

Stoichiometry

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Stoichiometry is the quantitative arithmetic of chemical reactions. It tells you how much product forms from given amounts of reactants, which reactant runs out first, what the percent yield is, and how to scale a reaction up or down. Every working chemist uses stoichiometry constantly, in the lab, in industrial chemistry, in pharmaceutical manufacturing, and in any quantitative reaction analysis.

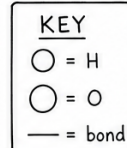
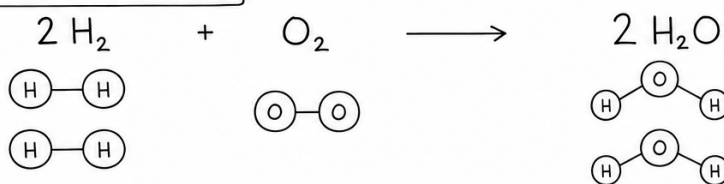
The whole framework rests on conservation of mass (atoms aren't created or destroyed) and the mole concept (counting via Avogadro's number). The five-step recipe, balance the equation, convert mass to moles, apply mole ratio, convert moles back to mass, check limiting reagent, handles essentially every textbook stoichiometry problem.

This study note covers balanced equations, mole ratios, the conversion recipe, limiting reagents and percent yield, solutions and gas stoichiometry, worked examples, common pitfalls, and the practical applications across industries.

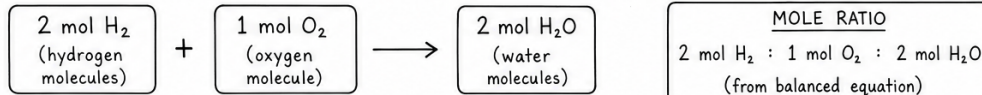
Balanced coefficients compare amounts in **moles**. A limiting-reactant calculation should compare the quantities required by the same reaction ratio.

STOICHIOMETRY

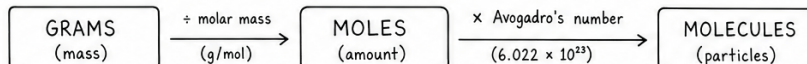
1. BALANCED CHEMICAL EQUATION



2. MOLE RATIO DIAGRAM

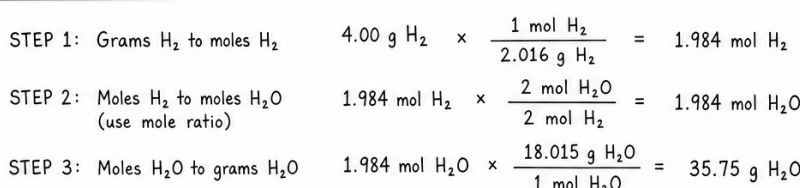


3. CONVERSIONS (ROADMAP)



4. WORKED EXAMPLE

Example: Calculate the mass of water (H₂O) produced from 4.00 g of hydrogen (H₂).



ANSWER: 35.75 g H₂O

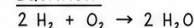
GIVEN:

Mass of H₂ = 4.00 g

FIND:

Mass of H₂O produced

EQUATION:



MOLAR MASSES:

H₂ = 2.016 g/mol

O₂ = 32.00 g/mol

H₂O = 18.015 g/mol

Stoichiometry balanced equation mole ratio modern textbook illustration

Balanced Chemical Equations

A chemical equation describes a reaction by listing reactants on the left and products on the right, with an arrow between them. Balancing requires equal numbers of each atom on both sides, reflecting conservation of mass (Lavoisier, 1774).

Example: hydrogen reacting with oxygen to form water. The skeleton equation $\text{H}_2 + \text{O}_2 \rightarrow \text{H}_2\text{O}$ isn't balanced (2 H and 2 O on left vs 2 H and 1 O on right). Balance by adjusting coefficients (never subscripts): $2\text{H}_2 + \text{O}_2 \rightarrow 2\text{H}_2\text{O}$. Now: 4 H and 2 O on each side. Done.

Balancing harder equations uses systematic methods: balance the most complex molecule first, save H and O for last, use fractional coefficients then multiply through. For **redox** reactions, balance using half-reactions tracking electron transfer separately.

Mole Ratios from the Equation

Coefficients in a balanced equation give mole ratios, not mass ratios. The equation $2\text{H}_2 + \text{O}_2 \rightarrow 2\text{H}_2\text{O}$ means: 2 moles of hydrogen plus 1 mole of oxygen produce 2 moles of water. The mass ratios are different ($4 \text{ g H}_2 + 32 \text{ g O}_2 \rightarrow 36 \text{ g H}_2\text{O}$, same total mass, but very different per-substance

proportions).

Mole ratios are the bridge between reactants and products in any stoichiometry calculation. To find moles of any species in the reaction, multiply moles of any other species by the appropriate ratio from the balanced equation.

This is the chemistry-specific step in every stoichiometry problem. Everything else is unit conversion. Reading mole ratios off a balanced equation is a foundational skill.

Mole Ratio Examples

Once you can extract these ratios on sight, stoichiometry becomes routine arithmetic.

$2\text{H}_2 + \text{O}_2 \rightarrow 2\text{H}_2\text{O}$: $\text{H}_2:\text{O}_2:\text{H}_2\text{O} = 2:1:2$. So 4 mol H_2 requires 2 mol O_2 and produces 4 mol H_2O . $\text{N}_2 + 3\text{H}_2 \rightarrow 2\text{NH}_3$: $\text{N}_2:\text{H}_2:\text{NH}_3 = 1:3:2$. So 5 mol N_2 requires 15 mol H_2 and produces 10 mol NH_3 . $\text{C}_3\text{H}_8 + 5\text{O}_2 \rightarrow 3\text{CO}_2 + 4\text{H}_2\text{O}$: $\text{C}_3\text{H}_8:\text{O}_2:\text{CO}_2:\text{H}_2\text{O} = 1:5:3:4$. So 2 mol propane requires 10 mol O_2 and produces 6 mol CO_2 and 8 mol H_2O .

The Five-Step Recipe

Every standard stoichiometry problem follows the same five steps:

This recipe scales from simple homework problems to industrial reactor design. The arithmetic gets longer for complex reactions, but the structure is always the same.

Balance the chemical equation. Without a balanced equation, the mole ratio is wrong. **Convert given mass (or volume, or particles) of starting substance to moles.** **Apply the mole ratio from the balanced equation** to get moles of the target substance. **Convert moles of target back to mass** (or volume, or particles). **Check for limiting reagent** if more than one reactant is given. The limiting reagent determines how much product can form; excess reactants leave leftovers.

Limiting Reagent

When you mix specified amounts of two or more reactants, one of them runs out before the others. That's the limiting reagent, it determines how much product can form. Excess reagents are whatever's left over.

To find the limiting reagent: convert each reactant to moles, divide by its coefficient, and the smallest result identifies the limiting reagent. The amount of product calculated using the limiting reagent's moles is the maximum theoretical yield.

Example: 4 mol H_2 + 1 mol O_2 in the reaction $2\text{H}_2 + \text{O}_2 \rightarrow 2\text{H}_2\text{O}$. H_2 : $4/2 = 2$. O_2 : $1/1 = 1$. O_2 is limiting

(smaller value). Theoretical yield: $1 \text{ mol O}_2 \times (2 \text{ mol H}_2\text{O} / 1 \text{ mol O}_2) = 2 \text{ mol H}_2\text{O}$. Excess H_2 : $4 \text{ mol provided} - (2 \times 1 \text{ mol used}) = 2 \text{ mol H}_2$ left over.

STOICHIOMETRY · THE FIVE-STEP RECIPE

Every stoichiometry problem follows the same five steps

Mass A $\rightarrow$ moles A $\rightarrow$ moles B (via balanced equation) $\rightarrow$ mass B. The path is always the same; only the numbers change.

1

Balance the chemical equation

Conservation of mass requires equal numbers of each atom on both sides. Use coefficients, never subscripts.

Example: $2\text{H}_2 + \text{O}_2 \rightarrow 2\text{H}_2\text{O}$

2

Convert given mass to moles

Divide by molar mass. The bridge from weighable amounts to atom-counted amounts.

$n_A = m_A / M_A$

3

Apply the mole ratio from the balanced equation

Coefficients give the stoichiometric ratio. This is the only step that uses chemistry; the rest is unit conversion.

$n_B = n_A \times (b/a)$

4

Convert moles of B back to mass

Multiply by molar mass of B. Inverse of step 2.

$m_B = n_B \times M_B$

5

Sanity-check with the limiting reagent

Identify which reactant runs out first; product yield is bounded by it. Excess reagents leave leftovers but don't increase yield.

Five-step stoichiometry recipe with mass-mole conversion

Theoretical and Percent Yield

The theoretical yield is the maximum amount of product that could form, based on stoichiometry from the limiting reagent. The actual yield is what you obtain experimentally, almost always less, due to incomplete reaction, side reactions, losses during purification.

Percent yield = $(\text{actual yield} / \text{theoretical yield}) \times 100\%$. The attainable yield depends on the reaction, equilibrium, side reactions and recovery losses; no single target range applies to every industrial process. Pharmaceutical syntheses sometimes accept much lower per-step yields when the products are valuable enough.

If percent yield exceeds 100%, your product is impure (contaminated with solvent or unreacted starting material). 100% yields are rare and slightly suspicious. Most synthetic chemists treat anything above 95% with skepticism unless they can verify purity.

Worked Example: Limiting Reagent

How much water forms when 6.0 g H_2 reacts with 32 g O_2 ? ($2\text{H}_2 + \text{O}_2 \rightarrow 2\text{H}_2\text{O}$)

Convert to moles. H_2 : $6.0 / 2.016 = 2.98$ mol. O_2 : $32 / 32.0 = 1.0$ mol.

Find limiting reagent. H_2 : $2.98 / 2 = 1.49$. O_2 : $1.0 / 1 = 1.0$. O_2 is limiting.

Use limiting reagent to find moles of product. Mol $\text{H}_2\text{O} = 1.0 \times (2/1) = 2.0$ mol.

Convert moles to mass. Mass $\text{H}_2\text{O} = 2.0 \times 18.02 = 36$ g.

So 36 g of water form. Excess H_2 : 2.98 mol provided $- (2 \times 1.0$ mol used) $= 0.98$ mol $\text{H}_2 = 1.97$ g unreacted.

Solution Stoichiometry

For solutions, use molarity to convert between volume and moles. $n = M \times V$ (moles = molarity $\times$ volume in liters). The rest of the recipe is the same.

Example: how much 0.5 M NaOH solution neutralizes 25 mL of 1.0 M HCl? Reaction: $\text{HCl} + \text{NaOH} \rightarrow \text{NaCl} + \text{H}_2\text{O}$ (1:1 mole ratio).

Moles HCl = $0.025 \text{ L} \times 1.0 \text{ M} = 0.025$ mol. Moles NaOH needed = 0.025 mol (1:1 ratio). Volume NaOH = $0.025 / 0.5 = 0.050 \text{ L} = 50 \text{ mL}$.

Titration calculations follow this same pattern. Acid-base neutralizations, redox titrations, and complexation titrations all reduce to solution stoichiometry.

Gas Stoichiometry

For gases at STP, use 22.4 L/mol to convert between volume and moles. For other conditions, use the ideal gas law $PV = nRT$. The rest of the recipe is unchanged.

Example: how many liters of CO_2 form at STP when 100 g of propane (C_3H_8 , molar mass 44.1 g/mol) burns completely? Reaction: $\text{C}_3\text{H}_8 + 5\text{O}_2 \rightarrow 3\text{CO}_2 + 4\text{H}_2\text{O}$.

Mol $\text{C}_3\text{H}_8 = 100 / 44.1 = 2.27$ mol. Mol $\text{CO}_2 = 2.27 \times 3 = 6.80$ mol. Volume = $6.80 \times 22.4 = 152 \text{ L}$.

Combustion calculations like this are central to chemical engineering, environmental analysis (CO_2 emissions), and energy calculations.

Empirical and Molecular Formulas Revisited

Stoichiometry also goes the other direction, given combustion analysis or elemental composition, work out the formula. Burn a sample, measure CO₂ and H₂O produced, and back-calculate moles of each element.

Example: 0.30 g of an unknown compound burns to produce 0.44 g CO₂ and 0.18 g H₂O. Mol C = 0.44 / 44 = 0.010 (one C per CO₂). Mol H = 2 × (0.18 / 18) = 0.020 (two H per H₂O). Mass of C = 0.010 × 12 = 0.12 g. Mass of H = 0.020 × 1 = 0.020 g. Total = 0.14 g, so the rest (0.30 – 0.14 = 0.16 g) is some other element, often O. Mol O = 0.16 / 16 = 0.010.

Mole ratio C : H : O = 0.010 : 0.020 : 0.010 = 1 : 2 : 1. Empirical formula: CH₂O. Combined with the molar mass (from a separate measurement), the molecular formula follows.

Excess Reagent Tracking

Stoichiometry isn't just about the limiting reagent. You often need to know how much excess remains, since excess reagents have to be removed, recycled, or disposed of.

A reagent may be supplied in excess to improve conversion of another reagent, but cost, equilibrium and separation requirements determine whether that choice is useful. In ammonia synthesis, the balanced reaction $\text{N}_2 + 3\text{H}_2 \rightleftharpoons 2\text{NH}_3$ establishes a 1:3 molar ratio of nitrogen to hydrogen. Unreacted gases can be recycled; this is not a general example of using cheap excess hydrogen.

Tracking excess matters for purification too. The unreacted excess reagent has to be separated from the product, which adds cost and complexity. Choosing the right reagent ratios is part of every reaction design.

Atom Economy and Green Chemistry

Atom economy measures the fraction of the stoichiometric reactant mass incorporated into the desired product. Include the coefficients from the balanced equation: $\text{AE} = \frac{\nu_P M_P}{\sum_i \nu_i M_i} \times 100\%$. Here ν is a stoichiometric coefficient and M is molar mass. The calculation describes the reaction equation; actual yield, excess reagent, solvent use and work-up losses are separate measures.

High atom economy means little waste. Low atom economy means lots of byproducts to dispose of. Green chemistry favors high-atom-economy reactions because they generate less waste, use less raw material, and produce cleaner products.

Example: an addition reaction ($\text{A} + \text{B} \rightarrow \text{AB}$) has 100% atom economy, every atom of A and B ends up in the product. A substitution reaction ($\text{A-X} + \text{Y} \rightarrow \text{A-Y} + \text{X}$) has lower atom economy because X is lost as waste. Modern industrial chemistry is increasingly evaluated on atom economy alongside

yield.

Common Mistakes with Stoichiometry

Forgetting to balance the equation. Unbalanced equations give wrong mole ratios. Always balance first; never skip this step. **Using mass ratios instead of mole ratios.** Coefficients are mole ratios, not mass ratios. You can't reason with grams directly; convert to moles first. **Skipping the limiting reagent check.** When two or more reactants are given, identify the limiting reagent before computing yields. Otherwise you may overestimate. **Confusing molar mass and atomic mass.** Molar mass is g/mol; atomic mass is amu. Numerically equal but different units in different contexts. **Misapplying the ideal gas law.** The 22.4 L/mol value is only at STP. For non-STP conditions, use the full $PV = nRT$. **Not tracking units.** Stoichiometry has many unit conversions. Track units explicitly through every step; they catch most calculation errors.

Industrial and Lab Stoichiometry

Industrial chemistry: reactor sizing, raw material orders, byproduct disposal, all stoichiometry-driven. **Pharmaceuticals:** drug synthesis routes optimized for atom economy, yield, and minimal protecting-group steps. Multi-step syntheses compound stoichiometric losses; high per-step yields matter enormously. **Combustion analysis:** determines C, H, O composition of organic compounds via CO_2 and H_2O production. **Titration:** determines unknown concentration by stoichiometric reaction with known concentration. Universal in analytical chemistry. **Environmental monitoring:** emission inventories, fertilizer runoff calculations, ozone-depletion potential, all stoichiometry. **Food chemistry:** nutritional labeling, fermentation calculations, baking soda vs baking powder ratios, stoichiometric balancing in everyday cooking.

Worked Example: Combustion Analysis

A 1.00 g sample of an organic compound containing only C, H, and O burns completely to produce 2.40 g CO_2 and 0.984 g H_2O . Find the empirical formula.

$\text{Mol C} = 2.40 / 44.0 = 0.0545 \text{ mol}$. $\text{Mass C} = 0.0545 \times 12.01 = 0.655 \text{ g}$. $\text{Mol H} = 2 \times 0.984 / 18.02 = 0.109 \text{ mol}$. $\text{Mass H} = 0.109 \times 1.008 = 0.110 \text{ g}$. $\text{Mass O} = 1.00 - 0.655 - 0.110 = 0.235 \text{ g}$. $\text{Mol O} = 0.235 / 16.00 = 0.0147 \text{ mol}$.

Mole ratio C : H : O = 0.0545 : 0.109 : 0.0147. Divide by smallest (0.0147): 3.71 : 7.41 : 1. Multiply by 4: 14.8 : 29.6 : 4 $\approx \text{C}_{15}\text{H}_{30}\text{O}_4$ (rough). Actually with cleaner experimental numbers, you'd typically land on a clean integer ratio like $\text{C}_3\text{H}_6\text{O}$, a propanol or related compound.

Combustion analysis was the standard method for organic compound formulas before mass spec-

trometry. It still gives elemental composition reliably.

Industrial Stoichiometry

Industrial chemistry runs on stoichiometric calculations at scale. The Haber process produces ammonia ($\text{N}_2 + 3\text{H}_2 \rightarrow 2\text{NH}_3$) at industrial scale. Stoichiometric calculations determine: how much N_2 and H_2 to feed the reactor, how much NH_3 comes out, how much energy is released, what reactor volume is needed for given throughput, and what byproducts (if any) form.

Other major industrial stoichiometric processes: contact process (sulfuric acid, 250M tons/year), Solvay process (sodium carbonate), Ostwald process (nitric acid), and steel production. Each uses stoichiometry-based design at every step. Engineers regularly compute byproduct flows, recycle streams, and conversion efficiencies, all stoichiometry applied at industrial scale.

Worked Example: Solution Stoichiometry

How many mL of 0.50 M H_2SO_4 are needed to neutralize 25.0 mL of 0.40 M NaOH?

Reaction: $\text{H}_2\text{SO}_4 + 2\text{NaOH} \rightarrow \text{Na}_2\text{SO}_4 + 2\text{H}_2\text{O}$. Mole ratio $\text{H}_2\text{SO}_4:\text{NaOH} = 1:2$.

Mol NaOH = $0.0250 \times 0.40 = 0.0100$ mol. Mol H_2SO_4 needed = $0.0100 \times (1/2) = 0.00500$ mol. Volume $\text{H}_2\text{SO}_4 = 0.00500 / 0.50 = 0.0100$ L = 10.0 mL.

About 10 mL of sulfuric acid neutralize 25 mL of sodium hydroxide solution at the given concentrations. Titration calculations like this run constantly in analytical chemistry labs to determine unknown concentrations or to prepare specific solutions.

The [OpenStax chemistry textbooks](#) provide further explanations of the chemical principles used in this note.

A Limiting-Reactant Problem

For $2\text{H}_2 + \text{O}_2 \rightarrow 2\text{H}_2\text{O}$, 2.0 mol of H_2 requires 1.0 mol of O_2 . If only 0.50 mol of O_2 is available, oxygen is limiting. The reaction can produce 1.0 mol of water and consume 1.0 mol of hydrogen, leaving 1.0 mol of H_2 .

Using a rounded water molar mass of 18.0 g/mol, the theoretical water yield is 18.0 g. If 15.3 g is recovered, percent yield is $15.3/18.0 \times 100 = 85.0\%$. A yield calculation should use the limiting reactant's theoretical product, rather than whichever starting mass looks smallest.

The points to check in this example are:

- Balance the reaction before applying mole ratios.
- Compare each available reactant with its required coefficient ratio.
- Theoretical yield and measured recovered yield are distinct.

Check Your Understanding

Try the questions before reading the answers.

Question 1. Which reactant limits the example?

Answer. Oxygen.

Question 2. Does the smaller mass always identify the limiting reactant?

Answer. No. Molar masses and reaction coefficients must be included.

FAQs

What is stoichiometry?

The quantitative arithmetic of chemical reactions, calculating amounts of reactants needed, products formed, limiting reagents, and percent yields from balanced chemical equations. Every quantitative chemistry problem involving reactions uses stoichiometry.

How do I balance a chemical equation?

Adjust coefficients (never subscripts) so each element has equal counts on both sides. Strategy: balance most complex molecule first, save H and O for last, use fractional coefficients then multiply through. For redox reactions, balance using half-reactions tracking electron transfer.

What is a mole ratio?

The ratio of coefficients in a balanced equation, giving how many moles of one substance correspond to each mole of another. The equation $2\text{H}_2 + \text{O}_2 \rightarrow 2\text{H}_2\text{O}$ has $\text{H}_2:\text{O}_2:\text{H}_2\text{O} = 2:1:2$. Mole ratios are the chemistry-specific bridge between reactants and products in stoichiometric calculations.

What is a limiting reagent?

The reactant that runs out first in a reaction with two or more reactants, determining the maximum amount of product. To find it: convert each reactant to moles, divide by its coefficient, and the smallest result identifies the limiting reagent. Excess reagents leave leftovers.

What is theoretical yield?

The maximum amount of product that could form, calculated from the limiting reagent's moles using stoichiometry. Actual yield is what you obtain experimentally; percent yield = actual / theoretical $\times 100\%$. Most reactions yield less than theoretical due to incomplete reaction, side reactions, or losses during purification.

How do I calculate percent yield?

Percent yield = (actual yield / theoretical yield) $\times 100\%$. The theoretical yield comes from stoichiometric calculation using the limiting reagent. Percent yields above 100% indicate impure product (contaminated with solvent or starting material).

What's the five-step stoichiometry recipe?

(1) Balance the equation. (2) Convert given mass to moles via molar mass. (3) Apply the mole ratio from the balanced equation. (4) Convert moles back to mass. (5) Check for limiting reagent if multiple reactants given. Every standard stoichiometry problem follows this pattern.

How do I do gas stoichiometry?

At STP (0°C, 1 atm), one mole of any ideal gas occupies 22.4 L. So convert between gas volume and moles using $V = n \times 22.4 \text{ L}$. For non-STP conditions, use the ideal gas law $PV = nRT$. The rest of the stoichiometry recipe is unchanged.

How do I do solution stoichiometry?

Use molarity to convert between volume and moles: $n = M \times V$ (volume in liters). Otherwise the recipe is the same. Titration calculations are the most common example: find unknown concentration by stoichiometric reaction with known concentration.

What is atom economy?

The fraction of reactant atoms that end up in the desired product. Higher atom economy means less waste. Calculated as (molar mass of product) / (sum of molar masses of all reactants) $\times 100\%$. Green chemistry favors high-atom-economy reactions.

Why is balancing the equation so important?

Because the mole ratio comes from the coefficients, and unbalanced equations give wrong ratios. Without correct stoichiometric coefficients, every subsequent calculation is wrong. Balance first, always; never skip the step even when the answer 'looks obvious.'

Can I do stoichiometry in mass units instead of moles?

Only by converting through moles. The mole is the universal counting unit; mass ratios depend on molar masses and don't follow the equation coefficients directly. Trying to skip the mole step is the most common source of stoichiometric errors.

How does stoichiometry apply to industrial chemistry?

It drives reactor sizing, raw material procurement, byproduct disposal, energy calculations, and economic analysis. Industrial chemists optimize stoichiometric ratios to maximize yield, minimize waste, and reduce cost. The Haber process for ammonia and the contact process for sulfuric acid are textbook examples of stoichiometric optimization.

What's the difference between empirical and molecular formula calculation?

Both use stoichiometry but in opposite directions. Empirical formula: from elemental composition (mass percentages), find the simplest integer ratio of atoms. Molecular formula: combine empirical formula with molar mass to get the actual atom count, which is an integer multiple of empirical.

How is stoichiometry used in industrial chemistry?

To determine reactor sizes, raw material orders, energy requirements, byproduct flows, and economic viability of chemical processes. The Haber process for ammonia, contact process for sulfuric acid, and Solvay process for sodium carbonate all rely on detailed stoichiometric analysis at industrial scale to optimize yield and cost.