# Acids and Bases

AP Chemistry ▪ Unit 8

AP Chemistry




## What we will cover




# 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 ⇒ conjugate base soaks up $\mathrm{H^+}$
- added base ⇒ weak acid soaks up $\mathrm{OH^-}$

**Henderson-Hasselbalch**

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

Equal amounts ⇒ $\mathrm{pH} = \mathrm{p}K_a$. **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.**

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

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 $\mathrm{H^+}$ reacts with the anion.

So pH can control whether an ionic solid dissolves.

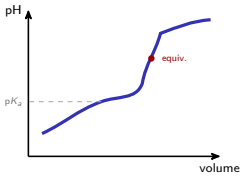
# 5 Titrations



## Acid-base titrations

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

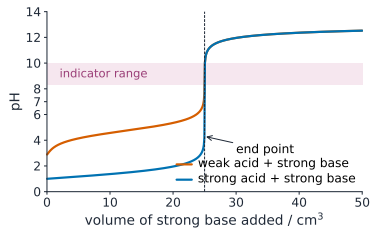
- the **equivalence point** 等当点 is the steep jump
- at the **half-equivalence point** 半等当点, $\mathrm{pH} = \mathrm{p}K_a$




> 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.

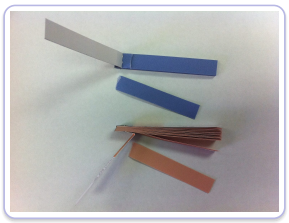


## Why one acid is stronger than another

| | |
|---|---|
| **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 $\mathrm{p}K_a$ 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)



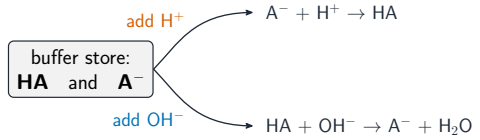
## Bronsted-Lowry, and the autoionization of water

| | |
|---|---|
| **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 °C |

$K_w$ is temperature-dependent, so neutral pH is 7.00 only at 25 °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

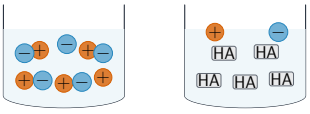
$$\text{pH} = -\log[\text{H}^+]$$

0 — acidic | 7 — neutral | 14 — 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**

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

**2 Weak acids**

$K_a$ equilibrium; $K_aK_b = K_w$

**3 Buffers**

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

**4 Titration**

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



## Check yourself

1. What is the pH of $0.001\ \text{M}$ 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.** $\text{p}K_a$