).

Writing Equations for Precipitation Reactions

If all of the potentially insoluble products are soluble, there will be no precipitate. Write NO REACTION next to the arrow.

—Not the case here

Writing Equations for Precipitation Reactions

One of the *potentially* insoluble products is *indeed* insoluble, so write its formula as the product of the reaction, using (*s*) to indicate solid.

The soluble cation and anion remaining in solution will be the other product and can be written as the aqueous solution of the ionic compound that would produce them if dissolved in aqueous solution, which we show as (*aq*).

$$Na_2CO_3(aq) + CuCl_2(aq) \rightarrow CuCO_3(s) + NaCl(aq)$$

Writing Equations for Precipitation Reactions

Balance the equation.
Remember to adjust only the coefficients, not the subscripts.

$$Na_2CO_3(\textit{aq}) + CuCl_2(\textit{aq}) \rightarrow CuCO_3(\textit{s}) + \mathbf{2}\ NaCl(\textit{aq})$$

Practice:

Predict the products and balance the equation.

Check your solution on next page

$KCl(aq) + AgNO_3(aq) \rightarrow ??$

Practice – Predict the products and balance the equation

$KCl(aq) + AgNO_3(aq) \rightarrow KNO_3(aq) + AgCl(s)$

Practice:

Predict the products and balance the equation

Check your solution on next page

$(NH_4)_2SO_4$ is mixed with an aqueous solution of $Pb(C_2H_3O_2)_2$

Practice – Predict the products and balance the equation

$(NH_4)_2SO_4$ is mixed with an aqueous solution of $Pb(C_2H_3O_2)_2$

$(NH_4)_2SO_{4(aq)} + Pb(C_2H_3O_2)_{2(aq)} \rightarrow PbSO_{4(s)} + 2NH_4C_2H_3O_{2(aq)}$

Writing Equations for Precipitation Reactions

The kind of reaction equations we just produced is known as the “molecular equation” of a precipitation reaction, where the ions are written as compound formulas (resembling molecular formulas, hiding the ionic nature of the constituent parts).

There are other ways to write the reaction equation as well.

Writing Chemical Equations for Reactions in Solution: Molecular and Complete Ionic Equations

- A **molecular equation** is an equation showing the complete neutral formulas for every compound in the reaction.
  -- Remember that these are ionic compounds, so they are NOT really molecules, just written in "molecular" form

$$AgNO_3(aq) + NaCl(aq) \rightarrow AgCl(s) + NaNO_3(aq)$$

- **Complete ionic equations** show aqueous ionic compounds that normally dissociate in solution as they are actually present in solution.

$$Ag^+(aq) + NO_3^-(aq) + Na+(aq) + Cl^-(aq) \rightarrow AgCl(s) + Na^+(aq) + NO_3^-(aq)$$

- When writing complete ionic equations, separate only aqueous ionic compounds into their constituent ions.
- Do NOT separate solid, liquid, or gaseous compounds.

Writing Chemical Equations for Reactions in Solution: Complete Ionic and Net Ionic Equations

- In the complete ionic equation, some of the ions in solution appear unchanged on both sides of the equation.
- These ions are called **spectator ions** because they do not participate in the reaction.

$$Ag^{+}(aq) + NO_3^{-}(aq) + Na^{+}(aq) + Cl^{-}(aq) \longrightarrow AgCl(s) + Na^{+}(aq) + NO_3^{-}(aq)$$

Spectator ions

Writing Chemical Equations for Reactions in Solution: Complete Ionic and Net Ionic Equations

- To simplify the equation, and to more clearly show what is happening, spectator ions can be omitted.
- Equations such as this one, which show only the *species* that actually participate in the reaction, are called **net ionic equations.**

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

Molecular, Complete Ionic, and Net Ionic Equations

To summarize:

A ***molecular equation*** is a chemical equation showing the complete, neutral formulas for every compound in a reaction.

A ***complete ionic equation*** is a chemical equation showing all of the species as they are actually present in solution.

A ***net ionic equation*** is an equation showing only the species that actually participate in the reaction.

-- what's left after eliminating the spectator ions from the complete ionic equation

Practice:
Check your solution on next page

Write the complete ionic and net ionic equation, and identify the spectator ions and balance

$$K_2SO_4(aq) + AgNO_3(aq) \rightarrow$$

Practice – Write the ionic and net ionic equation for each

$$K_2SO_4(aq) + 2\ AgNO_3(aq) \rightarrow 2\ KNO_3(aq) + Ag_2SO_4(s)$$

$$2K^+_{(aq)} + SO_4^{2-}{}_{(aq)} + 2Ag^+_{(aq)} + 2NO_3^-{}_{(aq)} \rightarrow 2K^+_{(aq)} + 2NO_3^-{}_{(aq)} + Ag_2SO_{4(s)}$$

K^+ and NO_3^- appear on both sides: "spectator ions"

We remove the spectator ions from the complete ionic equation to obtain the net ionic equation:

$$2\ Ag^+(aq) + SO_4^{2-}(aq) \rightarrow Ag_2SO_4(s)$$

Practice:
Check your solution on next page

Write the ionic and net ionic equation for each, identify the spectator ions and balance

$Na_2S(aq) + CaCl_2(aq) \rightarrow$

Practice

Write the ionic and net ionic equation for each, identify the spectator ions and balance

$Na_2S(aq) + CaCl_2(aq) \rightarrow$

$(Na^+ + S^{2-}) + (Ca^{2+} + Cl^-) \rightarrow (Na^+ + Cl^-) + (Ca^{2+} + S^{2-})$

$Na_2S(aq) + CaCl_2(aq) \rightarrow NaCl(aq) + CaS(aq)$

No reaction because all products are aqueous

All ions are spectator ions

Acid–Base Reactions

Acid–base reactions are reactions that form water upon mixing of an acid and a base.

$$\underset{\text{Acid}}{HCl(aq)} + \underset{\text{Base}}{NaOH(aq)} \longrightarrow \underset{\text{Water}}{H_2O(l)} + \underset{\text{Salt}}{NaCl(aq)}$$

Some Common Acids and Bases

TABLE 7.3 Some Common Acids and Bases

Acid	Formula	Base	Formula
hydrochloric acid	HCl	sodium hydroxide	$NaOH$
hydrobromic acid	HBr	lithium hydroxide	$LiOH$
nitric acid	HNO_3	potassium hydroxide	KOH
sulfuric acid	H_2SO_4	calcium hydroxide	$Ca(OH)_2$
perchloric acid	$HClO_4$	barium hydroxide	$Ba(OH)_2$
acetic acid	$HC_2H_3O_2$		

Acids produce H^+ cations when dissolved in water

- The "acidic" H atoms are put in front of the formula

Bases produce OH^- anions when dissolved in water

- Metal hydroxides are bases

Acid–Base Reactions

Acid–base reactions (also called **neutralization reactions**) generally form water and an ionic compound—called a **salt**—that usually remains dissolved in the solution.

The net ionic equation for many acid–base reactions is as follows:

$$H^+(aq) + OH^-(aq) \rightarrow H_2O(l)$$

Practice:
Check your solution on next page

Write the molecular, ionic, and net-ionic equation for the reaction of aqueous nitric acid with aqueous calcium hydroxide, and identify the spectator ions

1. Write the formulas of the reactants

$$HNO_3(aq) + Ca(OH)_2(aq) \rightarrow$$

2. Determine the possible products

a) determine the ions present when each reactant dissociates or ionizes

$$(H^+ + NO_3^-) \; + \; (Ca^{2+} + OH^-) \rightarrow$$

b) exchange the ions, H^+ combines with OH^- to make $H_2O(l)$

$$(H^+ + NO_3^-) + (Ca^{2+} + OH^-) \rightarrow (Ca^{2+} + NO_3^-) + H_2O(l)$$

c) write the formula of the salt

$$(H^+ + NO_3^-) + (Ca^{2+} + OH^-) \rightarrow Ca(NO_3)_2 + H_2O(l)$$

3. Determine the solubility of the salt

   $Ca(NO_3)_2$ is soluble

4. Write an (*s*) after the insoluble products and an (*aq*) after the soluble products

$$HNO_3(aq) + Ca(OH)_2(aq) \rightarrow Ca(NO_3)_2(aq) + H_2O(l)$$

5. Balance the equation

$$2\ HNO_3(aq) + Ca(OH)_2(aq) \rightarrow Ca(NO_3)_2(aq) + 2\ H_2O(l)$$

6. Dissociate all aqueous strong* electrolytes to get complete ionic equation

$$2H^+{}_{(aq)} + 2NO_3^-{}_{(aq)} + Ca^{2+}{}_{(aq)} + 2OH^-{}_{(aq)} \rightarrow Ca^{2+}{}_{(aq)} + 2NO_3^-{}_{(aq)} + H_2O_{(l)}$$

7. Eliminate spectator ions (Ca^{2+} and NO_3^-)to get net-ionic equation

$$2\,H^+(aq) + 2\,OH^-(aq) \rightarrow 2\,H_2O(l)$$

$$H^+(aq) + OH^-(aq) \rightarrow H_2O(l)$$

* In other courses you will probably be expected to **not** dissociate weak acids and write them in molecular form. I dislike that convention and I do not use it. But you should be aware of that expectation in the future.

Practice:
Check your solution on next page

Predict the products and balance the equation

$$HCl(\textit{aq}) + Ba(OH)_2(\textit{aq}) \rightarrow$$

$$H_2SO_4(\textit{aq}) + Sr(OH)_2(\textit{aq}) \rightarrow$$

Practice – Predict the products and balance the equation

$HCl(aq) + Ba(OH)_2(aq) \rightarrow$

$2\ HCl(aq) + Ba(OH)_2(aq)$ ® $2\ H_2O(l) + BaCl_2(aq)$

$H_2SO_4(aq) + Sr(OH)_2(aq) \rightarrow$

$H_2SO_4(aq) + Sr(OH)_2(aq) \rightarrow 2\ H_2O(l) + SrSO_4(s)$

Gas Evolution Reactions:

Strong acid or base pushes out a weaker one from a salt, when the weak acid or base is volatile or unstable, thus leaving the solution one way or another.

$$HNO_3(aq) + Na_2CO_3(aq) \rightarrow H_2CO_3(aq) + NaNO_3(aq)$$

Strong acid

CO_3^{2-} is the anion of a weak acid, H_2CO_3

Unstable:

$$H_2CO_3(aq) \rightarrow H_2O(l) + CO_2(g)$$

Bubbles out as gas

Gas Evolution Reactions in volving salts containing CO_3^{2-}

Begin by writing a skeletal equation that includes the reactants and products that form when the cation of each reactant combines with the anion of the other.

$$HNO_3(aq) + Na_2CO_3(aq) \rightarrow H_2CO_3(aq) + NaNO_3(aq)$$

You must recognize that $H_2CO_3(aq)$ decomposes into $H_2O(l)$ and $CO_2(g)$ and write the corresponding equation.

$$HNO_3(aq) + Na_2CO_3(aq) \rightarrow H_2O(l) + CO_2(g) + NaNO_3(aq)$$

Finally, balance the equation.

$$\mathbf{2}\ HNO_3(aq) + Na_2CO_3(aq) \rightarrow H_2O(l) + CO_2(g) + \mathbf{2}\ NaNO_3(aq)$$

Gas Evolution Reactions in volving salts containing CO_3^{2-}

Notice that a salt, such as Na_2CO_3, or $NaHCO_3$, is able to react with an acid and consume it (creating a new salt), even if it's not a "base" in the usual sense we have seen when we looked at acid-base neutralization reactions.

Chemistry and Health:
Neutralizing Excess Stomach Acid

- Your stomach normally contains acids that are involved in food digestion.
- Antacids are over-the-counter medicines that work by reacting with and neutralizing stomach acid.
- Antacids contain bases such as $Mg(OH)_2$, $Al(OH)_3$, and $NaHCO_3$.
- The base in an antacid neutralizes excess stomach acid, relieving heartburn and acid stomach.

Gas Evolution Reactions

TABLE 7.4 Types of Compounds That Undergo Gas Evolution Reactions

Reactant Type	Intermediate Product	Gas Evolved
sulfides	none	H_2S
carbonates and bicarbonates	H_2CO_3	CO_2
sulfites and bisulfites	H_2SO_3	SO_2
ammonium	NH_4OH	NH_3

H_2CO_3 and H_2SO_3 are weak and unstable acids.

H_2S is a weak acid that is a gas at room temperature.

Their salts make gas-evolution reactions when mixed with stronger acids.

NH_3 is a weak base that is a gas at room temperature. The formal intermediate $NH_4OH(aq)$ actually exists mostly as unreacted NH_3 and H_2O, and NH_3 leaves as gas.

Gas Evolution Reactions

TABLE 7.4 Types of Compounds That Undergo Gas Evolution Reactions

Example
$2\,HCl(aq) + K_2S(aq) \longrightarrow H_2S(g) + 2\,KCl(aq)$
$2\,HCl(aq) + K_2CO_3(aq) \longrightarrow H_2O(l) + CO_2(g) + 2\,KCl(aq)$
$2\,HCl(aq) + K_2SO_3(aq) \longrightarrow H_2O(l) + SO_2(g) + 2\,KCl(aq)$
$NH_4Cl(aq) + KOH(aq) \longrightarrow H_2O(l) + NH_3(g) + KCl(aq)$

Oxidation–Reduction Reactions

- Reactions involving the transfer of electrons are called **oxidation–reduction reactions** or **redox reactions.**
- Redox reactions are responsible for the rusting of iron, the bleaching of hair, and the production of electricity in batteries.
- Many redox reactions involve the reaction of a substance with oxygen.

Oxidation–Reduction Reactions

- A fundamental definition of oxidation is *the loss of electrons.*
- A fundamental definition of reduction is *the gain of electrons.*
- Helpful mnemonics:

OIL RIG

Oxidation **I**s **L**oss; **R**eduction **I**s **G**ain

LEO says GER

Lose **E**lectrons **O**xidation; **G**ain **E**lectrons **R**eduction

Oxidation–Reduction Reactions

- Notice that oxidation and reduction must occur together.
- If one substance loses electrons (oxidation), then another substance must gain electrons (reduction).
- For now, you simply need to be able to identify easy-to-identify redox reactions.

Oxidation–Reduction Reactions

We can readily recognize a redox reaction if:

- An elemental substance such as Na(s), O_2(g), H_2(g), S_8(s), etc. is a reactant or product

- Not all redox reactions involve elemental substances (far from it)
- More generally, in a redox reaction one substance transfers electrons to another substance.

$4\ Fe(s) + 3\ O_2(g) \rightarrow 2\ Fe_2O_3(s)$

$Fe(s) + CuSO_4(aq) \rightarrow Cu(s) + FeSO_4(aq)$

Combustion Reactions

Combustion reactions are a type of redox reaction.

Combustion reactions are characterized by the reaction of a substance with O_2 to form one or more oxygen-containing compounds, often including water.

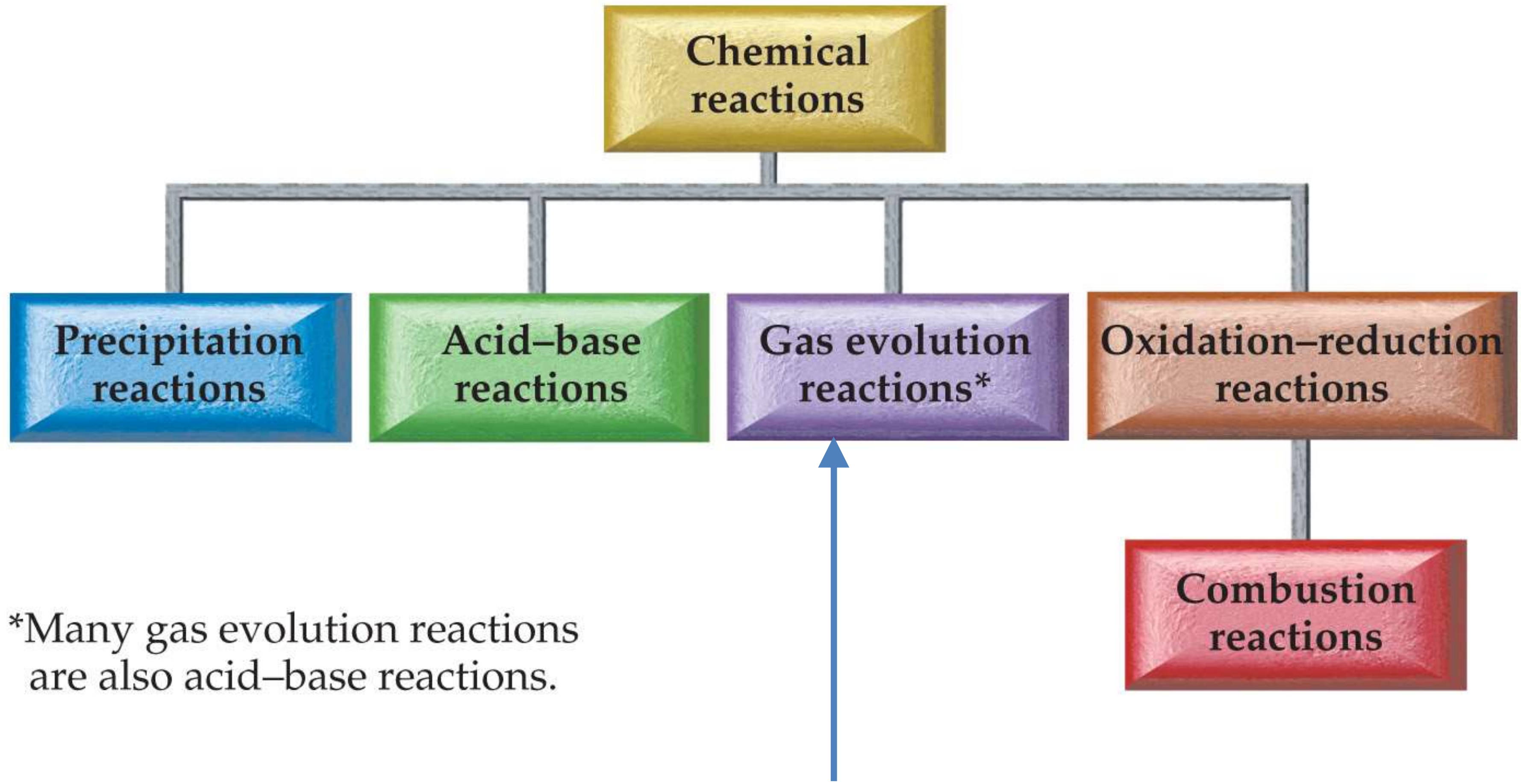
Combustion reactions are exothermic (they emit heat).

Compounds containing carbon and hydrogen—or carbon, hydrogen, and oxygen—always form carbon dioxide and water upon combustion.

Other combustion reactions include the reaction of carbon with oxygen to form carbon dioxide.

$$CH_4(g) + 2\,O_2(g) \longrightarrow CO_2(g) + 2\,H_2O(g)$$

So far we classified chemical reactions according to the "type of chemistry"



Not a fundamental reaction type like the others

We also classify chemical reactions according to what atoms (or groups of atoms) do

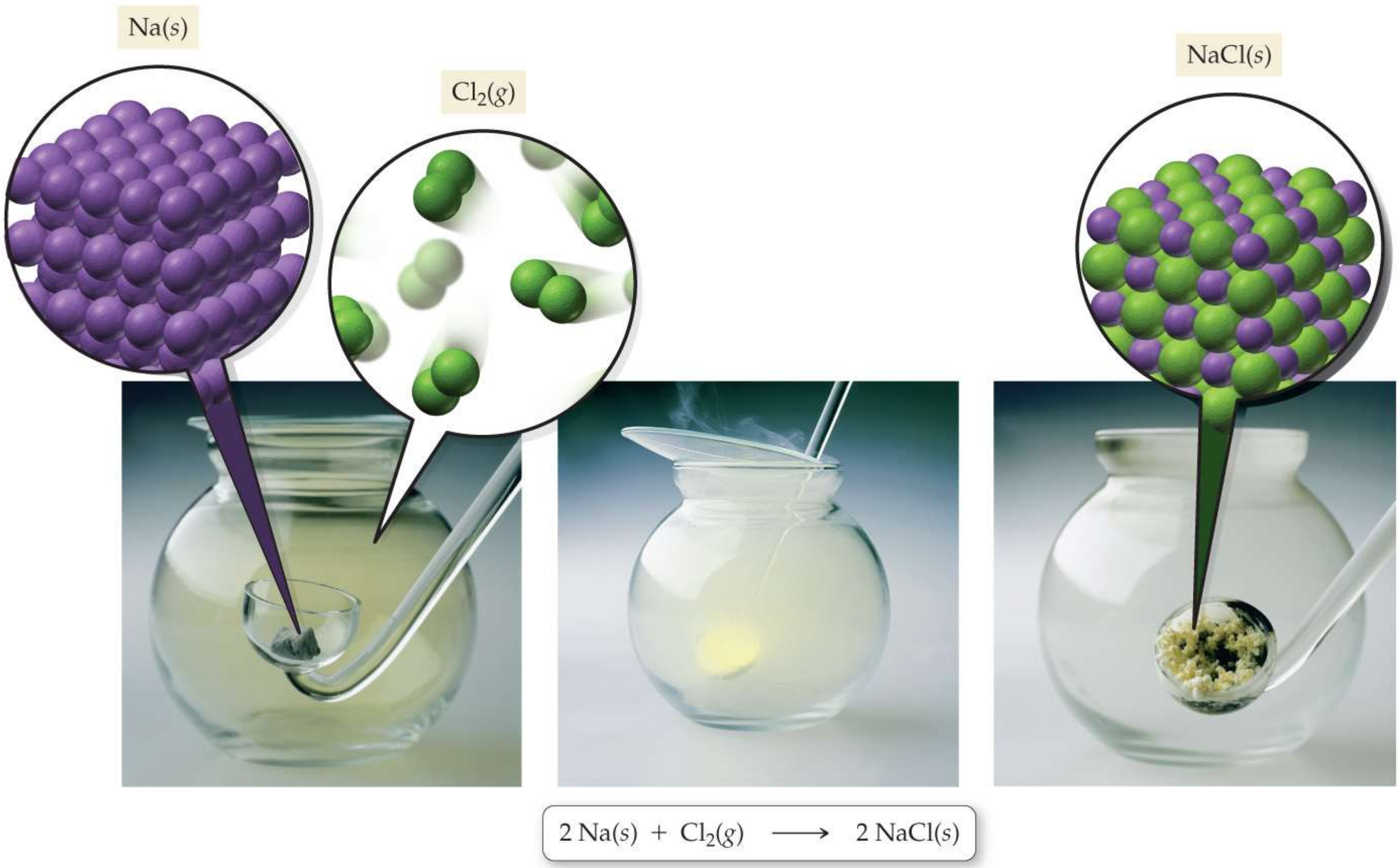
Type of Reaction	Generic Equation
synthesis or combination	$A + B \longrightarrow AB$
decomposition	$AB \longrightarrow A + B$
single-displacement	$A + BC \longrightarrow AC + B$
double-displacement	$AB + CD \longrightarrow AD + CB$

A type of redox rxn

Precipitation & acid-base rxns

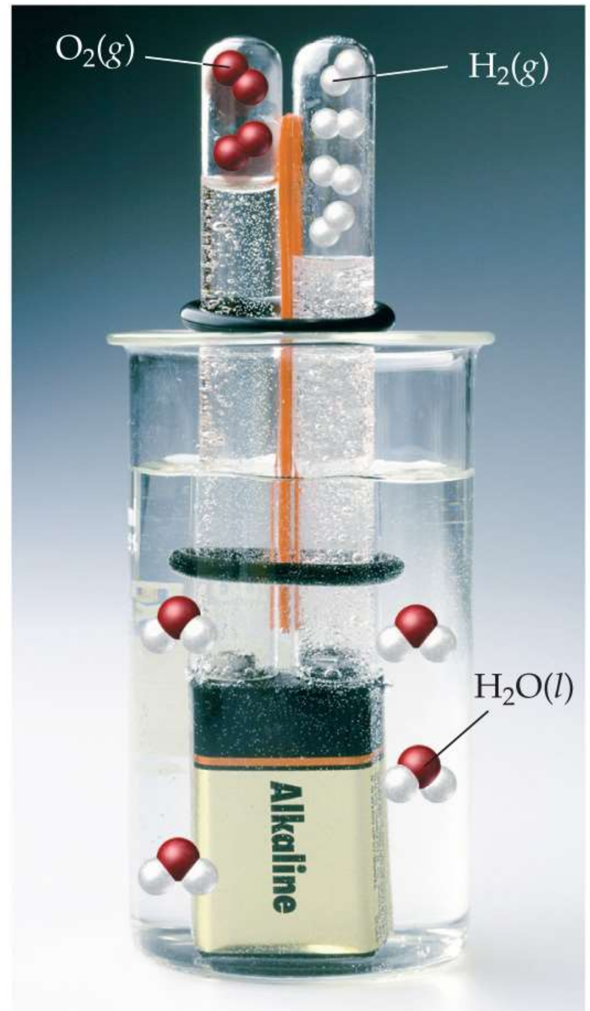
Synthesis Reactions

In a *synthesis* reaction, two simpler substances combine to make a more complex substance. In this series of photographs we see sodium metal and chlorine gas. When they combine, a chemical reaction occurs that forms sodium chloride.



Decomposition Reactions

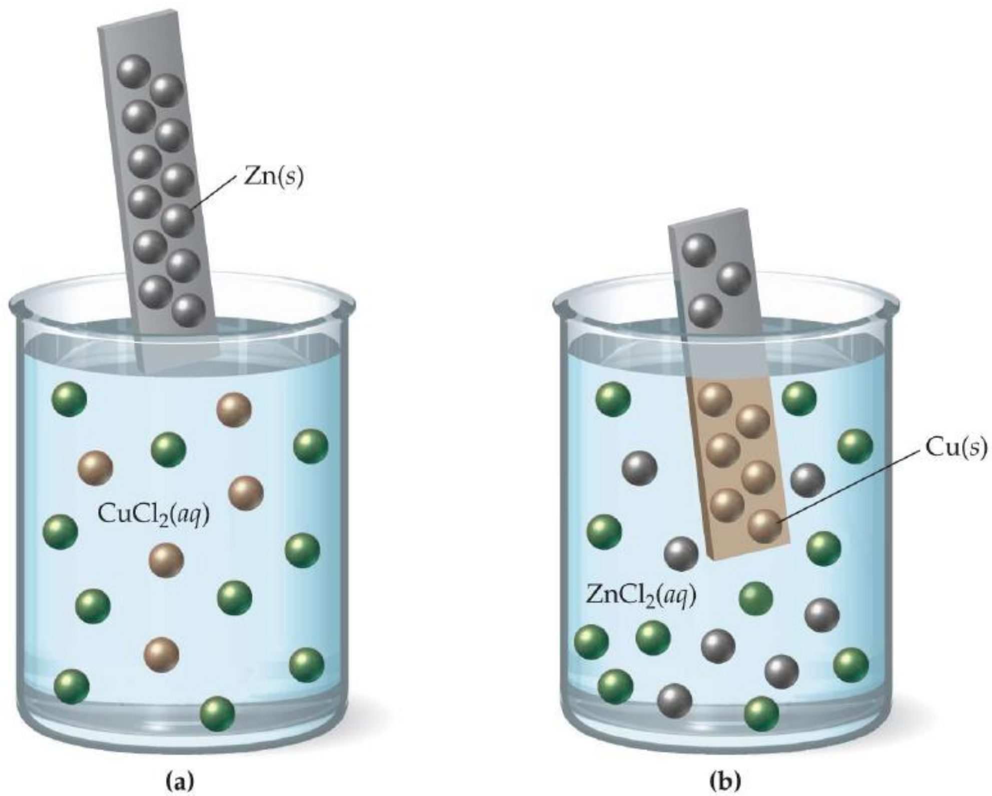
- In a *decomposition* reaction, a complex substance decomposes to form simpler substances.
- When electrical current is passed through water, the water undergoes a decomposition reaction to form hydrogen gas and oxygen gas.



$$2\,H_2O(l) \longrightarrow 2\,H_2(g) + O_2(g)$$

Single-Displacement Reactions

- In a displacement or *single-displacement* reaction, one element displaces another in a compound.
- When metallic zinc is immersed in a solution of copper(II) chloride, the zinc atoms displace the copper ions in solution.



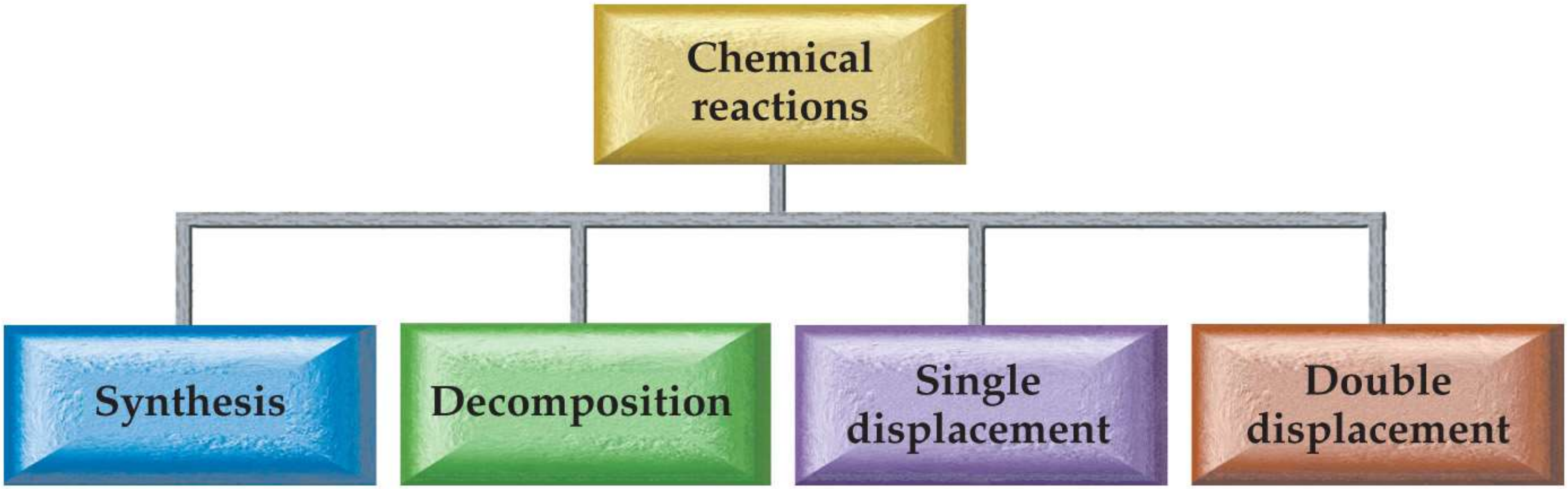
Double-displacement Reactions

- In a *double-displacement* reaction, two elements or groups of elements in two different compounds exchange places to form two new compounds.
- A double-displacement reaction follows the general form

$$AB + CD \rightarrow AD + BC$$

- Double displacement reactions include:
  - precipitation reactions
  - acid–base reactions
  - gas evolution reactions

Classification Flow Chart



Single Displacement reactions

$2Li(s) + 2HCl(aq) \rightarrow 2LiCl\ (aq) + H_2(g)$

Li(s) is more "active" than $H_2(g)$, so it can displace H^+ into its elemental form, $H_2(g)$

In this course, we will not need to predict if a proposed single-displacement is possible.

We only need to be able to recognize a reaction as a single-displacement reaction.

What other classification do single-displacement reactions belong to? 🤔

Chapter 7 in Review

- Chemical reactions: One or more substances—either elements or compounds—change into a different substance.
- Evidence of a chemical reaction: The only absolute evidence for a chemical reaction is chemical analysis showing that one or more substances have changed into another substance.
- However, one or more of the following often accompanies a chemical reaction: a color change; the formation of a solid or precipitate; the formation of a gas; the emission of light; and the emission or absorption of heat.

Chapter 7 in Review

- Chemical equations: Chemical equations must be balanced to reflect the conservation of matter in nature.
- Aqueous solutions and solubility: If a substance dissolves in water, it is soluble.
- Some specific types of reactions are precipitation reaction, acid–base reaction, gas evolution reaction, redox reaction, and combustion reaction.
- Chemical reaction classifications are synthesis reaction, decomposition reaction, single-displacement reaction, and double-displacement reaction.

Chemical Skills Learning Objectives

1. LO: Identify a chemical reaction.
2. LO: Write balanced chemical equations.
3. LO: Determine if a compound is soluble.
4. LO: Predict and write equations for precipitation reactions.
5. LO: Write molecular, complete ionic, and net ionic equations.
6. LO: Identify and write equations for acid–base reactions.
7. LO: Identify and write equations for gas evolution reactions.
8. LO: Identify redox reactions.
9. LO: Identify and write equations for combustion reactions.
10. LO: Classify chemical reactions.