Skills to Develop

- Write and balance chemical equations in molecular, total ionic, and net ionic formats.
- Define precipitation reactions
- Classify chemical reactions as one of these three types given appropriate descriptions or chemical equations
- Predict the solubility of common inorganic compounds by using solubility rules

Humans interact with one another in various and complex ways, and we classify these interactions according to common patterns of behavior. When two humans exchange information, we say they are communicating. When they exchange blows with their fists or feet, we say they are fighting. Faced with a wide range of varied interactions between chemical substances, scientists have likewise found it convenient (or even necessary) to classify chemical interactions by identifying common patterns of reactivity. This module will provide an introduction to three of the most prevalent types of chemical reactions: precipitation, acid-base, and oxidation-reduction.

Equations for Ionic Reactions

Given the abundance of water on earth, it stands to reason that a great many chemical reactions take place in aqueous media. When ions are involved in these reactions, the chemical equations may be written with various levels of detail appropriate to their intended use. To illustrate this, consider a reaction between ionic compounds taking place in an aqueous solution. When aqueous solutions of $\text{CaCl}_2$ and $\text{AgNO}_3$ are mixed, a reaction takes place producing aqueous $\text{Ca(NO}_3)_2$ and solid $\text{AgCl}$:

$$\text{CaCl}_2(\text{aq}) + 2\text{AgNO}_3(\text{aq}) \rightarrow \text{Ca(NO}_3)_2(\text{aq}) + 2\text{AgCl}(\text{s})$$

This balanced equation, derived in the usual fashion, is called a molecular equation because it doesn’t explicitly represent the ionic species that are present in solution. When ionic compounds dissolve in water, they may dissociate into their constituent ions, which are subsequently dispersed homogenously throughout the resulting solution (a thorough discussion of this important process is provided in the chapter on solutions). Ionic compounds dissolved in water are, therefore, more realistically represented as dissociated ions, in this case:

$$\text{CaCl}_2(\text{aq}) \rightarrow \text{Ca}^{2+}(\text{aq}) + 2\text{Cl}^-(\text{aq})$$
$$2\text{AgNO}_3(\text{aq}) \rightarrow 2\text{Ag}^+(\text{aq}) + 2\text{NO}_3^-(\text{aq})$$
$$\text{Ca(NO}_3)_2(\text{aq}) \rightarrow \text{Ca}^{2+}(\text{aq}) + 2\text{NO}_3^-(\text{aq})$$

Unlike these three ionic compounds, AgCl does not dissolve in water to a significant extent, as signified by its physical state notation, (s).

Explicitly representing all dissolved ions results in a complete ionic equation. In this particular case, the formulas for the dissolved ionic compounds are replaced by formulas for their dissociated ions:

$$\text{Ca}^{2+}(\text{aq}) + 2\text{Cl}^-(\text{aq}) + 2\text{Ag}^+(\text{aq}) + 2\text{NO}_3^-(\text{aq}) \rightarrow \text{Ca}^{2+}(\text{aq}) + 2\text{NO}_3^-(\text{aq}) + 2\text{AgCl}(\text{s})$$
Examining this equation shows that two chemical species are present in identical form on both sides of the arrow, \(\ce{Ca^{2+}(aq)}\) and \(\ce{NO3-}(aq)\). These spectator ions—ions whose presence is required to maintain charge neutrality—are neither chemically nor physically changed by the process, and so they may be eliminated from the equation to yield a more succinct representation called a net ionic equation:

\[
\cancel{\ce{Ca^{2+}(aq)}} + \ce{2Cl-}(aq) + \ce{2Ag+}(aq) + \cancel{\ce{2NO3-}(aq)} \rightarrow \cancel{\ce{Ca^{2+}(aq)}} + \cancel{\ce{2NO3-}(aq)} + \ce{2AgCl}(s)
\]

Following the convention of using the smallest possible integers as coefficients, this equation is then written:

\[
\ce{Cl-}(aq) + \ce{Ag+}(aq) \rightarrow \ce{AgCl}(s)
\]

This net ionic equation indicates that solid silver chloride may be produced from dissolved chloride and silver(I) ions, regardless of the source of these ions. These molecular and complete ionic equations provide additional information, namely, the ionic compounds used as sources of \(\ce{Cl^-}\) and \(\ce{Ag^+}\).

Example \(\PageIndex{1}\): Molecular and Ionic Equations

When carbon dioxide is dissolved in an aqueous solution of sodium hydroxide, the mixture reacts to yield aqueous sodium carbonate and liquid water. Write balanced molecular, complete ionic, and net ionic equations for this process.

**Solution**

Begin by identifying formulas for the reactants and products and arranging them properly in chemical equation form:

\[
\ce{CO2(aq) + NaOH(aq) → Na2CO3(aq) + H2O(l)}
\]  
(unbalanced)

Balance is achieved easily in this case by changing the coefficient for NaOH to 2, resulting in the molecular equation for this reaction:

\[
\ce{CO2(aq)+2NaOH(aq)→Na2CO3(aq) + H2O(l)} \nonumber
\]

The two dissolved ionic compounds, NaOH and Na\(_2\)CO\(_3\), can be represented as dissociated ions to yield the complete ionic equation:

\[
\ce{CO2 (aq) + 2Na+ (aq) + 2OH- (aq) → 2Na+ (aq) + CO3^{2-} (aq) + H2O (l)}} \nonumber
\]

Finally, identify the spectator ion(s), in this case Na\(^+\)(aq), and remove it from each side of the equation to generate the net ionic equation:

\[
\begin{align*}
\ce{CO2(aq) + 2OH- (aq) &→ CO3^{2-} (aq) + H2O (l)}
\end{align*}
\]

Exercise \(\PageIndex{1}\)
Diatomic chlorine and sodium hydroxide (lye) are commodity chemicals produced in large quantities, along with diatomic hydrogen, via the electrolysis of brine, according to the following unbalanced equation:

\[
\text{\ce{NaCl(aq) + H2O(l) ->[ electricity] NaOH(aq) + H2(g) + Cl2(g)}}
\]

Write balanced molecular, complete ionic, and net ionic equations for this process.

**Answer**

Balanced molecular equation: \[
\text{\ce{2NaCl (aq) + 2H2O(l) \rightarrow 2NaOH (aq) + H2(g) + Cl2(g)}}
\]

Balanced ionic equation: \[
\text{\ce{2Na^+(aq) + 2Cl^-(aq) + 2H2O(l) \rightarrow 2Na^+(aq) + 2OH^-(aq) + H2(g) + Cl2(g)}}
\]

Balanced net ionic equation: \[
\text{\ce{2Cl^-(aq) + 2H2O(l) \rightarrow 2OH^-(aq) + H2(g) + Cl2 (g)}}
\]

**Precipitation Reactions and Solubility Rules**

A precipitation reaction is one in which dissolved substances react to form one (or more) solid products. Many reactions of this type involve the exchange of ions between ionic compounds in aqueous solution and are sometimes referred to as double displacement, double replacement, or metathesis reactions. These reactions are common in nature and are responsible for the formation of coral reefs in ocean waters and kidney stones in animals. They are used widely in industry for production of a number of commodity and specialty chemicals. Precipitation reactions also play a central role in many chemical analysis techniques, including spot tests used to identify metal ions and gravimetric methods for determining the composition of matter (see the last module of this chapter).

The extent to which a substance may be dissolved in water, or any solvent, is quantitatively expressed as its solubility, defined as the maximum concentration of a substance that can be achieved under specified conditions. Substances with relatively large solubilities are said to be soluble. A substance will precipitate when solution conditions are such that its concentration exceeds its solubility. Substances with relatively low solubilities are said to be insoluble, and these are the substances that readily precipitate from solution. More information on these important concepts is provided in the text chapter on solutions. For purposes of predicting the identities of solids formed by precipitation reactions, one may simply refer to patterns of solubility that have been observed for many ionic compounds (Table \ref{solubility_rules}).

**Table \ref{solubility_rules}: Solubilities of Common Ionic Compounds in Water**

<table>
<thead>
<tr>
<th>Soluble compounds</th>
<th>Exceptions to these solubility rules include</th>
</tr>
</thead>
<tbody>
<tr>
<td>• group 1 metal cations (Li(^+), Na(^+), K(^+), Rb(^+), and Cs(^+)) and ammonium ion (\ce{NH4+})</td>
<td>• halides of Ag(^+), \ce{Hg2^2+}, and Pb(^{2+})</td>
</tr>
<tr>
<td>• the halide ions (Cl(^-), Br(^-), and I(^-))</td>
<td>• sulfates of Ag(^+), Ba(^{2+}), Ca(^{2+}), \ce{Hg2^2+}, Pb(^{2+}), and Sr(^{2+})</td>
</tr>
<tr>
<td>• the acetate (\ce{C2H3O2-}) and bicarbonate (\ce{HCO3-}) ions</td>
<td></td>
</tr>
<tr>
<td>• the nitrate (\ce{NO3-}) ions</td>
<td></td>
</tr>
<tr>
<td>• the sulfate (\ce{SO4-}) ion</td>
<td></td>
</tr>
</tbody>
</table>
Insoluble compounds contain exceptions to these insolubility rules include:

- carbonate (\(\text{CO}_3^{2-}\)), chromate (\(\text{CrO}_4^{2-}\)), phosphate (\(\text{PO}_4^{3-}\)), and sulfide (\(\text{S}^{2-}\)) ions
- hydroxide ion (\(\text{OH}^-\))
- compounds of these anions with group 1 metal cations and ammonium ion
- hydroxides of group 1 metal cations and \(\text{Ba}^{2+}\)

A vivid example of precipitation is observed when solutions of potassium iodide and lead nitrate are mixed, resulting in the formation of solid lead iodide:

\[
\text{K}_2\text{SO}_4(aq) + \text{Pb(NO}_3)_2(aq) \rightarrow \text{PbSO}_4(s) + 2\text{KNO}_3(aq)
\]

This observation is consistent with the solubility guidelines: The only insoluble compound among all those involved is lead iodide, one of the exceptions to the general solubility of iodide salts.

The net ionic equation representing this reaction is:

\[
\text{Pb}^{2+}(aq) + 2\text{I}^-(aq) \rightarrow \text{PbI}_2(s)
\]

Lead iodide is a bright yellow solid that was formerly used as an artist's pigment known as iodine yellow. The properties of pure \(\text{PbI}_2\) crystals make them useful for fabrication of X-ray and gamma ray detectors.
The solubility guidelines in Table may be used to predict whether a precipitation reaction will occur when solutions of soluble ionic compounds are mixed together. One merely needs to identify all the ions present in the solution and then consider if possible cation/anion pairing could result in an insoluble compound. For example, mixing solutions of silver nitrate and sodium fluoride will yield a solution containing Ag\(^{+}\), Na\(^{+}\), and F\(^{-}\) ions. Aside from the two ionic compounds originally present in the solutions, AgNO\(_3\) and NaF, two additional ionic compounds may be derived from this collection of ions: NaNO\(_3\) and AgF. The solubility guidelines indicate all nitrate salts are soluble but that AgF is one of the exceptions to the general solubility of fluoride salts. A precipitation reaction, therefore, is predicted to occur, as described by the following equations:

\[
\ce{NaF}(aq)+\ce{AgNO3}(aq)\rightarrow \ce{AgF}(s)+\ce{NaNO3}(aq)\hspace{20px}\text{(molecular)}
\]

\[
\ce{Ag+}(aq)+\ce{F-}(aq)\rightarrow \ce{AgF}(s)\hspace{20px}\text{(net: ionic)}
\]

**Example (PageIndex{2})**: Predicting Precipitation Reactions

Predict the result of mixing reasonably concentrated solutions of the following ionic compounds. If precipitation is expected, write a balanced net ionic equation for the reaction.

a. potassium sulfate and barium nitrate
b. lithium chloride and silver acetate
c. lead nitrate and ammonium carbonate

**Solution**

(a) The two possible products for this combination are KNO\(_3\) and BaSO\(_4\). The solubility guidelines indicate BaSO\(_4\) is insoluble, and so a precipitation reaction is expected. The net ionic equation for this reaction, derived in the manner detailed in the previous module, is

\[
\ce{Ba^2+}(aq)+\ce{SO4^2-}(aq)\rightarrow \ce{BaSO4}(s)\nonumber
\]

(b) The two possible products for this combination are LiC\(_2\)H\(_3\)O\(_2\) and AgCl. The solubility guidelines indicate AgCl is insoluble, and so a precipitation reaction is expected. The net ionic equation for this reaction, derived in the manner detailed in the previous module, is

\[
\ce{Ag+}(aq)+\ce{Cl-}(aq)\rightarrow \ce{AgCl}(s)\nonumber
\]

(c) The two possible products for this combination are PbCO\(_3\) and NH\(_4\)NO\(_3\). The solubility guidelines indicate PbCO\(_3\) is insoluble, and so a precipitation reaction is expected. The net ionic equation for this reaction, derived in the manner detailed in the previous module, is

\[
\ce{Pb^2+}(aq)+\ce{CO3^2-}(aq)\rightarrow \ce{PbCO3}(s)\nonumber
\]

**Exercise (PageIndex{2})**

Which solution could be used to precipitate the barium ion, Ba\(^{2+}\), in a water sample: sodium chloride, sodium hydroxide, or sodium sulfate? What is the formula for the expected precipitate?
Chemical reactions are classified according to similar patterns of behavior. A large number of important reactions are included in three categories: precipitation, acid-base, and oxidation-reduction (redox). Precipitation reactions involve the formation of one or more insoluble products.

**Glossary**

**combustion reaction**
vigorous redox reaction producing significant amounts of energy in the form of heat and, sometimes, light

**complete ionic equation**
chemical equation in which all dissolved ionic reactants and products, including spectator ions, are explicitly represented by formulas for their dissociated ions

**insoluble**
of relatively low solubility; dissolving only to a slight extent

**molecular equation**
chemical equation in which all reactants and products are represented as neutral substances
net ionic equation
chemical equation in which only those dissolved ionic reactants and products that undergo a chemical or physical change are represented (excludes spectator ions)

precipitate
insoluble product that forms from reaction of soluble reactants

precipitation reaction
reaction that produces one or more insoluble products; when reactants are ionic compounds, sometimes called double-displacement or metathesis

single-displacement reaction
(also, replacement) redox reaction involving the oxidation of an elemental substance by an ionic species

soluble
of relatively high solubility; dissolving to a relatively large extent

solubility
the extent to which a substance may be dissolved in water, or any solvent

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• Adelaide Clark, Oregon Institute of Technology

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