Skills to Develop

- To identify a precipitation reaction and predict solubilities.

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**Exchange (Double-Displacement) Reactions**

A precipitation reaction is a reaction that yields an insoluble product—a precipitate—when two solutions are mixed. We described a precipitation reaction in which a colorless solution of silver nitrate was mixed with a yellow-orange solution of potassium dichromate to give a reddish precipitate of silver dichromate:

\[
\text{AgNO}_3(\text{aq}) + \text{K}_2\text{Cr}_2\text{O}_7(\text{aq}) \rightarrow \text{Ag}_2\text{Cr}_2\text{O}_7(\text{s}) + \text{KNO}_3(\text{aq}) \label{4.2.1}
\]

This unbalanced equation has the general form of an exchange reaction:

\[
\underbrace{\text{AC}}_{\text{soluble}} + \overbrace{\text{BD}}^{\text{soluble}} \rightarrow \underbrace{\text{AD}}_{\text{insoluble}} + \overbrace{\text{BC}}^{\text{soluble}} \label{4.2.2}
\]

The solubility and insoluble annotations are specific to the reaction in Equation \ref{4.2.1} and not characteristic of all exchange reactions (e.g., both products can be soluble or insoluble). *Precipitation reactions* are a subclass of exchange reactions that occur between ionic compounds when one of the products is insoluble. Because both components of each compound change partners, such reactions are sometimes called *double-displacement reactions*. Two important uses of precipitation reactions are to isolate metals that have been extracted from their ores and to recover precious metals for recycling.
Video \[\text{PageIndex{1}}\]: Mixing Potassium Chromate and Silver Nitrate together to initiate a precipitation reaction \((Equation \ \ref{4.2.1})\).

While full chemical equations show the identities of the reactants and the products and give the stoichiometries of the reactions, they are less effective at describing what is actually occurring in solution. In contrast, equations that show only the hydrated species focus our attention on the chemistry that is taking place and allow us to see similarities between reactions that might not otherwise be apparent.

Let’s consider the reaction of silver nitrate with potassium dichromate above. When aqueous solutions of silver nitrate and potassium dichromate are mixed, silver dichromate forms as a red solid. The overall balanced chemical equation for the reaction shows each reactant and product as undissociated, electrically neutral compounds:

\[
\ce{2AgNO_3(aq) + K_2Cr_2O_7(aq) \rightarrow Ag_2Cr_2O_7(s) + 2KNO_3(aq)} \label{4.2.1a}
\]

Although Equation \(\ref{4.2.1a}\) gives the identity of the reactants and the products, it does not show the identities of the actual species in solution. Because ionic substances such as \(\ce{AgNO3}\) and \(\ce{K2Cr2O7}\) are strong electrolytes (i.e., they dissociate completely in aqueous solution to form ions). In contrast, because \(\ce{Ag2Cr2O7}\) is not very soluble, it separates from the solution as a solid. To find out what is actually occurring in solution, it is more informative to write the reaction as a complete ionic equation showing which ions and molecules are hydrated and which are present in other forms and phases:

\[
\ce{2Ag^+(aq) + 2NO_3^-(aq) + 2K^+(aq) + Cr_2O_7^{2-}(aq) \rightarrow Ag_2Cr_2O_7(s) + 2K^+(aq) + 2NO_3^-(aq)} \label{4.2.2a}
\]
Note that \(\ce{K^+ (aq)}\) and \(\ce{NO3^{−} (aq)}\) ions are present on both sides of Equation \(\ref{4.2.2a}\) and their coefficients are the same on both sides. These ions are called spectator ions because they do not participate in the actual reaction. Canceling the spectator ions gives the net ionic equation, which shows only those species that participate in the chemical reaction:

\[
2\ce{Ag^+(aq)} + \ce{Cr_2O_7^{2-}(aq)} \rightarrow \ce{Ag_2Cr_2O_7(s)} \tag{4.2.3}
\]

Both mass and charge must be conserved in chemical reactions because the numbers of electrons and protons do not change. For charge to be conserved, the sum of the charges of the ions multiplied by their coefficients must be the same on both sides of the equation. In Equation \(\ref{4.2.3}\), the charge on the left side is \(2(+1) + 1(−2) = 0\), which is the same as the charge of a neutral \(\ce{Ag2Cr2O7}\) formula unit on the right side.

By eliminating the spectator ions, we can focus on the chemistry that takes place in a solution. For example, the overall chemical equation for the reaction between silver fluoride and ammonium dichromate is as follows:

\[
2\ce{AgF(aq)} + (\ce{NH_4})_2\ce{Cr_2O_7(aq)} \rightarrow \ce{Ag_2Cr_2O_7(s)} + 2\ce{NH_4F(aq)} \tag{4.2.4}
\]

The complete ionic equation for this reaction is as follows:

\[
\ce{2Ag^+(aq)} + \cancel{\ce{2F^-(aq)}} + \cancel{\ce{2NH_4^+(aq)}} + \ce{Cr_2O_7^{2-}(aq)} \rightarrow \ce{Ag_2Cr_2O_7(s)} + \cancel{\ce{2NH_4^+(aq)}} + \cancel{\ce{2F^-(aq)}} \tag{4.2.5}
\]

Because two \(\ce{NH4^+(aq)}\) and two \(\ce{F^-(aq)}\) ions appear on both sides of Equation \(\ref{4.2.5}\), they are spectator ions. They can therefore be canceled to give the net ionic equation (Equation \(\ref{4.2.6}\)), which is identical to Equation \(\ref{4.2.3}\):

\[
\ce{2Ag^+(aq) + Cr_2O_7^{2-}(aq) \rightarrow Ag_2Cr_2O_7(s)} \tag{4.2.6}
\]

If we look at net ionic equations, it becomes apparent that many different combinations of reactants can result in the same net chemical reaction. For example, we can predict that silver fluoride could be replaced by silver nitrate in the preceding reaction without affecting the outcome of the reaction.

Example \(\PageIndex{1}\): Balancing Precipitation Equations

Write the overall chemical equation, the complete ionic equation, and the net ionic equation for the reaction of aqueous barium nitrate with aqueous sodium phosphate to give solid barium phosphate and a solution of sodium nitrate.

**Given:** reactants and products

**Asked for:** overall, complete ionic, and net ionic equations

**Strategy:**

Write and balance the overall chemical equation. Write all the soluble reactants and products in their dissociated form to give the complete ionic equation; then cancel species that appear on both sides of the complete ionic equation to give the net ionic equation.
Solution:

From the information given, we can write the unbalanced chemical equation for the reaction:

\[
\ce{Ba(NO_3)_2(aq) + Na_3PO_4(aq) \rightarrow Ba_3(PO_4)_2(s) + NaNO_3(aq)}
\]

Because the product is \(\text{Ba}_3\text{(PO}_4\text{)}_2\), which contains three \(\text{Ba}^{2+}\) ions and two \(\text{PO}_4^{3-}\) ions per formula unit, we can balance the equation by inspection:

\[
\ce{3Ba(NO_3)_2(aq) + 2Na_3PO_4(aq) \rightarrow Ba_3(PO_4)_2(s) + 6NaNO_3(aq)}
\]

This is the overall balanced chemical equation for the reaction, showing the reactants and products in their undissociated form. To obtain the complete ionic equation, we write each soluble reactant and product in dissociated form:

\[
\ce{3\text{Ba}^{2+}(aq) + 6\text{NO}_3^-(aq) + 6\text{Na}^+(aq) + 2\text{PO}_4^{3-}(aq) \rightarrow \text{Ba}_3(\text{PO}_4)_2(s) + 6\text{Na}^+(aq) + 6\text{NO}_3^-(aq)}
\]

The six \(\text{NO}_3^-\)(aq) ions and the six \(\text{Na}^+(aq)\) ions that appear on both sides of the equation are spectator ions that can be canceled to give the net ionic equation:

\[
\ce{3\text{Ba}^{2+}(aq) + 2\text{PO}_4^{3-}(aq) \rightarrow \text{Ba}_3(\text{PO}_4)_2(s)}
\]

Exercise \(\PageIndex{1}\): Mixing Silver Fluoride with Sodium Phosphate

Write the overall chemical equation, the complete ionic equation, and the net ionic equation for the reaction of aqueous silver fluoride with aqueous sodium phosphate to give solid silver phosphate and a solution of sodium fluoride.

**Answer:**

overall chemical equation:

\[
\ce{3\text{AgF}(aq) + Na_3PO_4(aq) \rightarrow Ag_3PO_4(s) + 3\text{NaF}(aq)}
\]

complete ionic equation:

\[
\ce{3\text{Ag}^+(aq) + 3\text{F}^-(aq) + 3\text{Na}^+(aq) + \text{PO}_4^{3-}(aq) \rightarrow Ag_3PO_4(s) + 3\text{Na}^+(aq) + 3\text{F}^-(aq)}
\]

net ionic equation:

\[
\ce{3\text{Ag}^+(aq) + \text{PO}_4^{3-}(aq) \rightarrow Ag_3PO_4(s)}
\]

So far, we have always indicated whether a reaction will occur when solutions are mixed and, if so, what products will form. As you advance in chemistry, however, you will need to predict the results of mixing solutions of compounds, anticipate what kind of reaction (if any) will occur, and predict the identities of the products. Students tend to think that this means they are supposed to "just know" what will happen when two substances are mixed. Nothing could be further from the truth: an infinite number of chemical reactions is possible, and neither you nor anyone else could possibly memorize
them all. Instead, you must begin by identifying the various reactions that could occur and then assessing which is the most probable (or least improbable) outcome.

The most important step in analyzing an unknown reaction is to write down all the species—that are actually present in the solution (not forgetting the solvent itself) so that you can assess which species are most likely to react with one another. The easiest way to make that kind of prediction is to attempt to place the reaction into one of several familiar classifications, refinements of the five general kinds of reactions (acid–base, exchange, condensation, cleavage, and oxidation–reduction reactions). In the sections that follow, we discuss three of the most important kinds of reactions that occur in aqueous solutions: precipitation reactions (also known as exchange reactions), acid–base reactions, and oxidation–reduction reactions.

### Predicting Solubilities

Table \(\PageIndex{1}\) gives guidelines for predicting the solubility of a wide variety of ionic compounds. To determine whether a precipitation reaction will occur, we identify each species in the solution and then refer to Table \(\PageIndex{1}\) to see which, if any, combination(s) of cation and anion are likely to produce an insoluble salt. In doing so, it is important to recognize that soluble and insoluble are relative terms that span a wide range of actual solubilities. We will discuss solubilities in more detail later, where you will learn that very small amounts of the constituent ions remain in solution even after precipitation of an “insoluble” salt. For our purposes, however, we will assume that precipitation of an insoluble salt is complete.

**Table \(\PageIndex{1}\): Guidelines for Predicting the Solubility of Ionic Compounds in Water**

<table>
<thead>
<tr>
<th>Soluble</th>
<th>Exceptions</th>
</tr>
</thead>
<tbody>
<tr>
<td>Rule 1</td>
<td>most salts that contain an alkali metal (Li(^+), Na(^+), K(^+), Rb(^+), and Cs(^+)) and ammonium (NH(_4)^+)</td>
</tr>
<tr>
<td>Rule 2</td>
<td>most salts that contain the nitrate (NO(_3^-)) anion</td>
</tr>
<tr>
<td>Rule 3</td>
<td>most salts of anions derived from monocarboxylic acids (e.g., CH(_3)CO(_2)^-)</td>
</tr>
<tr>
<td>Rule 4</td>
<td>most chloride, bromide, and iodide salts</td>
</tr>
<tr>
<td></td>
<td>Insoluble</td>
</tr>
<tr>
<td>Rule 5</td>
<td>most salts that contain the hydroxide (OH(^-)) and sulfide (S(^2-)) anions</td>
</tr>
<tr>
<td>Rule 6</td>
<td>most carbonate (CO(_3^{2-})) and phosphate (PO(_4^{3-})) salts</td>
</tr>
</tbody>
</table>
Soluble

Rule 7

most sulfate (SO$_4^{2-}$) salts that contain main group cations with a charge ≥ +2

but not salts of +1 cations, Mg$^{2+}$, and dipositive transition metal cations (e.g., Ni$^{2+}$)

Exceptions

Just as important as predicting the product of a reaction is knowing when a chemical reaction will not occur. Simply mixing solutions of two different chemical substances does not guarantee that a reaction will take place. For example, if 500 mL of a 1.0 M aqueous NaCl solution is mixed with 500 mL of a 1.0 M aqueous KBr solution, the final solution has a volume of 1.00 L and contains 0.50 M Na$^+$ (aq), 0.50 M Cl$^-$ (aq), 0.50 M K$^+$ (aq), and 0.50 M Br$^-$ (aq). As you will see in the following sections, none of these species reacts with any of the others. When these solutions are mixed, the only effect is to dilute each solution with the other (Figure 1).

Example 2

Using the information in Table 1, predict what will happen in each case involving strong electrolytes. Write the net ionic equation for any reaction that occurs.

a. Aqueous solutions of barium chloride and lithium sulfate are mixed.

b. Aqueous solutions of rubidium hydroxide and cobalt(II) chloride are mixed.

c. Aqueous solutions of strontium bromide and aluminum nitrate are mixed.

d. Solid lead(II) acetate is added to an aqueous solution of ammonium iodide.

Given: reactants

Asked for: reaction and net ionic equation

Strategy:

A. Identify the ions present in solution and write the products of each possible exchange reaction.
B. Refer to Table \(\PageIndex{1}\) to determine which, if any, of the products is insoluble and will therefore form a precipitate. If a precipitate forms, write the net ionic equation for the reaction.

**Solution:**

**A** Because barium chloride and lithium sulfate are strong electrolytes, each dissociates completely in water to give a solution that contains the constituent anions and cations. Mixing the two solutions *initially* gives an aqueous solution that contains \(\text{Ba}^{2+}, \text{Cl}^-, \text{Li}^+, \text{and SO}_4^{2-}\) ions. The only possible exchange reaction is to form LiCl and BaSO₄:

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

**B** We now need to decide whether either of these products is insoluble. Table \(\PageIndex{1}\) shows that LiCl is soluble in water (rules 1 and 4), but BaSO₄ is not soluble in water (rule 5). Thus BaSO₄ will precipitate according to the net ionic equation:

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

Although soluble barium salts are toxic, BaSO₄ is so insoluble that it can be used to diagnose stomach and intestinal problems without being absorbed into tissues. An outline of the digestive organs appears on x-rays of patients who have been given a “barium milkshake” or a “barium enema”—a suspension of very fine BaSO₄ particles in water.

An x-ray of the digestive organs of a patient who has swallowed a “barium milkshake.” A barium milkshake is a suspension of very fine BaSO₄ particles in water; the high atomic mass of barium makes it opaque to x-rays. Image used with permission from Wikipedia.
1. A Rubidium hydroxide and cobalt(II) chloride are strong electrolytes, so when aqueous solutions of these compounds are mixed, the resulting solution initially contains Rb\(^+\), OH\(^-\), Co\(^{2+}\), and Cl\(^-\) ions. The possible products of an exchange reaction are rubidium chloride and cobalt(II) hydroxide:

\[
\text{Rb}^+(aq) + \text{OH}^-(aq) + \text{Co}^{2+}(aq) + 2\text{Cl}^-(aq)
\]

B According to Table \(\PageIndex{1}\), RbCl is soluble (rules 1 and 4), but Co(OH)\(_2\) is not soluble (rule 5). Hence Co(OH)\(_2\) will precipitate according to the following net ionic equation:

\[
\text{(Co}^{2+}(aq) + 2\text{OH}^-(aq) \rightarrow \text{Co(OH)}_2(s))
\]

2. A When aqueous solutions of strontium bromide and aluminum nitrate are mixed, we initially obtain a solution that contains Sr\(^{2+}\), Br\(^-\), Al\(^{3+}\), and NO\(_3^-\) ions. The two possible products from an exchange reaction are aluminum bromide and strontium nitrate:

\[
\text{Sr}^{2+}(aq) + 2\text{Br}^-(aq) + \text{Al}^{3+}(aq) + 3\text{NO}_3^-(aq)
\]

B According to Table \(\PageIndex{1}\), both AlBr\(_3\) (rule 4) and Sr(NO\(_3\))\(_2\) (rule 2) are soluble. Thus no net reaction will occur.

1. A According to Table \(\PageIndex{1}\), lead acetate is soluble (rule 3). Thus solid lead acetate dissolves in water to give Pb\(^{2+}\) and CH\(_3\)CO\(_2^-\) ions. Because the solution also contains NH\(_4^+\) and I\(^-\) ions, the possible products of an exchange reaction are ammonium acetate and lead(II) iodide:

\[
\text{NH}_4^+(aq) + \text{I}^-(aq) + \text{Pb}^{2+}(aq) + 2\text{CH}_3\text{CO}_2^-(aq)
\]

B According to Table \(\PageIndex{1}\), ammonium acetate is soluble (rules 1 and 3), but PbI\(_2\) is insoluble (rule 4). Thus Pb(C\(_2\)H\(_3\)O\(_2\))\(_2\) will dissolve, and PbI\(_2\) will precipitate. The net ionic equation is as follows:

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

Exercise \(\PageIndex{2}\)

Using the information in Table \(\PageIndex{1}\), predict what will happen in each case involving strong electrolytes. Write the net ionic equation for any reaction that occurs.

a. An aqueous solution of strontium hydroxide is added to an aqueous solution of iron(II) chloride.
b. Solid potassium phosphate is added to an aqueous solution of mercury(II) perchlorate.
c. Solid sodium fluoride is added to an aqueous solution of ammonium formate.
d. Aqueous solutions of calcium bromide and cesium carbonate are mixed.

Answer a
\[
\text{Fe}^{2+}(aq) + 2\text{OH}^-(aq) \rightarrow \text{Fe(OH)}_2(s)
\]

Answer b
\[
2\text{PO}_4^{3-}(aq) + 3\text{Hg}^{2+}(aq) \rightarrow \text{Hg}_3(\text{PO}_4)_2(s)
\]

Answer c
\[
\text{NaF}(s) \text{ dissolves; no net reaction}
\]

Answer d
\[
\text{Ca}^{2+}(aq) + \text{CO}_3^{2-}(aq) \rightarrow \text{CaCO}_3(s)
\]

Precipitation Reactions in Photography

Precipitation reactions can be used to recover silver from solutions used to develop conventional photographic film. Although largely supplanted by digital photography, conventional methods are often used for artistic purposes. Silver bromide is an off-white solid that turns black when exposed to light, which is due to the formation of small particles of silver metal. Black-and-white photography uses this reaction to capture images in shades of gray, with the darkest areas of the film corresponding to the areas that received the most light. The first step in film processing is to enhance the black/white contrast by using a developer to increase the amount of black. The developer is a reductant: because silver atoms catalyze the reduction reaction, grains of silver bromide that have already been partially reduced by exposure to light react with the reductant much more rapidly than unexposed grains.

After the film is developed, any unexposed silver bromide must be removed by a process called “fixing”; otherwise, the entire film would turn black with additional exposure to light. Although silver bromide is insoluble in water, it is soluble in a dilute solution of sodium thiosulfate (Na$_2$S$_2$O$_3$; photographer’s hypo) because of the formation of [Ag(S$_2$O$_3$)$_2$]$^{3-}$ ions. Thus washing the film with thiosulfate solution dissolves unexposed silver bromide and leaves a pattern of metallic silver granules that constitutes the negative. This procedure is summarized in Figure \(\PageIndex{2}\). The negative image is then projected onto paper coated with silver halides, and the developing and fixing processes are repeated to give a positive image. (Color photography works in much the same way, with a combination of silver halides and organic dyes superimposed in layers.) "Instant photo" operations can generate more than a hundred gallons of dilute silver waste solution per day. Recovery of silver from thiosulfate fixing solutions involves first removing the thiosulfate by oxidation and then precipitating Ag$^+$ ions with excess chloride ions.
Example \((\PageIndex{3})\)

A silver recovery unit can process 1500 L of photographic silver waste solution per day. Adding excess solid sodium chloride to a 500 mL sample of the waste (after removing the thiosulfate as described previously) gives a white precipitate that, after filtration and drying, consists of 3.73 g of AgCl. What mass of NaCl must be added to the 1500 L of silver waste to ensure that all the Ag\(^+\) ions precipitate?

**Given:** volume of solution of one reactant and mass of product from a sample of reactant solution

**Asked for:** mass of second reactant needed for complete reaction

**Strategy:**

A. Write the net ionic equation for the reaction. Calculate the number of moles of AgCl obtained from the 500 mL sample and then determine the concentration of Ag\(^+\) in the sample by dividing the number of moles of AgCl formed by the volume of solution.

B. Determine the total number of moles of Ag\(^+\) in the 1500 L solution by multiplying the Ag\(^+\) concentration by the total volume.

C. Use mole ratios to calculate the number of moles of chloride needed to react with Ag\(^+\). Obtain the mass of NaCl by multiplying the number of moles of NaCl needed by its molar mass.

**Solution:**

We can use the data provided to determine the concentration of Ag\(^+\) ions in the waste, from which the number of moles of Ag\(^+\) in the entire waste solution can be calculated. From the net ionic equation, we can determine how many moles of Cl\(^-\) are needed, which in turn will give us the mass of NaCl necessary.

A. The first step is to write the net ionic equation for the reaction:

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

We know that 500 mL of solution produced 3.73 g of AgCl. We can convert this value to the number of moles of AgCl as follows:
\[ \text{moles} \cdot \text{AgCl} = \frac{\text{grams} \cdot \text{AgCl}}{\text{molar mass} \cdot \text{AgCl}} = 3.73 \cdot \cancel{\text{g} \cdot \text{AgCl}} \left( \frac{1 \cdot \text{mol} \cdot \text{AgCl}}{143.32 \cdot \cancel{\text{g} \cdot \text{AgCl}}} \right) = 0.0260 \cdot \text{mol} \cdot \text{AgCl} \]

Therefore, the 500 mL sample of the solution contained 0.0260 mol of Ag\(^+\). The Ag\(^+\) concentration is determined as follows:

\[
[\text{Ag}^+] = \frac{\text{moles} \cdot \text{Ag}^+}{\text{liters soln}} = \frac{0.0260 \cdot \text{mol} \cdot \text{AgCl}}{0.500 \cdot \text{L}} = 0.0520 \cdot \text{M} \]

B The total number of moles of Ag\(^+\) present in 1500 L of solution is as follows:

\[
\text{moles} \cdot \text{Ag}^+ = 1500 \cdot \cancel{\text{L}} \left( \frac{0.520 \cdot \text{mol}}{1 \cdot \cancel{\text{L}}} \right) = 78.1 \cdot \text{mol} \cdot \text{Ag}^+ \]

C According to the net ionic equation, one Cl\(^-\) ion is required for each Ag\(^+\) ion. Thus 78.1 mol of NaCl are needed to precipitate the silver. The corresponding mass of NaCl is

\[
\text{mass} \cdot \text{NaCl} = 78.1 \cdot \cancel{\text{mol} \cdot \text{NaCl}} \left( \frac{58.44 \cdot \text{g} \cdot \text{NaCl}}{1 \cdot \cancel{\text{mol} \cdot \text{NaCl}}} \right) = 4560 \cdot \text{g} \cdot \text{NaCl} \]

Note that 78.1 mol of AgCl correspond to 8.43 kg of metallic silver, which is worth about $7983 at 2011 prices ($32.84 per troy ounce). Silver recovery may be economically attractive as well as ecologically sound, although the procedure outlined is becoming nearly obsolete for all but artistic purposes with the growth of digital photography.

Exercise \(\PageIndex{3}\):

Because of its toxicity, arsenic is the active ingredient in many pesticides. The arsenic content of a pesticide can be measured by oxidizing arsenic compounds to the arsenate ion (AsO\(_4^{3-}\)), which forms an insoluble silver salt (Ag\(_3\)AsO\(_4\)). Suppose you are asked to assess the purity of technical grade sodium arsenite (NaAsO\(_2\)), the active ingredient in a pesticide used against termites. You dissolve a 10.00 g sample in water, oxidize it to arsenate, and dilute it with water to a final volume of 500 mL. You then add excess AgNO\(_3\) solution to a 50.0 mL sample of the arsenate solution. The resulting precipitate of Ag\(_3\)AsO\(_4\) has a mass of 3.24 g after drying. What is the percentage by mass of NaAsO\(_2\) in the original sample?

Answer

91.0%

Precipitation reactions are a subclass of double displacement reactions.

Summary

A complete ionic equation consists of the net ionic equation and spectator ions. Predicting the solubility of ionic compounds in water can give insight into whether or not a reaction will occur. The chemical equation for a reaction in solution can be written in three ways. The \textit{overall chemical equation} shows all the substances present in their undisassociated forms; the \textit{complete ionic equation} shows all the substances present in the form in which they actually exist in solution; and the \textit{net ionic equation} is derived from the complete ionic equation by omitting all \textit{spectator ions},
ions that occur on both sides of the equation with the same coefficients. Net ionic equations demonstrate that many different combinations of reactants can give the same net chemical reaction. In a precipitation reaction, a subclass of exchange reactions, an insoluble material (a precipitate) forms when solutions of two substances are mixed. To predict the product of a precipitation reaction, all species initially present in the solutions are identified, as are any combinations likely to produce an insoluble salt.

Contributors

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