6.4: Percentage Yield

When we use stoichiometric calculations to predict quantities in a reaction, our results are based on the assumption that everything happens in reality exactly as described by the chemical equation. Unfortunately, that is not always the case. When you work in the laboratory, things often go wrong. When you are weighing out reactants and transferring the materials to reaction vessels, some material will often remain on the spatula or in the weighing vessel. As you collect product, some may spill, or in a vigorous reaction, material may escape from the reaction vessel. Stoichiometric calculations will give you a **theoretical yield** for a reaction; the yield that you should obtain assuming that the reaction proceeds with 100% efficiency and that no material is lost in handling. The amount of material that you isolate from a given reaction is called the **actual yield** and it is always less than the theoretical yield. The *percentage* of the theoretical yield that you actually isolate is called the **percentage yield**.

Consider the reaction between silver nitrate and sodium chloride to form solid silver chloride. If we react 10.00 grams silver nitrate with excess sodium chloride, we would predict that we would obtain:

We set up the problem to solve for mol product; the general equation is:

$$(molproduct) = (molreactant) \times \left(\frac{molproduct}{molreactant}\right)$$

For *mol product* and *mol reactant*, we use the expressions for (mass)/(molar mass), as shown in the scheme above. The stoichiometric mole ratio is set up so that mol reactant will cancel, giving a solution in mol product. Substituting,

$$(x\ g\ AgCl)\left(\frac{1\ mol\ AgCl}{143.32\ g\ AgCl}\right) = (10.00\ g\ AgNO_3)\left(\frac{1\ mol\ AgNO_3}{169.88\ g\ AgNO_3}\right) \times \left(\frac{1\ mol\ AgCl}{1\ mol\ AgNO_3}\right)$$

$$(x\ g\ AgCl) = (10.00\ g\ AgNO_3)\left(\frac{1\ mol\ AgNO_3}{169.88\ g\ AgNO_3}\right) \times \left(\frac{1\ mol\ AgCl}{1\ mol\ AgNO_3}\right) \times \left(\frac{143.32\ g\ AgCl}{1\ mol\ AgCl}\right) = 8.440\ g$$

This mass, calculated from the masses of starting materials and the stoichiometry of the equation is the theoretical yield. Because the silver chloride is a precipitate from an aqueous solution, however, we must filter it, dry it, transfer it to our balance and weigh it, before we can measure how much product we obtain as our actual yield. As we filter it, a small amount of solid is likely to remain stuck to the sides of the flask. When it is dry and we transfer it to the balance, some solid will remain on the filter paper, some on the spatula, and (most likely) your lab partner will sneeze at an inopportune moment and blow some of it all over the desktop. Considering all of this, it is very unlikely that we will end up with an actual yield of 8.044 grams of solid AgCl.

Let's assume that we have done all of these operations (including the sneeze) and when we weigh our solid AgCl, we actually obtain 7.98 grams of solid. We know that we obtain 7.98 grams of product (our actual yield) and we calculated that we should obtain 8.440 grams of product (the theoretical yield). The percentage of the theoretical yield that we obtain is called the **percentage yield**, and is calculated as (actual yield)/(theoretical yield) 100. In the present case:

$$\left(\frac{7.98g}{8.440g}\right) \times 100 = 94.5\%$$

The percentage yield of solid AgCl that we obtained in this reaction is therefore 94.5% (not bad, actually, considering your lab partner). The concept of percentage yield is generally applied to all experimental work in chemistry.

? Exercise 6.4.1

Powdered zinc and solid sulfur combine explosively to form zinc sulfide. You and your lab partner carefully mix 0.010 mole of solid zinc powder with exactly 0.010 mole of powdered sulfur in a small porcelain crucible. Knowing your lab partner, you allow your instructor to ignite the mixture. The explosion forms a cloud of ZnS, scatters some all over the ground, and leaves a crusty pile of product in the crucible. You transfer this and determine that 0.35 grams of solid product has been recovered. Calculate the percentage yield.

Exercise 6.4.2

The Harber process is used making ammonia from nitrogen and hydrogen according to the equation shown below. The yield of the reaction, however, is not 100%.

$$N_2 + 3\ H_2 \rightarrow 2\ NH_3$$

a. Suppose you end up with 6.2 moles of ammonia, but the reaction stoichiometry predicts that you should have 170.0 grams of ammonia. What is the percent yield for this reaction?
b. If you started with 6.2 grams of nitrogen and you produce 6.2 grams of ammonia what would be the percent yield?
