GLOBAL EDITION

Introductory Chemistry

SIXTH EDITION in SI Units

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Pearson

Periodic Table of the Elements

1 H 1.01 hydrogen	
Atomic number	1
Element symbol	H
Atomic mass*	1.01
Element name	hydrogen

PERIOD	GROUP 1 1A	2 2A	3 3B	4 4B	5 5B	6 6B	7 7B	8 8B	9 8B
1	1 **H** 1.01 hydrogen								
2	3 **Li** 6.94 lithium	4 **Be** 9.01 beryllium							
3	11 **Na** 22.99 sodium	12 **Mg** 24.31 magnesium							
4	19 **K** 39.10 potassium	20 **Ca** 40.08 calcium	21 **Sc** 44.96 scandium	22 **Ti** 47.88 titanium	23 **V** 50.94 vanadium	24 **Cr** 52.00 chromium	25 **Mn** 54.94 manganese	26 **Fe** 55.85 iron	27 **Co** 58.93 cobalt
5	37 **Rb** 85.47 rubidium	38 **Sr** 87.62 strontium	39 **Y** 88.91 yttrium	40 **Zr** 91.22 zirconium	41 **Nb** 92.91 niobium	42 **Mo** 95.95 molybdenum	43 **Tc** (99) technetium	44 **Ru** 101.07 ruthenium	45 **Rh** 102.91 rhodium
6	55 **Cs** 132.91 cesium	56 **Ba** 137.33 barium	57 **La** 138.91 lanthanum	72 **Hf** 178.49 hafnium	73 **Ta** 180.95 tantalum	74 **W** 183.85 tungsten	75 **Re** 186.21 rhenium	76 **Os** 190.23 osmium	77 **Ir** 192.22 iridium
7	87 **Fr** (223) francium	88 **Ra** (226) radium	89 **Ac** (227) actinium	104 **Rf** (261) rutherfordium	105 **Db** (262) dubnium	106 **Sg** (263) seaborgium	107 **Bh** (262) bohrium	108 **Hs** (265) hassium	109 **Mt** (266) meitnerium

Lanthanide series	58 **Ce** 140.12 cerium	59 **Pr** 140.91 praseodymium	60 **Nd** 144.24 neodymium	61 **Pm** (147) promethium	62 **Sm** 150.36 samarium	63 **Eu** 151.97 europium
Actinide series	90 **Th** (232) thorium	91 **Pa** (231) protactinium	92 **U** (238) uranium	93 **Np** (237) neptunium	94 **Pu** (244) plutonium	95 **Am** (243) americium

*The mass number of an important radioactive isotope—not the atomic mass—is shown in parentheses for those elements with no stable isotopes.

5. Check to make certain the equation is balanced by summing the total number of each type of atom on both sides of the equation.	$SiO_2(s) + 3\,C(s) \longrightarrow SiC(s) + 2\,CO(g)$	$2\,C_8H_{18}(l) + 25\,O_2(g) \longrightarrow 16\,CO_2(g) + 18\,H_2O(g)$

Reactants		Products
1 Si atom	⟶	1 Si atom
2 O atoms	⟶	2 O atoms
3 C atoms	⟶	3 C atoms

The equation is balanced.

Reactants		Products
16 C atoms	⟶	16 C atoms
36 H atoms	⟶	36 H atoms
50 O atoms	⟶	50 O atoms

The equation is balanced.

▶ **SKILLBUILDER 7.2** | Write a balanced equation for the reaction between solid chromium(III) oxide and solid carbon to produce solid chromium and carbon dioxide gas.

▶ **SKILLBUILDER 7.3** | Write a balanced equation for the combustion reaction of gaseous C_4H_{10} and gaseous oxygen to form gaseous carbon dioxide and gaseous water.

▶ **FOR MORE PRACTICE** Example 7.17; Problems 35, 36, 37, 38.

EXAMPLE 7.4 Balancing Chemical Equations

Write a balanced equation for the reaction of solid aluminum with aqueous sulfuric acid to form aqueous aluminum sulfate and hydrogen gas.

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.	**SOLUTION** $Al(s) + H_2SO_4(aq) \longrightarrow Al_2(SO_4)_3(aq) + H_2(g)$
Since both aluminum and hydrogen occur as free 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 also 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.	$Al(s) + 3\,H_2SO_4(aq) \longrightarrow Al_2(SO_4)_3(aq) + H_2(g)$
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.	$2\,Al(s) + 3\,H_2SO_4(aq) \longrightarrow Al_2(SO_4)_3(aq) + H_2(g)$
Balance H next. Since there are 6 H atoms on the left side, place a 3 in front of $H_2(g)$ on the right side.	$2\,Al(s) + 3\,H_2SO_4(aq) \longrightarrow Al_2(SO_4)_3(aq) + 3\,H_2(g)$
Finally, 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

▶ **SKILLBUILDER 7.4** | **Balancing Chemical Equations**

Write a balanced equation for the reaction of aqueous lead(II) acetate with aqueous potassium iodide to form solid lead(II) iodide and aqueous potassium acetate.

▶ **FOR MORE PRACTICE** Problems 39, 40, 41, 42, 43, 44.

EXAMPLE 7.5 Balancing Chemical Equations

Balance the chemical equation.

$Fe(s) + HCl(aq) \longrightarrow FeCl_3(aq) + H_2(g)$

Since Cl occurs in only one compound on each side of the equation, balance it first. One Cl atom is on the left side of the equation, and 3 Cl atoms are on the right side. To balance Cl, place a 3 in front of HCl.	**SOLUTION** $Fe(s) + 3\ HCl(aq) \longrightarrow FeCl_3(aq) + H_2(g)$
Since H and Fe occur as free elements, balance them last. There is 1 Fe atom on the left side of the equation and 1 Fe atom on the right, so Fe is balanced. There are 3 H atoms on the left and 2 H atoms on the right. Balance H by placing a $\frac{3}{2}$ in front of H_2. (That way you don't alter other elements that are already balanced.)	$Fe(s) + 3\ HCl(aq) \longrightarrow FeCl_3(aq) + \frac{3}{2} H_2(g)$
The equation now contains a coefficient fraction; clear it by multiplying the entire equation (both sides) by 2.	$[Fe(s) + 3\ HCl(aq) \longrightarrow FeCl_3(aq) + \frac{3}{2} H_2(g)] \times 2$ $2\ Fe(s) + 6\ HCl(aq) \longrightarrow 2\ FeCl_3(aq) + 3\ H_2(g)$
Finally, sum the number of atoms on each side to check that the equation is balanced.	$2\ Fe(s) + 6\ HCl(aq) \longrightarrow 2\ FeCl_3(aq) + 3\ H_2(g)$

Reactants		Products
2 Fe atoms	⟶	2 Fe atoms
6 Cl atoms	⟶	6 Cl atoms
6 H atoms	⟶	6 H atoms

▶ **SKILLBUILDER 7.5 | Balancing Chemical Equations**

Balance the chemical equation.

$HCl(g) + O_2(g) \longrightarrow H_2O(l) + Cl_2(g)$

▶ **FOR MORE PRACTICE** Problems 45, 46, 47, 48, 49, 50.

CONCEPTUAL CHECKPOINT 7.3

Which quantity must always be the same on both sides of a balanced chemical equation?

(a) the number of each type of atom

(b) the number of each type of molecule

(c) the sum of all of the coefficients

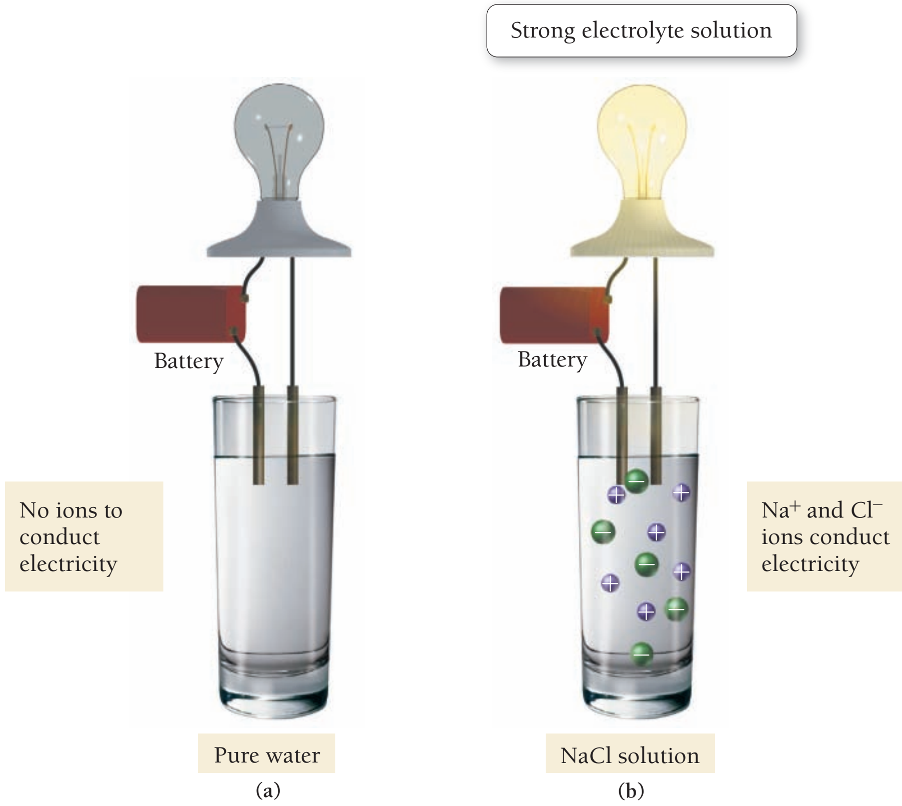
7.5 Aqueous Solutions and Solubility: Compounds Dissolved in Water

▶ Determine whether a compound is soluble.

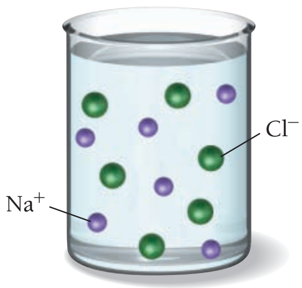
In the previous section, we balanced chemical equations that represent chemical reactions. We now turn to investigating several types of reactions.

Aqueous Solutions

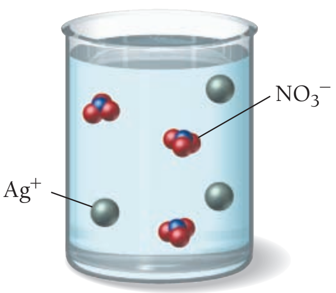
Since many of these reactions occur in water, we must first understand *aqueous solutions*. Reactions occurring in aqueous solutions are among the most common and important. An **aqueous solution** is a homogeneous mixture of a substance with water. For example, a sodium chloride (NaCl) solution (also called a saline solution) is composed of sodium chloride dissolved in water. Sodium chloride



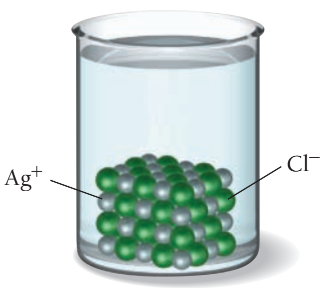
▲ **FIGURE 7.6 Ions as conductors** (a) Pure water does not conduct electricity. (b) Ions in a sodium chloride solution conduct electricity, causing the bulb to light. Solutions such as NaCl are called strong electrolyte solutions.



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



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



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

solutions are common both in the oceans and in living cells. You can form a sodium chloride solution yourself by adding table salt to water. As you stir the salt into the water, it seems to disappear. However, you know the salt is still there because if you taste the water, it has a salty flavor. How does sodium chloride dissolve in water?

When ionic compounds such as NaCl dissolve in water, they usually dissociate into their component ions. A sodium chloride solution, represented as NaCl(*aq*), does not contain any NaCl units; only dissolved Na^+ ions and Cl^- ions are present.

We know that NaCl is present as independent sodium and chloride ions in solution because sodium chloride solutions conduct electricity, which requires the presence of freely moving charged particles. Substances (such as NaCl) that completely dissociate into ions in solution are *strong electrolytes*, and the resultant solutions are **strong electrolyte solutions** (▲ FIGURE 7.6). Similarly, a silver nitrate solution, represented as $AgNO_3(aq)$, does not contain any $AgNO_3$ units, but only dissolved Ag^+ ions and NO_3^- ions. It, too, is a strong electrolyte solution. When compounds containing polyatomic ions such as NO_3^- dissolve, the polyatomic ions dissolve as intact units.

Not all ionic compounds, however, dissolve in water. AgCl, for example, does not. If we add AgCl to water, it remains as solid AgCl and appears as a white solid at the bottom of the beaker.

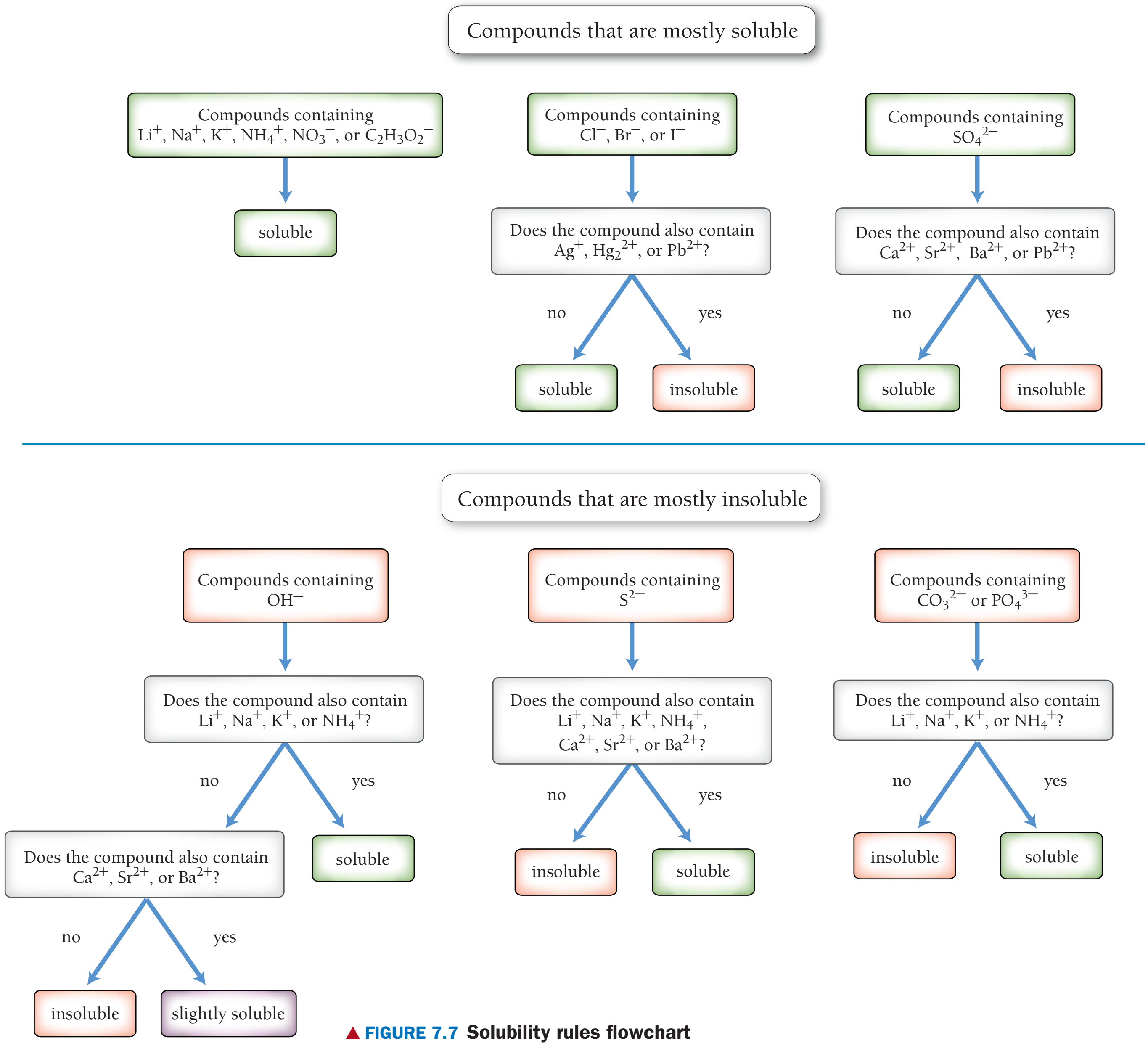
Solubility

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. NaCl, for example, is soluble in water. If we mix solid sodium chloride into water, it dissolves and forms a strong electrolyte solution. AgCl, on the other hand, is insoluble in water. If we mix solid silver chloride into water, it remains as a solid within the liquid water.

There is no easy way to predict whether a particular compound will be soluble or insoluble in water. For ionic compounds, however, empirical rules have been deduced from observations of many compounds. These **solubility rules** are summarized in Table 7.2 and ▼ **FIGURE 7.7**. For example, the solubility rules indicate that compounds containing the lithium ion are *soluble*. That means that compounds such as LiBr, $LiNO_3$, Li_2SO_4, LiOH, and Li_2CO_3 all dissolve in water to form strong electrolyte solutions. If a compound contains Li^+, it is soluble. Similarly, the solubility rules state that compounds containing the NO_3^- ion are soluble. Compounds such as $AgNO_3$, $Pb(NO_3)_2$, $NaNO_3$, $Ca(NO_3)_2$, and $Sr(NO_3)_2$ all dissolve in water to form strong electrolyte solutions.

The solubility rules apply only to the solubility of the compounds in water.

The solubility rules also state that, with some exceptions, compounds containing the CO_3^{2-} ion are *insoluble*. Compounds such as $CuCO_3$, $CaCO_3$, $SrCO_3$, and $FeCO_3$ do not dissolve in water. Note that the solubility rules have 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 all soluble.



▲ **FIGURE 7.7 Solubility rules flowchart**

TABLE 7.2 Solubility Rules

Compounds Containing the Following Ions Are Mostly Soluble	Exceptions
Li^+, Na^+, K^+, NH_4^+	None
NO_3^-, $C_2H_3O_2^-$	None
Cl^-, Br^-, I^-	When any of these ions pair with Ag^+, Hg_2^{2+}, or Pb^{2+}, the compound is insoluble.
SO_4^{2-}	When SO_4^{2-} pairs with Ca^{2+}, Sr^{2+}, Ba^{2+}, or Pb^{2+}, the compound is insoluble.
Compounds Containing the Following Ions Are Mostly Insoluble	**Exceptions**
OH^-, S^{2-}	When either of these ions pairs with Li^+, Na^+, K^+, or NH_4^+, the compound is soluble. When S^{2-} pairs with Ca^{2+}, Sr^{2+}, or Ba^{2+}, the compound is soluble. When OH^- pairs with Ca^{2+}, Sr^{2+}, or Ba^{2+}, the compound is slightly soluble.*
CO_3^{2-}, PO_4^{3-}	When either of these ions pairs with Li^+, Na^+, K^+, or NH_4^+, the compound is soluble.

*For many purposes these can be considered insoluble.

EXAMPLE 7.6 Determining Whether a Compound is Soluble

Is each compound soluble or insoluble?

(a) AgBr **(b)** $CaCl_2$ **(c)** $Pb(NO_3)_2$ **(d)** $PbSO_4$

SOLUTION

(a) Insoluble; compounds containing Br^- are normally soluble, but Ag^+ is an exception.
(b) Soluble; compounds containing Cl^- are normally soluble, and Ca^{2+} is not an exception.
(c) Soluble; compounds containing NO_3^- are always soluble.
(d) Insoluble; compounds containing SO_4^{2-} are normally soluble, but Pb^{2+} is an exception.

▶ SKILLBUILDER 7.6 | Determining Whether a Compound Is Soluble

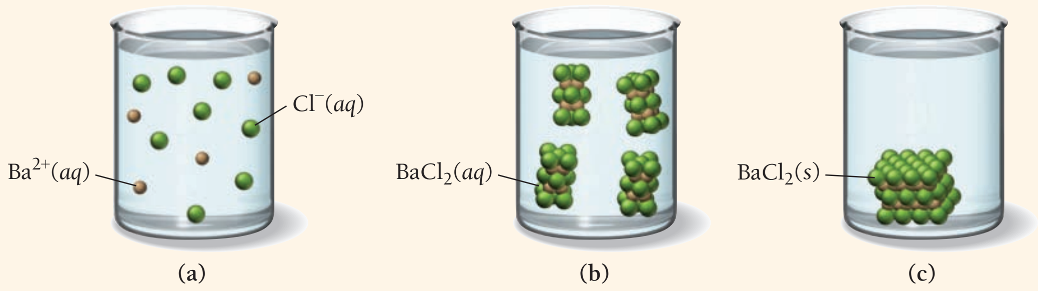
Is each compound soluble or insoluble?

(a) CuS **(b)** $FeSO_4$ **(c)** $PbCO_3$ **(d)** NH_4Cl

▶ FOR MORE PRACTICE Example 7.18; Problems 57, 58, 59, 60, 61, 62.

CONCEPTUAL CHECKPOINT 7.4

Which image best depicts a mixture of $BaCl_2$ and water?



7.6 Precipitation Reactions: Reactions in Aqueous Solution That Form a Solid

▶ Predict and write equations for precipitation reactions.

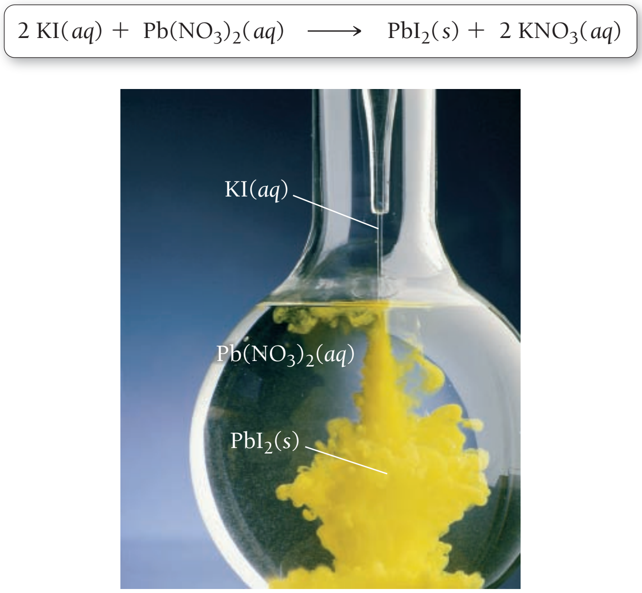
Recall from Section 7.1 that sodium carbonate in laundry detergent reacts with dissolved Mg^{2+} and Ca^{2+} ions to form solids that precipitate (come out of) solution. This reaction is an example of a **precipitation reaction**—a reaction that forms a solid, called a **precipitate**, when two aqueous solutions are mixed.

Precipitation reactions are common in chemistry. Potassium iodide and lead nitrate, for example, both form colorless, strong electrolyte solutions when dissolved in water (see the solubility rules in Section 7.5). When the two solutions are combined, however, a brilliant yellow precipitate forms (▼ FIGURE 7.8). We describe this precipitation reaction with the chemical equation:

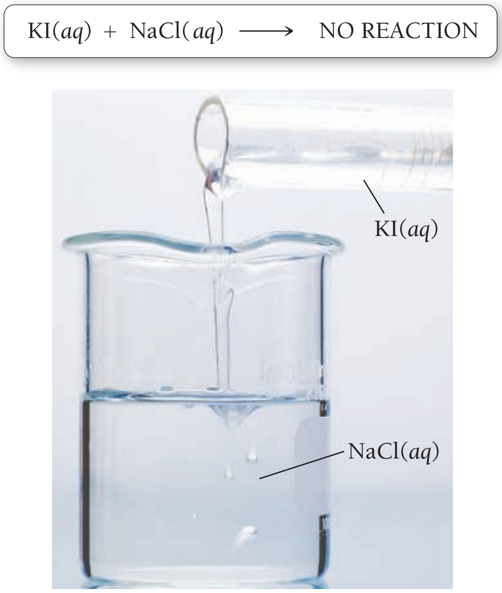
$$2\,KI(aq) + Pb(NO_3)_2(aq) \longrightarrow PbI_2(s) + 2\,KNO_3(aq)$$

Precipitation reactions do not always occur when two aqueous solutions mix. For example, when we combine solutions of KI(*aq*) and NaCl(*aq*), nothing happens (▼ FIGURE 7.9).

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



▲ FIGURE 7.8 **Precipitation** When we mix a potassium iodide solution with a lead(II) nitrate solution, a brilliant yellow precipitate of $PbI_2(s)$ forms.



▲ FIGURE 7.9 **No reaction** When we mix a potassium iodide solution with a sodium chloride solution, no reaction occurs.

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. Consider the precipitation reaction from Figure 7.8:

$$\underset{\text{soluble}}{2\,KI(aq)} + \underset{\text{soluble}}{Pb(NO_3)_2(aq)} \longrightarrow \underset{\text{insoluble}}{PbI_2(s)} + \underset{\text{soluble}}{2\,KNO_3(aq)}$$