

CHEMFILE MINI-GUIDE TO PROBLEM SOLVING

Percent Yield

Although we can write perfectly balanced equations to represent perfect reactions, the reactions themselves are often not perfect. A reaction does not always produce the quantity of products that the balanced equation seems to guarantee. This happens not because the equation is wrong but because reactions in the real world seldom produce perfect results.

As an example of an imperfect reaction, look again at the equation that shows the industrial production of ammonia.



In the manufacture of ammonia, it is nearly impossible to produce 2 mol (34.08 g) of NH_3 from the simple reaction of 1 mol (28.02 g) of N_2 and 3 mol (6.06 g) of H_2 because some ammonia molecules begin breaking down into N_2 and H_2 molecules as soon as they are formed.

There are several reasons that real-world reactions do not produce products at a yield of 100%. Some are simple mechanical reasons, such as:

- Reactants or products leak out, especially when they are gases.
- The reactants are not 100% pure.
- Some product is lost when it is purified.

There are also many chemical reasons, including:

- The products decompose back into reactants (as with the ammonia process).
- The products react to form different substances.
- Some of the reactants react in ways other than the one shown in the equation. These are called *side reactions*.
- The reaction occurs very slowly. This is especially true of reactions involving organic substances.

Chemists are very concerned with the yields of reactions because they must find ways to carry out reactions economically and on a large scale. If the yield of a reaction is too small, the products may not be competitive in the marketplace. If a reaction has only a 50% yield, it produces only 50% of the amount of product that it theoretically should. In this chapter, you will learn how to solve problems involving real-world reactions and percent yield.

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