

Limiting Reactants and Percent Yield

How do you find the limiting reactant?

●●○ Intermediate

Reactions & Stoichiometry

Last reviewed 3 October 2026

Before you start: it helps to understand [Stoichiometry](#).

What is it?

In most real reactions the reactants are **not** mixed in exactly the ratio of the balanced equation. One of them runs out first.

- The **limiting reactant** is the reactant that is **used up first**. It limits how much product can form.
- Any other reactant is **in excess**: some of it is left over when the reaction stops.
- The **theoretical yield** is the maximum amount of product, calculated from the limiting reactant.
- The **actual yield** is the amount of product really obtained in an experiment.
- The **percent yield** compares the two.

KEY IDEA

The amount of product depends only on the reactant that runs out first, the limiting reactant, not on how much of the other reactants you add.

Why does it matter?

- **Chemists choose the limiting reactant on purpose.** The most expensive or hardest-to-make reactant is often the limiting one, while a cheap one (such as oxygen from air) is used in excess so that none of the valuable reactant is wasted.
- **Percent yield measures efficiency.** In the lab and in industry, a low yield means wasted materials, money and energy.
- **Excess reactant must be dealt with.** Leftover reactant often has to be separated from the product, recycled or disposed of safely.

How does it work?

Finding the limiting reactant

1. Convert the amount of **each** reactant into moles.

2. For **each** reactant, calculate how much product it could make on its own, using the mole ratio from the balanced equation.
3. The reactant that gives the **smaller** amount of product is the **limiting reactant**. That smaller amount is the **theoretical yield**.

An equivalent method: take one reactant, calculate how much of the other it needs, and compare that with how much you actually have.

Excess reactant left over

Use the limiting reactant to calculate how much of the excess reactant is **used up**, then subtract that from the amount you started with.

Percent yield

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

Both yields must be in the same units (usually grams).

∞ THINK OF IT LIKE THIS

Making sandwiches needs **2 slices of bread + 1 slice of cheese** → **1 sandwich**. You have 10 slices of bread but only 3 slices of cheese. The cheese runs out after 3 sandwiches, with 4 slices of bread left over. The cheese is the limiting “reactant”, even though you have fewer slices of it. If you drop one sandwich on the floor and only 2 are eaten, your “percent yield” is $2 \div 3 \times 100\% = 67\%$.

🔍 MORE PRECISELY

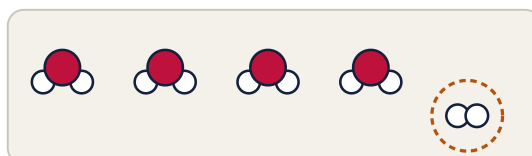
Percent yields are usually below 100% because reactions may not go to completion, side reactions can make unwanted products, and some product is lost while it is separated and purified. A yield **above** 100% doesn't mean extra product was created. It signals a problem, for example a product that is still wet or contains impurities.

Visualise it

Before 5 H₂ + 2 O₂



After 4 H₂O + 1 H₂ left over



O₂ runs out first: it is the limiting reactant.

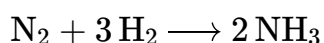
H₂ is in excess (1 molecule is left over).

Two O₂ molecules can react with only four H₂ molecules, so one H₂ is left over.

Worked example

WORKED EXAMPLE: Limiting reactant, theoretical yield and excess

Question: 28.0 g of nitrogen and 9.00 g of hydrogen react to make ammonia. Which is the limiting reactant? What mass of ammonia can form, and what mass of the excess reactant is left over? (N = 14.01, H = 1.008)



Step 1 Moles: N₂: $\frac{28.0}{28.02} = 0.9993$ mol; H₂: $\frac{9.00}{2.016} = 4.464$ mol

Step 2 Hydrogen needed by all the nitrogen: $0.9993 \times \frac{3}{1} = 2.998$ mol. We have 4.464 mol, which is more than enough, so **nitrogen is the limiting reactant**.

Step 3 Theoretical yield: $0.9993 \times \frac{2}{1} = 1.999$ mol NH₃, and $1.999 \times 17.03 = 34.04$ g, which is **34.0 g**

Step 4 Hydrogen left over: $4.464 - 2.998 = 1.466$ mol, and $1.466 \times 2.016 = 2.96$ g

Check: 28.0 g + 9.00 g = 37.0 g of reactants, and 34.0 g + 2.96 g = 37.0 g of products and leftovers.

 WORKED EXAMPLE: Percent yield

Question: In the reaction above, the chemist actually collects 27.2 g of ammonia. What is the percent yield?

Step 1 Actual yield = 27.2 g; theoretical yield = 34.04 g (use the unrounded value from the previous calculation)

Step 2 Percent yield = $\frac{27.2}{34.04} \times 100\% = 79.9\%$

Answer: the percent yield is **79.9%**.

Common mistake

 COMMON MISTAKE: Picking the reactant with the smaller mass

In the example, there are 28.0 g of nitrogen but only 9.00 g of hydrogen, yet **nitrogen** is the limiting reactant. Mass alone tells you nothing; you must compare **moles**, using the mole ratio.

 COMMON MISTAKE: Picking the reactant with fewer moles

Even fewer moles doesn't automatically mean limiting. In $2\text{H}_2 + \text{O}_2 \rightarrow 2\text{H}_2\text{O}$ with 1.0 mol of H_2 and 0.6 mol of O_2 , the hydrogen needs only 0.50 mol of oxygen, so **hydrogen** is limiting, even though there are fewer moles of oxygen. Always use the ratio from the balanced equation.

 COMMON MISTAKE: Calculating the yield from the excess reactant

The theoretical yield must come from the **limiting** reactant. Using the excess reactant overestimates it (in the example, hydrogen alone could make 50.7 g of ammonia, but only 34.0 g can form).

 NOTATION NOTE

- The limiting reactant is also called the **limiting reagent**; an excess reactant is also called an **excess reagent**.
- "Yield" can be given as a mass, an amount in moles, or a percentage. Check which one a question asks for.

Remember this

REMEMBER THIS

- The **limiting reactant** runs out first and decides how much product forms.
- Compare reactants in **moles** using the mole ratio, never by mass alone.
- Theoretical yield comes from the limiting reactant.
- Percent yield = $\text{actual} \div \text{theoretical} \times 100\%$.

References

1. Brown, T. L.; LeMay, H. E., Jr.; Bursten, B. E.; Murphy, C. J.; Woodward, P. M.; Stoltzfus, M. W. Chemistry: The Central Science, 15th ed.; Pearson, 2022.

Practise this topic with flashcards and a quiz at chemistryclarity.com/chemistry/limiting-reactant/