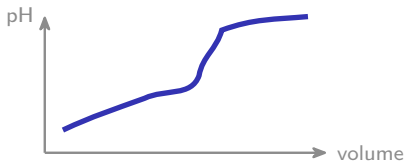


Acids and Bases

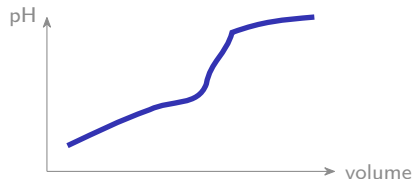
AP Chemistry ■ Unit 8

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





Bronsted-Lowry and pH

Bronsted-Lowry -: an acid **donates** a proton, a base **accepts** one – forming **conjugate pairs** .

Water and pH

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

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



pH measures acidity of a solution



Strong acids and bases

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

Two logarithms

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

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

Worked example

0.010 M HCl:

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

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



Weak acid and base equilibria

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

Ionization constant

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

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

Worked example

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

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

$$\text{pH} = 3$$

A strong acid at 0.10 M would give pH 1.



Buffers and Henderson-Hasselbalch

A **buffer** 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

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

Equal amounts $\Rightarrow \text{pH} = \text{p}K_a$.
Capacity grows with total concentration.

Worked example – the Henderson-Hasselbalch equation

A buffer holds 0.20 M acetic acid ($pK_a = 4.74$) and 0.30 M acetate. Find its pH.

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

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

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.

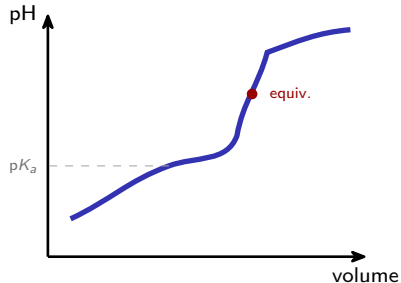


Acid-base titrations

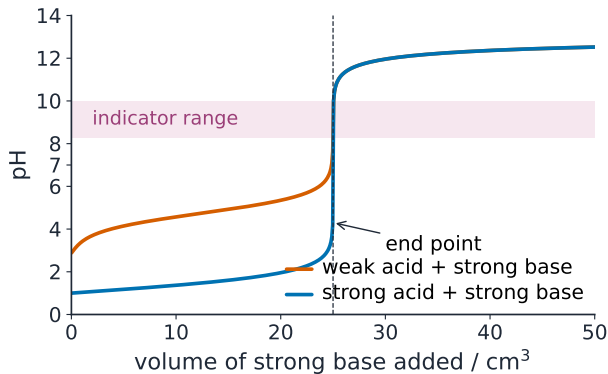
A **titration curve** plots pH against titrant volume:

- the **equivalence point** is the steep jump
- at the **half-equivalence point** ,
 $\text{pH} = \text{p}K_a$

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



Structure and pK_a

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

- **binary acids** : stronger *down* a group (weaker bond)
- **oxyacids** : more oxygens $\Rightarrow$ stronger

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

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

$\text{pH} < \text{p}K_a$: the acid form dominates; $\text{pH} > \text{p}K_a$: 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

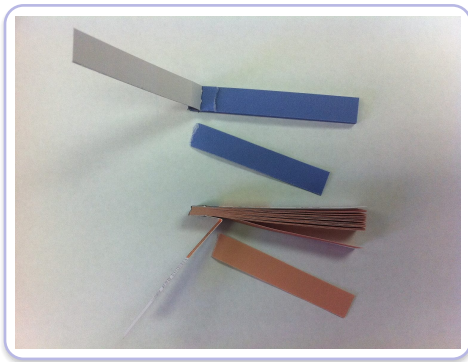
down a group, a larger atom spreads the charge better – HI is stronger than HF

an indicator

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

a proton (H^+) donor

a base

a proton acceptor

conjugate pairs

when an acid donates, what remains is its conjugate base;
a base that accepts becomes its conjugate acid

the hydronium ion

H_3O^+ – what a proton actually forms in water

autoionization

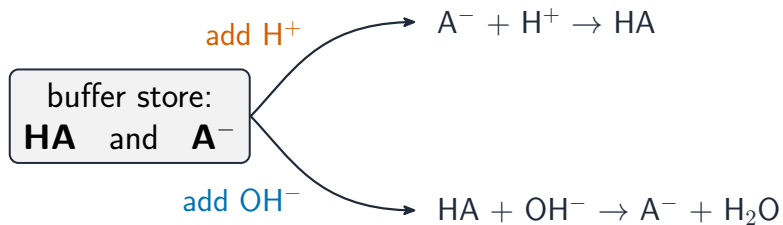
$2\text{H}_2\text{O} \rightleftharpoons \text{H}_3\text{O}^+ + \text{OH}^-$, which is why pure water conducts
a tiny current

K_w

the ion-product constant of water : $[\text{H}_3\text{O}^+][\text{OH}^-] =$
 1.0×10^{-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

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

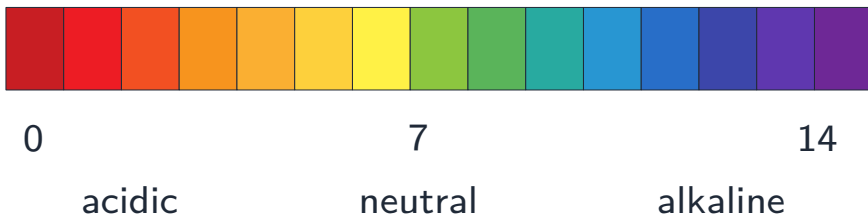
when it fails

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.

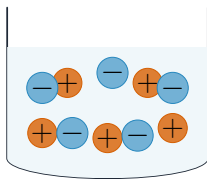
pH and pOH of Strong Acids and Bases

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

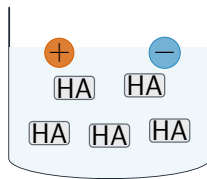


The pH scale: pH is the negative log of the hydrogen-ion concentration

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 $[\text{H}_3\text{O}^+] = [\text{OH}^-]$; since K_w rises with temperature, neutral pH drifts below 7 above 25 °C. Use $K_w = [\text{H}_3\text{O}^+][\text{OH}^-] = 1.0 \times 10^{-14}$ (25 °C) to convert between $[\text{OH}^-]$ and $[\text{H}_3\text{O}^+]$.
- $\text{pH} = -\log[\text{H}^+]$ is **logarithmic** – each unit is a ten-fold change in $[\text{H}^+]$.
- A **strong** acid/base dissociates completely (so $[\text{H}^+] = \text{concentration}]$; 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.



Unit 8 – key takeaways

1 pH

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

2 Weak acids

$$K_a \text{ equilibrium; } K_a K_b = K_w$$

3 Buffers

weak acid + conjugate base;

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

4 Titration

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

Check yourself

- 1 What is the pH of 0.001 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