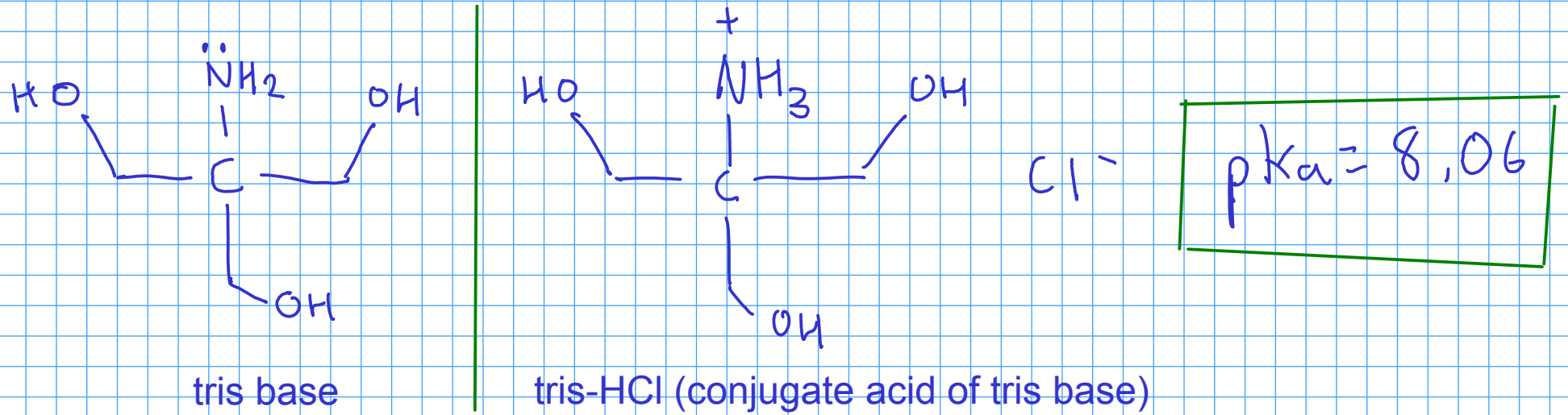


# Buffer calculation: Tris buffer - Tris(hydroxymethyl)-aminomethane



Calculate the pH of a buffer made from 50 mL of 0.10M tris and 50 mL of 0.15M tris-HCl

$$pH = pK_{a, \text{acidic}} + \log \left( \frac{[\text{basic species}]}{[\text{acidic species}]} \right)$$

$$\text{tris} : M_1 V_1 = M_2 V_2 \quad (0.10)(50) = M_2 (100) ; [\text{tris}] = 0.050 \text{ M}$$

$$[\text{tris-HCl}] = 0.075 \text{ M}$$

$$pH = 8,06 + \log \left( \frac{0,050}{0,075} \right) = 7,88$$

Take 100. mL of the 0.10 M tris / 0.15 M tris-HCL buffer, and add 5.0 mL of .10 M HCl.  
What is the pH of the mixture?

The hydrochloric acid will react with the tris base:



Use H-H equation, but to do that ...

$$\text{pH} = \text{pK}_{a, \text{tris-HCl}} + \log \left( \frac{[\text{tris}]}{[\text{tris-HCl}]} \right)$$

... I need to calculate the new [tris] and [tris-HCL]

$$\text{tris: } 100 \text{ mL} \times 0.10 \text{ M} = 10 \text{ mmol}$$

$$\text{tris-HCl: } 100 \text{ mL} \times 0.15 \text{ M} = 15 \text{ mmol}$$

$$\text{HCl: } 5 \text{ mL} \times 0.10 \text{ M} = 0.5 \text{ mmol}$$

$$\text{tris: } 10 - 0.5 = 9.5 \text{ mmol}$$

$$\text{tris-HCl: } 15 + 0.5 = 15.5 \text{ mmol}$$

$$\text{HCl: } 0.5 - 0.5 = 0 \text{ mmol}$$

$$[\text{tris}] = \frac{9.5 \text{ mmol}}{105 \text{ mL}} = 0.0905 \text{ M}$$

$$[\text{tris-HCl}] = \frac{15.5 \text{ mmol}}{105 \text{ mL}} = 0.148 \text{ M}$$

$$\text{pH} = 8.06 + \log \left( \frac{0.0905}{0.148} \right) = \boxed{7.85}$$

pH decreased by 0.03  
pH units!

Compare this 0.03 pH change with adding 5.0 mL of 0.10 M HCl to 100. mL of pure water.

$$M_1 V_1 = M_2 V_2$$

$$(0.10 \text{ M})(5.0 \text{ mL}) = M_2 (105 \text{ mL})$$

$$0.00476 \text{ M} = M_2 = [\text{H}_3\text{O}^+] \dots \text{strong acid!}$$

$$\text{pH} = -\log(0.00476) = 2.32$$

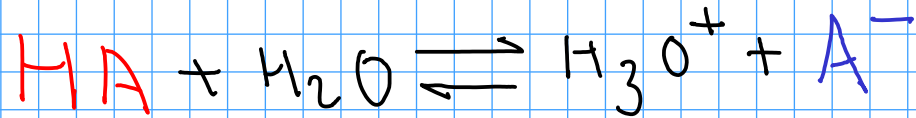
This is a >4 pH unit change from water's original pH of 7!

## INDICATORS

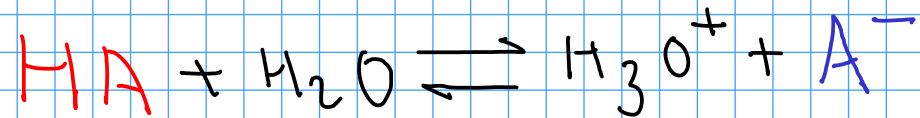
- Instead of using a pH meter to monitor acidity, we may choose to use an acid-base INDICATOR.

- Acid-base indicators are weak acids or weak bases which are highly colored.

- The color of the undissociated indicator MUST BE DIFFERENT than the color of the dissociated form!



The indicator must be present in very low concentrations - so that the indicator's equilibrium DOES NOT CONTROL the pH of the solution!



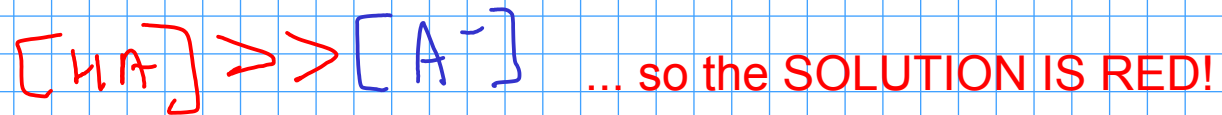
Look at the Henderson-Hasselbalch equation - we want to know how much of the red form and how much of the blue form are present!

$$\text{pH} = \text{pK}_{a,\text{ind}} + \log \left( \frac{[\text{A}^-]}{[\text{HA}]}\right)$$

When does the color of the indicator change?

IF the pH is  $\ll$  pKa, then the log term above must be both large AND negative!

- What color is the solution?



If the pH is  $\gg$  pKa, then the log term above must be both large AND positive!

- What color is the solution?



- So, the color changes when the pH of the solution is near the pKa of the indicator, BUT we can only DETECT the change when enough of the other form is present.