Precipitates

There is usually a limit to the amount of solute that can be dissolved in a solvent. A solution is called **saturated** when no more solute can be dissolved in it. If more solute is added to a saturated solution, it will not dissolve.

A solution that has less than the maximum amount of solute that can be dissolved is called an **unsaturated** solution. A solution that contains a greater amount of solute than that needed to form a saturated solution is said to be **supersaturated**. A supersaturated solution is an unstable, nonequilibrium state achieved by manipulating conditions such as temperature.

When a solution is supersaturated, the excess solute will leave the solution and form a **precipitate**.

Predicting the Formation of a Precipitate

Precipitates will form in supersaturated solutions, but not in unsaturated or saturated solutions. This fact can be used to predict whether or not a precipitate will form in a given solution. In order to make such a prediction, we will calculate the reaction quotient (Q), which is called the **ion product** in this case.

Recall that the reaction quotient is calculated by substituting the actual concentrations measured at a given point into the equilibrium expression. This result is then compared to the value of the solubility product (K_{sp}) to determine if the solution is unsaturated, saturated, or supersaturated.

- If $Q > K_{sp}$, then the solution is supersaturated and a precipitate will form.
- If $Q = K_{sp}$, then the solution is saturated and a precipitate will not form.
- If $Q < K_{sp}$, then the solution is unsaturated and a precipitate will not form.

Example 1

Suppose 0.01 mol of $PbCl_2(s)$ is dissolved in 150 mL of hot water, and the solution is cooled slowly to $25°C$. Is the solution supersaturated?

Precipitation Reactions

So far we have been discussing single solid substances added to pure water to form aqueous solutions. However, it is also possible to generate a precipitate by mixing two aqueous solutions.

A reaction in which two solutions are mixed and a precipitate forms is called a **precipitation reaction**. As with other reactions, precipitation reactions are described by balanced equations, such as the example shown below.

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

On the left side of the equation are the original solutions. On the right side of the equation is the solid precipitate and the remaining solution.

When the formula for an ionic compound is written followed by (*aq*), it is actually being described as a solution of its ions. Thus, when two solutions are mixed, two sets of ions are being mixed. The final solution contains all of the ions from each of the individual solutions. The precipitate that forms must be a combination of these ions.

Consider an example in which a solution of $AgNO_3$ is mixed with a solution of KBr. The combined solution will contain the ions

$$Ag^+,\ NO_3^-,\ K^+, \text{ and } Br^-$$

Ions can combine only if they form an electrically neutral compound. Thus, the possible compounds in this mixture are

$$AgNO_3,\ KBr,\ AgBr, \text{ and } KNO_3$$

The first two possibilities are the original reactants, and can be disregarded. That leaves two possibilities for the precipitate. In this particular reaction, the precipitate that forms is $AgBr$, as shown in the equation below.

$$AgNO_3(aq) + KBr(aq) \rightarrow AgBr(s) + KNO_3(aq)$$

Determining which product is the precipitate is no simple task, and must typically be done experimentally. In many cases, even chemists can only make educated guesses. There is, however, a set of guiding information that can be used to make such predictions. These **solubility rules** have been provided for you on an attached handout.

The solubility rules classify a substance as *low solubility* if less than $0.01\ mol$ dissolves in a liter of water. Because it dissolves so little, a substance classified as *low solubility* will usually precipitate in a precipitation reaction.

If we return to our example, the solubility rules indicate that a compound containing NO_3^- is soluble and will thus dissolve in water. On the other hand, compounds containing Ag^+ are generally insoluble in water. Thus, KNO_3 must dissolve (and dissociate into ions) while $AgBr$ must be the precipitate.

Notice that a precipitation reaction is a double displacement reaction in which the positive and negative ions of the reactants exchange partners. Not all double displacement reactions result in the formation of a precipitate.

It is possible to predict whether or not a given double displacement reaction will result in the formation of a precipitate by using the solubility rules. To do so, first predict the products of the reaction (if they are not given to you), then determine the solubility of the products using the solubility rules. There are two possible outcomes:

1. If both products are soluble in water, then no reaction will be observed.

2. If one of the products has low solubility, then it will precipitate.

Example 2
Predict whether the following reactants will result in a precipitation reaction. If there is a reaction, write the balanced chemical equation and identify the precipitate.

$Cu(NO_3)_2(aq) + NaOH(aq) \rightarrow$

Net Ionic Equations

The equations used to describe precipitation reactions have been **formula equations**, which show the reactants and products of the reaction. For example

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

Formula equations do not give a true picture of what is occurring in a solution because the soluble ionic compounds actually dissociate into ions. A clearer picture is formed when you write an equation in which all the soluble ionic compounds are shown as dissociated ions in solution. For example

$$Pb^{2+}(aq) + 2NO_3^-(aq) + 2K^+(aq) + 2I^-(aq) \rightarrow PbI_2(s) + 2K^+(aq) + 2NO_3^-(aq)$$

Such an equation is called a **complete ionic equation**.

Notice that only the iodide and lead(II) ions undergo a chemical change. The potassium and nitrate ions are unchanged in this reaction. Ions that do not take part in a chemical reaction and are found in solution both before and after the reaction are called **spectator ions**.

If we remove the spectator ions from the complete ionic equation, a much simpler equation called the **net ionic equation** results.

$$Pb^{2+}(aq) + 2I^-(aq) \rightarrow PbI_2(s)$$

Example 3
Write the complete ionic and net ionic equations for the reaction from Example 2.

Worksheet

1. A solution is prepared by dissolving 0.0015 mol of silver acetate ($AgC_2H_3O_2$) in 50 mL of hot water. Silver acetate dissolves according to the equation

$$AgC_2H_3O_2(s) \Leftrightarrow Ag^+(aq) + C_2H_3O_2^-(aq)$$

   Will a precipitate form if the solution is cooled to $25°C$? ($K_{sp} = 2.3 \times 10^{-3}$)

2. A solution is prepared by dissolving 0.01 g of gallium trifluoride (GaF_3) in 100 mL of hot water. Will a precipitate form if the solution is cooled to $25°C$? ($K_{sp} = 1.6 \times 10^{-4}$)

3. A solution is prepared by dissolving 0.957 mol of SrF_2 in 1000 L of hot water. Will a precipitate form if the solution is cooled to $18°C$? ($K_{sp} = 2.8 \times 10^{-9}$ at $18°C$)

4. A solution is prepared by dissolving 0.0001 mol of $Zn(OH)_2$ in 50 L of hot water. Will a precipitate form if the solution is cooled to $25°C$? ($K_{sp} = 4.5 \times 10^{-17}$)

5. A solution is prepared by dissolving 0.03 mol of potassium perchlorate ($KClO_4$) in 75 mL of hot water. potassium perchlorate dissolves according to the equation

$$KClO_4(s) \Leftrightarrow K^+(aq) + ClO_4^-(aq)$$

   Will a precipitate form if the solution is cooled to $25°C$? ($K_{sp} = 8.9 \times 10^{-3}$)

6. A solution is prepared by dissolving 1.4 g of silver sulfate (Ag_2SO_4) in 100 mL of hot water. Will a precipitate form if the solution is cooled to $25°C$? ($K_{sp} = 1.2 \times 10^{-5}$)

For each of the following reactants, predict whether a precipitation reaction will take place between them. If there is no reaction, write "no reaction." If there is a reaction, write the formula equation, complete ionic equation, and net ionic equations that describe the reaction.

7. $Li_2CO_3(aq) + Co(C_2H_3O_2)_2(aq) \rightarrow$
8. $Fe(NO_3)_3(aq) + K_2S(aq) \rightarrow$
9. $Pb(NO_3)_2(aq) + Li_2SO_4(aq) \rightarrow$
10. $NH_4Cl(aq) + Cu(C_2H_3O_2)_2(aq) \rightarrow$
11. $K_3PO_4(aq) + Cd(C_2H_3O_2)_2(aq) \rightarrow$