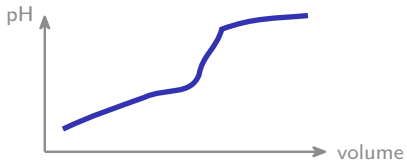
Acids and Bases

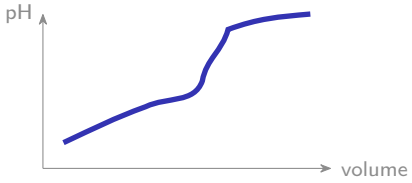
AP Chemistry ▪ Unit 8

AP Chemistry



What we will cover

1. Bronsted-Lowry and pH
2. Strong acids and bases
3. Weak acid and base equilibria
4. Buffers and Henderson-Hasselbalch
5. Acid-base titrations
6. Structure and pKa



1 pH

Bronsted-Lowry and pH

Bronsted-Lowry 布朗斯特-劳里: an acid donates a proton, a base accepts one – forming **conjugate pairs** 共轭酸碱对.

Water and pH

$$K_w = [\mathrm{H}^+][\mathrm{OH}^-] = 10^{-14}$$

$$\mathrm{pH} = -\log[\mathrm{H}^+]$$

pH measures acidity of a solution

2

Strong

Strong acids and bases

A **strong acid** 强酸 ionizes completely, so $[\mathrm{H}^+]$ equals its concentration.

Two logarithms

$$\mathrm{pH} = -\log[\mathrm{H}^+],\ \mathrm{pOH} = -\log[\mathrm{OH}^-]$$

$$\mathrm{pH} + \mathrm{pOH} = 14$$

Worked example

$0.010\,\mathrm{M}$ HCl:

$$[\mathrm{H}^+] = 0.010 \Rightarrow \mathrm{pH} = 2$$

$$\mathrm{pOH} = 12,\ [\mathrm{OH}^-] = 10^{-12}$$

3

Weak

Weak acid and base equilibria

A **weak acid** 弱酸 only partly ionizes:

Ionization constant

$$K_a = \frac{[\mathrm{H^+}][\mathrm{A^-}]}{[\mathrm{HA}]}$$

Larger K_a = stronger. For a conjugate pair, $K_a K_b = K_w$.

Worked example

$0.10\,\mathrm{M}$, $K_a = 10^{-5}$:

$$K_a \approx \frac{x^2}{0.10} \Rightarrow x = 10^{-3}$$

$$\mathrm{pH} = 3$$

A strong acid at $0.10\,\mathrm{M}$ would give pH 1.

4 Buffers

Buffers and Henderson-Hasselbalch

A **buffer** 缓冲溶液 holds pH steady with both a weak acid and its conjugate base:

- added acid $\Rightarrow$ conjugate base soaks up H^+
- added base $\Rightarrow$ weak acid soaks up OH^-

Henderson-Hasselbalch

$$\mathrm{pH} = \mathrm{p}K_a + \log \frac{[\mathrm{A}^-]}{[\mathrm{HA}]}$$

Equal amounts $\Rightarrow \mathrm{pH} = \mathrm{p}K_a$.
Capacity 缓冲容量 grows with total concentration.

Worked example – the Henderson-Hasselbalch equation

A buffer holds 0.20 M acetic acid ($\mathrm{p}K_a = 4.74$) and 0.30 M acetate. Find its pH.

- $\mathrm{pH} = \mathrm{p}K_a + \log \dfrac{[\mathrm{A}^-]}{[\mathrm{HA}]}$
- $= 4.74 + \log \dfrac{0.30}{0.20} = 4.74 + 0.18$
- $= 4.92$

When the two concentrations are equal the log is zero and $\mathrm{pH} = \mathrm{p}K_a$ exactly. That is the point of maximum **buffer capacity** 缓冲容量.

pH and Solubility

The solubility of a salt with a basic anion **rises** in acidic solution: the added H^+ reacts with the anion.

> So pH can control whether an ionic solid dissolves.

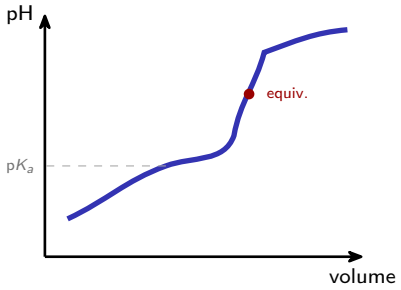
5 Titrations

Acid-base titrations

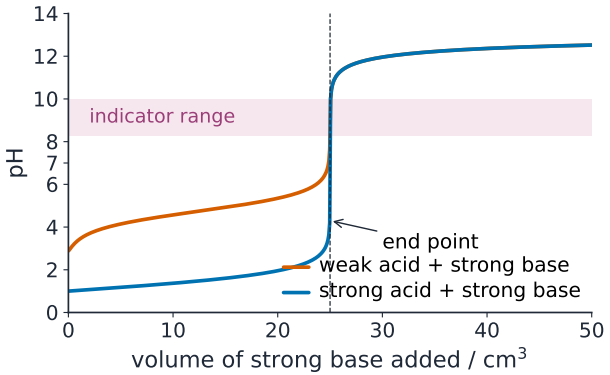
A **titration curve** 滴定曲线 plots pH against titrant volume:

- the **equivalence point** 等当点 is the steep jump
- at the **half-equivalence point** 半等当点, $\mathrm{pH} = \mathrm{p}K_a$

Strong-strong: equivalence at pH 7.
Weak acid-strong base: above 7.



A titration curve has a steep jump



A titration curve has a steep jump near the equivalence point

6 Structure

Structure and pKa

Acid strength comes down to how easily the acidic H leaves, and how stable the **conjugate base** 共轭碱 is:

- **binary acids** 无氧酸: stronger *down* a group (weaker bond)
- **oxyacids** 含氧酸: more oxygens ⇒ stronger

$\mathrm{p}K_a = -\log K_a$

Smaller $\mathrm{p}K_a$ = stronger acid.

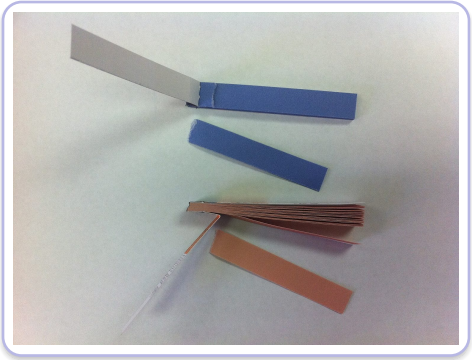
$\mathrm{pH} < \mathrm{p}K_a$: the acid form dominates; $\mathrm{pH} > \mathrm{p}K_a$: the conjugate base dominates.

Why one acid is stronger than another

the principle	an acid is stronger when its conjugate base is more stable
electronegativity	more electronegative atoms nearby pull electron density away and stabilise the negative charge
resonance	a charge spread over several atoms is far more stable – why carboxylic acids beat alcohols
atomic size	down a group, a larger atom spreads the charge better – HI is stronger than HF
an indicator 酸碱指示剂	itself a weak acid whose protonated and deprotonated forms have different colours, so it changes colour over its own pK_a range

Never argue from the acid; argue from the anion left behind. Every structural rule here is about stabilising that charge.

Introduction to Acids and Bases



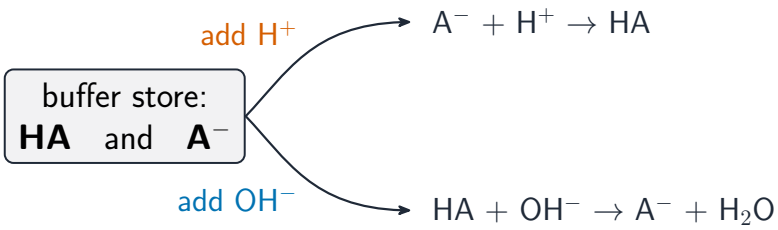
Litmus paper: acid–base indicators change colour with H concentration (pH)

Bronsted-Lowry, and the autoionization of water

an acid 酸	a proton (H^+) donor
a base 碱	a proton acceptor
conjugate pairs	when an acid donates, what remains is its conjugate base; a base that accepts becomes its **conjugate acid** 共轭酸
the hydronium ion 水合氢离子	H_3O^+ – what a proton actually forms in water
autoionization 自偶电离	$2H_2O \rightleftharpoons H_3O^+ + OH^-$, which is why pure water conducts a tiny current
K_w	the **ion-product constant of water** 水的离子积: $[H_3O^+][OH^-] = 1.0 \times 10^{-14}$ at $25\,^\circ C$

K_w is temperature-dependent, so neutral pH is 7.00 only at $25\,^\circ C$. Warm water is still neutral at a lower pH.

Properties of Buffers



either way, the pH barely changes

A buffer resists pH change by mopping up added acid or base

Choosing a buffer, and what limits it

the best ratio	a buffer works best when the weak acid and its conjugate base are at similar concentration – pH near pK_a
choosing one	pick a weak acid whose pK_a is close to the pH you want, ideally within one unit
buffer capacity 缓冲容量	how much acid or base it can absorb before the pH changes sharply
what raises it	higher total concentration of both components, and a ratio near 1:1
when it fails	once one component is consumed, the pH moves rapidly – the buffer is exhausted

Capacity and pH are separate choices: the ratio sets the pH, the total concentration sets how much abuse it will take.

pH and pOH of Strong Acids and Bases

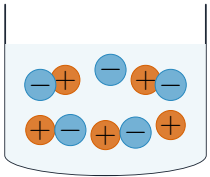
$$\text{pH} = -\log[\text{H}^+]$$

0 7 14

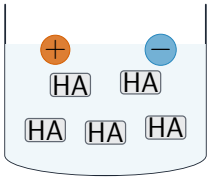
acidic neutral alkaline

The pH scale: pH is the negative log of the hydrogen-ion concentration

Introduction to Acids and Bases, continued



strong acid
fully split into ions (e.g. HCl)



weak acid
mostly molecules, few ions (e.g. ethanoic acid)

A strong acid is fully dissociated; a weak acid is only partly dissociated

The Autoionization of Water

pH strips: water autoionizes so pure water is pH 7; acids and bases shift H+ and OH-

Exam tips

- **Neutral means equal ions, not pH 7.** In pure water $[H_3O^+] = [OH^-]$; since K_w rises with temperature, neutral pH drifts below 7 above 25 °C. Use $K_w = [H_3O^+][OH^-] = 1.0 \times 10^{-14}$ (25 °C) to convert between $[OH^-]$ and $[H_3O^+]$.
- $\text{pH} = -\log[H^+]$ is **logarithmic** – each unit is a ten-fold change in $[H^+]$.
- A **strong** acid/base dissociates completely (so $[H^+]$ = concentration); a **weak** one only partly, needing K_a.

Exam tips (continued)

- A **buffer** contains a weak acid **and** its conjugate base together; it resists pH change by mopping up added acid or base.
- On a titration curve the **equivalence point** is the steep jump; the **half-equivalence point** has $\text{pH} = \text{p}K_a$.
- A smaller $\text{p}K_a$ means a stronger acid.

7 Recap

Unit 8 – key takeaways

1 pH

$\mathrm{pH} = -\log[\mathrm{H}^+]$; $\mathrm{pH} + \mathrm{pOH} = 14$

2 Weak acids

K_a equilibrium; $K_a K_b = K_w$

3 Buffers

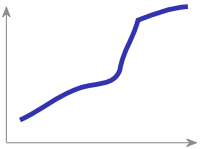
weak acid + conjugate base;
$\mathrm{pH} = \mathrm{p}K_a + \log \frac{[\mathrm{A}^-]}{[\mathrm{HA}]}$

4 Titration

equivalence jump; half-equivalence $\mathrm{pH} = \mathrm{p}K_a$

Check yourself

1. What is the pH of $0.001\,\mathrm{M}$ HCl?
2. In Bronsted-Lowry terms, what does a base do?
3. A buffer needs which two species present?
4. At the half-equivalence point, pH equals what?



Answers **1.** 3 **2.** accepts a proton **3.** a weak acid and its conjugate base **4.** pK_a