Learning Objectives

- Outline the differences between strong electrolyte, weak electrolyte, and a nonelectrolyte
- Predict the solubility of ionic compounds in water using solubility rules
- Memorize seven strong acids that are strong electrolytes
- Distinguish ways of writing aqueous reaction equations
  - General (molecular) ionic
  - Total (complete) ionic
  - Net Ionic

The objective of this section is to predict what will happen for single and double replacement reactions that occur in aqueous solutions. Note, if in a double displacement reaction two solutions combine and form a solid, you have a precipitation reaction. Many texts treat this as a type of reaction, but we will treat it as a subset of the double displacement reactions.

Electrolytes

Pure water does not conduct electricity, and it has been observed that when a substance dissolves in water, it may produce mobile ions that allow the water to conduct electricity, and we call that compound an electrolyte, or it may not, in which case we call it a nonelectrolyte.

![Electrolyte images](image)

Figure \(\PageIndex{1}\): Ethanol on the left is a nonelectrolyte and does not conduct electricity. KCl is a strong electrolyte and the bulb is very bright. Acetic acid is a weak electrolyte, and although the image may not show it, if the concentrations are the same, the light is dimmer than for the KCl.

Exercise \(\PageIndex{1}\)

Both salt (NaCl) and table sugar, glucose (C₆H₁₂O₆) dissolve in water. Why is salt water a strong electrolyte, while sugar...
water is a non-electrolyte?

**Answer**

NaCl is an ionic compound that dissociates into ions that conduct electricity. $\text{C}_6\text{H}_{12}\text{O}_6$ is a molecular compound that does not break up into ions, and therefore, does not conduct electricity. The sugar molecule remains intact, but each sugar molecule is separated from the other when added to water. The reason it dissolves in water is because of the term "the like dissolves the like", meaning both sugar and water are polar molecules. We will discuss this in more depth later in the text.

There are two basic ways an aqueous compound can be an electrolyte; being a soluble ionic compound or a strong acid.

**Soluble Ionic Compounds**

- If ionic compounds dissolve and form a solution, the ions separate and are free to move about and conduct the electricity. But not all ionic compounds dissolve, and so they can be weak, strong or even nonelectrolytes. Typically, the nondissolved ionic compound forms a solid that falls to the bottom as a precipitate.

Figure \(\PageIndex{2}\): In the above image, the solid KCl is being surrounded by water molecules which cause the ions to leave the crystal and enter the solution. Once they enter the solution they are mobile and can conduct electricity, so KCl is an electrolyte (it is actually a strong electrolyte)
Covalent Compound that React with Water

Just as in the case with ionic compounds, covalent compounds can be weak, strong or nonelectrolytes. There are two types we will look at, acids and a type of bases called the amines.

**Acids**

The second way to produce an electrolyte happens when certain types of covalent molecules that react with the water. Acids give a proton to the water and so form ions as in the image below where HCl reacts with water to form chloride and hydronium ions (figure \(\PageIndex{3}\)).

![Figure \(\PageIndex{3}\): In the above image the gaseous molecule HCl dissolves in water (a), where it then gives a proton to the water and forms the electrolyte with chloride and hydronium ions.](image)

Strong acids are acids that completely react with water to form hydronium and the acid’s anion.

**Note**

You need to memorize 7 strong acids, which are acids that completely react with water to form hydronium ions and the anion of the acid, and so these are strong electrolytes.

1. HCl, HBr, HI (heavier binary halides, HF is weak)
2. H\(_2\)SO\(_4\) (sulfuric acid)
3. HNO\(_3\) (nitric acid)
4. HClO\(_3\), HClO\(_4\).

In this class other acids can be assumed weak unless called strong.
Amine Bases

Some bases are covalent compounds and the amines are an important group. They all have a nitrogen that can extract a proton from the water and form ions, as in the case below of ammonia, which grabs a proton from the water forming the weak electrolyte ammonium hydroxide (figure \(\PageIndex{4}\)).

\[
\text{NH}_3 + \text{H}_2\text{O} \rightleftharpoons \text{NH}_4^+ + \text{OH}^-
\]

Figure \(\PageIndex{4}\): In the above image, ammonia grabs a proton from water forming ammonium hydroxide.

Types of Electrolytes

Compounds can be Strong, Weak, or Nonelectrolytes

1. **Strong Electrolytes** - Strong Conductors of Electricity due to formation of a large number of Mobile Ions
2. **Weak Electrolytes** - Weak Conductors of Electricity due to formation of a few Mobile Ions
3. **NonElectrolytes** – Non Conductors of Electricity as they do not form Ions in aqueous solutions

**Strong Electrolytes**

**Ionic** - Soluble Salts and Strong Bases

NaCl(aq) --> Na\(^+\) (aq) + Cl\(^-\) (aq)

NaOH(aq) --> Na\(^+\) (aq) + OH\(^-\) (aq)

**Covalent** - Strong Acids (protonate water)

HCl(aq) + H\(_2\)O --> H\(_3\)O\(^+\) (aq) + Cl\(^-\) (aq)

H\(_2\)SO\(_4\) (aq) + H\(_2\)O --> H\(_3\)O\(^+\) (aq) + HSO\(_4\)\(^-\) (aq)

**Weak Electrolytes**

**Ionic** - Slightly Soluble Salts

CoCl\(_2\) (s) \(\rightleftharpoons\) Co\(^{2+}\) (aq) + 2Cl\(^-\) (aq)

**Covalent** - Weak Acids & Amine Bases (hydrolyze water)

HF(aq) + H\(_2\)O \(\rightleftharpoons\) H\(_3\)O\(^+\) (aq) + F\(^-\) (aq)

NH\(_3\) (aq) + H\(_2\)O \(\rightleftharpoons\) NH\(_4\)\(^+\) (aq) + OH\(^-\)
NonElectrolytes

Ionic - Insoluble Salts
\[ \text{CoS(aq)} \rightleftharpoons \text{Co}^{+2}(aq) + \text{S}^{-2}(aq) \]

Covalent - Molecules which do not hydrolyze or protonate water
\[ \text{C}_{12}\text{H}_{22}\text{O}_{11}(s) + \text{H}_2\text{O} \rightarrow \text{C}_{12}\text{H}_{22}\text{O}_{11}(aq) \]

The following animations gives an atomic scale visualization of strong, weak and nonelectrolytes

Video \( \PageIndex{1} \): 2'37" YouTube animation uploaded by Georgianna Allen giving an explanation for how some compounds can be weak, strong or non-electrolytes (https://youtu.be/B5mSet29qxA).
Video (PageIndex(2)): 1'29" YouTube animation uploaded by BerkeleyChemDemos showing conductivity of various solutes. Note, the solid ionic compounds do not conduct because their ions are not mobile. You should pause the video before they add the light bulbs and try and predict if the bulb will come on.

Predicting Products for Displacement Reactions

Determine what the products will be for the following reactions. Initially we will ignore phases, but by the time we finish this exercise you will need to be able to identify phases as you write the equation. This process requires that you identify the ions in the reactants and for double displacement reactions swap partners and use the principle of charge neutrality to determine what the product are. For single replacement reaction you need to figure which species (anion or cation) is gaining or losing charge and the other species is a spectator ion. Once that is done you need to balance the equations.

Students frequently swap ions without thinking about the product formula and this leads to mistakes. It is suggested you follow these steps:

1. Identify reactant ions and their charge
2. Swap Ions
3. Determine Product Formula based on principle of charge neutrality
4. Balance Equation
Example:

\[\text{Pb}^{2+} \text{(ClO)}_{2}\text{(aq)} + \text{Na}_{2}\text{SO}_{4}\text{(aq)} \rightarrow \text{PbSO}_{4}\text{(aq)} + 2\text{NaClO}\text{(aq)}\]

1. **Identify reactant ions and their charge**

   \[
   \begin{array}{c|c}
   \text{Pb}^{2+} & \text{ClO}^{-} \\
   \hline
   \text{Na}^{+} & \text{SO}_{4}^{2-} \\
   \end{array}
   \]

   Table \(\PageIndex{1}\): Matrix showing the cations and ions for each of the two ionic compounds in the double displacement reaction

2. **Swap Ions**

   \[
   \begin{array}{c|c}
   \text{Pb}^{2+} & \text{ClO}^{-} \\
   \hline
   \text{Na}^{+} & \text{SO}_{4}^{2-} \\
   \end{array}
   \]

   Table \(\PageIndex{2}\): Matrix of Table \(\PageIndex{1}\) showing the cations and ions swapping partners.

   \[
   \text{Pb}^{2+} + \text{SO}_{4}^{2-} \rightarrow \text{?} \\
   \text{Na}^{+} + \text{ClO}^{-} \rightarrow \text{?}
   \]

   Note: If dealing with a strong acid, treat the "H" as a cation (break the acid \(\text{HA}\) into \(\text{H}^{+}\) and \(\text{A}^{-}\), where \(\text{A}\) is any anion, noting the number of hydrogens equals the charge of the anion, as the acid must be neutral). We will cover acids in the next section.

3. **Determine Product Formula based on principle of charge neutrality**

   \[\text{Pb}^{2+} \text{(ClO)}_{2}\text{(aq)} + \text{Na}_{2}\text{SO}_{4}\text{(aq)} \rightarrow \text{PbSO}_{4}\text{(aq)} + 2\text{NaClO}\text{(aq)}\]

   If you need help determining formulas review section 2.6

4. **Balance Equation**

   \[\text{Pb}^{2+} \text{(ClO)}_{2}\text{(aq)} + \text{Na}_{2}\text{SO}_{4}\text{(aq)} \rightarrow \text{PbSO}_{4}\text{(aq)} + 2\text{NaClO}\text{(aq)}\]
Exercise \(\PageIndex{2}\)

Predict products for the following reactions

a. \(\text{Mg}_3(\text{PO}_4)_2 + \text{NaClO}_3 \rightarrow ?\)

b. \(\text{AlPO}_4 + \text{Na}_2\text{SO}_3 \rightarrow ?\)

**Answer a**

\(\text{Mg}_2(\text{ClO}_3)_2 + \text{Na}_3\text{PO}_4\)

**Answer b**

\(\text{Al}_2(\text{SO}_3)_3 + \text{Na}_3\text{PO}_4\)

The next step is figuring out the phase of the product. There are 3 common phases for aqueous reactions, solid (precipitate), aqueous and gas. In the case of aqueous products resulting from mixing two solutions, we state there is no reaction. In general chemistry 2 we will use mathematical formula of solubility products, but in this class we will use the Solubility rules, which are a good "rule of thumb" to determine if a salt will dissolve or not. Now note, these are all relative values, and even a soluble salt will form a precipitate, if you add a lot of it. For example, you can dissolve around 140 grams of potassium iodide in 100 g of water at 32\(\degree\)F, which makes it a very soluble salt, but a precipitate will occur if you try and add 200 g of KI to 100 g of water.

**Solubility Rules**

We saw that some compounds are soluble, other are partially soluble, and some are completely insoluble. How do we know determine this? We will start off with the simplest types, which are ionic compounds, and we will base these on the nature of the ions. There are sort of two opposing processes going on. Are the attractions of the ions in the crystal stronger than their attraction towards the water, or weaker? If the ionic attractions within the crystal are stronger, they do not dissolve and they form a precipitate. If on the other hand, the ions are more attracted to the water, they leave the crystal and the compound is soluble. We will use the solubility rules to determine if a salt is soluble or not.

**What are the Solubility Rules?**

These are what I am calling a "rule of thumb," and allow us to roughly predict if a salt will dissolve or not. It must be understood that the concept is relative, for example, table salt is considered a soluble salt and if you add table salt to water it will dissolve a lot, up to 359g per liter, but at that point it becomes saturated, and any more will form a precipitate. On the other hand silver chloride is an insoluble salt, and you can only dissolve 0.0019g into 1 liter, but any more will fall to the bottom as a precipitate.
Note, some textbooks give slightly different rules, and this set is incomplete. If your text is different, please discuss this with your instructor. When you get to general chemistry 2 you will learn a different approach, where we can quantify the amount dissolved for an insoluble salt, like the 1.9 mg/liter for silver chloride.

Is there a strategy to using the solubility rules?
Yes, for the ones I have set up below.

- We first look at the [+] cation, and ask if it is any one of the cations listed in step 1A. If yes, we say it is soluble, and the question is answered. After this step we focus on the [-] anions. If it is not soluble from step 1A, we go to 1B, and if the anion is from this list, it is soluble.
- We now go to the compounds that are usually soluble, step II, and you need to memorize the exceptions, which are insoluble.
- We now go to the compounds that are usually insoluble, step III, and you need to memorize the exceptions, which are soluble.

### Solubility Rules

I. Soluble
   a. Group 1A & Ammonium (the only cations in this list)
   b. NO$_3^-$, ClO$_4^-$, ClO$_3^-$, CH$_3$CO$_2^-$

II. Usually Soluble
   a. Cl$^-$, Br$^-$, I$^-$, (Except those with Ag$^+$, Hg$_2$+2, & Pb$^+$2)
   b. F$^-$ (Except Mg$^{+2}$, Ca$^{+2}$, Sr$^{+2}$, Ba$^{+2}$ & Pb$^{+2}$)
   c. SO$_4^{2-}$ (Except those with Ca$^{+2}$, Sr$^{+2}$, Ba$^{+2}$, Ag$^{+}$ & Pb$^{+2}$)

III. Insoluble (Except with cations from 1.a)
   a. OH$^-$ (Except those with Sr$^{+2}$ & Ba$^{+2}$)
   b. Everything else (this is not true, but will work in this class)

Table\PageIndex{3}: Solubility rules with schema showing flow diagram. NOTE: the exceptions in steps II and III have opposite meanings.

In the following video we will first predict the products for the double displacement reaction of aluminum sulfate and barium chloride, and then apply the solubility rules to determine the phases of the products.

\[ \text{Al}_2(\text{SO}_4)_3 + \text{BaCl}_2 \rightarrow? \]
Exercise \(\PageIndex{3}\)

Use the solubility rules to determine if the following compounds are soluble or insoluble. Indicate answer by writing formula followed by (aq) for soluble and (s) for insoluble. (You may also want to write the names of the species)

a. \(\text{PbSO}_4\)

b. \(\text{NaClO}_4\)

**Answer a**

**Insoluble.** Sulfate ions are typically soluble, but lead(II) is an exception, so lead(II) sulfate is an insoluble solid. This means that \(\text{PbSO}_4\) (s) is not broken up into ions in water, but is a precipitate.

**Answer b**

**Soluble.** Perchlorate ions and alkali metals are soluble, so sodium perchlorate is soluble, meaning that it breaks up into ions, \(\text{Na}^+\) (aq) + \(\text{ClO}_4^-\) (aq) in water.
Identify the following as a strong electrolyte, a weak electrolyte, or a non-electrolyte.

a. Ba(NO$_3$)$_2$

b. H$_3$PO$_4$

c. C$_6$H$_{12}$O$_6$

d. HNO$_3$

e. AgBr

**Answer a**

strong electrolyte - Rule 1B

**Answer b**

weak electrolyte - weak acid (not one of the seven strong acids)

**Answer c**

non-electrolyte - molecular compound, not ionic

**Answer d**

strong electrolyte- strong acid (one of the seven strong acids)

**Answer e**

non-electrolyte - insoluble precipitate - Exception to Rule 2A

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**Mixtures of Multiple Solutions**

Consider 3 aqueous salt solutions each with different cations and anions. For example, consider NaCl, AgNO$_3$ and (NH$_4$)$_2$CO$_3$. There are 3 cations and 3 anions which results in $3^2$ or 9 combinations (of which 3 are the reactants). This seems like a very complicated problem (4 salts would result in 16 potential combinations and 5 salts would result in 25) and so we need to develop a technique to see the problem. This can be done through a matrix, where the rows represent the cations, the columns the anions and the cells the potential combinations:

<table>
<thead>
<tr>
<th>Cl$^-$</th>
<th>NO$_3^-$</th>
<th>CO$_3^{2-}$</th>
</tr>
</thead>
</table>

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Setting up a matrix with cation in the column and anions in the row, and if you sequentially fill this, the diagonal is your reactants, and the off diagonal are the possible products.

Now think about it for a minute. If you look at each row it starts with a cation, and if rule 1A says it is soluble, you can knock out the whole row, giving it XXX in Table 5 to indicate no precipitate (all salts sodium and ammonium are solube), then then look at each column and according to rule 1B, you can knock out the middle column because all salts of nitrate are soluble. This reduced your workload to just two cells.

<table>
<thead>
<tr>
<th></th>
<th>Cl(^-)</th>
<th>NO(_3)(^-)</th>
<th>CO(_3)(^{2-})</th>
</tr>
</thead>
<tbody>
<tr>
<td>Na(^+)</td>
<td>XXX</td>
<td>XXX</td>
<td>XXX</td>
</tr>
<tr>
<td>Ag(^+)</td>
<td>AgCl (s)</td>
<td>XXX</td>
<td>Ag(_2)CO(_3)((s))</td>
</tr>
<tr>
<td>NH(_4)(^+)</td>
<td>XXX</td>
<td>XXX</td>
<td>XXX</td>
</tr>
</tbody>
</table>

Using the matric method to reduce the workload when calculating potential precipitates for mixtures of multiple salts. Note, all sodium, nitrate and ammonium salts are soluble and so we can ignore them in identifying potential precipitates, so the answer is AgCl(s) and Ag\(_2\)CO\(_3\)(s).

The following video shows how to solve these problems in the shortest amount of time for the following reaction:

\[
\text{NaCl} + \text{Na}_3\text{PO}_4 + \text{AgNO}_3 + \text{Pb(NO}_3\text{)}_2 + \text{Na}_2\text{SO}_4
\]
Video shader Video shader: 4’31” YouTube on calculating any precipitates formed when 5 soluble salts are mixed (https://youtu.be/hdO0hp6cDUk)

Exercise shader Exercise shader

Identify if any precipitates that would form if the following solutions are mixed:

a. Pb(ClO$_4$)$_2$(aq), K$_2$SO$_4$(aq), AgCH$_3$CO$_2$ (aq) + KCl(aq)

b. Ba(ClO$_3$)$_2$(aq), Li$_2$SO$_4$(aq), NH$_4$CH$_3$CO$_2$ (aq) + BaCl$_2$

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**Answer a**

<table>
<thead>
<tr>
<th></th>
<th>ClO$_4^-$</th>
<th>SO$_4^-$</th>
<th>CH$_3$CO$_2^-$</th>
<th>Cl$^-$</th>
</tr>
</thead>
<tbody>
<tr>
<td>Pb$^{+2}$</td>
<td>XXX</td>
<td>PbSO$_4$ (s)</td>
<td>XXX</td>
<td>PbCl$_2$ (s)</td>
</tr>
<tr>
<td>K$^+$</td>
<td>XXX</td>
<td>XXX</td>
<td>XXX</td>
<td>XXX</td>
</tr>
<tr>
<td>Ag$^+$</td>
<td>XXX</td>
<td>Ag$_2$SO$_4$ (aq)</td>
<td>XXX</td>
<td>AgCl (s)</td>
</tr>
<tr>
<td>K$^+$</td>
<td>XXX</td>
<td>XXX</td>
<td>XXX</td>
<td>XXX</td>
</tr>
</tbody>
</table>
So the answer is PbSO\(_4\) (s), PbCl\(_2\) (s) and AgCl (s), everything else is aqueous.

**Answer b**

d

<table>
<thead>
<tr>
<th></th>
<th>ClO(_4)^-</th>
<th>SO(_4^{2-})</th>
<th>CH(_3)CO(_2)^-</th>
<th>Cl(^-)</th>
</tr>
</thead>
<tbody>
<tr>
<td>Ba(^{+2})</td>
<td>XXX</td>
<td>BaSO(_4) (s)</td>
<td>XXX</td>
<td>BaCl(_2) (aq)</td>
</tr>
<tr>
<td>Li(^+)</td>
<td>XXX</td>
<td>XXX</td>
<td>XXX</td>
<td>XXX</td>
</tr>
<tr>
<td>NH(_4)^+</td>
<td>XXX</td>
<td>XXX</td>
<td>XXX</td>
<td>XXX</td>
</tr>
<tr>
<td>Ba(^{+2})</td>
<td>XXX</td>
<td>BaSO(_4) (s)</td>
<td>XXX</td>
<td>BaCl(_2) (aq)</td>
</tr>
</tbody>
</table>

So the answer is BaSO\(_4\), everything else is aqueous.

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### Describing Aqueous Equations

There are three basic ways you will need to know that are used to describe aqueous chemical reactions. These are typically single and double displacement reactions. The difference between these lies in how you describe the ions of dissolved ionic compounds. Do you describe them as the separate ions, which are really what is floating around, or do you describe them by their neutral compound formula? That is, you can represent the solution as either the salt(aq) or its ions

\[NaCl(aq) \text{ means } Na^+(aq) + Cl^-(aq)\]

The General (Molecular) Equation uses the formula of the salt, the total ionic equations separates soluble salts (and strong acids) into their ions, while the net ionic equation cancels any ions that appear on both the reactant and product sides of the balanced equation, the "spectator ions", as they do not participate in the reaction. The net ionic equation typically describes the actual chemistry of the reaction, and what is going on.

1. **General (Molecular) Equation**
2. **Total Ionic Equation**
3. **Net Ionic Equation**

In solving these problems remember, (aq) = aqueous, (s) = solid, (g) = gas and (l) = liquid, and be sure the final equation is balanced.
**General (Molecular) Equation**

The General or Molecular Equation describes the reactions in terms of the formula of the salts, indicating the phase. These are often (incorrectly) called "molecular equations" because they use the formula of the neutral compound as if it was a molecule. The following equation is an example of a general double displacement reaction, and given the reactants, you need to be able to predict product compounds, balance the equation and predict the phases.

\[ \text{Pb} \left( \text{ClO}_4 \right)_2 \left( \text{aq} \right) + \text{Na}_2\text{SO}_4 \left( \text{aq} \right) \rightarrow \text{PbSO}_4 \left( \text{s} \right) + 2\text{NaClO}_4 \left( \text{aq} \right) \]

As indicated above, the aqueous ions separate on dissolution, and so we could write them out as separate entities, which is the Total or Complete Ionic Equation.

**Total Ionic Equation:**

The Total or complete ionic equation writes soluble ionic compounds in terms of their ions. So if it is an ionic compound with (aq), you break it into its ions. If it is covalent, you do not break it into ions. If it is an acid, you only break it into its ions if it is strong, and that will be covered in the next section, but weak acids that are dissolved, like HF(aq) do not break into ions, while strong ones like HCl do.

Describing eq. 3.4.6 in terms of a net ionic equation gives:

\[ \text{Pb}^{+2} \left( \text{aq} \right) + 2\text{ClO}_4^{-} \left( \text{aq} \right) + 2\text{Na}^{+} \left( \text{aq} \right) + \text{SO}_4^{2-} \left( \text{aq} \right) \rightarrow \text{PbSO}_4 \left( \text{s} \right) + 2\text{Na}^{+} \left( \text{aq} \right) + 2\text{ClO}_4^{-} \left( \text{aq} \right) \]

This is often called a precipitate reaction as the lead (II) sulfate crashes to the bottom of the container, and some textbooks classify precipitation reactions as a class of reaction.

**Net Ionic Equation**

The net ionic equation tells you what is happening, that is, what bonds (covalent or ionic) are being created. If you look at equation 3.4.7 you see there are two perchlorate and two sodium ions on both sides of the equation, and so we call them spectator ions because they did not react. If we cancel them the equation is still balanced and this gives the net ionic equation, which in this reaction results in the formation of the ionic compound, lead (II) sulfate.

\[ \text{Pb}^{+2} \left( \text{aq} \right) + \text{SO}_4^{2-} \left( \text{aq} \right) \rightarrow \text{PbSO}_4 \left( \text{s} \right) \]

Example \(\langle\text{PageIndex}(1)\rangle\)
Write the general, complete and net ionic equations for the reaction of barium chloride and potassium sulfate

Solution

\[ \text{General (Molecular Equation)} \]

\[ \text{Complete Ionic Equation} \]

\[ \text{Net Ionic Equation} \]

Exercise PageIndex(1)

Write balanced net ionic equations to describe any reaction which occur when the following solutions below are mixed.

a. \((Na_2CO_3 + Sr(NO_3)_2)\)

b. \((Cu_2SO_4 + BaCl_2)\)

Answer a

\[ \text{Sr}^{2+}(aq) + \text{CO}_3^{-2} (aq) \rightarrow \text{SrCO}_3 (s) \]

Answer b

\[ \text{Ba}^{2+}(aq) + \text{SO}_4^{2-} (aq) \rightarrow \text{Ba(SO}_4 (s) \]

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- November Palmer, Ronia Kattoum & Emily Chaote (UALR)