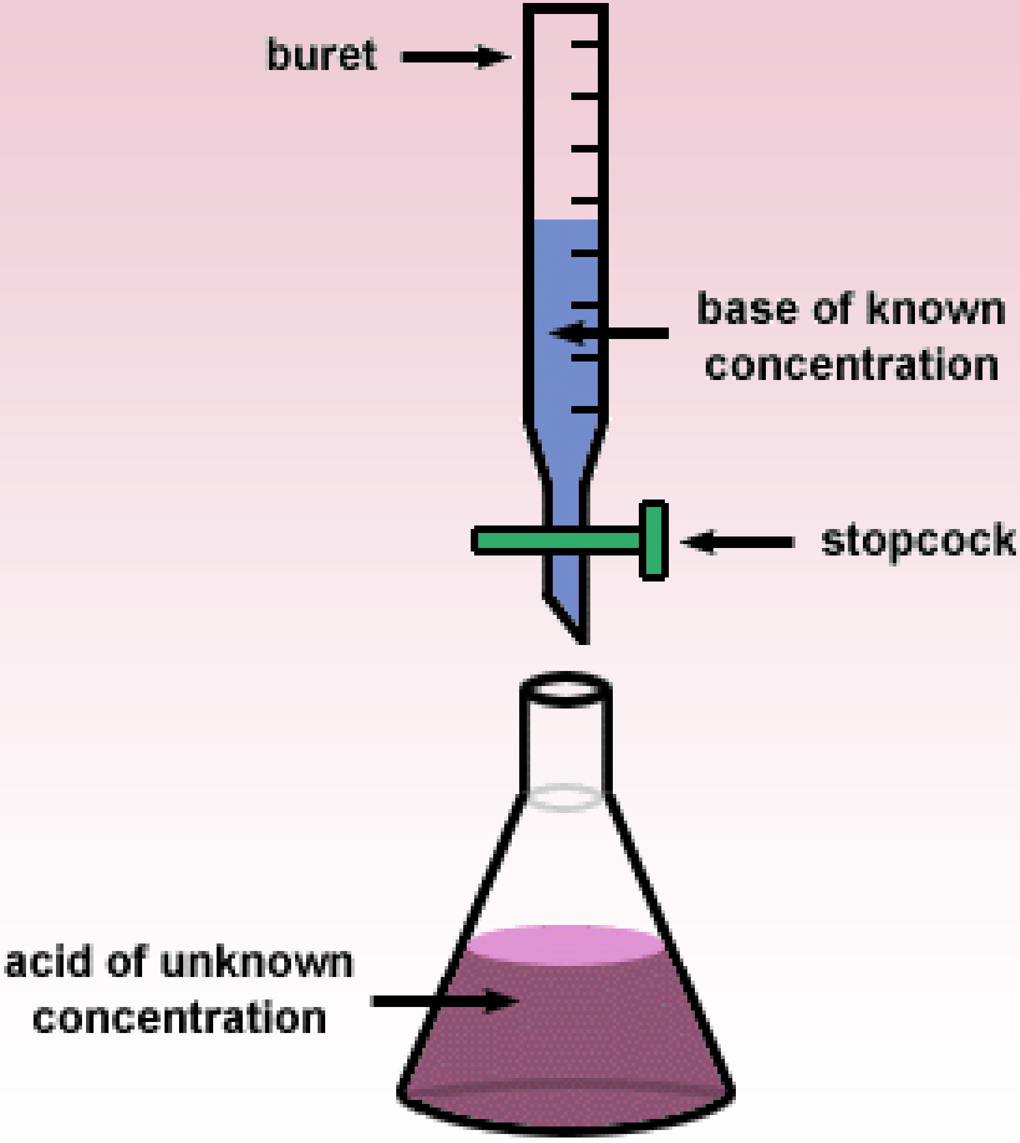
TITRATIONS & BUFFERS

TITRATIONS



Purpose:

To determine the unknown concentration of the acid.

TITRATIONS

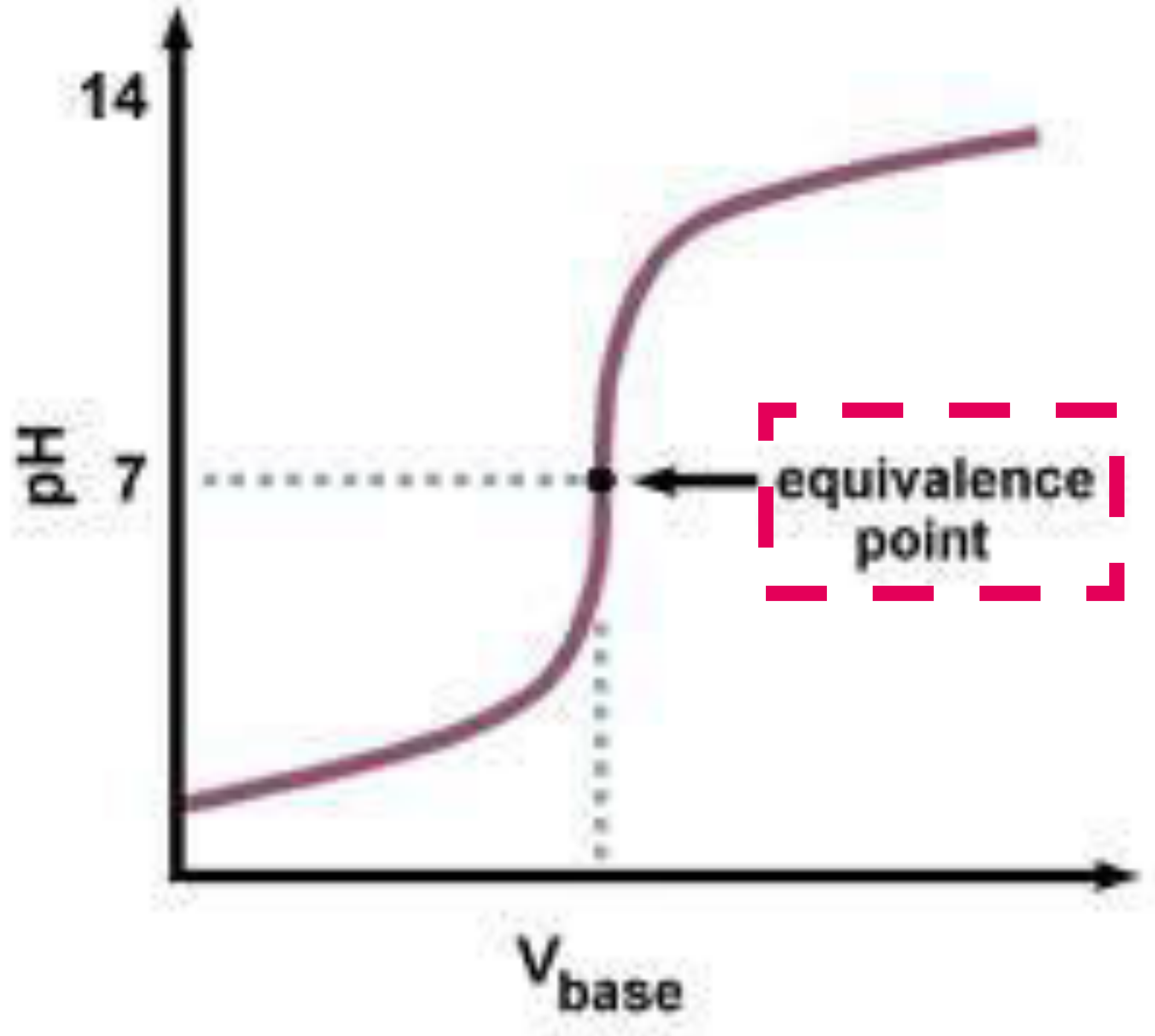
Equivalence point: The point (pH) in the titration when an equal number of moles of acid and base have been added

End point: The point (pH) at which the indicator changes colour indicating an end to the titration

For a successful titration, choose an indicator that changes colour at a pH value close to the pH at the equivalence point.

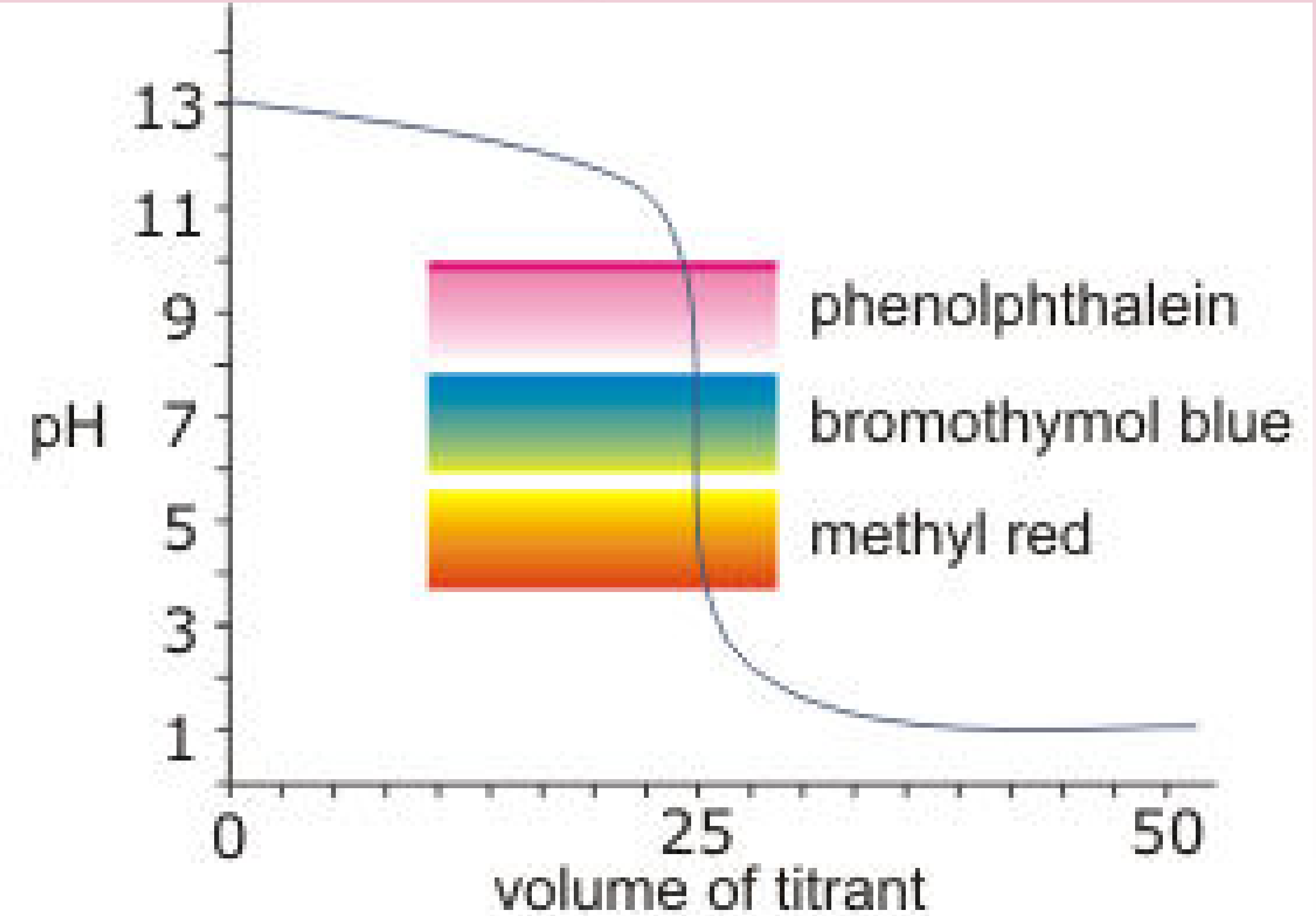
TITRATIONS

Titration Curve



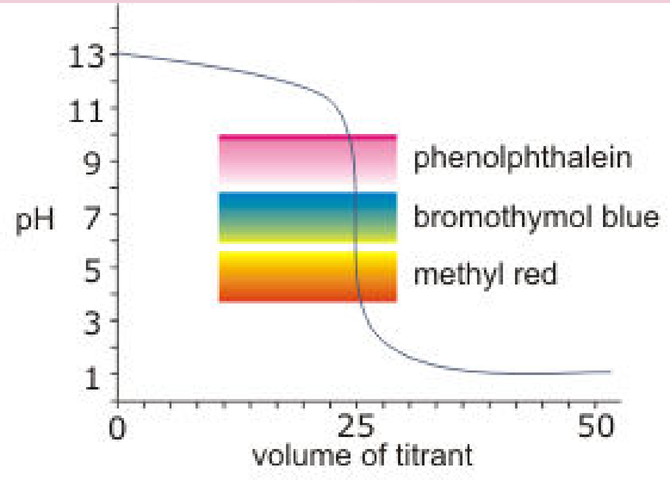
TITRATIONS

Indicator Ranges



TITRATIONS

Indicator Ranges



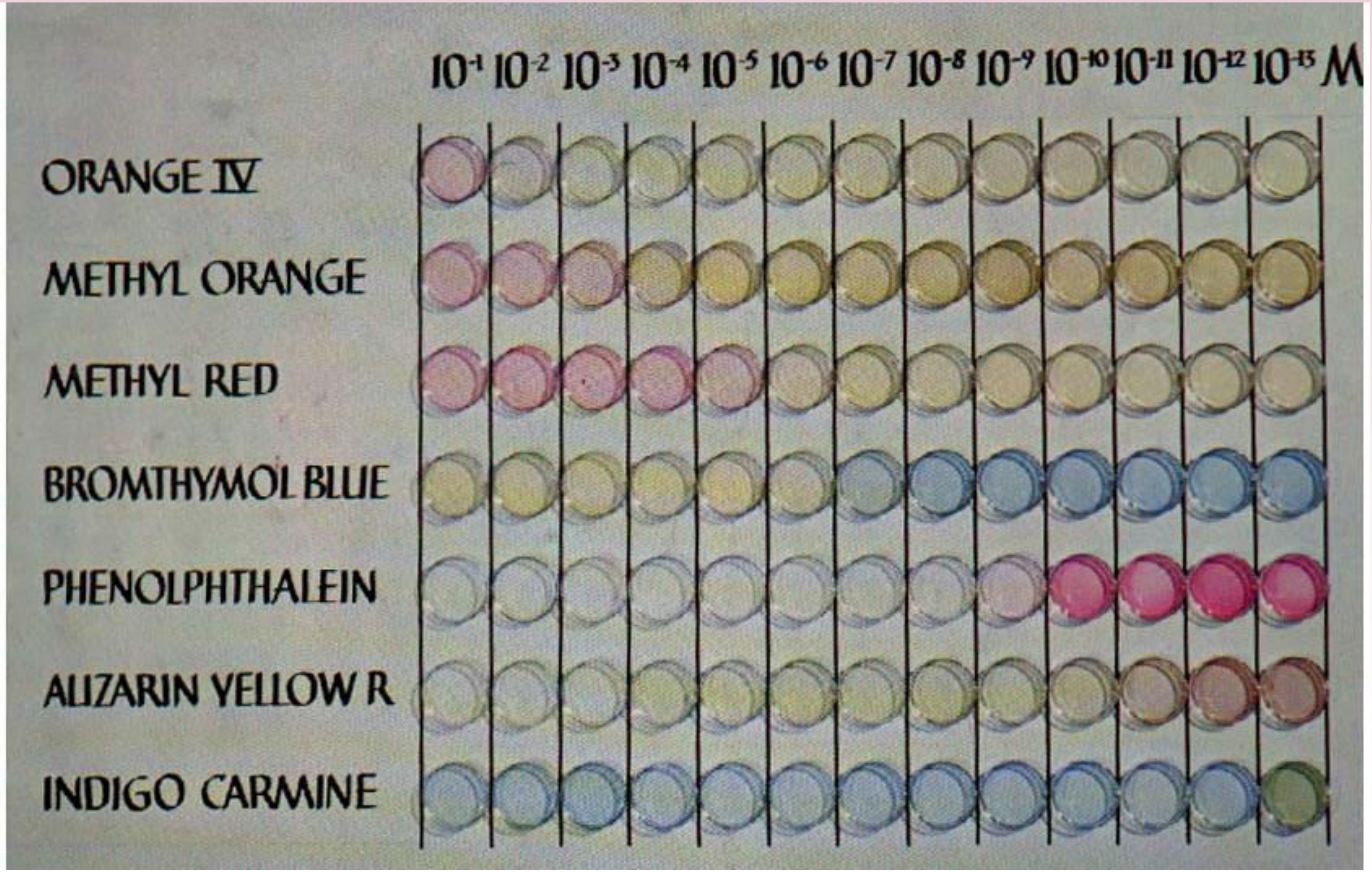
Strong acid + Strong base titration: resulting solution has a pH = 7, so bromothymol blue could be used (pH range is 6.0 – 7.6)

Weak acid + Strong base titration: resulting solution has a pH > 7 so phenolphthalein could be used (pH range is 8.2 – 10.0)

Strong acid + weak base titration: resulting solution has a pH < 7 so methyl orange could be used (pH range is 3.1 – 4.4)

TITRATIONS

Indicator Ranges

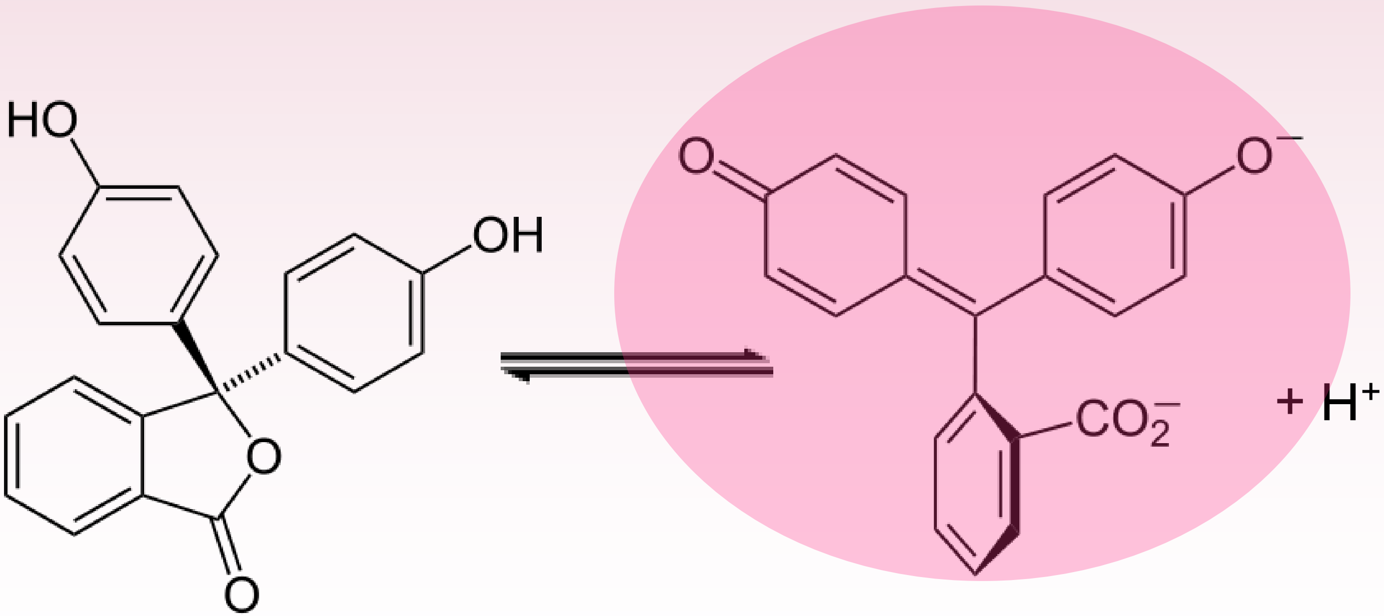


TITRATIONS

Indicators

How do indicators work?

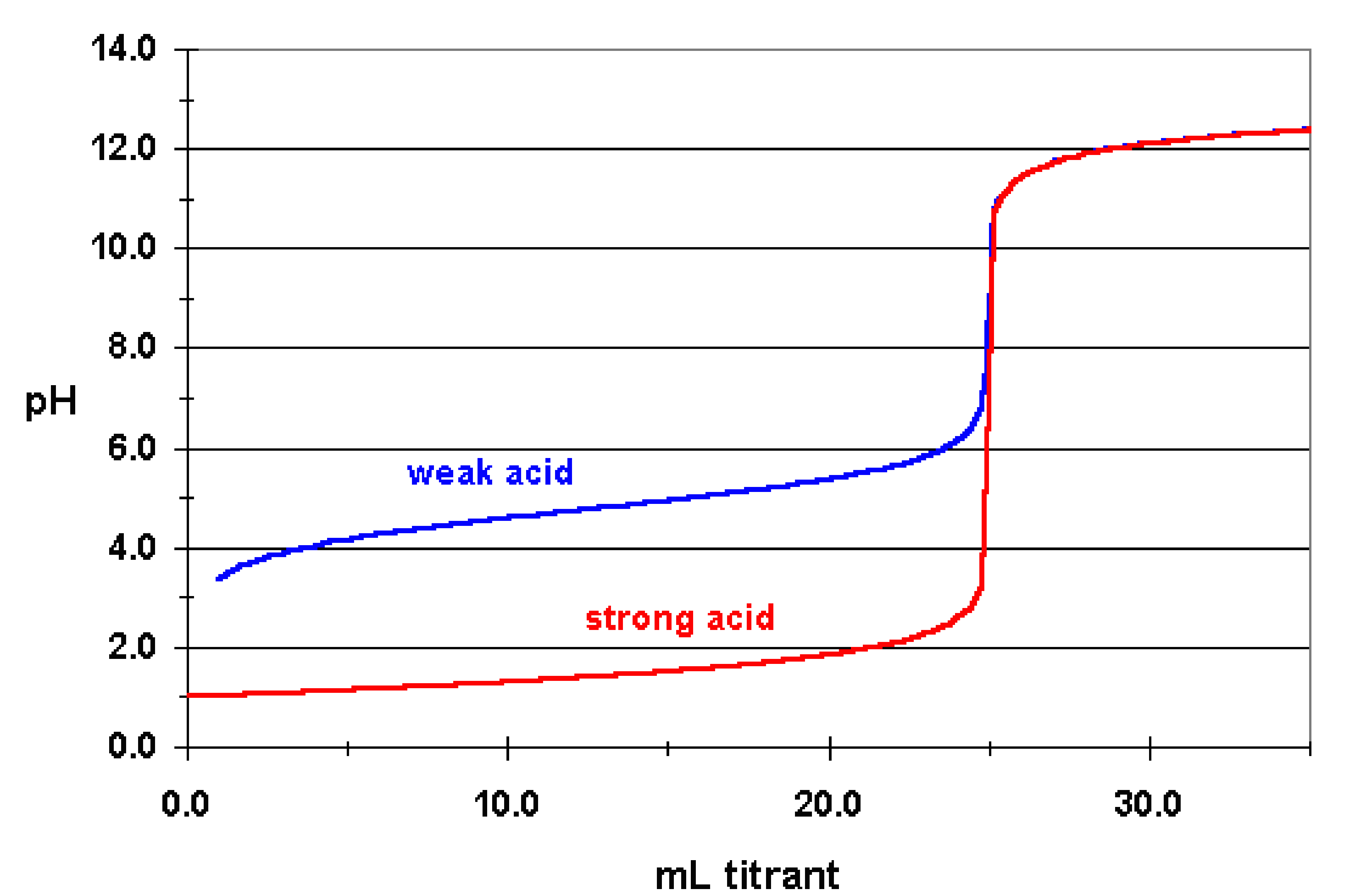
Through equilibrium!



Phenolphthalein

STRONG-WEAK TITRATIONS & BUFFERS

Strong-Weak Titration Curves



Buffers

Buffer solutions contain a mixture of an acid and its conjugate base or a base with its conjugate acid.

Due to the presence of the conjugate acid-base pair, the addition of more acid or base will not cause the pH to drastically change.

Buffers

Buffers are important in biological systems where large changes in pH could be detrimental to the organism.

Almost all organic substances in our body act as a buffer.

$$H_2O + CO_2 \Longleftrightarrow H_2CO_3 \Longleftrightarrow H^+ + HCO_3^-$$

Weak acid
(CO_2 controlled by lungs)

Weak base
(Controlled by kidneys)

Buffers

Buffers in Food, Drink & Medicines

- Sodium citrate + citric acid : buffer in food that acts like a preservative
- Citric acid + calcium citrate: buffer that deters bacterial growth
- Aspartame: decomposition is pH dependent
- Medicines: coating reduces the risk of upset stomach

Strong-Weak Titrations

What happens when you add a strong base to a weak acid?

$$CH_3COOH <===> CH_3COO^- + H^+$$

$$NaOH \rightarrow Na^+ + OH^-$$

Strong-Weak Titrations

Two-step process:

1. All OH^- ions from the strong base will react with H^+ to produce H_2O.
2. Equilibrium is shifted to the
3. pH is then calculated from the new equilibrium based on the remaining $[H^+]$.

Strong Weak Titrations

What happens when you add a strong acid to a weak base?

$$NH_3 + H_2O \Longleftrightarrow NH_4^+ + OH^-$$

$$HCl \rightarrow H^+ + Cl^-$$

Strong-Weak Titrations

Two-step process:

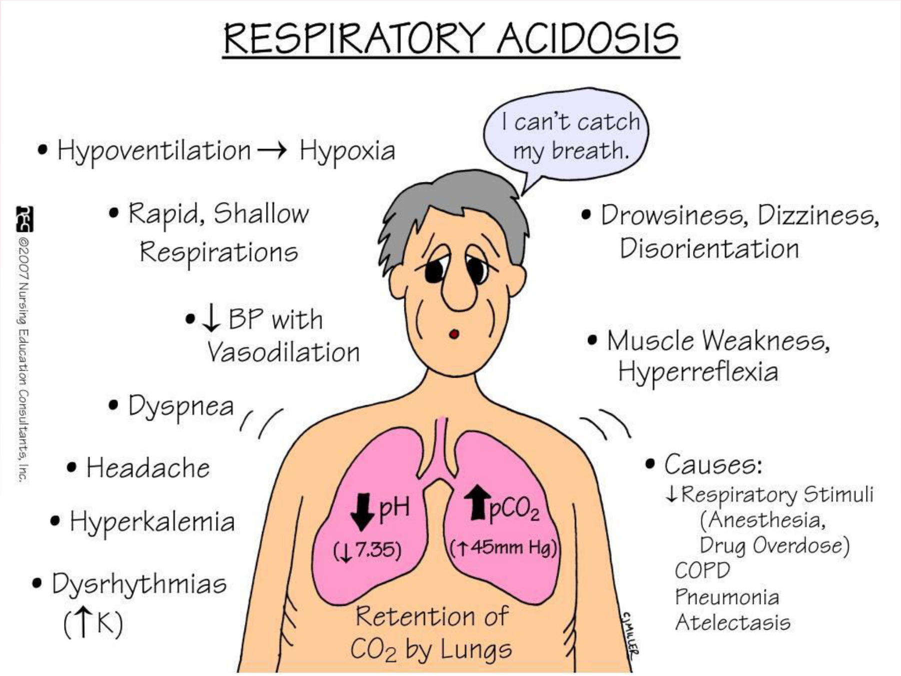
1. All H^+ ions from the strong acid will react with OH^- to produce H_2O.
2. Equilibrium is shifted to the $\boxed{}$
3. pH is then calculated from the new equilibrium based on the remaining $[H^+]$.

Buffers

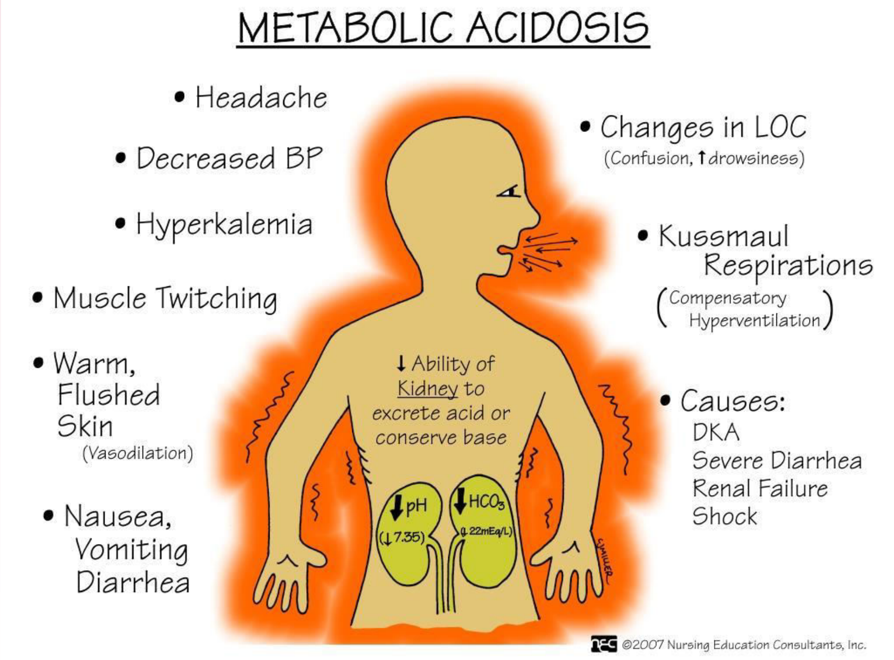
Too much acid = **acidosis**

$$H_2O + CO_2 <===> H_2CO_3 <===> H^+ + HCO_3^-$$

Respiratory (Lungs) Acidosis:



Metabolic (Kidneys) Acidosis:

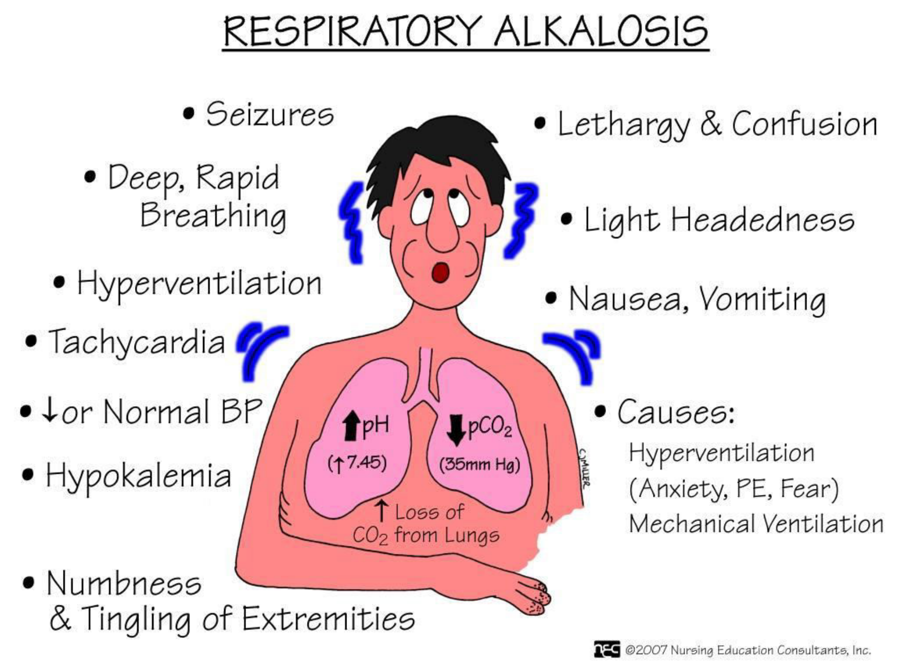


Buffers

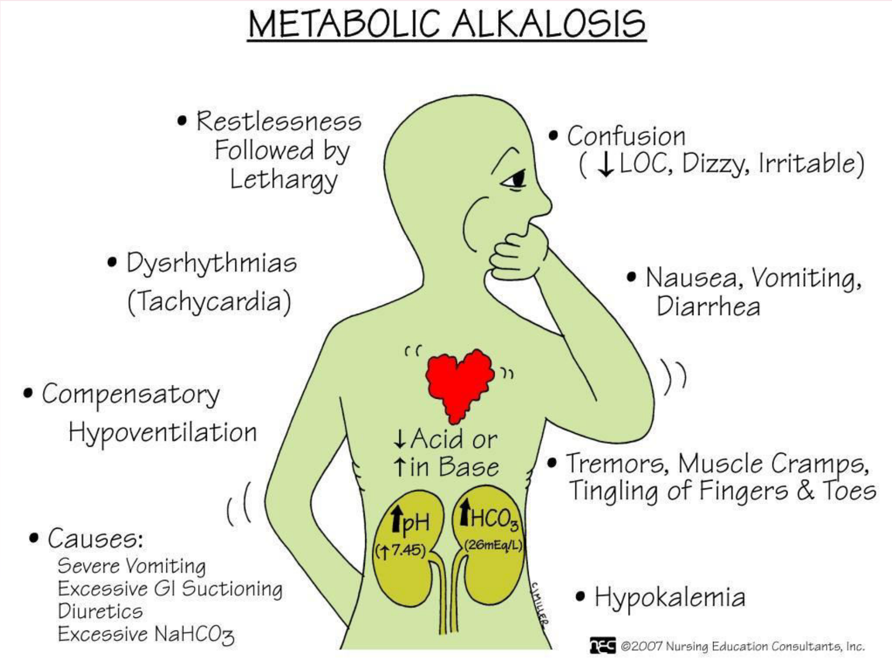
Too much base = **alkalosis**

$$H_2O + CO_2 <===> H_2CO_3 <===> H^+ + HCO_3^-$$

Respiratory (Lungs) Alkalosis:



Metabolic (Kidneys) Alkalosis:



Buffer Equations

The following can be used to determine ion concentrations in buffers:

Acid Buffer:

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

HA = acid

A^- = conjugate base

Base Buffer:

$$[OH^-] = K_b \times \frac{[B]}{[HB^+]}$$

B = base

HB^+ = conjugate acid

Buffer Equations

Alternatively, you can use moles instead of concentration to get the same answer:

Acid Buffer:

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

HA = acid

A^- = conjugate base

Base Buffer:

$$[OH^-] = K_b \times \frac{\text{moles of B}}{\text{moles of HB}^+}$$

B = base

HB^+ = conjugate acid

Buffer Equations

A 1.00 L sample of an aqueous solution contains 0.200 mol of acetic acid and 0.100 mol of acetate. Calculate:

a) The pH of the solution

b) The pH of the solution after adding 1.00 mL of 12.0 M HCl

Buffer Equations

A 1.00 L sample of an aqueous solution contains 0.200 mol of acetic acid and 0.100 mol of acetate. Calculate:

a) The pH of the solution

b) The pH of the solution after adding 1.00 mL of 12.0 M HCl

a) $$HC_2H_3O_{2(aq)} \Longleftrightarrow H^+_{(aq)} + C_2H_3O_2^-{}_{(aq)}$$

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

$$[H^+] = (1.8 \times 10^{-5}) \times \frac{(0.200 \text{ mol})}{(0.100 \text{ mol})}$$

$$[H^+] = 0.000036 \text{ mol/L}$$

$$pH = -\log [H^+] \; 0.000036 \text{ mol/L}$$

$$pH = -\log (0.000036 \text{ mol/L})$$

$$pH = 4.44$$

.: pH = 4.44

Buffer Equations

A 1.00 L sample of an aqueous solution contains 0.200 mol of acetic acid and 0.100 mol of acetate. Calculate:

a) The pH of the solution

b) The pH of the solution after adding 1.00 mL of 12.0 M HCl

b) moles HCl: $n = C \times V$

$n = (12\ M)(0.001L)$

$n = 0.012\ mol$

Assume this produces 0.012 mol of acetic acid and uses up 0.012 mol of acetate

moles acetic acid: $0.200 + 0.012 = 0.212\ mol$

moles acetate: $0.100 - 0.012 = 0.088\ mol$

$$[H^+] = K_a \times \frac{\text{mols } HC_2H_3O_2}{\text{mols } C_2H_3O_2^-}$$

$$[H^+] = (1.8 \times 10^{-5}) \times \frac{(0.212\ mol)}{(0.088\ mol)}$$

$$[H^+] = 4.3363 \times 10^{-5}$$

$pH = -\log [H^+]$

$pH = -\log (4.3363 \times 10^{-5})$

$pH = 4.36$

.: $pH = 4.36$

The Effect of Adding a Strong Acid or Base to a Buffer

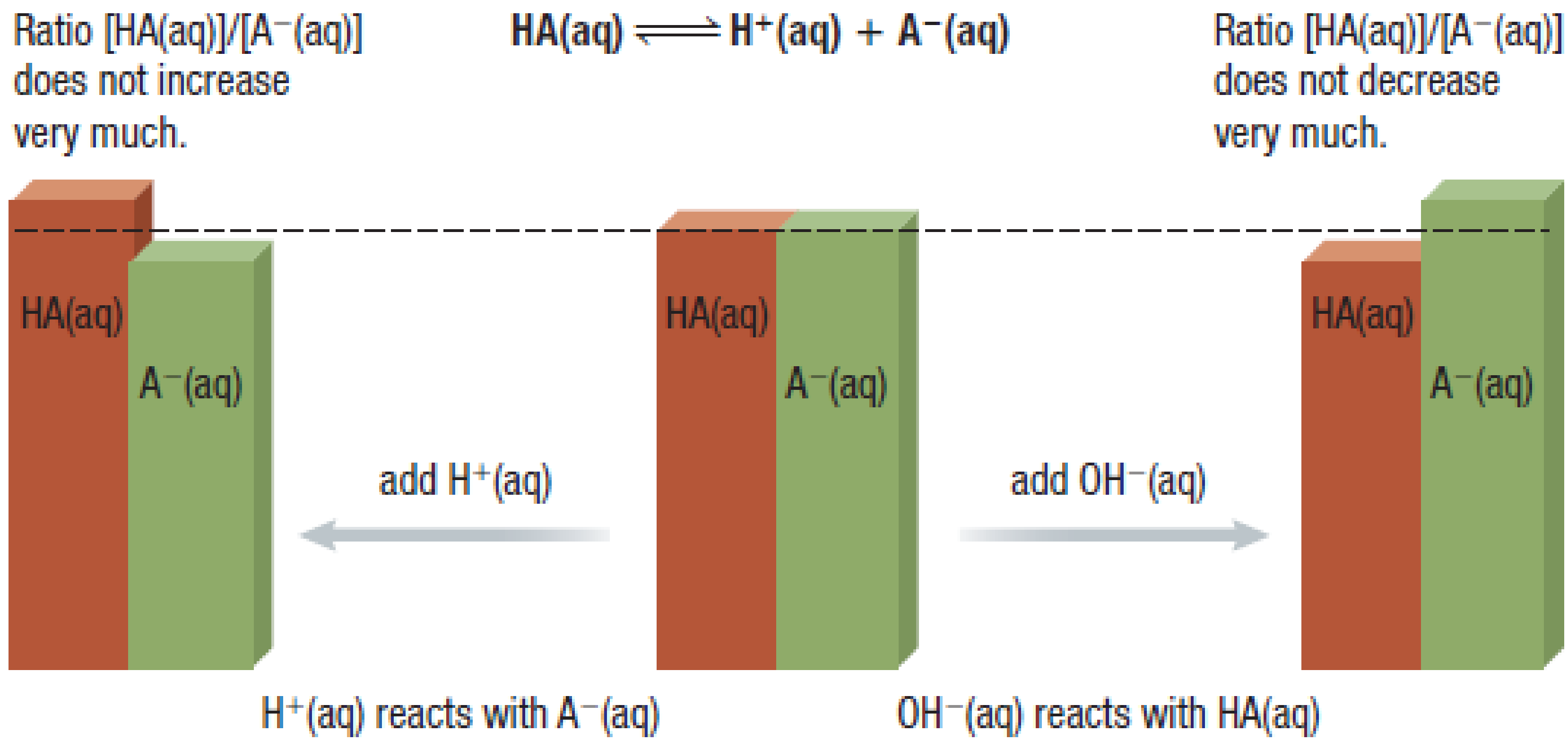
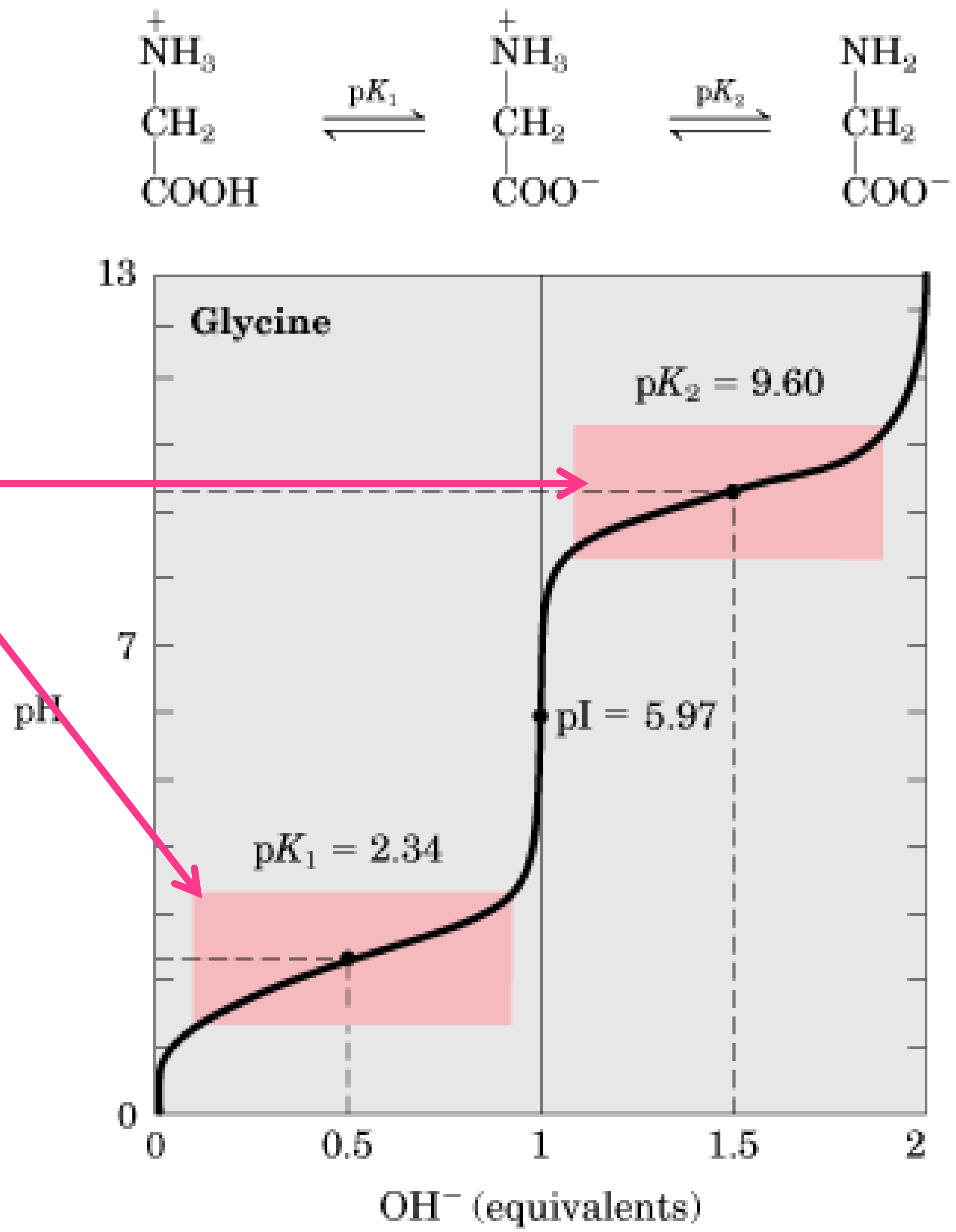


Figure 7 The effect of adding a strong acid or base to a buffer system

Buffers

Notice the buffer prevents the pH from rising too quickly at these two ranges *(slope of the curve is <u>decreased</u> at these points)*



Buffers

Also notice that the pKa of the buffer matches the pH that the buffer works best in.

This is because

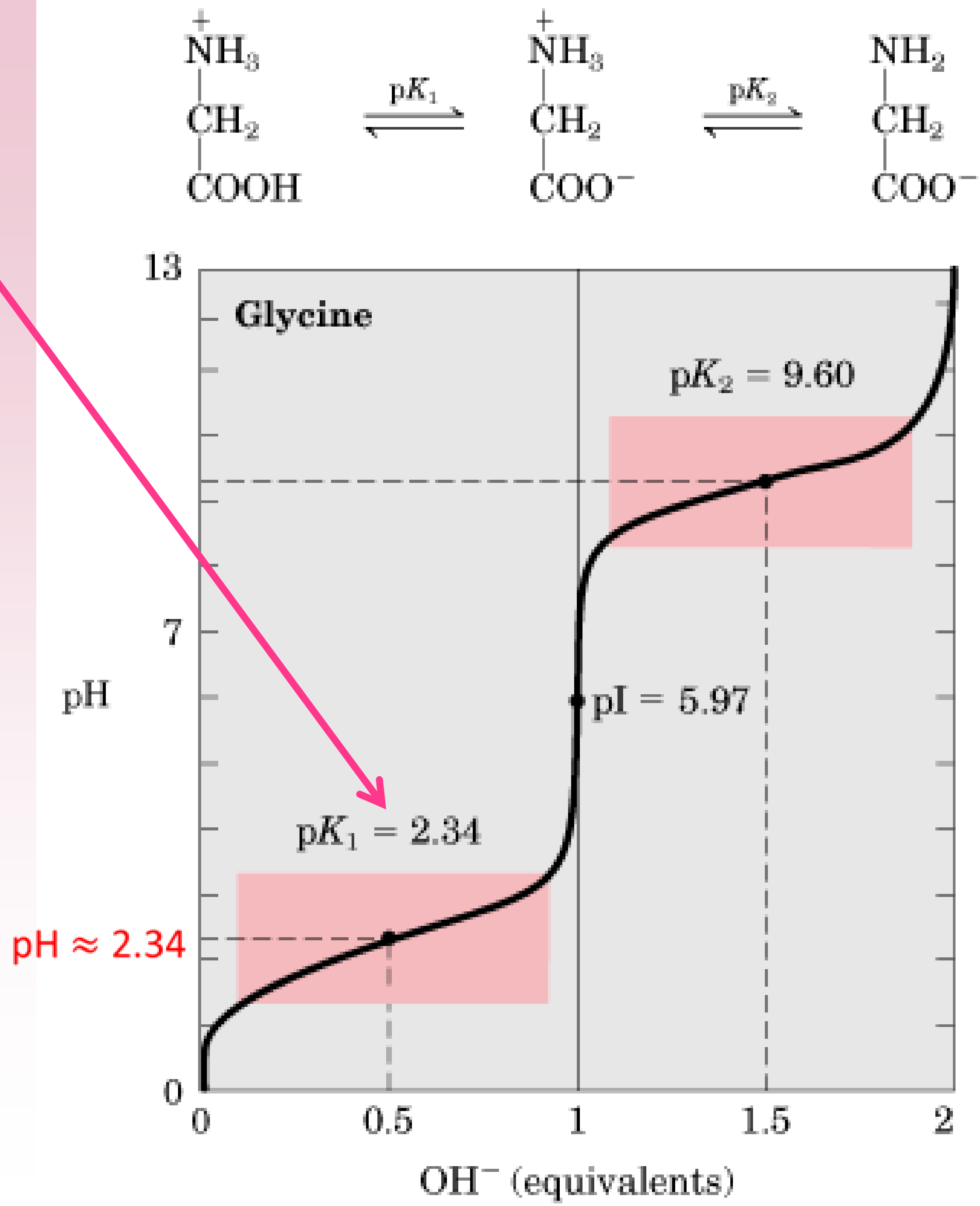
$$pH \approx pK_a + log\frac{[Base]}{[Acid]}$$

and when the concentration of base and acid are equal (i.e. in a buffer system), it becomes

$$pH \approx pK_a + \log[1]$$

$$pH \approx pK_a + 0$$

$$pH \approx pK_a$$



The Capacity of a Buffer

- The amount of added H_3O^+ or OH^- that a buffer can absorb without a significant change in pH

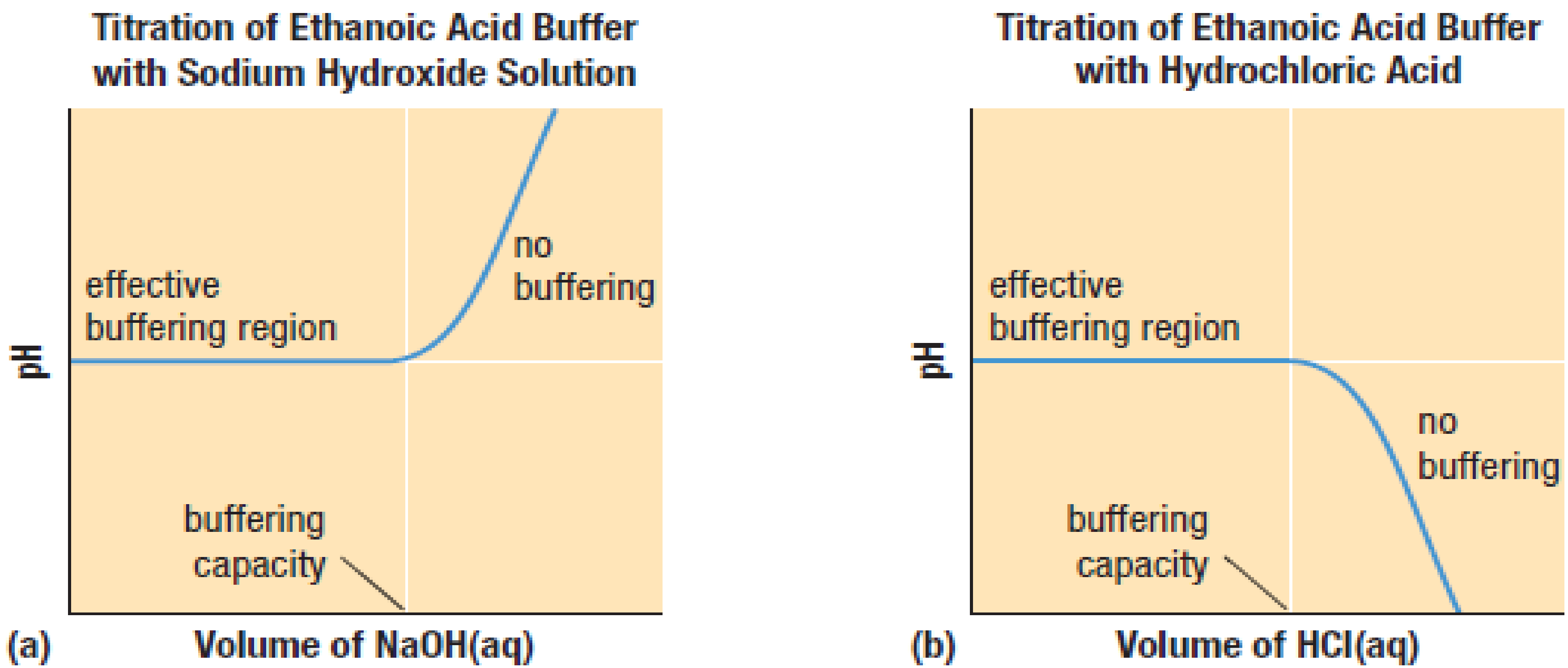


Figure 6 (a) Ethanoic acid buffer with a strong base added (b) Ethanoic acid buffer with a strong acid added. The pH changes quickly once all of the available buffer is depleted.

The Capacity of a Buffer

- To calculate the buffer region, find the equivalence point volume and divide that by 2.