Foundations: Water, pH & Buffers

BCH 100 — Introductory Biochemistry · Dr. Radi

By the end of this unit, you can…

- Explain water's polarity and hydrogen bonding, and identify H-bond donors and acceptors
- Describe the hydrophobic effect and its role in protein folding and membranes
- Work with acids, bases, pH, pKa, the Henderson–Hasselbalch equation, titration curves, and buffer capacity
- Explain how physiological buffers (bicarbonate) hold blood at pH 7.4

Today's route

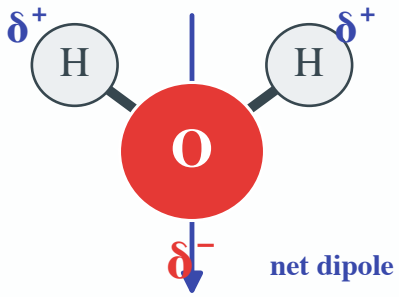
1. **Water – Polarity & Hydrogen Bonding**
2. **The Hydrophobic Effect**
3. **pH, Acids, Bases & pKa**
4. **The Henderson–Hasselbalch Equation**
5. **Titration Curves & Buffers**
6. **Buffer Capacity & Blood**

1 · Water — Polarity & Hydrogen Bonding

"See why the bent, polar water molecule — and the hydrogen bonds it makes — is the stage every reaction in you plays out on."

Water is bent, and that changes everything

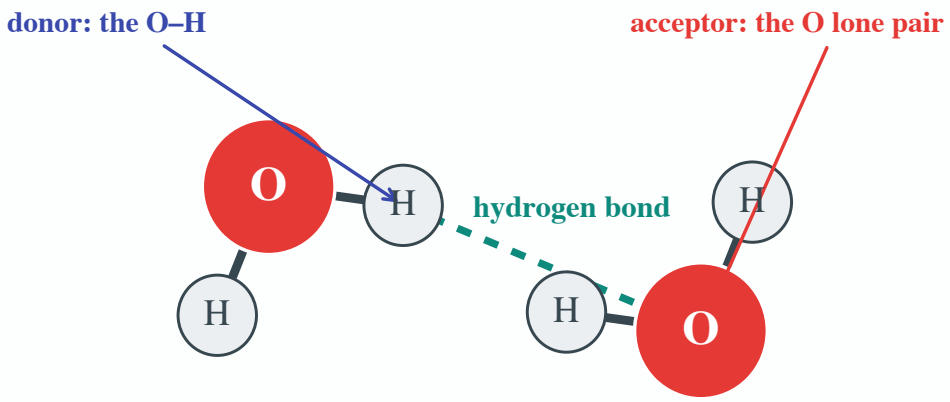
Life happens in water. The molecule is **bent**, and oxygen **hogs the shared electrons** — so the O end goes slightly **negative (δ−)** and the H's go **positive (δ+)**. A tiny magnet with a + end and a − end.



bent shape + unequal sharing =
a molecule with a + end and a − end

The hydrogen bond

Point one water's **δ+ hydrogen** at another water's **δ− oxygen** and they stick — a **hydrogen bond**. The O–H is the **donor**; the O with its lone pairs is the **acceptor**. Each water can make **four**.



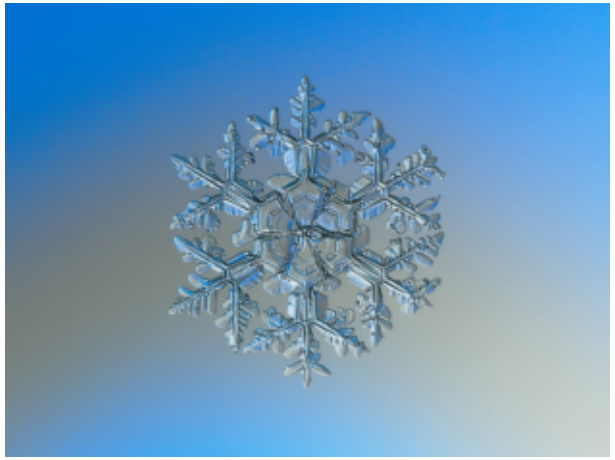
each water can make FOUR — that's why water is so cohesive

💡 Spot the donor and acceptor — you'll see the exact same bond holding DNA's strands together and folding every protein.

Water's superpowers all come from H-bonds

That one bond, multiplied by trillions, gives water its whole personality: **ice floats** (the open lattice is less dense), it's **cohesive** (surface tension, capillary rise), and it's the **universal solvent** that dissolves the salts and sugars of life.

Ice floats



the open H-bonded lattice
is less dense than liquid water

Cohesion

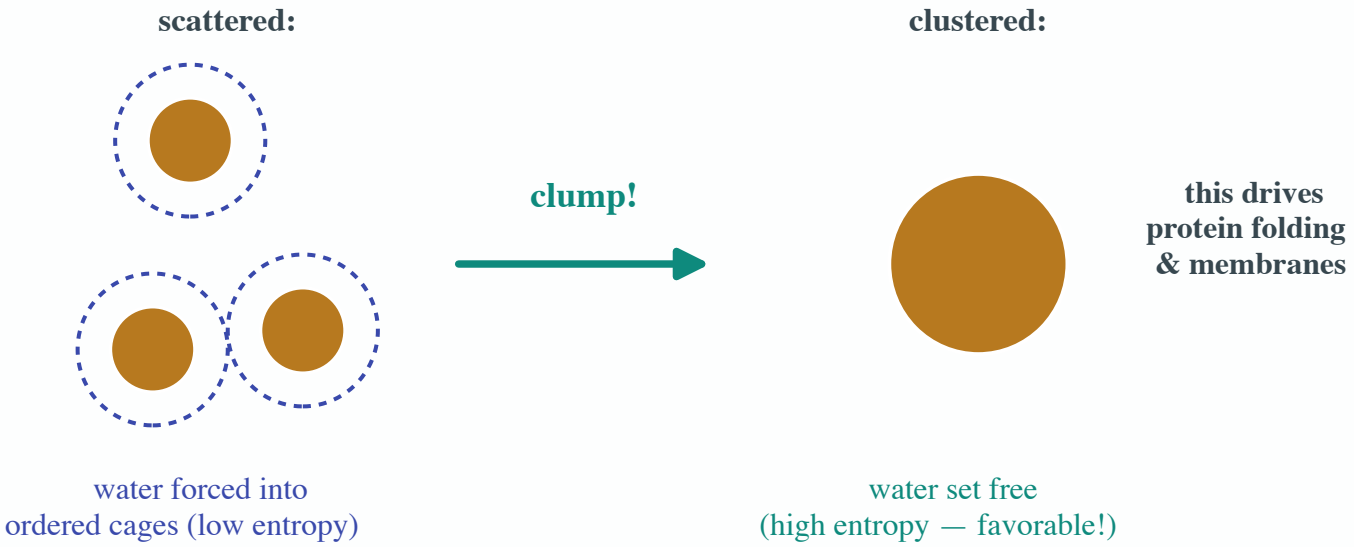
surface tension is strong
enough to hold a strider up

2 · The Hydrophobic Effect

"See why oil and water don't mix — and how that same push folds every protein and builds every membrane."

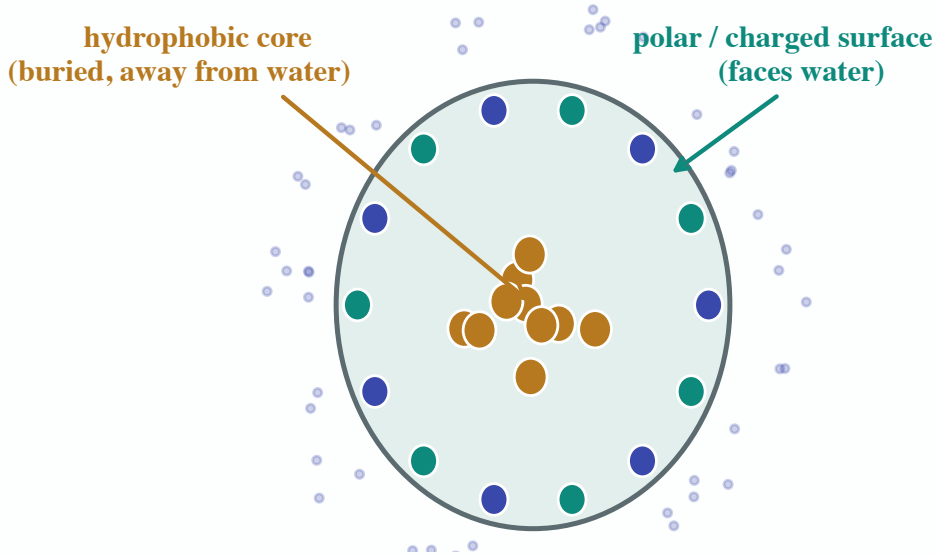
Why oil and water refuse to mix

Drop something **nonpolar** in water and the water can't hydrogen-bond to it — so it wraps the intruder in an **ordered cage**. Cages cost **entropy**. So the nonpolar bits **clump together**, shrinking the surface and setting the trapped water **free**. Entropy wins.



This is what folds a protein

That same shove is the **main driver of protein folding**: the greasy side chains bury themselves in a **hydrophobic core**, away from water, while the polar ones face out. Water isn't a bystander here — it's the **force doing the folding**.



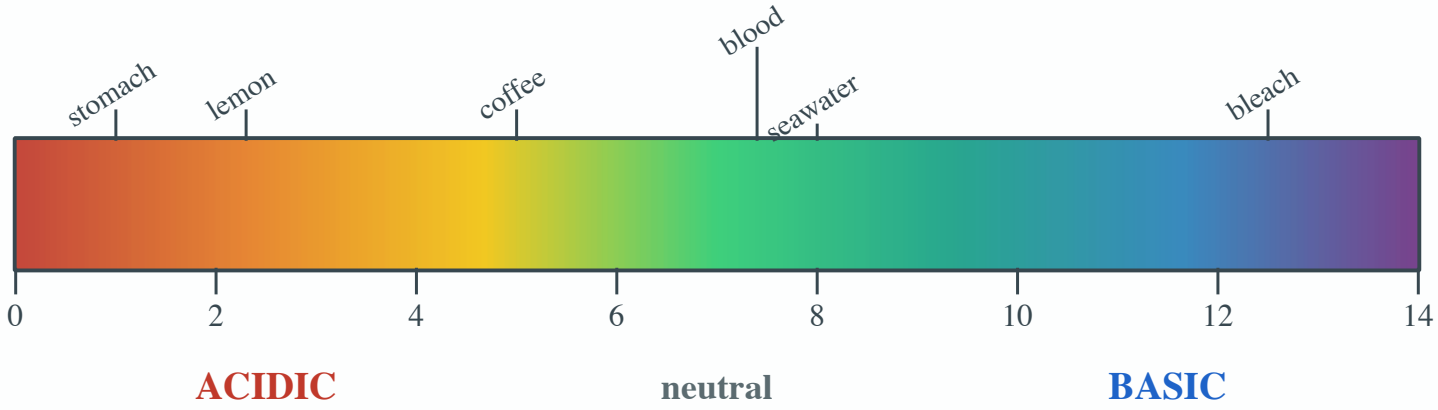
💡 Same logic builds every membrane — lipid tails hide from water, heads face it.

3 · pH, Acids, Bases & pKa

"Set up the language of acids and bases — the pH scale, conjugate pairs, and what pKa really tells you."

The pH scale

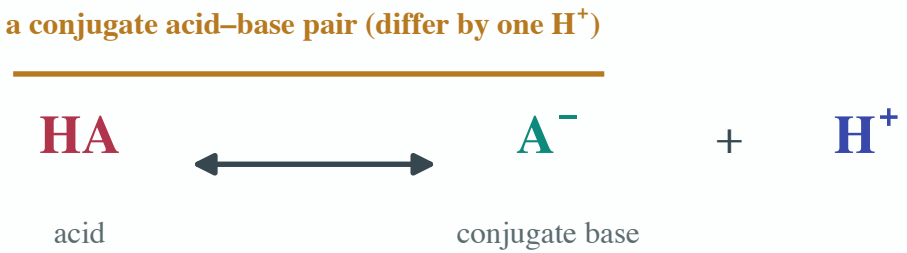
pH = –log[H⁺], so a **low** number means **more** H⁺ and more acidic. Each step is a **10×** change — pH 2 has a hundred times the H⁺ of pH 4. Your blood sits tight at **7.4**.



pH = –log[H⁺] · each step = 10× the H⁺

Acids, bases & conjugate pairs

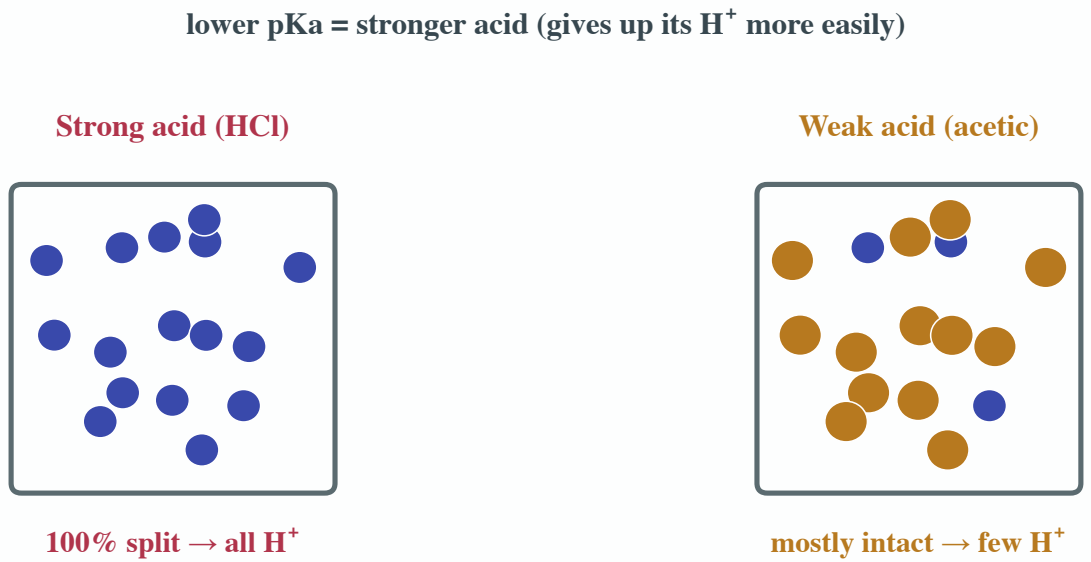
An **acid** gives up a **proton (H^+)**; a **base** takes one. Lose the proton and an acid becomes its **conjugate base** — the two differ by exactly one H^+, so they always come as a **pair**.



$$e.g.\ acetic\ acid\ (CH_3COOH) \rightleftharpoons acetate\ (CH_3COO^-) + H^+$$

Strong vs weak – and what pKa means

A **strong** acid dissociates **completely**; a **weak** acid only **partly**, leaving most molecules intact. **pKa** measures that eagerness: a **lower pKa = a stronger acid**. Biochemistry runs almost entirely on **weak** acids and bases.



💡 Weak is the point – only a partial giver can act as a buffer.

4 · The Henderson–Hasselbalch Equation

"Turn the acid–base ratio into a pH with one equation — and see why pH = pKa is the magic point."

One equation ties pH, pKa, and the ratio

The **Henderson–Hasselbalch** equation says the pH of a buffer is just its **pKa** plus the **log of the ratio** of base to acid. The punchline: when **$[A^-] = [HA]$**, that log is zero, so **pH = pKa**.

$$\mathbf{pH = pKa + \log\left([A^-] / [HA] \right)}$$

$[A^-] = [HA]$ (1:1)	$\log 1 = 0$	**pH = pKa**
10× more A^-	$\log 10 = +1$	**pH = pKa + 1**
10× more HA	$\log 0.1 = -1$	**pH = pKa – 1**

💡 **KNOW HOW TO USE THIS.** Given any two of pH, pKa, and the ratio, you can always find the third.

Try one: an acetate buffer

Acetic acid has **pKa 4.76**. Mix in **ten times** as much acetate as acid — what's the pH?

$$\text{pH} = 4.76 + \log(10/1) = 4.76 + 1 = 5.76$$

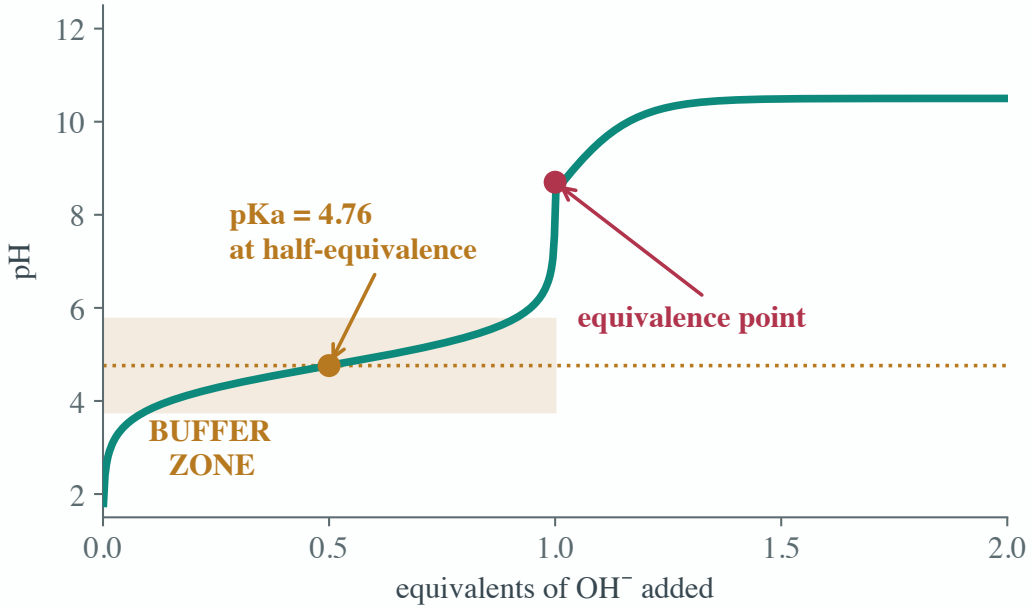
more conjugate base than acid → pH climbs one unit above the pKa

5 · Titration Curves & Buffers

"Read a titration curve — find the pKa at half-equivalence, spot the buffer zone, and see how a buffer soaks up acid or base."

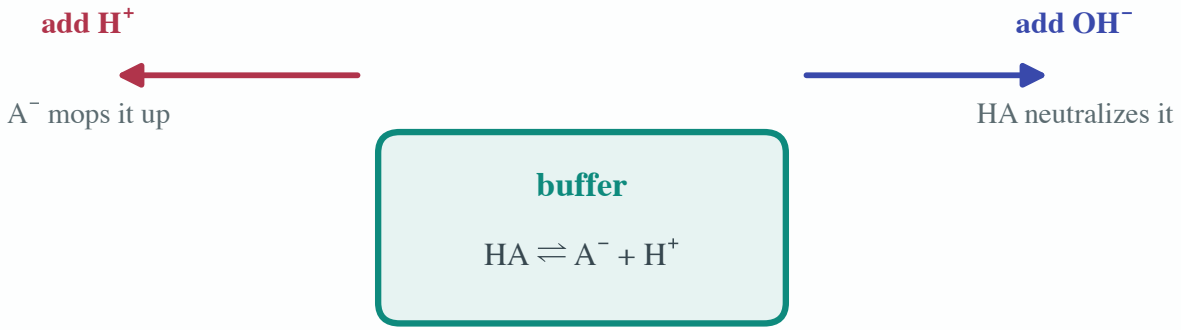
The titration curve tells you the pKa

Add base to a weak acid and the pH climbs — but not evenly. At **half-equivalence**, half the acid is deprotonated so **[A⁻] = [HA]**, which means **pH = pKa** right there. The flat stretch around it is the **buffer zone**; the steep jump is the **equivalence point**.



What a buffer actually is

A **buffer** is a weak acid and its conjugate base sitting together. Throw in **H^+** and the **A^-** mops it up; throw in **OH^-** and the **HA** neutralizes it. Either way the pair absorbs the hit and the **pH barely moves**.



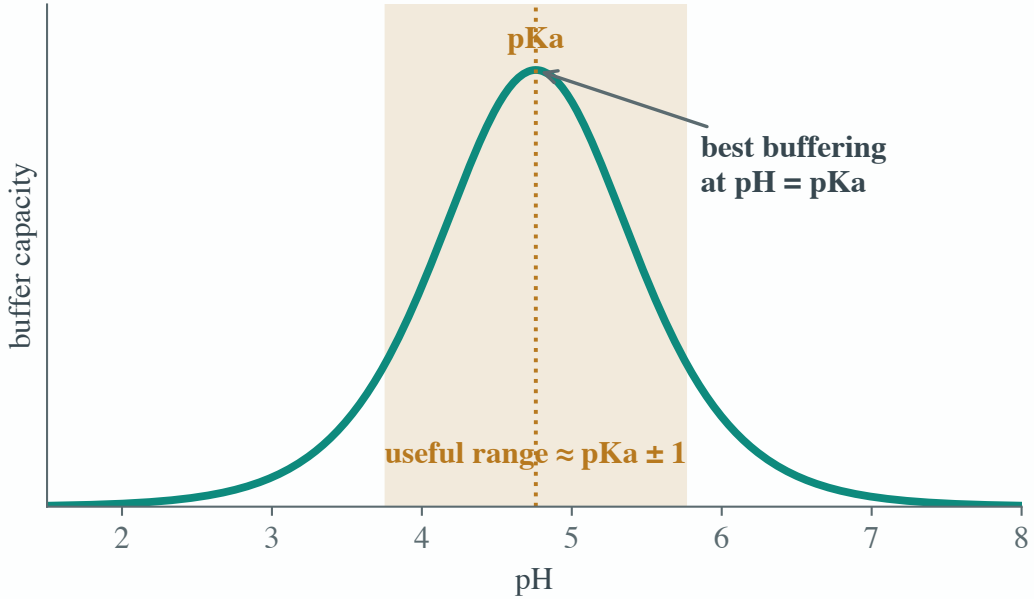
either way, the conjugate pair absorbs the shock — pH barely budges

6 · Buffer Capacity & Blood

"See when a buffer works best (pH near its pKa) and how your blood holds 7.4 with the bicarbonate system."

A buffer works best near its pKa

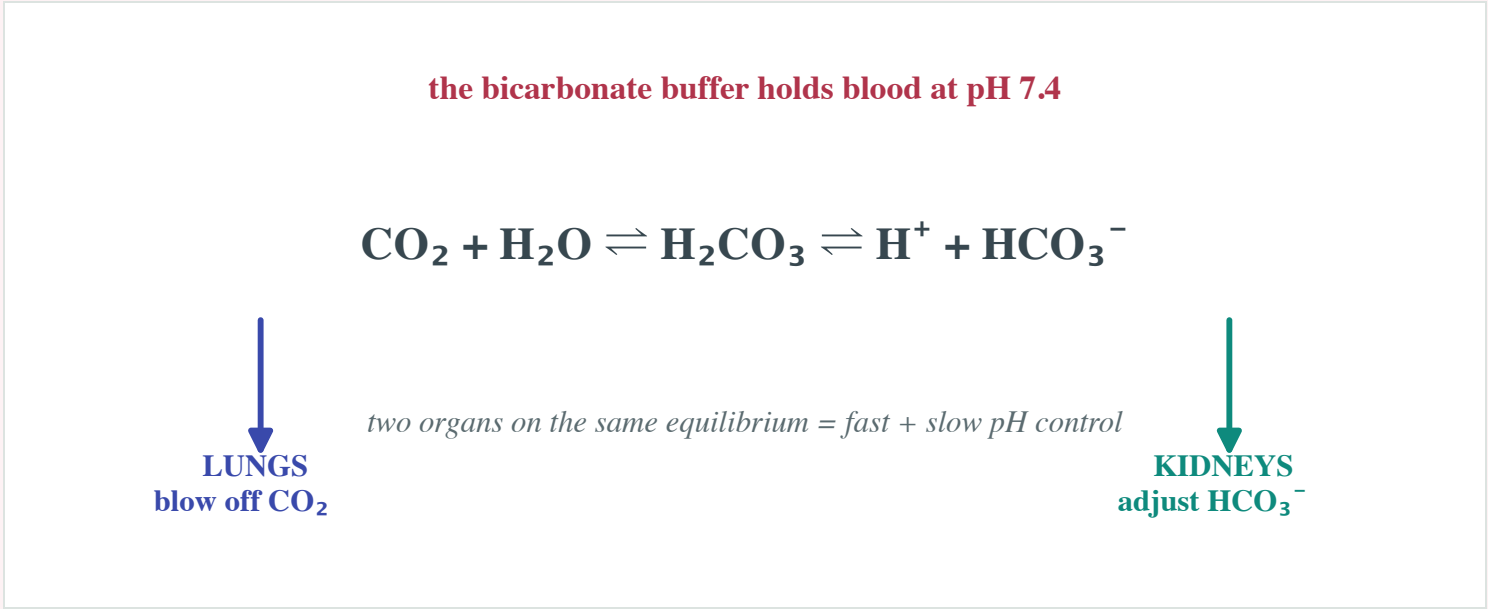
A buffer's **capacity** — how much acid or base it can swallow — peaks when **pH = pKa** and fades fast on either side. The **useful range is about pKa ± 1**. Outside it, one side of the pair runs out and the pH lurches.



💡 So you pick a buffer whose pKa is close to the pH you want to hold.

How your blood holds pH 7.4

Blood runs on the **bicarbonate buffer**: $CO_2 + H_2O \rightleftharpoons H_2CO_3 \rightleftharpoons H^+ + HCO_3^-$. Two organs steer the same equilibrium — the **lungs** blow off CO_2 in seconds, the **kidneys** tune HCO_3^- over hours. Together they pin blood at **pH 7.4**.



💡 Inside cells the **phosphate** buffer (pKa ~6.8) and proteins do the job. Bicarbonate's low pKa (6.1) works only because it's an **open system** — lungs and kidneys keep topping it up.

Can you…?

- □ explain water's polarity and hydrogen bonding, and identify H-bond donors and acceptors?
- □ describe the hydrophobic effect and its role in protein folding and membranes?
- □ work with acids, bases, pH, pKa, the Henderson–Hasselbalch equation, titration curves, and buffer capacity?
- □ explain how physiological buffers (bicarbonate) hold blood at pH 7.4?

If any box stays empty, the practice site has a drill for it.
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