

Equilibria

A-Level Chemistry ■ Topic 25

A-Level Chemistry



What we will cover

- 1 Conjugate acids & bases
- 2 pH & the constants
- 3 Calculating pH
- 4 Buffer solutions
- 5 Solubility & partition



pH paper estimates acidity



Conjugate acids and bases

An acid loses H^+ to give its **conjugate base** ; a base gains H^+ to give its **conjugate acid** .

A **conjugate pair** differs by just one H^+ – e.g. CH_3COOH and CH_3COO^- .

conjugate pair: differ by one H^+

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pH and the constants

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

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

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

A larger K_a (smaller $\text{p}K_a$) is a stronger acid; $K_w = 1.0 \times 10^{-14}$ at 298 K.

Calculating pH

strong acid – $[\text{H}^+] = \text{conc.}$; $\text{HCl } 0.050 \Rightarrow \text{pH} = 1.30$

strong base – $[\text{H}^+] = K_w/[\text{OH}^-]$; $\text{NaOH } 0.10 \Rightarrow \text{pH} = 13.0$

weak acid – $[\text{H}^+] = \sqrt{K_a[\text{HA}]}$; ethanoic $0.10 \Rightarrow \text{pH} = 2.87$

The four pH calculations

strong acid

fully ionised, so $[\text{H}^+]$ is the acid's concentration; $\text{pH} = -\log[\text{H}^+]$

weak acid

only partly ionised, so use $[\text{H}^+] = \sqrt{K_a \times [\text{HA}]}$, where K_a is the **acid dissociation constant**

strong alkali

find $[\text{OH}^-]$, then use the **ionic product of water** $K_w = [\text{H}^+][\text{OH}^-] = 10^{-14}$

a buffer

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

a conjugate pair

an acid and the base left after it donates its proton – the two components of every buffer

A buffer works because the mixture holds **a store of both partners**: the acid mops up added alkali, and the conjugate base mops up added acid.

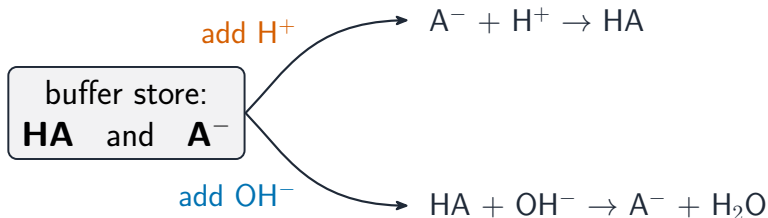


Buffer solutions

A **buffer** (a weak acid + its conjugate base) resists pH change:

- added H^+ removed by A^-
- added OH^- removed by HA

HCO_3^- keeps blood near pH 7.4.



either way, the pH barely changes

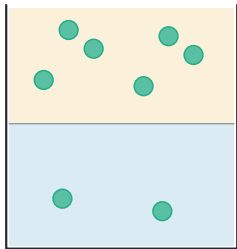
Solubility product



The **solubility product** $K_{sp} = [Ag^+][Cl^-]$ (each ion to the power of its number).

The **common ion effect** : a salt is **less** soluble in a solution already containing one of its ions (Le Chatelier).

Partition coefficients



$$K_{pc} = \frac{\text{solvent 1 (non-polar)} \left[\text{solute in 1} \right]}{\text{solvent 2 (water)} \left[\text{solute in 2} \right]}$$

The **partition coefficient** is the ratio of a solute's concentrations in two immiscible solvents.

$$K_{pc} = \frac{[\text{solute in 1}]}{[\text{solute in 2}]}$$

It depends on the **polarity** of the solute & solvents.

When a partition coefficient applies

what it is

the ratio of a **solute** 's concentrations in two immiscible solvents at equilibrium

the condition

it holds only when the solute is in the **same physical state** in both solvents – no dimerising or dissociating

what it depends on

the **polarity** of the solute and of the two solvents

the rule

“like dissolves like” – a polar solute favours the polar solvent

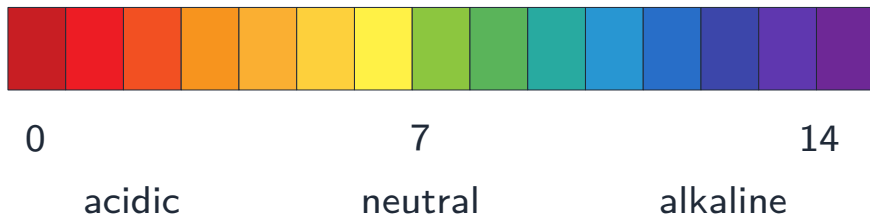
where it is used

solvent extraction, and chromatography, where the same partition runs continuously along the plate

Repeated extraction with **small** portions removes more solute than one extraction with the same total volume. That is the calculation the exam sets.

pH and the equilibrium constants

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



The pH scale: $\text{pH} = -\log$ of the hydrogen-ion concentration

Exam tips

- $\text{pH} = -\log[\text{H}^+]$; for a **strong acid** $[\text{H}^+] = \text{concentration}$, for a **weak acid** use $[\text{H}^+] = \sqrt{K_a[\text{HA}]}$.
- For a base, get $[\text{H}^+]$ from $K_w = [\text{H}^+][\text{OH}^-] = 1.0 \times 10^{-14}$.
- For a **buffer** use $[\text{H}^+] = K_a \times [\text{acid}]/[\text{salt}]$ and explain how it removes added H^+/OH^- .
- Write the K_{sp} expression with the correct powers; the number of decimal places in a pH equals the significant figures of $[\text{H}^+]$.



Worked example – pH of a strong acid

Find the pH of $0.050 \text{ mol dm}^{-3}$ hydrochloric acid.

- HCl is a **strong** acid: it dissociates completely.
- So $[\text{H}^+] = 0.050 \text{ mol dm}^{-3}$ – the concentration itself.
- $\text{pH} = -\log_{10}[\text{H}^+] = -\log_{10}(0.050)$
- $= 1.30$.

Strong acid: $[\text{H}^+]$
is the concentration.

Weak acid: you must
go through K_a – never as-
sume full dissociation.

Topic 25 – key takeaways

1

Conjugate pairs

differ by one H^+

2

pH

$-\log[\text{H}^+]$; K_a , K_w for
weak/strong

3

Buffers

weak acid + its salt resist pH
change

4

Solubility

K_{sp} ; the common ion effect

Check yourself

- 1 Write the pH formula.
- 2 What does a buffer resist?
- 3 What two things make a buffer?
- 4 What does the common ion effect do to solubility?

Answers 1. $\text{pH} = -\log[\text{H}^+]$ 2. a change in pH 3. a weak acid & its conjugate base
4. lowers it