

Titration Calculations

How do you turn titration results into a concentration, a mass or a purity?

●●○ Intermediate

Acids & Bases

Last reviewed 6 October 2026

Before you start: it helps to understand [Molarity and Solution Stoichiometry](#) and [Acid–Base Titration Curves](#).

What is it?

In a **titration**, a solution of known concentration (the **standard solution** or **titrant**, in the burette) is added to a measured volume of an unknown solution (the **analyte**, delivered by pipette as an **aliquot**) until the reaction is exactly complete: the **equivalence point**, shown by an indicator colour change (the **end point**).

The volume delivered from the burette is the **titre**. A rough titre is followed by accurate repeats until at least two or three **concordant** titres agree closely (usually within 0.10 mL). Their **mean** is used in the calculation, and the **standard deviation** shows how precise the titrations were.

KEY IDEA

Every titration calculation follows the same path:

known solution (volume × concentration) → **moles** → **mole ratio** from the balanced equation → **moles of unknown** → concentration, mass or purity.

Why does it matter?

- **Quality control.** Titrations check the acidity of vinegar and wine, the vitamin C in juice and the purity of medicines.
- **Precision chemistry.** With a burette and pipette, results to 3 or 4 significant figures are routine.
- **Exams and lab work.** Titration calculations appear in almost every chemistry course and practical assessment.

How does it work?

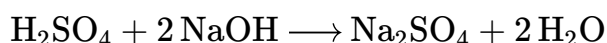
1. Choosing the titres

Leave out the rough titre and any titre that does not agree with the others. Average the concordant ones, and calculate the standard deviation to report precision.

2. The basic calculation

$$n = c \times V \quad c = \frac{n}{V}$$

with V in **litres**. Always write the balanced equation: the mole ratio is often not 1 : 1. For example:



3. Dilutions and aliquots

If the original sample was diluted before titrating, scale the moles back up: a 25.00 mL aliquot from a 100.0 mL flask contains one quarter of the moles in the flask.

4. Back titration

When a substance doesn't react quickly or doesn't dissolve (such as limestone, CaCO_3), add a **known excess** of acid, let it react completely, then titrate the **leftover** acid with a base. The acid that reacted = acid added – acid left over.

THINK OF IT LIKE THIS

A back titration is like finding out how much a friend ate from a full bag of sweets. You can't count what they ate directly, but you know how many were in the bag and you can count what is left. Eaten = started with – left over.

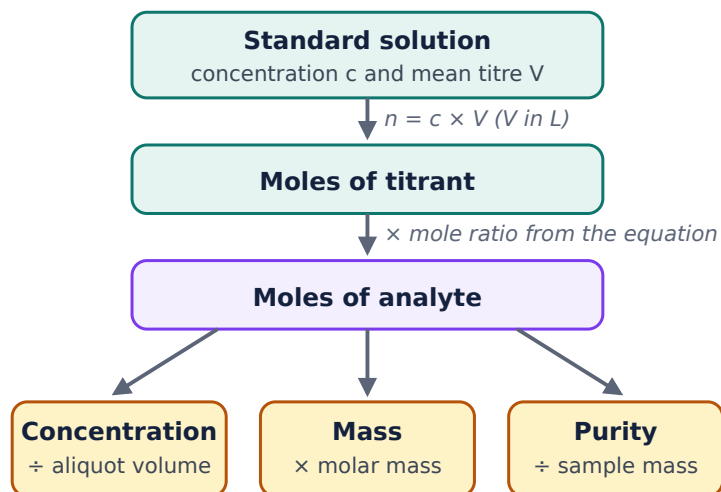
MORE PRECISELY

The end point (indicator change) and the equivalence point (exact stoichiometric amount) are not quite the same; a well-chosen indicator makes the difference smaller than the burette's reading uncertainty. A titre is the difference of two burette readings, each uncertain by about ± 0.05 mL, so the titre is uncertain by about ± 0.10 mL; larger titres give a smaller percent uncertainty. Standard solutions are made from **primary standards**, very pure, stable solids such as sodium carbonate or potassium hydrogen phthalate.

Visualise it

The titration calculation path

Work from what you know to what you want



The calculation path for any titration.

Worked example

WORKED EXAMPLE: Concentration from concordant titres

Question: 25.00 mL of NaOH is titrated with 0.1000 mol/L HCl. Titres: rough 24.10 mL, then 23.45 mL, 23.50 mL and 23.40 mL. Find the concentration of the NaOH.

Step 1 Mean of the three accurate titres:

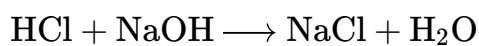
$$\begin{aligned}\bar{V} &= \frac{1}{3}(23.45 + 23.50 \\ &\quad + 23.40) \text{ mL} \\ &= 23.45 \text{ mL}\end{aligned}$$

The standard deviation is 0.05 mL: the titres are precise.

Step 2 Moles of HCl:

$$\begin{aligned}n &= 0.1000 \frac{\text{mol}}{\text{L}} \\ &\quad \times 0.02345 \text{ L} \\ &= 2.345 \times 10^{-3} \text{ mol}\end{aligned}$$

Step 3 The mole ratio is 1 : 1:



so $n(\text{NaOH}) = 2.345 \times 10^{-3} \text{ mol}$.

Step 4 Concentration:

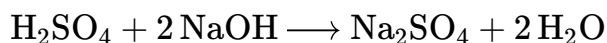
$$\begin{aligned}c \\ &= \frac{2.345 \times 10^{-3} \text{ mol}}{0.02500 \text{ L}} \\ &= 0.09380 \text{ mol/L}\end{aligned}$$

 WORKED EXAMPLE: A diprotic acid

Question: 20.00 mL of sulfuric acid needs 24.80 mL of 0.1500 mol/L NaOH. Find the concentration of the acid.

Step 1 $n(\text{NaOH}) = 0.1500 \text{ mol/L} \times 0.02480 \text{ L} = 3.720 \times 10^{-3} \text{ mol}$

Step 2 The mole ratio is 1 : 2:



so $n(\text{H}_2\text{SO}_4) = \frac{1}{2} \times 3.720 \times 10^{-3} \text{ mol} = 1.860 \times 10^{-3} \text{ mol}$

Step 3 Concentration:

$$\begin{aligned} c &= \frac{1.860 \times 10^{-3} \text{ mol}}{0.02000 \text{ L}} \\ &= 0.09300 \text{ mol/L} \end{aligned}$$

 WORKED EXAMPLE: Dilution and an aliquot: vinegar

Question: 10.00 mL of vinegar is diluted to 100.0 mL. A 25.00 mL aliquot needs 17.60 mL of 0.1000 mol/L NaOH. Find the concentration of ethanoic acid in the original vinegar, and the mass of acid in 100 mL of vinegar.

Step 1 Moles in the aliquot (1 : 1 reaction):

$$\begin{aligned} n &= 0.1000 \frac{\text{mol}}{\text{L}} \\ &\quad \times 0.01760 \text{ L} \\ &= 1.760 \times 10^{-3} \text{ mol} \end{aligned}$$

Step 2 The flask holds $\frac{100.0}{25.00} = 4$ aliquots: $4 \times 1.760 \times 10^{-3} \text{ mol} = 7.040 \times 10^{-3} \text{ mol}$, all from the 10.00 mL of vinegar.

Step 3 Concentration:

$$\begin{aligned} c &= \frac{7.040 \times 10^{-3} \text{ mol}}{0.01000 \text{ L}} \\ &= 0.7040 \text{ mol/L} \end{aligned}$$

Step 4 Mass in 100 mL (0.1000 L): $0.7040 \text{ mol/L} \times 0.1000 \text{ L} \times 60.05 \text{ g/mol} = 4.23 \text{ g}$, so the vinegar is about 4.2 % acid.

 WORKED EXAMPLE: A back titration: purity of limestone

Question: 0.500 g of limestone is added to 50.00 mL of 0.2000 mol/L HCl (an excess). The leftover HCl needs 12.40 mL of 0.1000 mol/L NaOH. Find the percent of CaCO_3 in the limestone.

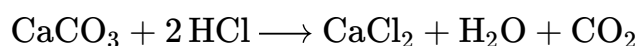
Step 1 HCl added: $0.2000 \text{ mol/L} \times 0.05000 \text{ L} = 1.000 \times 10^{-2} \text{ mol}$

Step 2 HCl left over (= NaOH used): $0.1000 \text{ mol/L} \times 0.01240 \text{ L} = 1.240 \times 10^{-3} \text{ mol}$

Step 3 HCl that reacted with the limestone:

$$\begin{aligned} &1.000 \times 10^{-2} \text{ mol} - 1.240 \\ &\quad \times 10^{-3} \text{ mol} \\ &= 8.760 \times 10^{-3} \text{ mol} \end{aligned}$$

Step 4 The mole ratio is 1 : 2:



$$\text{so } n(\text{CaCO}_3) = \frac{1}{2} \times 8.760 \times 10^{-3} \text{ mol} = 4.380 \times 10^{-3} \text{ mol}$$

Step 5 Mass and purity:

$$\begin{aligned} m &= 4.380 \times 10^{-3} \text{ mol} \\ &\quad \times 100.09 \frac{\text{g}}{\text{mol}} \\ &= 0.4384 \text{ g} \\ \text{purity} &= \frac{0.4384 \text{ g}}{0.500 \text{ g}} \\ &\quad \times 100\% = 87.7\% \end{aligned}$$

Common mistake

 COMMON MISTAKE: Including the rough titre

The rough titre is only a guide to where the end point is. Average only the accurate, concordant titres.

 COMMON MISTAKE: Forgetting the mole ratio

Sulfuric acid is diprotic: 1 mol H_2SO_4 reacts with 2 mol NaOH. Using 1 : 1 doubles the answer.

 **COMMON MISTAKE:** Leaving volumes in mL

$n = c \times V$ needs V in **litres** when c is in mol/L: 23.45 mL = 0.02345 L.

 NOTATION NOTE

- Record burette readings to two decimal places (e.g. 23.45 mL), and report the mean titre the same way.
- Concentrations from careful titrations are usually given to 4 significant figures.

Remember this

 REMEMBER THIS

- Use the mean of concordant titres (within about 0.10 mL); leave out the rough titre.
- Path: $n = cV$ (known) $\rightarrow$ mole ratio $\rightarrow$ moles of unknown $\rightarrow c = n/V$, mass or purity.
- Scale up for dilutions: moles in flask = moles in aliquot $\times$ (flask volume $\div$ aliquot volume).
- Back titration: reacted = added – left over.

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/titration-calculations/