5.11: Limiting Reactant and Percent Yield

Mixing of exact amounts of reactants such that all are consumed and none left over in a chemical reaction almost never occurs. Instead, one of the reactants is usually a **limiting reactant**. Suppose, for example that 100 g of elemental zinc (atomic mass 65.4) and 80 g of elemental sulfur (atomic mass 32.0) are mixed and heated undergoing the following reaction:

$$\mathrm{Zn} + \mathrm{S} \rightarrow \mathrm{ZnS} \tag{5.11.1}$$

What mass of ZnS, formula mass 97.4 g/mol, is produced? If 100 g of zinc react completely, the mass of S reacting and the mass of ZnS produced would be given by the following calculations:

$\text{Mass s} = 100.0 \text{ G Zn} \times \frac{1 \text{ mol Zn}}{65.4 \text{ g Zn}} \times \frac{1 \text{ mol S}}{1 \text{ mol Zn}} \times \frac{32.0 \text{ g S}}{1 \text{ mol S}} = 48.9 \text{ g S}$

$\text{Mass ZnS} = 100.0\text{g Zn} \times \frac{1 \text{ mol Zn}}{65.4 \text{ g Zn}} \times \frac{1 \text{ mol S}}{1 \text{ mol Zn}} \times \frac{97.4 \text{ g ZnS}}{1 \text{ mol ZnS}} = 149 \text{ g S}$

Only 48.9 g of the available S react, so sulfur is in excess and zinc is the limiting reactant. A similar calculation for the amount of Zn required to react with 80 g of sulfur would show that 164g of Zn would be required, but only 100 g is available.

Exercise

A solution containing 10.0 g of HCl dissolved in water (a solution of hydrochloric acid) was mixed with 8.0 g of Al metal undergoing the reaction

$2\,\mathrm{Al} + 6\,\mathrm{HCl} \rightarrow 2\,\mathrm{AlCl_3} + 3\,\mathrm{H_2}$

Given atomic masses H 1.0, Al 27.0, and Cl 35.5, which reactant was left over? How much? What mass of $AlCl_3$ was produced?

Answer

HCl was the limiting reactant. Only 2.47 g of Al were consumed leaving 5.53 g of Al unreacted. The mass of $AlCl_3$ produced was 12.2 g

Percent Yield

The mass of product calculated from the mass of limiting reactant in a chemical reaction is called the **stoichiometric yield** of a chemical reaction. By measuring the actual mass of a product produced in a chemical reaction and comparing it to the mass predicted from the stoichiometric yield it is possible to calculate the *percent yield*. This concept is illustrated by the following example.

Suppose that a water solution containing 25.0 g of $CaCl_2$ was mixed with a solution of excess sodium sulfate,

$$\mathrm{CaCl_2(aq)} + \mathrm{Na_2SO_4(aq)} \rightarrow \mathrm{CaSO_4(s)} + 2\,\mathrm{NaCl(aq)} \tag{5.11.2}$$

to produce a solid precipitate of $CaSO_4$, the desired product of the reaction. (Recall that a *precipitate* is a solid formed by the reaction of species in solution; such a solid is said to *precipitate* from the solution.) Removed by filtration and dried, the precipitate was found to have a mass of 28.3 g, the **measured yield**. What was the percent yield?

Using atomic masses Ca 40.0, Cl 35.5, Na 23.0, and O, 16.0 gives molar masses of 111 g/mol for $CaCl_2$ and 136 g/mol for $CaSO_4$. Furthermore, 1 mole of $CaCl_2$ yields 1 mol of $CaSO_4$. The stoichiometric yield of $CaSO_4$ is given by the following calculation

$$\text{Mass } \mathrm{CaSO_4} = 25.0 \text{ g } \mathrm{CaCl_2} \times \frac{1 \text{ mol } \mathrm{CaCl_2}}{111 \text{ g } \mathrm{CaCl_2}} \times \frac{1 \text{ mol } \mathrm{CaSO_4}}{1 \text{ mol } \mathrm{CaCl_2}} \times \frac{136 \text{ g } \mathrm{CaSO_4}}{1 \text{ mol } \mathrm{CaSO_4}} = 30.6 \text{ g } \mathrm{CaSO_4} \tag{5.11.3}$$

The percent yield is calculated by the following:

$$\text{Percent yield} = \frac{\text{Measured yield}}{\text{Stoichiometric yield}} \times 100 \tag{5.11.4}$$

$$\text{Percent yield} = \frac{28.3 \text{ g}}{30.6 \text{ g}} \times 100 = 92.5\% \tag{5.11.5}$$