Mole Ratios of Precipitates

Problems involving insoluble precipitates are stoichiometry problems in which you must calculate some amount when given some other amounts involved in a precipitation reaction. Precipitation reactions are all double displacement reactions.

$$AB + CD \leftrightarrow AD + CD(s)$$

You will answer questions involving the mixing of two solutions. If you know the concentration and volume, you can calculate the number of moles by multiplying them.

$$n\ (mol) = C\ (mol/l) \times v\ (l)$$

You should balance the reactions before you solve the problems so you know the ratios of the numbers of moles of the various compounds. To solve some of the problems, you will manipulate molar amounts and then convert from moles to grams (by using molecular mass) or to milliliters (by using the concentration or molarity). Please refer to Chapter 3 to review stoichiometry problems.

Sample problems:

1. **Fourteen grams (14.0 g) of silver are dissolved in nitric acid. An excess of sodium chloride is added to form AgCl precipitate. What will be the mass of the dried precipitate?**

   The atomic mass of silver is 107.87 g/mol. The atomic mass of chlorine is 35.45 g/mol. However you decide to do this problem, it ends up being related to the ratio of the atomic mass of silver and the molecular mass of silver chloride: 107.87/(107.87+35.45) = 0.7527.

   You quickly calculate the result by dividing the mass of silver by this ratio.

   14.0 g Ag ÷ 0.7527 g Ag/g AgCl = **18.6 g** AgCl

   You can take the step-by-step approach.

   14 g Ag ÷ 107.87 g/mol = 0.130 mol Ag = 0.130 mol AgCl
   0.103 mol AgCl × 143.32 g/mol AgCl = **18.6 g** AgCl

2. **You dissolve 5.00 g of silver alloy in nitric acid. After adding an excess of potassium iodide, you filter and dry the precipitate. It weighs 8.69 g. What is the percent of silver in the alloy?**

   The atomic mass of silver is 107.87 g/mol. The atomic mass of iodine is 126.9 g/mol. Calculate the number of moles of AgI to get the number of moles of Ag. Use the moles of Ag to get the mass of Ag and compare with the sample's mass.

   8.69 g AgI ÷ 234.8 g/mol AgI = 0.037 mol AgI = 0.037 mol Ag
   0.037 mol Ag × 107.87 g/mol = 3.99 g Ag
   3.99 g Ag ÷ 5.00 g alloy × 100 = **79.8%**

3. **You combine an excess of an unknown iron salt with 10.0 g of potassium hydroxide to form a precipitate. The filtered and dried precipitate weighs 17.9 g. What is the nature of the iron in the iron salt? Assume all of the iron precipitates as hydroxide.**

Potassium hydroxide, KOH, has a molecular mass of 56.11 g/mol.

$10.0 \text{ g} \div 56.11 \text{ g/mol} = 0.178 \text{ mol}$

The iron hydroxide formed $Fe(OH)_x$ with a molecular mass of 55.85 + 18.02x g/mol. Using the mass of the precipitate, the number of moles of hydroxide and this formula, you can write the following equation.

$17.9 \text{ g} = (55.85 + 18.02x) \text{ g/mol} \times 0.178 \text{ mol}$
$17.9 \text{ g} = 9.94 \text{ g} + 3.21x \text{ g}$
$8.06 \text{ g} = 3.21x \text{ g}$
$x = 8.06/3.21 = 2.5$

The iron salt has an oxidation state of **2.5**, which means half is ferrous [Fe(II)] and half is ferric [Fe(III)].