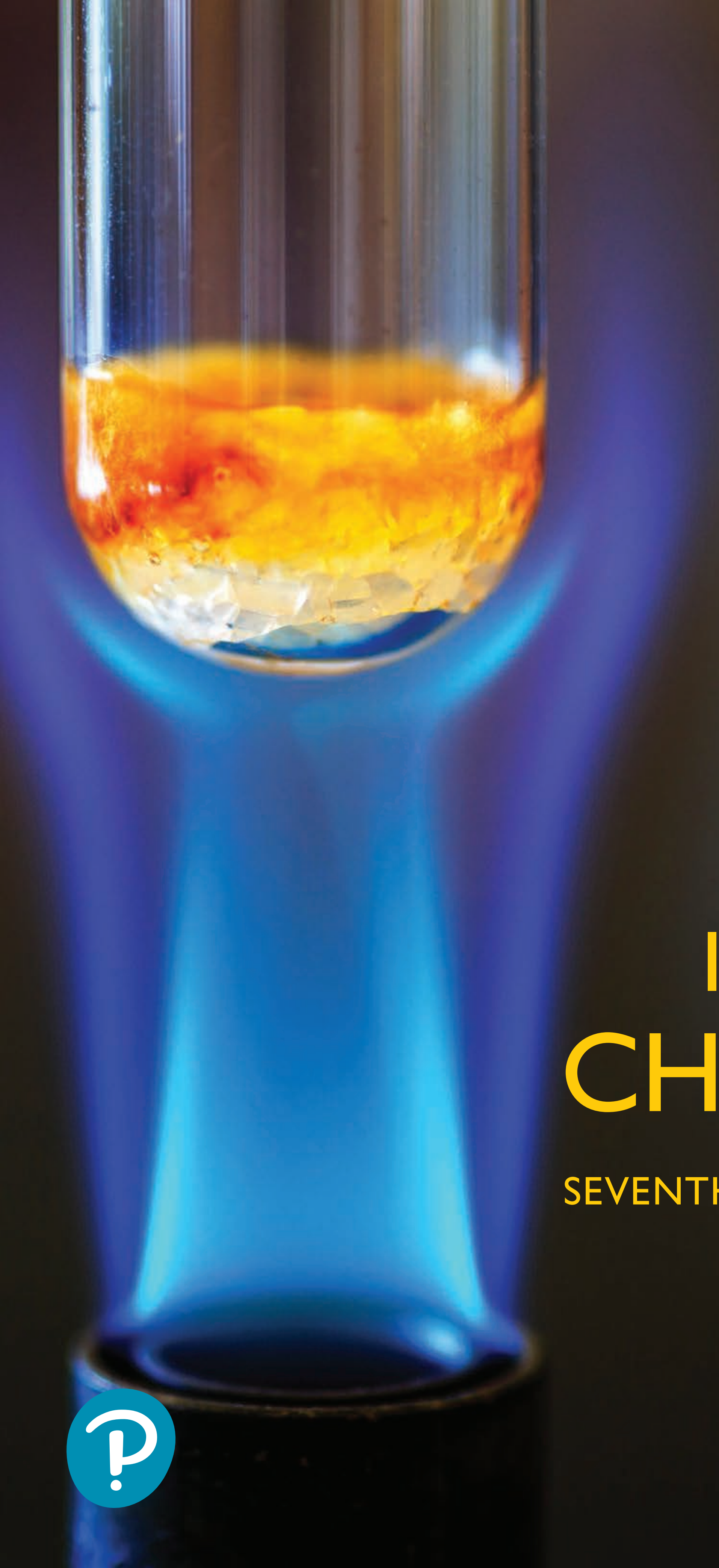


GLOBAL EDITION

Introductory CHEMISTRY

SEVENTH EDITION IN SI UNITS

Nivaldo J. Tro

INTRODUCTORY CHEMISTRY

SEVENTH EDITION IN SI UNITS

Nivaldo J. Tro

Pearson

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.

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

SOLUTION

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.	$Fe(s) + 3\,HCl(aq) \longrightarrow FeCl_3(aq) + H_2(g)$
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, 51, 52.

CONCEPTUAL ✔ CHECKPOINT 7.3

ANSWER NOW!

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.

WATCH NOW! Key Concept Video 7.5

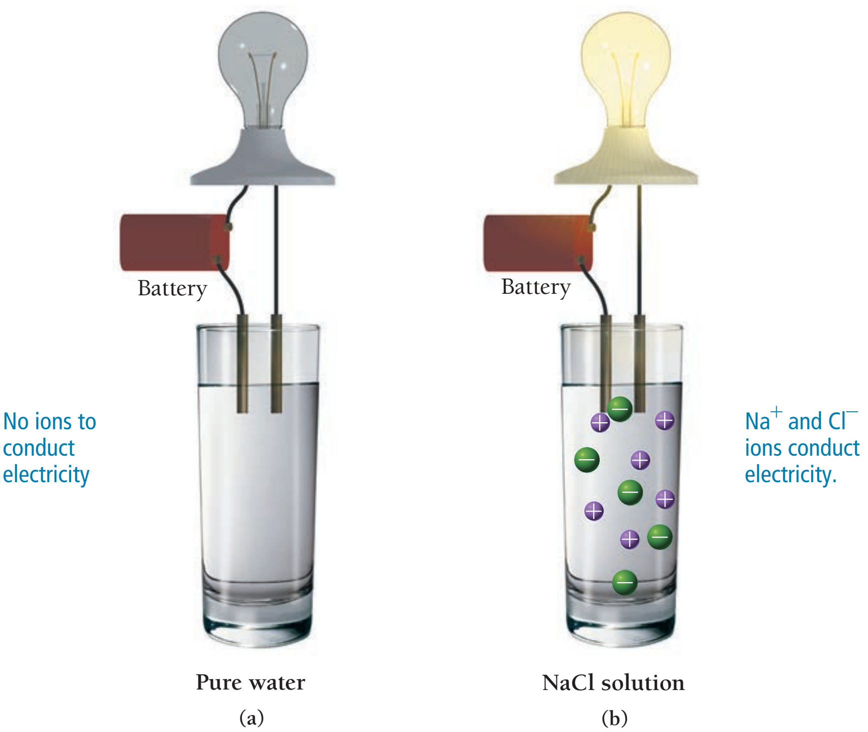
Types of Aqueous Solutions and Solubility

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 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?

Strong Electrolyte Solution



▲ **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.

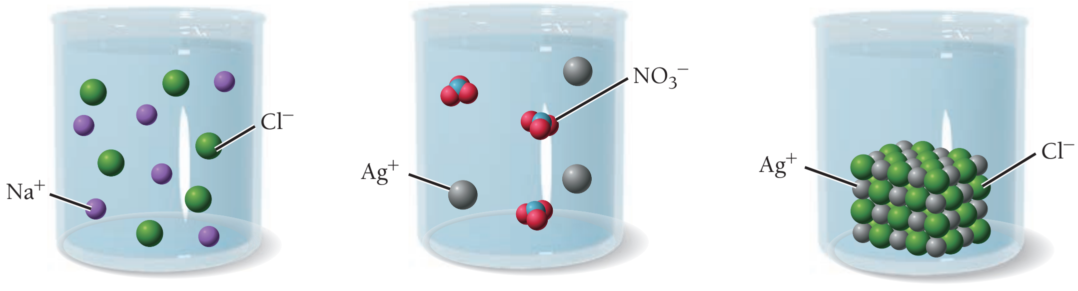
When ionic compounds such as NaCl dissolve in water, they usually dissociate into their component ions. A sodium chloride solution, represented as $NaCl(aq)$, does not contain any NaCl units; only dissolved Na^+ ions and Cl^- ions are present.

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.



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.

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 7.3 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.

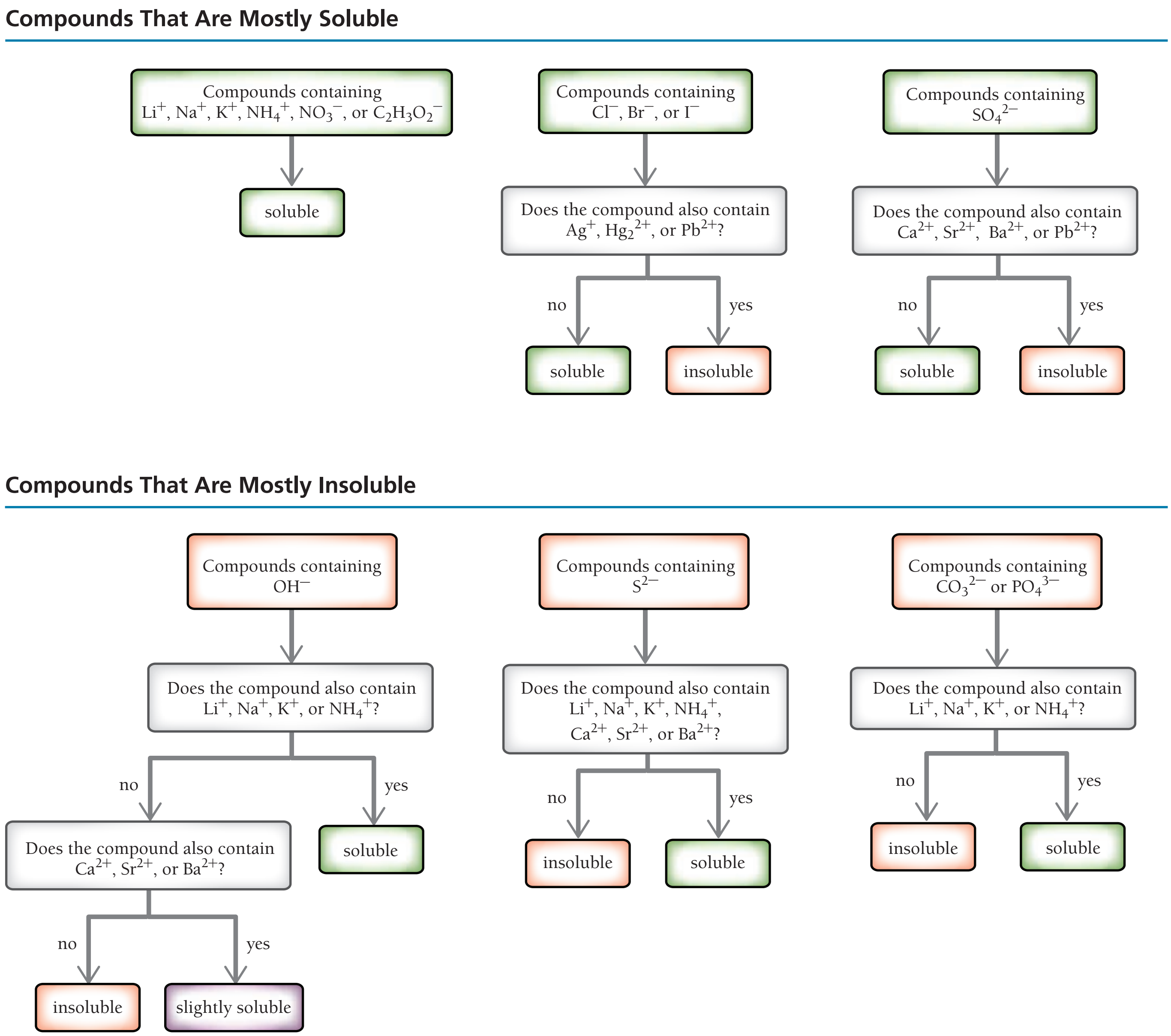
The solubility rules apply only to the solubility of the compounds in water.

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.

TABLE 7.3 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.



▲ **FIGURE 7.7 Solubility rules flowchart**

WATCH NOW! Interactive Worked Example Video 7.6

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?

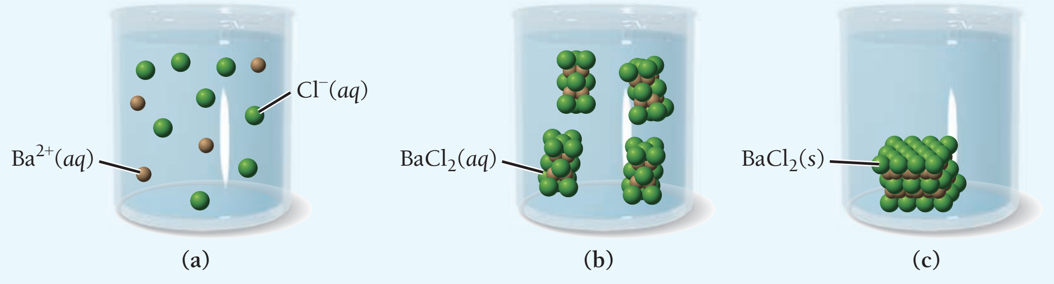
(a) CuS **(b)** $FeSO_4$ **(c)** $PbCO_3$ **(d)** NH_4Cl

▶ **FOR MORE PRACTICE** Example 7.18; Problems 61, 62, 63, 64, 65, 66.

CONCEPTUAL ✔ CHECKPOINT 7.4

ANSWER NOW!

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.

WATCH NOW! Key Concept Video 7.6 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.

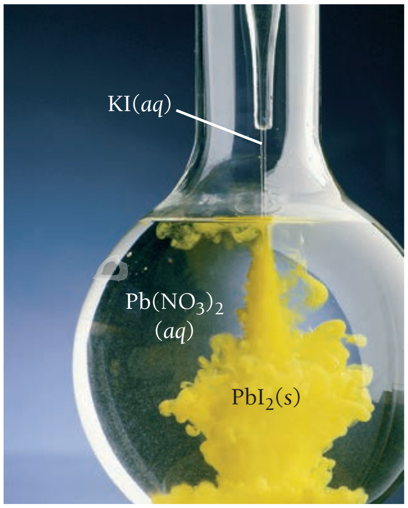
Precipitation reactions are common in chemistry. Potassium iodide and lead nitrate, for example, both form colorless, strong electrolyte solutions when dissolved in water (see the solubility rules in Section 7.5). When the two solutions are combined, however, a brilliant yellow precipitate forms (▼ FIGURE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

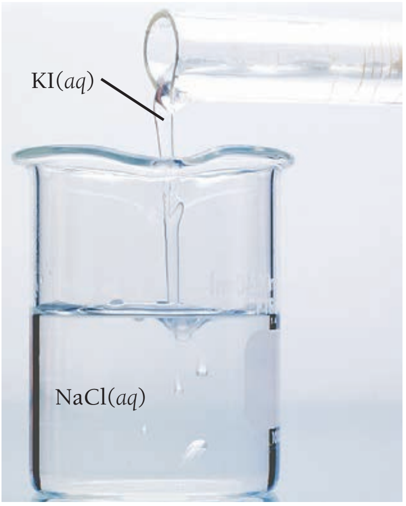
$$KI(aq) + NaCl(aq) \longrightarrow \text{NO REACTION}$$

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



▲ **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.

$KI(aq) + NaCl(aq) \longrightarrow$ NO REACTION

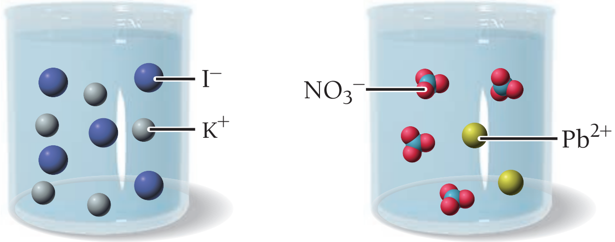


▲ **FIGURE 7.9 No reaction** When we mix a potassium iodide solution with a sodium chloride solution, no reaction occurs.

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:

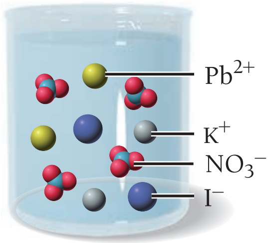
$$\underset{\text{soluble}}{2\,KI(aq)} + \underset{\text{soluble}}{Pb(NO_3)_2(aq)} \longrightarrow \underset{\text{insoluble}}{PbI_2(s)} + \underset{\text{soluble}}{2\,KNO_3(aq)}$$

KI and $Pb(NO_3)_2$ are both soluble, but the precipitate, PbI_2, is *insoluble*. Before mixing, $KI(aq)$ and $Pb(NO_3)_2(aq)$ are each dissociated in their respective solutions.



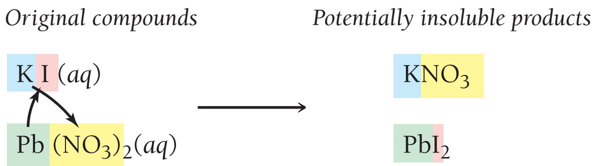
$KI(aq)$ $Pb(NO_3)_2(aq)$

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

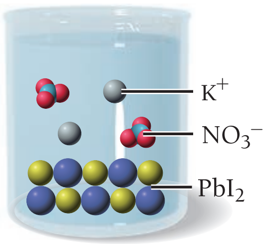


$KI(aq)$ and $Pb(NO_3)_2(aq)$

However, new compounds—potentially insoluble ones—are now possible. Specifically, the cation from one compound can pair with the anion from the other compound to form new (and potentially insoluble) products:



If, on the one hand, the *potentially insoluble* products are both *soluble*, no reaction occurs. If, on the other hand, one or both of the potentially insoluble products are *indeed insoluble*, a precipitation reaction occurs. In this case, KNO_3 is soluble, but PbI_2 is insoluble. Consequently, PbI_2 precipitates.



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

To predict whether a precipitation reaction occurs when two solutions are mixed and to write an equation for the reaction, we follow the steps in the procedure that accompanies Examples 7.7 and 7.8. As usual, the steps are shown in the left column, and two examples of applying the procedure are shown in the center and right columns.