

Acid–Base Titration Curves

What does a titration curve show?

●●● Advanced

Acids & Bases

Last reviewed 3 October 2026

Before you start: it helps to understand [Molarity and Solution Stoichiometry](#), [Weak Acids and \$K_a\$ and Buffers](#).

What is it?

In a **titration**, a solution of known concentration (the **titrant**, often NaOH) is added gradually to a measured volume of another solution (for example an acid). A **titration curve** is a graph of the **pH** of the mixture against the **volume of titrant added**.

Key points on the curve:

- **Equivalence point:** exactly enough base has been added to react with all the acid (moles of OH^- = moles of acid H^+ available).
- **Half-equivalence point** (weak acids): half the acid has been neutralized.
- **End point:** where an **indicator** changes colour. A good indicator makes the end point match the equivalence point.

KEY IDEA

Near the equivalence point the pH changes very sharply: one drop can move it by several units. That jump is what makes titrations so precise.

Why does it matter?

- **Finding unknown concentrations:** titrations are one of the most accurate ways to measure how much acid or base a sample contains, from vinegar to medicines.
- **Measuring pK_a :** the half-equivalence point of a weak acid titration gives its pK_a directly.
- **Choosing the right indicator:** the shape of the curve tells you which indicator will change colour at the right moment.

How does it work?

1. Strong acid + strong base

- The pH starts **low** (for 0.100 M HCl, pH 1.00) and rises slowly.

- Near equivalence, it rises **steeply**, passing through **pH 7.00** at the equivalence point (the salt formed, such as NaCl, is neutral).
- After equivalence, excess OH^- makes the pH level off at a high value.

2. Weak acid + strong base

- The pH starts **higher** (0.100 M acetic acid: pH 2.88), because a weak acid ionizes only partly.
- A flatter **buffer region** follows, where the solution contains both the weak acid and its conjugate base.
- At the **half-equivalence point**, $[\text{HA}] = [\text{A}^-]$, so **pH = pKa** (4.74 for acetic acid).
- The **equivalence point is above 7** (here pH 8.72), because the conjugate base (acetate) is weakly basic.

3. Choosing an indicator

Pick an indicator whose colour-change range lies within the **steep part** of the curve:

Indicator	Changes colour between pH (approx.)	Suits
Methyl orange	3.1 – 4.4	strong acid + weak base
Bromothymol blue	6.0 – 7.6	strong acid + strong base
Phenolphthalein	8.2 – 10.0	weak acid + strong base (also strong + strong)

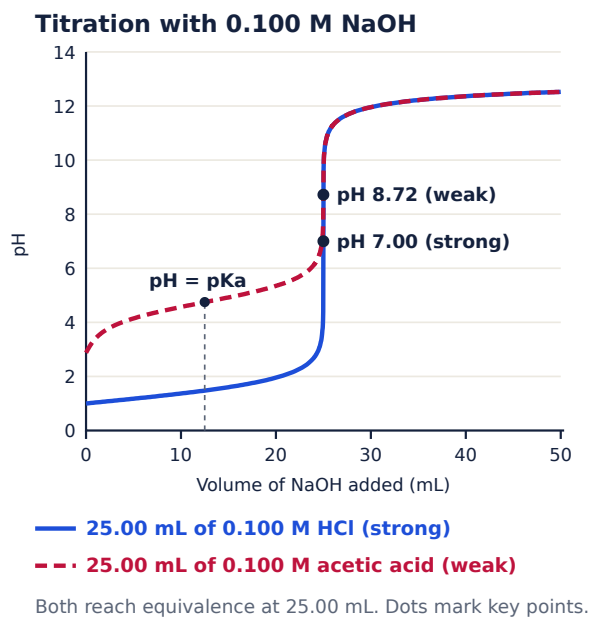
∞ THINK OF IT LIKE THIS

Filling a bucket that has a hidden tipping point: for a long time each cup of water barely changes anything, then suddenly one cup tips it over. In a titration, the equivalence point is the tipping point, and the indicator is the alarm that goes off when it tips.

🔍 MORE PRECISELY

For a strong acid–strong base titration the steep jump is so large that phenolphthalein and bromothymol blue both work well. For a weak acid–weak base titration there is no sharp jump, so indicators are unreliable and a pH meter is used. The equivalence point of a strong acid–weak base titration (e.g. HCl with ammonia) is **below 7**, because the conjugate acid (NH_4^+) is weakly acidic.

Visualise it



Strong vs weak acid titrated with NaOH (calculated curves).

Worked example

WORKED EXAMPLE: Equivalence volume and pH before equivalence

Question: 25.00 mL of 0.100 M HCl is titrated with 0.100 M NaOH. (a) What volume of NaOH reaches equivalence? (b) What is the pH after 12.50 mL of NaOH has been added?

Step 1 Moles of HCl = $0.100 \times 0.02500 = 0.00250$ mol, so NaOH needed = 0.00250 mol;
volume = $0.00250 \div 0.100 = \mathbf{25.0 \text{ mL}}$

Step 2 After 12.50 mL: NaOH added = 0.00125 mol; HCl left = $0.00250 - 0.00125 = 0.00125$ mol

Step 3 Total volume = 37.50 mL, so $[\text{H}_3\text{O}^+] = 0.00125 \div 0.03750 = 0.0333 \text{ M}$ and
pH = **1.48**

 WORKED EXAMPLE: pH at the half-equivalence and equivalence points of a weak acid

Question: 25.00 mL of 0.100 M acetic acid is titrated with 0.100 M NaOH. Find the pH (a) at the half-equivalence point and (b) at the equivalence point. ($K_a = 1.8 \times 10^{-5}$, $pK_a = 4.74$)

Step 1 **Half-equivalence** (12.50 mL): $[HA] = [A^-]$, so $pH = pK_a = 4.74$

Step 2 **Equivalence** (25.00 mL): all the acid has become acetate: 0.00250 mol in 50.00 mL = 0.0500 M

Step 3 Acetate is a weak base: $K_b = K_w/K_a = 5.6 \times 10^{-10}$, so $[OH^-] \approx \sqrt{5.6 \times 10^{-10} \times 0.0500} = 5.3 \times 10^{-6} \text{ M}$

Step 4 $pOH = 5.28$, $pH = 8.72$

Common mistake

 **COMMON MISTAKE:** Assuming every equivalence point is at pH 7

Only strong acid + strong base titrations reach equivalence at pH 7. Weak acid + strong base: **above** 7. Strong acid + weak base: **below** 7.

 **COMMON MISTAKE:** Confusing equivalence with neutral

“Equivalence” means the moles of acid and base match the balanced equation, not that the solution is neutral.

 **COMMON MISTAKE:** Choosing an indicator outside the steep region

Methyl orange changes colour around pH 3–4, well before the equivalence point of a weak acid titration (pH 8.72), so it would give an end point that is far too early.

 NOTATION NOTE

- The solution in the burette is the **titrant**; the solution being measured is the **analyte**.
- Some courses use “end point” and “equivalence point” interchangeably; strictly, the end point is where the indicator changes.

Remember this

REMEMBER THIS

- Equivalence: moles of base = moles of acid (from the balanced equation).
- Strong + strong: equivalence at pH 7. Weak acid + strong base: above 7. Strong acid + weak base: below 7.
- Half-equivalence of a weak acid: $\text{pH} = \text{pK}_a$.
- Choose an indicator that changes colour within the steep jump.

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.

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