ANSWER KEY. pH values match the simulation readouts for the settings given in each step.

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Titration Virtual Lab

Chemistry and AP Chemistry | ANSWER KEY

Open **activities.zornscience.org/titration-lab/** and follow each step. Show your work with units.

Objective

Titrate a **strong acid** and a **weak acid** with a strong base. Use the pH curves to find the **equivalence point**, the **pK_a** of the weak acid, and the best **indicator**.

NGSS alignment: HS-PS1-7. Use mathematical representations to support the claim that atoms, and therefore mass, are conserved during a chemical reaction.

Practices and concepts: Analyzing and interpreting data; Using mathematics and computational thinking; Patterns; Cause and effect.

AP alignment: 4.6 Introduction to Titration; 8.5 Acid-Base Titrations; 8.7 pH and pK_a. Skills 3.A (represent data with graphs) and 5.D (identify information presented graphically to solve a problem).

Meet the simulation

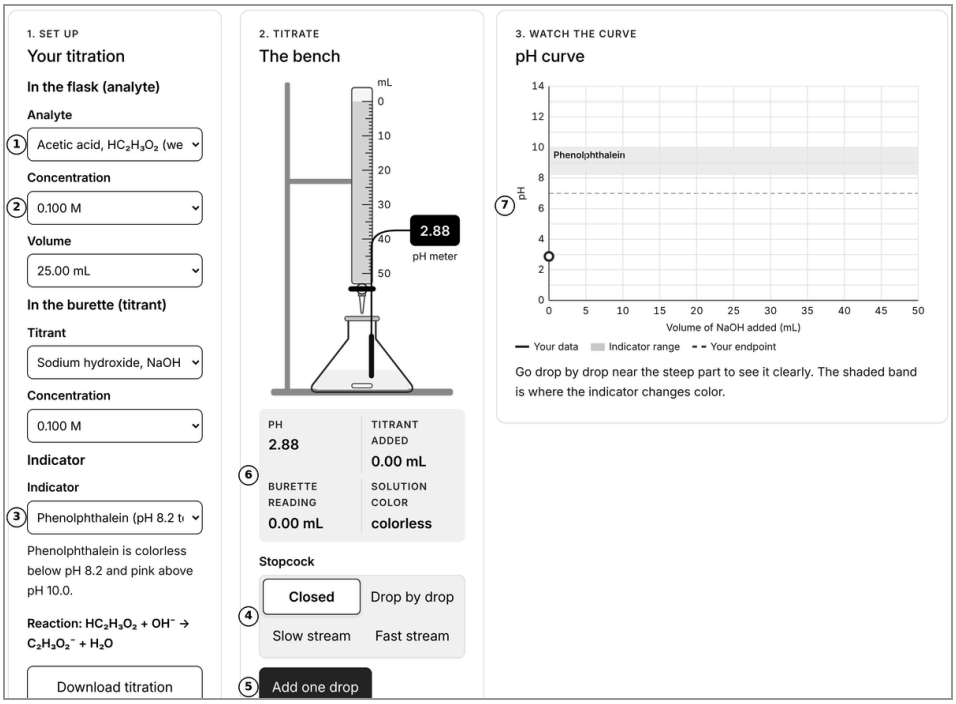


Figure 1. The Virtual Titration Simulator.

(1) Analyte: the acid or base in the flask.

(2) Concentration of the analyte. (Volume and the titrant are set below it.)

(3) Indicator. Its range is the shaded band on the graph.

(4) Stopcock: closed, drop by drop, slow stream or fast stream.

(5) Adds exactly one drop (0.05 mL).

(6) Readouts: pH, titrant added, burette reading, solution color.

(7) The pH curve your data draws.

Pre-lab

1. Define **analyte**, **titrant**, **equivalence point**, and **endpoint**.

Analyte: the solution in the flask whose amount is being found. Titrant: the solution of known concentration in the burette. Equivalence point: moles of titrant added equal moles of analyte. Endpoint: where the indicator changes color.

2. What volume of 0.100 M NaOH exactly neutralizes 25.00 mL of 0.100 M HCl?

mol HCl = (0.100 mol/L)(0.02500 L) = 0.00250 mol = mol NaOH needed
V = 0.00250 mol / 0.100 mol/L = 0.0250 L = <u>25.0 mL</u>

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3. Predict: will the pH at the equivalence point be 7, above 7, or below 7 for (a) HCl with NaOH and (b) acetic acid with NaOH?

(a) 7 (b) above 7. Any prediction with a reason is acceptable; Parts A and B test it.

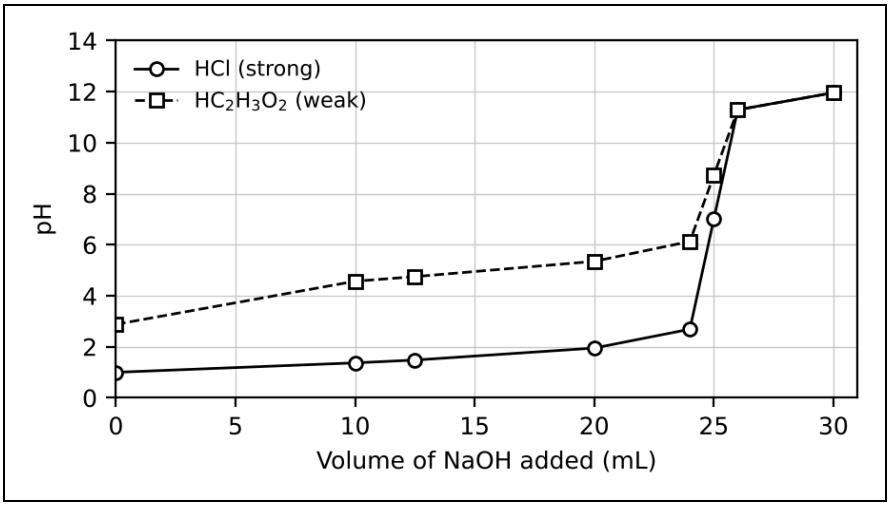
Part A and Part B. Build two pH curves

1. **Part A:** Set Analyte = **Hydrochloric acid**, Concentration = **0.100 M**, Volume = **25.00 mL**, Titrant = **Sodium hydroxide**, titrant Concentration = **0.100 M**, Indicator = **Phenolphthalein**.
2. Use **Slow stream** to get close to each volume in the table, select **Closed**, then select **Add one drop** until **Titrant added** is within 0.05 mL of the volume. Record the pH.
3. From 24 mL on, add one drop at a time. When the solution turns pale pink and stays pink, select **Record endpoint**. Then finish to 30 mL.
4. Select **Show key points** and record the pH at the equivalence point. Select **Hide key points**.
5. **Part B:** Change the Analyte to **Acetic acid** (all other settings the same). The flask starts over. Repeat steps 2 to 4, but leave the key points showing.

NaOH added (mL)	0.00	10.00	12.50	20.00	24.00	26.00	30.00	Endpoint (mL)	pH at equivalence
Part A pH, HCl	1.00	1.37	1.48	1.95	2.69	11.29	11.96	25.05	7.00
Part B pH, acetic acid	2.88	4.57	4.75	5.35	6.13	11.29	11.96	25.00	8.72

Values are at the exact volumes; a reading 0.05 mL away differs by 0.02 or less. Endpoint: the first drop that stays pale pink, anywhere from 24.99 to 25.05 mL, is correct.

4. Plot both curves on the grid (pH on the y-axis, NaOH added on the x-axis). Label each curve and both axes.



Both curves rise sharply at 25 mL. The acetic acid curve starts higher, has a flat buffer region near 12.5 mL, and jumps to 8.72 instead of 7.00. After equivalence the two curves match (excess OH^-).

5. Compare the two curves. Give three differences, using numbers from your table.

Starting pH: 1.00 (HCl) vs 2.88 (acetic). Before equivalence the acetic curve has a flat buffer region (4.57 to 5.35 from 10 to 20 mL). pH at equivalence: 7.00 vs 8.72. The jump at equivalence is smaller for the weak acid.

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Part C. Key points and indicators (acetic acid)

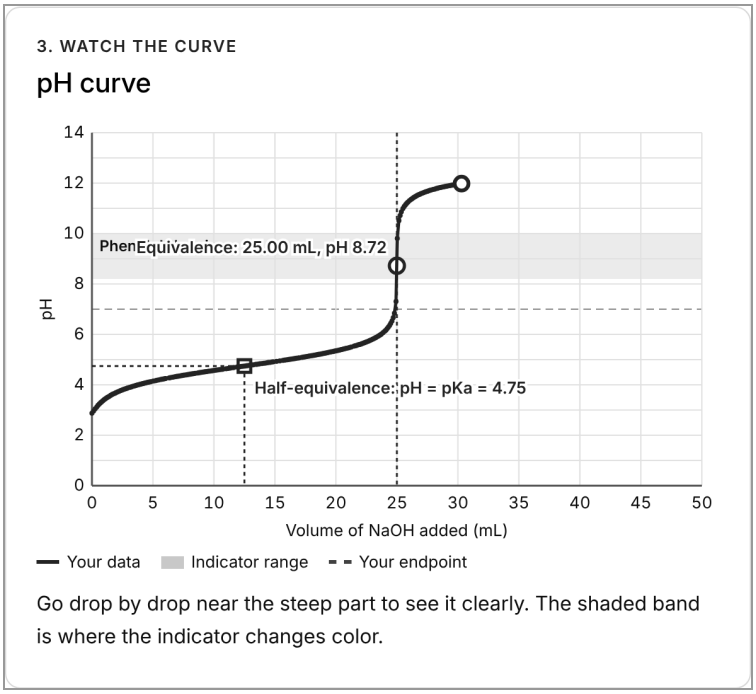


Figure 2. Show key points marks the equivalence point (circle) and the half-equivalence point (square).

1. Keep your acetic acid titration from Part B on screen.
2. Make sure **Show key points** is on. Read the equivalence and half-equivalence points on the graph and in the Key points list below it.
3. Change only the Indicator to each one in the table. Your data stay. Read the indicator sentence in the Key points list each time.

Indicator	Color change range (pH)	Changes color at about (mL)	Good or poor choice?
Phenolphthalein	8.2 to 10.0	**25.01**	**Good**
Bromothymol blue	6.0 to 7.6	**24.78**	**Poor (0.22 mL early)**
Methyl red	4.4 to 6.2	**19.55**	**Poor (5.45 mL early)**
Methyl orange	3.1 to 4.4	**2.25**	**Poor (22.75 mL early)**

6. Record the equivalence point and the half-equivalence point (volume and pH). What is the pK_a of acetic acid from your curve?

Equivalence: 25.00 mL, pH 8.72. Half-equivalence: 12.50 mL, pH 4.75. pK_a = pH at half-equivalence = 4.75 (the table value is 4.74).

7. Use your indicator table to state a rule for choosing an indicator. Why is methyl red fine for HCl but poor for acetic acid?

Choose an indicator whose color change range includes the pH at the equivalence point. For HCl the pH jumps from 3.70 at 24.90 mL to 10.00 at 25.05 mL, so methyl red (4.4 to 6.2) changes right at 25.00 mL. For acetic acid the pH passes 4.4 to 6.2 slowly in the buffer region, long before equivalence (8.72).

8. Write the net ionic equation for the Part B reaction, with states. Then explain why the pH at the equivalence point is above 7.

$HC_2H_3O_2(aq) + OH^-(aq) \rightarrow C_2H_3O_2^-(aq) + H_2O(l)$
At equivalence only the conjugate base $C_2H_3O_2^-$ is left. It takes H^+ from water and makes OH^-, so the solution is basic.

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Analysis

9. A student titrates 25.00 mL of acetic acid of unknown concentration with 0.100 M NaOH. The equivalence point is at 31.20 mL. Find the concentration of the acid.

mol NaOH = (0.100 mol/L)(0.03120 L) = 0.003120 mol = mol $HC_2H_3O_2$ (1:1)
M = 0.003120 mol / 0.02500 L = <u>0.125 M</u>

10. (AP extension) Use your pK_a to find K_a. Then calculate the starting pH of 0.100 M acetic acid and compare it to your reading at 0.00 mL.

$K_a = 10^{-4.75} = 1.8 \times 10^{-5}$
$[H^+] = \sqrt{(K_aC)} = \sqrt{((1.8 \times 10^{-5})(0.100))} = 1.3 \times 10^{-3}$ M, pH = 2.87. The simulation shows 2.88 (it solves exactly, without the approximation).

11. (AP extension) Use the Henderson-Hasselbalch equation to predict the pH after 20.00 mL of NaOH is added to the acetic acid. Compare with your table.

Start: (0.100)(0.02500 L) = 0.00250 mol HA. Added: (0.100)(0.02000 L) = 0.00200 mol OH^-. HA left = 0.00050 mol, A^- = 0.00200 mol. pH = 4.74 + log(0.00200/0.00050) = 4.74 + 0.60 = <u>5.34</u>, which matches 5.35 in the table.

12. Give two ways a real titration in a school lab could give a less accurate equivalence volume than this simulation.

Any two: overshooting the endpoint (a drop is often more than 0.05 mL, or a drop hangs on the tip); reading the burette meniscus wrong; a solution concentration that is not exactly 0.100 M (NaOH absorbs CO_2); poor mixing; judging the faint pink color by eye.

Conclusion

13. Write a claim that answers the objective: how do the titration curves of a strong acid and a weak acid differ, and how can a curve give the pK_a and the right indicator? Support it with evidence (numbers) and reasoning.

Claim: Both acids reach equivalence at the same volume, but the weak acid starts at a higher pH, has a buffer region, and has a basic equivalence point. Its pK_a is the pH at half-equivalence, and the indicator must change near the equivalence pH.
Evidence: Equivalence at 25.00 mL for both; equivalence pH 7.00 (HCl) and 8.72 (acetic); pH 4.75 at 12.50 mL; phenolphthalein changed at 25.01 mL while methyl red changed at 19.55 mL.
Reasoning: Equal moles of acid need equal moles of NaOH. At half-equivalence [HA] = $[A^-]$, so pH = pK_a. The conjugate base makes the equivalence point basic, so an indicator with a range near 8.7 marks it.

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