10.6: Writing Chemical Equations for Reactions in Solution- Complete Chemical, Complete Ionic, and Net Ionic Equations

A typical precipitation reaction occurs when an aqueous solution of barium chloride is mixed with one containing sodium sulfate. The **complete chemical equation** can be written to describe what happens, and such an equation is useful in making chemical calculations. (In some textbooks this is referred to as the molecular equation. This can be somewhat confusing, because the equation does not typically contain any molecules. For this reason, we will continue to refer to it as the complete chemical equation in this textbook.)

$$\underbrace{BaCl_2(aq) + Na_2SO_4(aq) \longrightarrow BaSO_4(s) + 2\,NaCl(aq)}_{\text{Complete Chemical Equation}} \tag{10.6.1}$$

However, Equation 10.6.1 does not really represent the microscopic particles (that is, the ions) present in the solution. Below is the complete ionic equation:

$$\underbrace{Ba^{2+}(aq) + \overbrace{2\,Cl^-(aq)}^{\text{spectator}} + \overbrace{2\,Na^+(aq)}^{\text{spectator}} + SO_4^{2-}(aq) \longrightarrow BaSO_4(s) + \overbrace{2\,Na^+(aq)}^{\text{spectator}} + \overbrace{2\,Cl^-(aq)}^{\text{spectator}}}_{\text{Complete Ionic Equation}} \tag{10.6.2}$$

Equation 10.6.2 is rather cumbersome and includes so many different ions that it may be confusing. In any case, we are often interested in the independent behavior of ions, not the specific compound from which they came. A precipitate of $BaSO_4(s)$ will form when *any* solution containing $Ba^{2+}(aq)$ is mixed with *any* solution containing $SO_4^{2-}(aq)$ (provided concentrations are not extremely small). This happens independently of the $Cl^-(aq)$ and $Na^+(aq)$ ions in Equation 10.6.2. These ions are called **spectator ions** because they do not participate in the reaction. When we want to emphasize the independent behavior of ions, a **net ionic equation** is written, omitting the spectator ions. For precipitation of $BaSO_4$ the net ionic equation is

$$\underbrace{Ba^{2+}(aq) + SO_4^{2-}(aq) \longrightarrow BaSO_4(s)}_{\text{Net Ionic Equation}} \tag{10.6.3}$$

✔ Example 10.6.1

a. When a solution of $AgNO_3$ is added to a solution of $CaCl_2$, insoluble $AgCl$ precipitates. Write three equations (complete chemical equation, complete ionic equation, and net ionic equation) that describe this process.
b. Write the balanced net ionic equation to describe any reaction that occurs when the solutions of Na_2SO_4 and NH_4I are mixed.

Solution

Equation Type	Example 10.6.1a	Example 10.6.1b
Complete Chemical Equation	$2\,AgNO_3(aq) + CaCl_2(aq) \longrightarrow$ $2\,AgCl(s) + Ca(NO_3)_2(aq)$ The proper states and formulas of all products are written and the chemical equation is balanced.	$Na_2SO_4(aq) + NH_4I_2(aq) \longrightarrow$ $2\,NaI(aq) + (NH_4)_2SO_4(aq)$ Both products are aqueous so there is no net ionic equation that can be written.
Complete Ionic Equation	$2\,Ag^+(aq) + 2\,NO_3^-(aq) + Ca^{2+}(aq) + Cl^-(aq) \longrightarrow$ $2\,AgCl(s) + Ca^{2+}(aq) + 2\,NO_3^-(aq)$ AgCl is a solid so it does not break up into ions in solution.	
Net Ionic Equation	$Ag^+(aq) + Cl^-(aq) \longrightarrow AgCl(s)$ All spectator ions are removed.	NaI and $(NH_4)_2SO_4$ are both soluble. No net ionic equation

The occurrence or nonoccurrence of precipitates can be used to detect the presence or absence of various species in solution. A $BaCl_2$ solution, for instance, is often used as a test for the presence of $SO_4^{2-}(aq)$ ions. There are several insoluble salts of Ba, but they all dissolve in dilute acid except for $BaSO_4$. Thus, if $BaCl_2$ solution is added to an unknown solution which has previously been acidified, the occurrence of a white precipitate is proof of the presence of the SO_4^{2-} ion.

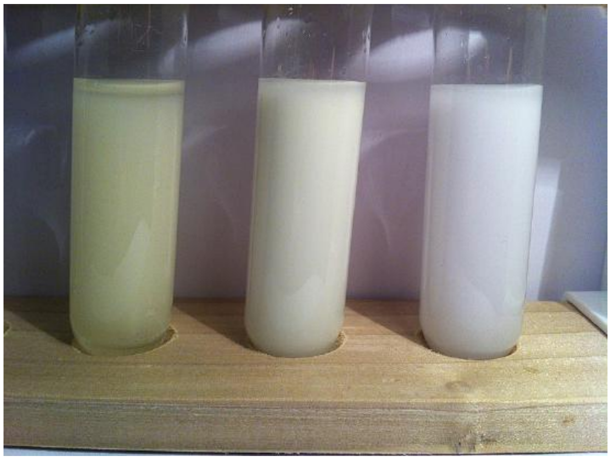

Figure 10.6.1: The three common silver halide precipitates: AgI, $AgBr$ and $AgCl$ (left to right). The silver halides precipitate out of solution, but often form suspensions before settling. (CC BY-SA 3.0; Cychr).

$AgNO_3$ solutions are often used in a similar way to test for halide ions. If $AgNO_3$ solution is added to an acidified unknown solution, a white precipitate indicates the presence of Cl^- ions, a cream-colored precipitate indicates the presence of Br^- ions, and a yellow precipitate indicates the presence of I^- ions (Figure 10.6.1). Further tests can then be made to see whether perhaps a mixture of these ions is present. When $AgNO_3$ is added to tap water, a white precipitate is almost always formed. The Cl^- ions in tap water usually come from the Cl_2 which is added to municipal water supplies to kill microorganisms.

Exercise 10.6.1

Write balanced net ionic equations to describe any reaction that occurs when the following solutions are mixed.

a. $K_2CO_3 + SrCl_2$
b. $FeSO_4 + Ba(NO_3)_2$

Answer a

$$Sr^{2+}(aq) + CO_3^{2-}(aq) \longrightarrow SrCO_3(s)$$

Answer b

$$Ba^{2+}(aq) + SO_4^{2-}(aq) \longrightarrow Ba(SO_4)(s)$$

Precipitates are also used for quantitative analysis of solutions, that is, to determine the amount of solute or the mass of solute in a given solution. For this purpose it is often convenient to use the first of the three types of equations described above. Then the rules of stoichiometry may be applied.

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