7.5: Solution Stoichiometry

As we learned in Chapter 5, double replacement reactions involve the reaction between ionic compounds in solution and, in the course of the reaction, the ions in the two reacting compounds are "switched" (they *replace* each other). As an example, silver nitrate and sodium chloride react to form sodium nitrate and the *insoluble* compound, silver chloride.

$$\mathrm{AgNO_3(aq) + NaCl(aq) \rightarrow AgCl(s) + NaNO_3(aq)}$$

Because these reactions occur in aqueous solution, we can use the concept of molarity to directly calculate the number of moles of products that will be formed, and hence the mass of precipitates. In the reaction shown above, if we mixed 123 mL of a 1.00 M solution of NaCl with 72.5 mL of a 2.71 M solution of $AgNO_3$, we could calculate the moles (and hence, the mass) of AgCl that will be formed as follows:

First, we must examine the reaction stoichiometry. In this reaction, one mole of $AgNO_3$ reacts with one mole of NaCl to give one mole of AgCl. Because our ratios are one, we don't need to include them in the equation. Next, we need to calculate the number of moles of each reactant:

$$0.123 L \times \left(\frac{1.00\ mole}{1.00\ L}\right) = 0.123\ moles\ NaCl$$

$$0.0725 L \times \left(\frac{2.71\ mole}{1.00\ L}\right) = 0.196\ moles\ AgNO_3$$

Because this is a *limiting reactant* problem, we need to recall that the moles of product that can be formed will equal the *smaller* of the number of moles of the two reactants. In this case, NaCl is limiting and $AgNO_3$ is in excess. Because our stoichiometry is one-to-one, we will therefore form 0.123 moles of AgCl. Finally, we can convert this to mass using the molar mass of AgCl:

$$0.0725 L \times \left(\frac{2.71\ mole}{1.00\ L}\right) = 0.196\ moles\ AgNO_3$$

In a reaction where the stoichiometry is not one-to-one, you simply need to include the stoichiometric ratio in you equations. Thus, for the reaction between lead (II) nitrate and potassium iodide, two moles of potassium iodide are required for every mole of lead (II) iodide that is formed.

\[\ce{Pb(NO3)2 (aq) + 2 KI (aq) → PbI2 (s) + 2 KNO3 (aq)} \nonumber \] For example: 1.78 grams of lead (II) nitrate are dissolved in 17.0 mL of water and then mixed with 25.0 mL of 2.5 M potassium iodide solution. What *mass* of lead (II) iodide will be formed and what will be the *final concentration* of potassium nitrate in the solution? Again, we need to look at this as a limiting reactant problem and first calculate the number of moles of each reactant:

$$1.78\ g \times \left(\frac{1.00\ mole}{331.2\ g}\right) = 5.37 \times 10^{-3}\ moles\ Pb(NO_3)_2$$

$$0.0025\ L \times \left(\frac{2.50\ mole}{1.00\ L}\right) = 6.25 \times 10^{-3}\ moles\ KI$$

The stoichiometry of this reaction is given by the ratios:

$$\left(\frac{1\ mole\ PbI_2}{2\ mole\ KI}\right)\ and\ \left(\frac{1\ mole\ PbI_2}{1\ mole\ Pb(NO_3)_2}\right)$$

so the number of moles of product that would be formed from each reactant is calculated as:

$$\left(\frac{1\ mole\ PbI_2}{1\ mole\ Pb(NO_3)_2}\right)$$

$$6.25 \times 10^{-3}\ moles\ KI \times \left(\frac{1\ mole\ PbI_2}{2\ moles\ KI}\right) = 3.12 \times 10^{-3}\ moles\ PbI_2$$

Potassium iodide produces the smaller amount of PbI_2 and hence, is *limiting* and lead (II) nitrate is in *excess*. The mass of lead (II) iodide that will be produced is then calculated from the number of moles and the molar mass:

$$3.12 \times 10^{-3}\ moles\ \times \left(\frac{461\ grams}{1\ mole}\right) = 1.44\ grams\ PbI_2$$

To determine the concentration of potassium nitrate in the final solution, we need to note that two moles of potassium nitrate are formed for every mole of PbI_2, or a stoichiometric ratio of

$$\left(\frac{2\ moles\ KNO_3}{1\ mole\ PbI_2}\right)$$

Our final volume is (17.0 + 25.0) = 42.0 mL, and the concentration of potassium nitrate is calculated as:

$$\frac{3.12 \times 10^{-3}\ moles\ PbI_2 \times \left(\frac{2\ moles\ KNO_3}{1\ mole\ PbI_2}\right)}{0.0420\ L} = 0.148\ moles\ KNO_3/L\ or\ 0.148\ M$$

Exercise 7.5.1

1. A sample of 12.7 grams of sodium sulfate (Na_2SO_4) is dissolved in 672 mL of distilled water.
   a. What is the molar concentration of sodium sulfate in the solution?
   b. What is the concentration of sodium ion in the solution?
2. How many moles of sodium sulfate must be added to an aqueous solution that contains 2.0 moles of barium chloride in order to precipitate 0.50 moles of barium sulfate?
3. If 1.0 g of NaN_3 reacts with 25 mL of 0.20 M $NaNO_3$ according to the reaction shown below, how many moles of N_2*(g)* are produced?

$$5NaN3(s) + NaNO3(aq) \rightarrow 3Na2O(s) + 8N2(g)$$

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