Percent Yield

Unit: Stoichiometry

NGSS Standards/MA Curriculum Frameworks (2016): HS-PS1-7

Mastery Objective(s): (Students will be able to…)

- Calculate the percent yield of a reaction.

Success Criteria:

- Theoretical yield calculated correctly using stoichiometry calculations.
- Algebra and rounding to appropriate number of significant figures is correct.

Tier 2 Vocabulary: yield

Language Objectives:

- Explain how to turn fractions into percentages.

Notes:

<u>theoretical yield</u>: the amount of a product predicted, based only on stoichiometry calculations.

<u>actual yield</u>: the actual amount of product recovered in the laboratory.

<u>percent yield</u>: the amount of product recovered, expressed as a percentage of the theoretical yield.

When you do a stoichiometry calculation, the answer to the question "how much product should be produced" is the theoretical yield.

Actual yield depends on several factors. Many reactions do not go to completion, but instead reach an equilibrium condition where the amount of reactants and products is constant. Sometimes it is not possible to recover all of the product because of challenges associated with separating it from the other reactants and products. *Etc.*

Because of these factors, the actual yield is determined by performing the reaction in a laboratory and measuring the amount of product you got.

Once you have the actual and theoretical yield numbers, the percent yield is:

$$\frac{\text{actual yield}}{\text{theoretical yield}} \times 100 = \text{percent yield}$$

Use this space for summary and/or additional notes:

Sample Problem:

Q: Suppose you perform the reaction:

$$3\text{ Al (s)} + 3\text{ CuCl}_2\text{ (aq)} \rightarrow 2\text{ AlCl}_3\text{ (aq)} + 3\text{ Cu (s)}$$

If you start with 9.0 g of Al and you recover 28 g of Cu, what was your percent yield?

A: First, calculate the theoretical (predicted) yield of Cu using stoichiometry:

$$\frac{9.0\text{ g Al}}{1} \times \frac{1\text{ mol Al}}{27.0\text{ g Al}} \times \frac{3\text{ mol Cu}}{2\text{ mol Al}} \times \frac{63.6\text{ g Cu}}{1\text{ mol Cu}} = 31.8\text{ g Cu}$$

Then calculate: $\frac{\text{actual yield}}{\text{theoretical yield}} \times 100 = \text{percent yield}:$

$$\frac{28\text{ g Cu recovered}}{31.8\text{ g Cu predicted}} = 0.88 \times 100 = 88\,\%$$

Note that the percent yield cannot exceed 100 %. Conservation of mass tells us that if we had enough aluminum and copper chloride in the above reaction to make 31.8 g of copper metal, there's no way we could actually make more than that. If you calculate a percent yield greater than 100 %, you should:

- Double-check your calculations to make sure you didn't make a mistake.
- Look for other compounds that might have gotten into the product that you measured. For example:
  - If you collect a precipitate on a piece of filter paper, it will be with with everything else that was in the beaker. Even if you evaporate the water, it will leave behind other compounds that had been dissolved. Also, if you let the product dry in the air on a humid day, the compound could be hygroscopic and/or form a hydrate.
  - If you collect a gas, the most common way is to use a eudiometer (gas collection tube) that starts filled with water, and to have the gas displace the water. However, there will also be water vapor in the tube, so you need to account for the water based on the vapor pressure of water at the temperature the gas was collected.

Use this space for summary and/or additional notes:

Homework Problems

In order to isolate percent yield problems, these questions refer to the mass-mass stoichiometry homework problems starting on page 410.

1. In problem #1, part b on page 410, suppose that 3.85 g KCl was recovered in the lab. What was the percent yield?

   Answer:

2. In problem #2, part a on page 410, suppose that 125 g of NaOH was recovered. What was the percent yield?

   Answer:

3. In problem #3 on page 411 suppose that 10 L of O_2 was recovered.
   a. What was the percent yield?

      Answer:

   b. What might have happened in the lab that could account for the fact that you got a percent yield higher than 100 %?

4. In problem #4 on page 411 suppose you started with 3.00 g of precipitate, which was still wet from the solution in the reaction. You let it dry in the lab until the mass stopped changing, and you recorded it to be 2.50 g. Is this a good answer? If not, what else could you do?

Use this space for summary and/or additional notes: