

Weak Acids and Ka

How do you calculate the pH of a weak acid?

●●● Advanced

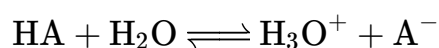
Acids & Bases

Last reviewed 3 October 2026

Before you start: it helps to understand [pH](#) and the [pH Scale](#).

What is it?

A **weak acid** gives up its proton to water only **partly**. At any moment most of its molecules are still intact, and only a small fraction are ionized:



The **acid dissociation constant**, K_a , measures how far this goes:

$$K_a = \frac{[\text{H}_3\text{O}^+][\text{A}^-]}{[\text{HA}]}$$

The larger K_a , the stronger the acid. Chemists often use $\text{p}K_a = -\log K_a$: the **smaller** the $\text{p}K_a$, the **stronger** the acid.

KEY IDEA

A weak acid sets up an **equilibrium**. Because only a little of it ionizes, its pH is higher (less acidic) than a strong acid of the same concentration.

Why does it matter?

- **Most acids are weak.** Vinegar (acetic acid), citrus fruits (citric acid), and the acids in many medicines and foods are weak acids.
- **Biology runs on weak acids and bases.** Amino acids, the phosphate groups in DNA and the carbonic acid in blood all behave as weak acids or bases.
- **K_a is the key to buffers and titrations**, the next two lessons.

How does it work?

1. Comparing acid strength

Acid	K_a (25 °C)	pKa
Hydrofluoric acid, HF	6.8×10^{-4}	3.17
Formic acid, HCOOH	1.8×10^{-4}	3.74
Acetic acid, CH ₃ COOH	1.8×10^{-5}	4.74
Hydrocyanic acid, HCN	4.9×10^{-10}	9.31

HF is the strongest of these weak acids; HCN is the weakest.

2. Calculating the pH of a weak acid

Use an **ICE table** (Initial, Change, Equilibrium). For an acid of initial concentration C , let x = amount that ionizes:

	HA	H ₃ O ⁺	A ⁻
Initial	C	0	0
Change	$-x$	$+x$	$+x$
Equilibrium	$C - x$	x	x

So $K_a = \frac{x^2}{C - x}$. When the acid is weak, x is tiny compared with C , so $C - x \approx C$, and:

$$x = [\text{H}_3\text{O}^+] \approx \sqrt{K_a \times C}$$

Check the approximation: if x is less than 5% of C , it is acceptable. Otherwise, solve the quadratic equation instead.

3. Percent ionization

$$\% \text{ ionized} = \frac{[\text{H}_3\text{O}^+]}{C} \times 100\%$$

4. Weak bases

Weak bases work the same way with K_b and OH^- . For ammonia, $\text{NH}_3 + \text{H}_2\text{O} \rightleftharpoons \text{NH}_4^+ + \text{OH}^-$, and $[\text{OH}^-] \approx \sqrt{K_b \times C}$. For a conjugate acid–base pair, $K_a \times K_b = K_w$.

∞ THINK OF IT LIKE THIS

Imagine a dance hall where couples (HA) can split up (H^+ and A^-) and pair up again. With a strong acid, every couple splits and nobody pairs again. With a weak acid, couples keep splitting and re-forming, but at any moment almost all are still together. K_a tells you how many are split at any moment.

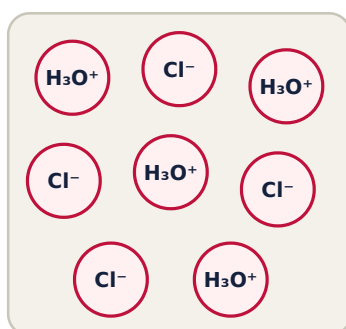
🔍 MORE PRECISELY

K_a values depend on temperature (the table is for 25 °C). The “5% rule” is a convention, not a law: it keeps the error in x small. A weak acid ionizes **more** (as a percentage) when it is more dilute, even though its pH rises. Strictly, equilibrium constants use activities, not concentrations, but concentrations work well for dilute solutions.

Visualise it

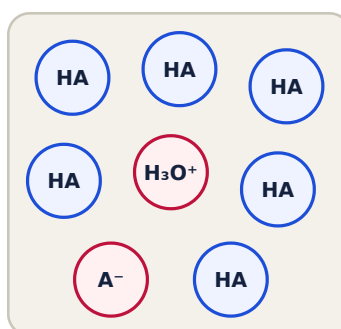
Strong acid (HCl)

ionizes completely



Weak acid (HA)

ionizes only slightly



ions



un-ionized acid molecules

Not to scale: in 0.20 M acetic acid only about 1 molecule in 100 is ionized. Most stay as HA.

A strong acid ionizes completely; a weak acid mostly stays as molecules.

Worked example

WORKED EXAMPLE: pH of a weak acid

Question: Calculate the pH and percent ionization of 0.20 M acetic acid. ($K_a = 1.8 \times 10^{-5}$)

Step 1 $[\text{H}_3\text{O}^+] \approx \sqrt{1.8 \times 10^{-5} \times 0.20} = 1.9 \times 10^{-3} \text{ M}$

Step 2 Check: 1.9×10^{-3} is about 0.9% of 0.20, below 5%, so the approximation is fine.

Step 3 $\text{pH} = -\log(1.9 \times 10^{-3}) = \mathbf{2.72}$; percent ionization $\approx \mathbf{0.9\%}$

Compare: 0.20 M HCl (strong) has pH 0.70.

WORKED EXAMPLE: K_a from a measured pH

Question: A 0.10 M solution of a weak acid has pH 3.00. Calculate K_a .

Step 1 $[\text{H}_3\text{O}^+] = [\text{A}^-] = 10^{-3.00} = 1.0 \times 10^{-3} \text{ M}$

Step 2 $[\text{HA}] = 0.10 - 0.0010 = 0.099 \text{ M}$

Step 3 $K_a = \frac{(1.0 \times 10^{-3})^2}{0.099} = \mathbf{1.0 \times 10^{-5}}$

WORKED EXAMPLE: pH of a weak base

Question: Calculate the pH of 0.20 M ammonia. ($K_b = 1.8 \times 10^{-5}$)

Step 1 $[\text{OH}^-] \approx \sqrt{1.8 \times 10^{-5} \times 0.20} = 1.9 \times 10^{-3} \text{ M}$

Step 2 $\text{pOH} = 2.72$, so $\text{pH} = 14.00 - 2.72 = \mathbf{11.28}$

Common mistake

COMMON MISTAKE: Treating a weak acid like a strong acid

For 0.20 M acetic acid, $[\text{H}_3\text{O}^+]$ is **not** 0.20 M. That would give pH 0.70 instead of 2.72. Only strong acids ionize completely.

COMMON MISTAKE: Mixing up K_a and $\text{p}K_a$

A **larger** K_a means a stronger acid, but a **smaller** $\text{p}K_a$ means a stronger acid. Check which one you are comparing.

 COMMON MISTAKE: Using the approximation without checking

If x is more than 5% of the starting concentration (for example with fairly strong weak acids such as HF, or very dilute solutions), the shortcut is not accurate enough, so solve the quadratic.

 NOTATION NOTE

- K_a is also called the **acidity constant** or **acid ionization constant**.
- Pure liquid water is left out of the K_a expression.
- Some courses write $\text{HA} \rightleftharpoons \text{H}^+ + \text{A}^-$ instead of including water; the K_a expression is the same.

Remember this

 REMEMBER THIS

- Weak acids ionize only partly: $\text{HA} + \text{H}_2\text{O} \rightleftharpoons \text{H}_3\text{O}^+ + \text{A}^-$.
- $K_a = [\text{H}_3\text{O}^+][\text{A}^-]/[\text{HA}]$; larger K_a (smaller $\text{p}K_a$) = stronger acid.
- Shortcut: $[\text{H}_3\text{O}^+] \approx \sqrt{K_a C}$, valid if under 5% ionizes.
- Weak bases: same method with K_b and OH^- ; $K_a K_b = K_w$.

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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