Henderson-Hasselbalch Equation

Essential Question

How can we calculate a buffer's pH?

Henderson-Hasselbalch Equation

The Equation

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

pH	$\mathrm{p}K_a$	$[\mathrm{A}^-]$	$[\mathrm{HA}]$
the buffer's pH	$-\log K_a$ of HA	conjugate base (M)	weak acid (M)

- Use it for buffers: HA mixed with A^-

Henderson-Hasselbalch Equation

Start with pK_a

- pK_a = −log K_a
- Stronger acid → smaller pK_a
- Buffer pH lands near pK_a

Weak acid	K_a	pK_a
HF	6.8×10^{-4}	**3.17**
CH_3COOH	1.8×10^{-5}	**4.74**
HCN	6.2×10^{-10}	**9.21**

smaller K_a → bigger pK_a

Which Is HA? Which Is A^-?

- HA: the weak acid
- A^-: its conjugate base
- Salts supply the A^-

HA (bottom)	A^- (top)
CH_3COOH	CH_3COO^- from $NaCH_3COO$
HF	F^- from NaF
NH_4^+ from NH_4Cl	NH_3

What the Ratio Tells You

- Equal amounts: pH = pK_a
- More base: pH above pK_a
- More acid: pH below pK_a



ratio = 1
pH = pK_a

ratio > 1
pH > pK_a

ratio < 1
pH < pK_a

Quick Log Values

- Ratio 10 adds 1 pH unit
- Ratio 0.1 subtracts 1
- Good buffers: pK_a ± 1

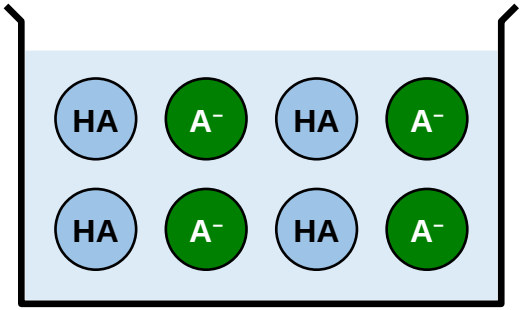
$[A^-]/[HA]$	log	pH
10	+1	**pK_a + 1**
1	0	**pK_a**
0.1	−1	**pK_a − 1**

good buffers stay in this range

Moles Work Too

- Same flask, same volume
- Volume cancels in the ratio
- Plug in moles directly

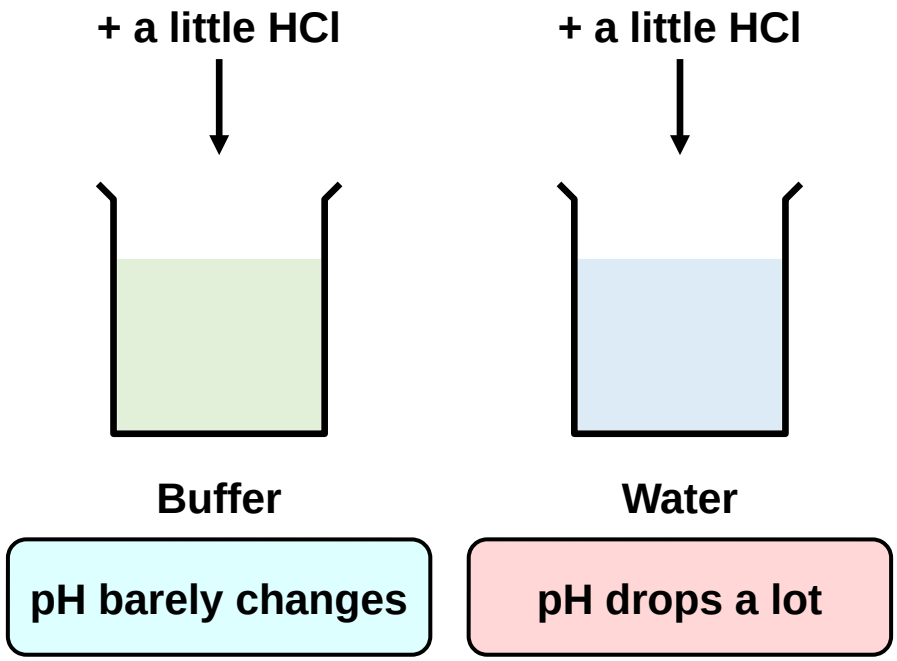
$$\frac{[A^-]}{[HA]} = \frac{\text{mol } A^-}{\text{mol HA}}$$



one flask, one volume

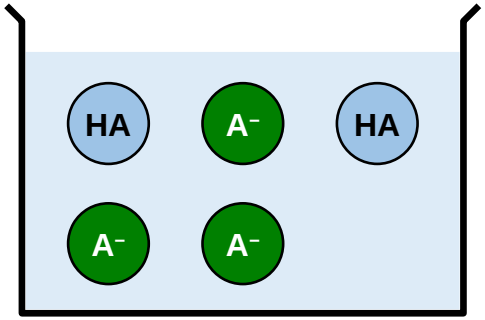
Small Additions, Small Change

- A little acid: ratio barely shifts
- So pH barely changes
- No buffer: big pH change



Example: Acetate Buffer pH

A buffer contains 0.20 M CH_3COOH and 0.30 M $NaCH_3COO$. K_a of acetic acid is 1.8×10^{-5}. Calculate the pH of the buffer.



0.20 M HA
0.30 M A^-

Example: Ammonia Buffer pH

A buffer contains 0.20 M NH_3 and 0.30 M NH_4Cl. The pK_a of NH_4^+ is 9.26. Calculate the pH of the buffer.

NH_4^+
0.30 M

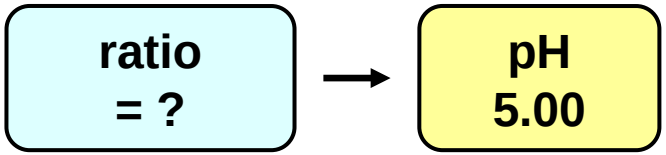
acid?

NH_3
0.20 M

base?

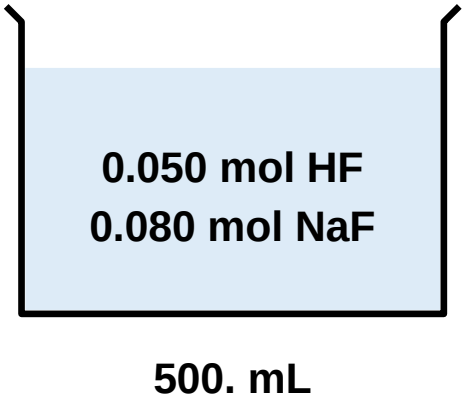
Example: Target pH

A student wants an acetic acid/sodium acetate buffer with a pH of 5.00. K_a of acetic acid is 1.8×10^{-5}. What $[A^-]/[HA]$ ratio is needed?



Example: HF Buffer from Moles

A buffer is made by dissolving 0.050 mol HF and 0.080 mol NaF in enough water to make 500. mL of solution. K_a of HF is 6.8×10^{-4}. Calculate the pH.



Henderson-Hasselbalch Equation

Big Ideas

- Base over acid, inside the log
- Equal amounts: pH = pK_a
- Small additions barely shift pH

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