25.1 ACIDS AND BASES

From the acid in your stomach to the buffer in your blood — pH chemistry is everywhere. Let's decode it.

https://subarnam.com.np/

🎯 Learning outcomes

Candidates should be able to:

1. understand and use the terms conjugate acid and conjugate base
2. define conjugate acid–base pairs, identifying such pairs in reactions
3. define the terms pH, K_a, pK_a and K_w mathematically and use them in calculations (K_b and the equation $K_w = K_a \times K_b$ will not be tested)
4. calculate $[H^+(aq)]$ and pH values for:
   (a) strong acids
   (b) strong alkalis
   (c) weak acids
5. (a) define a buffer solution
   (b) explain how a buffer solution can be made
   (c) explain how buffer solutions control pH; use chemical equations in these explanations
   (d) describe and explain the uses of buffer solutions, including the role of HCO_3^- in controlling pH in blood
6. calculate relevant concentrations and the pH of buffer solutions, given appropriate data
7. understand and use the term solubility product, K_{sp}
8. write an expression for K_{sp}
9. calculate K_{sp} from concentrations and vice versa
10. (a) understand and use the common ion effect to explain the different solubility of a compound in a solution containing a common ion
    (b) perform calculations using K_{sp} values and concentration of a common ion

Before you begin: Check your prerequisite knowledge

1. Write down the definition of a *Brønsted–Lowry acid* and *Brønsted–Lowry base*. Compare your definitions with those of another learner.
2. Explain the meaning of the terms: *equilibrium*, *equilibrium constant*, *dissociation*, *strong acid*, *strong base*, *weak acid*, *weak base*, *neutralisation*.
3. Make a list of strong acids and weak acids, and strong bases and weak bases
4. Write an equilibrium expression for this acid. Note that the process is the same as for writing equilibrium expressions for K_c. The only difference is that you put K_a instead of K_c.
5. Explain why the equation $H^+ + OH^- \rightarrow H_2O$ applies to any acid–alkali reaction.
6. Write a paragraph describing how to carry out an experiment to find the enthalpy change of neutralisation of an acid by an alkali.
7. Explain the meaning of the terms: *solute*, *solution*, *salt*, *sparingly soluble*, *polarisation*, *non-polar*, *enthalpy change*, *hydration*.

Introduction

Forget "acids are sour and dangerous." The real definition is about one thing only: protons (H^+ ions).

> 📖 **BRØNSTED–LOWRY THEORY**
>
> An acid is a **proton donor** (gives away H^+).
>
> A base is a **proton acceptor** (takes H^+).

🔁 Conjugate Acid–Base Pairs

Every time an acid loses a proton, what's left behind can accept a proton right back — that makes it a conjugate pair. **The two species in a pair differ by exactly one H^+.**

$$\underset{\text{acid}_1}{HA} + \underset{\text{base}_2}{B} \rightleftharpoons \underset{\text{base}_1}{A^-} + \underset{\text{acid}_2}{BH^+}$$

HA and A^- are one conjugate pair. B and BH^+ are the other conjugate pair.

EVERYDAY RELEVANCE

An antacid tablet works exactly like this: it contains a base (e.g. $CaCO_3$ or $Mg(OH)_2$) that accepts H^+ from excess stomach acid (HCl), calming heartburn in minutes.

Conjugate Acids & Bases | Acids, Bases & Alkali's | Chemistry | FuseSchool - YouTube

Strong vs Weak — It's Not About Concentration!

Common mix-up: "strong" and "concentrated" are NOT the same thing. Strength is about how completely an acid dissociates into ions — nothing to do with how much of it you dissolved.

EXAM TRAP

A dilute strong acid can still be "strong" (fully dissociated), and a concentrated weak acid is still "weak" (only partially dissociated). Strength = extent of dissociation.

Strong acids and bases — fully dissociated

Type	Example	Dissociation equation
Strong acid	Hydrochloric acid	$HCl \rightarrow H^+ + Cl^-$
Strong acid	Nitric acid	$HNO_3 \rightarrow H^+ + NO_3^-$
Strong acid (diprotic)	Sulfuric acid	$H_2SO_4 \rightarrow 2H^+ + SO_4^{2-}$
Strong base	Sodium hydroxide	$NaOH \rightarrow Na^+ + OH^-$
Strong base	Potassium hydroxide	$KOH \rightarrow K^+ + OH^-$

Weak acids and bases — only partly dissociated

Type	Example	Dissociation equation
Weak acid	Ethanoic acid	$CH_3COOH \rightleftharpoons H^+ + CH_3COO^-$
Weak acid	Nitrous acid	$HNO_2 \rightleftharpoons H^+ + NO_2^-$
Weak acid	Carbonic acid	$H_2CO_3 \rightleftharpoons H^+ + HCO_3^-$
Weak base	Ammonia	$NH_3 + H_2O \rightleftharpoons NH_4^+ + OH^-$

Notice the arrow: → for complete (strong) dissociation, ⇌ for partial, reversible (weak) dissociation.

EVERYDAY RELEVANCE

Vinegar is only about 5% ethanoic acid, yet it's a weak acid even at higher concentrations — that's why you can safely put it on chips, while battery acid (dilute sulfuric acid) is dangerously corrosive at similar concentrations because it's strong.

Ionic product of water

Pure water conducts electricity slightly, so it must be ionised to a small extent:

$$2H_2O(l) \rightleftharpoons H_3O^+(aq) + OH^-(aq)$$

For this equilibrium:

$$K_c = \frac{[H_3O^+(aq)][OH^-(aq)]}{[H_2O(l)]^2}$$

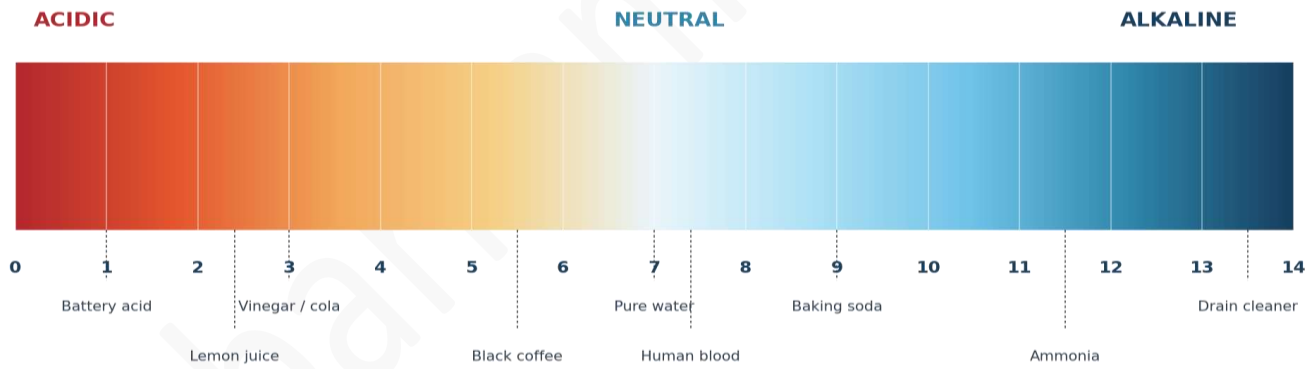
Because $[H_2O(l)]$ is a constant, we can define a new constant called the **ionic product of water**, K_w, such that $K_w = [H_3O^+(aq)][OH^-(aq)]$. Because $[H_3O^+] = [H^+(aq)]$ (hydrogen ions in water are all hydrated), this expression is often simplified to $K_w = [H^+][OH^-]$.

The ionic product of water is the **equilibrium constant** for the **dissociation of water at 298 K**

Experimentally, K_w is found to have a value of 1.0×10^{-14} mol^2 dm^{-6} at 25 °C.

Ionic product of water, Kw: the equilibrium constant for the ionisation of water. $K_w = [H^+]\,[OH^-]$

Calculation of pH



The pH scale goes from 0.0 to 14.0

Acids have a pH below 7.0

Pure water is **neutral** with a pH of 7.0

Bases and alkalis have a pH above 7.0

The pH equation

$$pH = -\log_{10}[H^+] \quad \text{or} \quad [H^+] = 10^{-pH}$$

pH = 7 does not always mean neutral.

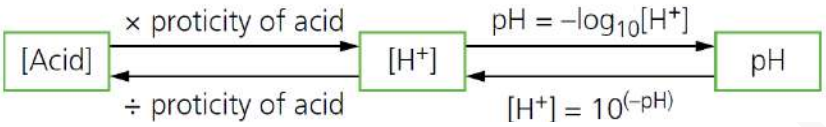
Neutral means: **$[H^+] = [OH^-]$**

At 25 °C, this happens at pH 7.00 because Kw = 1.00×10^{-14}.

At other temperatures, Kw changes, so the pH of neutral water may not be 7.00.

pH of strong acids

- They ionise completely.
- Therefore, for a strong monoprotic acid: $[H^+] \approx$ concentration of acid



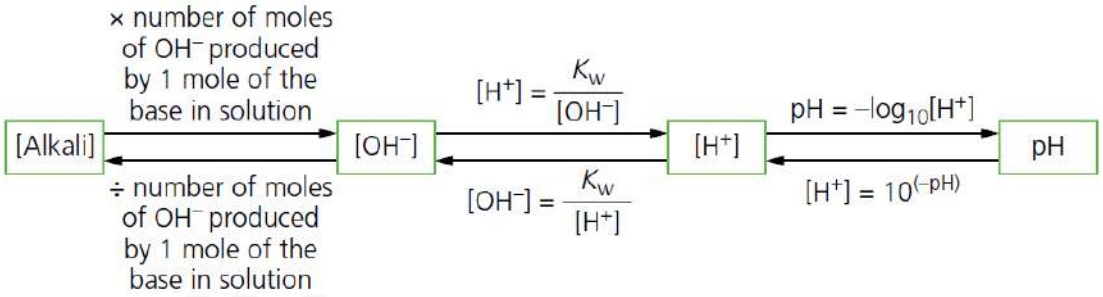
WORKED EXAMPLE — STRONG ACID

Find the pH of 0.010 mol dm^{-3} HCl.

HCl is a strong acid → fully dissociated, so $[H^+] = [HCl] = 0.010$ mol dm^{-3}.

$pH = -\log_{10}(0.010) = 2.00$

pH of strong alkalis



- They ionise completely to produce OH^-.
- Use:

$[H^+][OH^-] = Kw$

- Then calculate pH using **$pH = -\log[H^+]$**.

WORKED EXAMPLE — STRONG ALKALI

Find the pH of 0.0050 mol dm^{-3} NaOH.

NaOH is a strong base → $[OH^-] = 0.0050$ mol dm^{-3}.

$[H^+] = Kw \div [OH^-] = (1.0 \times 10^{-14}) \div 0.0050 = 2.0 \times 10^{-12}$ mol dm^{-3}

$pH = -\log_{10}(2.0 \times 10^{-12}) = 11.70$

pH of weak acids

Acid dissociation constant, Ka: the equilibrium constant for the dissociation of a weak acid.

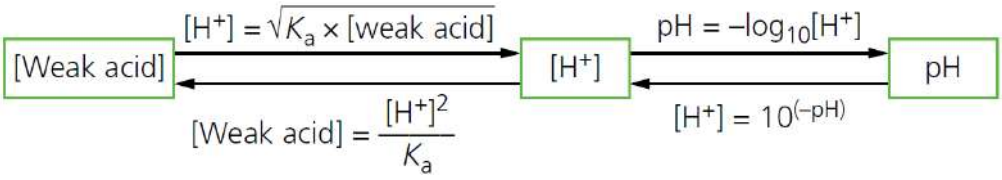
$$Ka = \frac{[H+][A-]}{[HA]}$$

pKa values: the values of Ka expressed as a logarithm of base 10.

$$pKa = -\log_{10} Ka$$

Remember that a small 'p' in front of a symbol related to acids and bases means the negative logarithm (to the base 10).

So, pH is $-\log_{10}[H^+]$ and p*K*a = $-\log_{10}K_a$



WORKED EXAMPLE — PH OF A WEAK ACID

Ethanoic acid, Ka = 1.7×10^{-5} mol dm^{-3}, concentration = 0.10 mol dm^{-3}. Find the pH.

Since dissociation is small, assume [HA] ≈ 0.10 mol dm^{-3} and $[H^+] = [A^-]$.

$$[H^+] = \sqrt{(Ka \times [HA])} = \sqrt{(1.7 \times 10^{-5} \times 0.10)} = 1.30 \times 10^{-3} \text{ mol dm}^{-3}$$

$$pH = -\log_{10}(1.30 \times 10^{-3}) = 2.89$$

Percentage ionisation

% ionisation= [H+]/[initial acid]×100

The greater the Ka value, the **more strongly acidic** the acid is

The greater the pKa value, the **less strongly acidic** the acid is.

Buffer solution

Buffer solution: a solution in which the pH does not change significantly when small amounts of acids and bases are added.

OR

A solution that minimises changes in pH when small amounts of acid or base are added.

How is a buffer made?

- **weak acid - conjugate base**(e.g. CH_3COOH + CH_3COO^- from CH_3COONa)
- a **weak base – conjugate acid**(e.g. NH_3 + NH_4^+)

Method 1

Mix a weak acid with its salt.

Example:

CH_3COOH + CH_3COONa

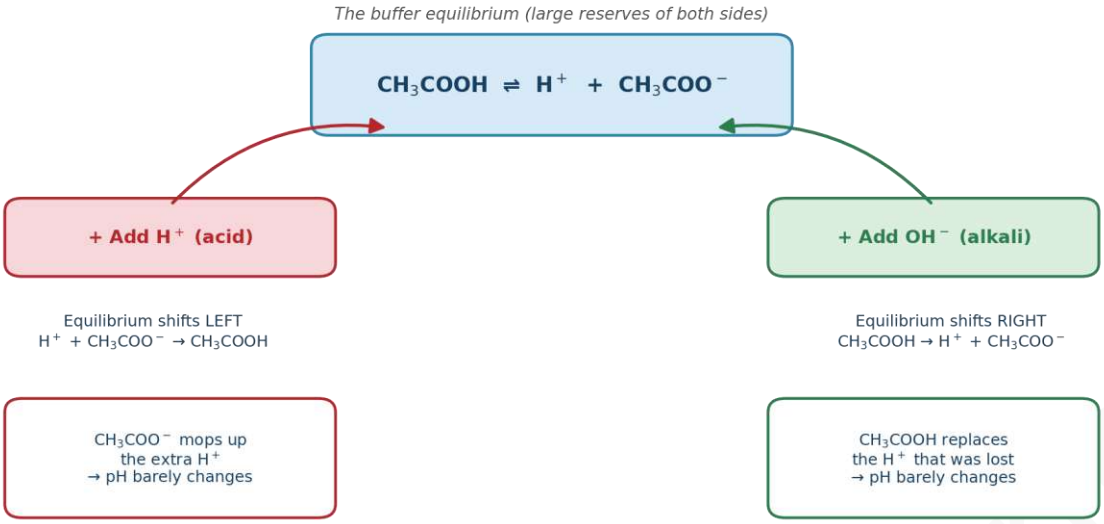
Method 2

Partially neutralise a weak acid with a strong alkali.

Example:

$CH_3COOH + NaOH \rightarrow CH_3COONa + H_2O$

How Buffer works?



Added acid

$CH_3COO^- + H^+ \rightarrow CH_3COOH$

Therefore:

- added H^+ is removed
- $[H^+]$ does not increase much
- pH changes only slightly

Added alkali

$CH_3COOH + OH^- \rightarrow CH_3COO^- + H_2O$

Therefore:

- OH^- is removed
- $[OH^-]$ does not increase much
- pH changes only slightly

How your blood stays at pH 7.35–7.45

Your cells constantly produce CO_2. It reacts with water in blood plasma to form the HCO_3^- / CO_2 buffer pair:

$$CO_2(g) + H_2O(l) \rightleftharpoons H^+(aq) + HCO_3^-(aq)$$

- If H^+ rises → equilibrium shifts LEFT, mopping up the excess H^+.
- If H^+ falls → equilibrium shifts RIGHT, releasing more H^+.

If this buffer fails, blood pH drops and acidosis can result — severe cases can cause organ malfunction and even coma.

Calculating buffer pH

$$[H^+] = Ka \times \frac{[acid]}{[salt]} \text{ or,}$$

$$pH = pKa + \log_{10} \frac{[salt]}{[acid]}$$

WORKED EXAMPLE — BUFFER PH

A buffer contains 0.20 mol dm^{-3} CH_3COOH and 0.10 mol dm^{-3} CH_3COONa. Ka = 1.7×10^{-5} mol dm^{-3}.

$$[H^+] = Ka \times ([acid]/[salt]) = 1.7\times10^{-5} \times (0.20/0.10) = 3.4 \times 10^{-5} \text{ mol dm}^{-3}$$

$$pH = -\log_{10}(3.4 \times 10^{-5}) = 4.47$$

K_{sp}

In a saturated solution, an equilibrium exists between the undissolved solid and its dissolved ions.

For example:

$Ca(OH)_2(s) \rightleftharpoons Ca^{2+}(aq) + 2OH^-(aq)$

At equilibrium, the rate at which ions leave the solid equals the rate at which ions return to the solid.

Solubility Product Constant (Ksp) - YouTube

Solubility product, Ksp: the product of the concentrations of each ion in a saturated solution of a sparingly soluble salt at 298K, raised to the power of their relative concentrations

e.g. for $PbCl_2 \rightleftharpoons Pb^{2+} + 2Cl^-$: $Ksp = [Pb^{2+}][Cl^-]^2$

Remember that the solubility product is only relevant for salts which are sparingly (very slightly) soluble. The concept cannot be used for soluble salts such as sodium chloride. Remember that soluble salts include salts of Group I elements, all nitrates and ammonium salts and many sulfates. Halides are generally soluble except for lead(II) halides and silver halides.

- For salt $C(s) \rightleftharpoons aA^{x+}(aq) + bB^{y-}(aq)$
- **$Ksp = [A^{x+}]^a[B^{y-}]^b$**
- **Smaller Ksp** generally means **a less soluble salt**

Solubility tells us how much salt dissolves.
Ksp tells us about the equilibrium between the solid and its ions.

The common ion effect

The common ion effect: the reduction of the solubility of a dissolved salt by adding a compound that has an ion in common with the dissolved salt, e.g, the addition of sodium chloride to a solution of very slightly soluble lead(II) chloride.

Example:

$AgCl(s) \rightleftharpoons Ag^+ + Cl^-$

Adding KCl increases $[Cl^-]$, shifting the equilibrium left and causing **AgCl to precipitate**.

- **Precipitate forms if ion product > Ksp**
- If: **Ion product < Ksp → no precipitate**

Demonstrate and discuss the common-ion effect as in this video:

www.youtube.com/watch?v=c9_Ab0BzlC4

Calculations:

www.youtube.com/watch?v=VVyseROOw7E www.youtube.com/watch?v=qawipem0Lw

The shell of this nautilus is composed mainly of calcium carbonate. The nautilus adjusts conditions so that shell material is formed when the concentration of calcium ions and carbonate ions in seawater is high enough to precipitate calcium carbonate.

> 🌍 **EVERYDAY RELEVANCE**
>
> Kidney stones often form from calcium oxalate once its Ksp is exceeded in urine.

Calculations involving titrations

Key Formulas at a Glance

Quantity	Formula
Ionic product of water	$K_w = [H^+][OH^-] = 1.0 \times 10^{-14}$ mol² dm^{-6} (25 °C)
pH	$pH = -\log_{10}[H^+]$ $[H^+] = 10^{-pH}$
Acid dissociation constant	$K_a = [H^+][A^-]/[HA]$
pKa	$pK_a = -\log_{10} K_a$
$[H^+]$ of a weak acid	$[H^+] = \sqrt{(K_a \times [HA])}$
Buffer $[H^+]$	$[H^+] = K_a \times ([acid]/[salt])$
Buffer pH	$pH = pK_a + \log_{10}([salt]/[acid])$

Quantity	Formula
Solubility product	Ksp = product of ionic concentrations, each to the power of its relative amount

💡 **BEFORE YOU CLOSE THIS DOC**

- ✅ Strong/weak = extent of dissociation, NOT concentration
- ✅ → means complete dissociation; ⇌ means partial, reversible dissociation
- ✅ Buffer explanations need equations AND equilibrium-shift reasoning
- ✅ Ksp only applies to sparingly soluble salts

⚠️ **Key Exam Tips**

1. **Kw only valid at 298K** unless stated otherwise
2. Strong acids/alkalis: **fully dissociated** (→ not ⇌)
3. Weak acids: use **Ka**, NEVER assume full dissociation
4. **pKa and Ka inverse relationship** (lower pKa = stronger acid)
5. Ksp only for **sparingly soluble salts** (not Group 1 salts, nitrates, ammonium salts)
6. Common ion effect: **equilibrium shifts left** when a common ion is added
7. Buffer pH formula: **learn both forms** ($[H^+] = Ka \times [acid]/[salt]$ AND $pH = pKa + \log([salt]/[acid])$)
8. For buffer calculations: **[salt] ≈ initial salt concentration**, **[acid] ≈ initial acid concentration**

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