Writing Net Ionic Equations for Precipitation Reactions

Here are the steps in writing equations for the reaction that occurs when aqueous solutions of two ionic compounds are mixed:

1. List all of the ions that are present when the solutions are mixed. Remember that ionic compounds dissociate into the cation and anion.

2. Write the two products of the exchange reaction that could occur, in which the cation of one compound joins with the anion of the other compound.

3. Use the solubility rules to determine the solubility of the two products in Step 2. If both are soluble, nothing will precipitate, so no reaction would occur. If one (or both) of the compounds is insoluble, that compound will precipitate.

4. If a reaction occurs, write the balanced equation for the reaction, putting (*s*) after the compound that would precipitate.

5. Now write an ionic equation for the reaction in Step 4, in which all aqueous compounds are dissociated into ions. Do not dissociate precipitates; leave as a single formula. Remember that a free ion will always have (*aq*) after it, never (*s*).

6. Identify any spectator ions, which appear on both sides of the equation. These will generally be the ions of the soluble product of the exchange reaction.

7. Rewrite the ionic equation without the spectator ions. This is the net ionic equation. In general this equation will simply have the precipitate as the product, and the ions that make it up as the reactants:

e.g. $A^+\,(aq) + B^-\,(aq) \rightarrow AB\,(s)$ [general format for the net ionic equation for a precipitation reaction]

Now here is an example to illustrate these steps:

Example: Write equations for the reaction between aqueous solutions of silver nitrate and potassium sulfide:

1. The reactants are $AgNO_3$ and K_2S, so the ions present in solution are Ag^+, NO_3^-, K^+ and S^{2-}.

2. If an exchange occurred, Ag^+ would join with S^{2-} and K^+ would join with NO_3^-. So the products would be AgCl and KNO_3:

$$AgNO_3 + K_2S \rightarrow Ag_2S + KNO_3 \qquad \text{(unbalanced)}$$

3. According to the solubility rules, Ag_2S is insoluble (sulfides are usually insoluble and Ag^+ is <u>not</u> one of the exceptions to that rule). Since compounds of K^+ and NO_3^- are always soluble, KNO_3 is soluble. So Ag_2S will be the precipitate.

4. The balanced equation for the reaction is:

$$2\ AgNO_3\ (aq)\ +\ K_2S\ (aq)\ \rightarrow\ Ag_2S\ (s)\ +\ 2\ KNO_3\ (aq)$$

5. Dissociate all compounds except Ag_2S to get the complete ionic equation:

$$2\ Ag^+\ (aq)\ +\ 2\ NO_3^-\ (aq)\ +\ 2\ K^+\ (aq)\ +\ S^{2-}\ (aq)\ \rightarrow\ Ag_2S\ (s)\ +\ 2\ K^+\ (aq)\ +\ 2\ NO_3^-\ (aq)$$

(remember that a subscript after an ion in the formula of a compound becomes the coefficient of that ion when the compound is shown as dissociated)

6. K^+ and NO_3^- are the spectator ions. These are the ions of the soluble product of the reaction, KNO_3.

7. Cancelling the spectator ions:

$$2\ Ag^+\ (aq)\ +\ \cancel{2\ NO_3^-\ (aq)\ +\ 2\ K^+\ (aq)}\ +\ S^{2-}\ (aq)\ \rightarrow\ Ag_2S\ (s)\ +\ \cancel{2\ K^+\ (aq)\ +\ 2\ NO_3^-\ (aq)}$$

Net ionic equation: $2\ Ag^+\ (aq)\ +\ S^{2-}\ (aq)\ \rightarrow\ Ag_2S\ (s)$

Notice the format of the net ionic equation:

cation of precipitate (*aq*) + anion of precipitate (*aq*) → precipitate (*s*)

In most cases, the net ionic equation for a precipitation reaction will follow this pattern.

Now practice with more examples from this week's recitation handout, Sapling, or your textbook!