# BUFFER CAPACITY

- A buffer is good only as long as there is a significant concentration of both the acidic and basic species

- buffer capacity: how much acid or base can a buffer resist before losing its ability to buffer

- Buffer pH depends on the RATIO of acid to base!

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

Henderson-Hasselbalch Equation

Ratio determines pH; the actual concentrations don't!

- So, if you make a buffer with 1.0M HA and 1.0M A- , it will have the same pH as a buffer with 2.0M HA and 2.0M A- .... but the 2M buffer will have a higher BUFFER CAPACITY - it will resist more additions of acid or base.

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

HO $\overset{..}{N}H_2$ OH

$C$

OH

tris base

$HO$ $\overset{+}{N}H_3$ OH

$C$

OH

$Cl^-$
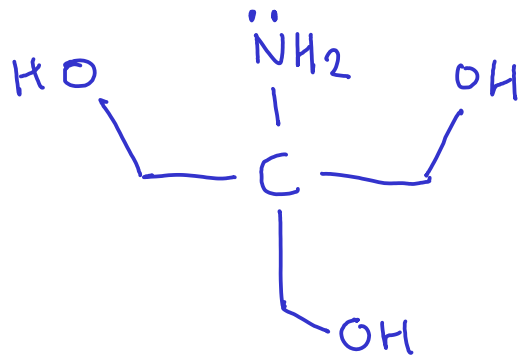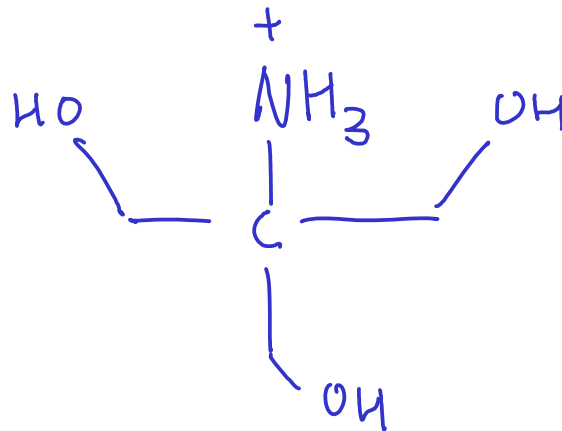
tris-HCl (conjugate acid of tris base)

$pKa = 8.06$

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

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

$[tris]$: $M_1V_1 = M_2V_2$; $(0.10\,M)(50.\,mL) = M_2(100.\,mL)$

$0.050\,M\ tris = M_2$

$[tris-H^+]$: $M_1V_1 = M_2V_2$; $(0.15\,m)(50.\,mL) = M_2(100.\,mL)$

$0.075\,M\ tris-H^+ = M_2$

$$pH = 8.06 + log\left(\frac{0.050\,m}{0.075\,m}\right) = 7.88$$

Take 100. mL of the previous buffer (0.050 M tris / 0.075 M tris-HCl), and add 5.0 mL of 0.10 M HCl. What is the pH of the mixture?

The HCl should react with basic component of the buffer (tris), and change it to its conjugate acid

$$\text{tris} + H_3O^+ \longrightarrow \text{tris-}H^+ + H_2O$$

(from HCl)
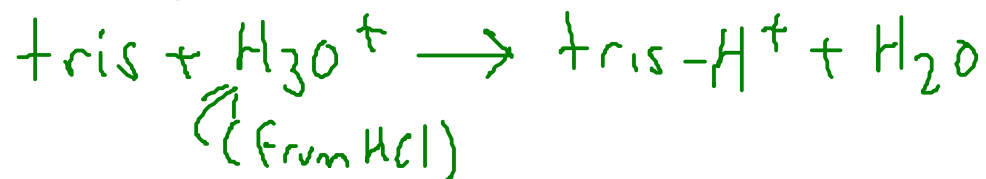
... so we need to find out the NEW concentrations of each species in the system.

| Species | Initial mmol | Δ in rxn | Final mmol | [conc.] |
|---------|-------------|----------|-----------|---------|
| tris | 100 mL × 0.050 m = 5.0 mmol | −0.5 mmol | 4.5 mmol | $\frac{4.5\ mmol}{105\ mL}$ = 0.0428571 m * |
| tris-H⁺ | 100 mL × 0.075 m = 7.5 mmol | +0.5 mmol | 8.0 mmol | $\frac{8.0\ mmol}{105\ mL}$ = 0.0761905 m |
| HCl | 5 mL × 0.10 M = 0.5 mmol | −0.5 mmol | 0 mmol | 0 |

* Solution volume is now 105 mL (100 mL of buffer plus 5 mL of HCl)

$$pH = 8.06 + \log\left(\frac{0.0428571\ m}{0.0761905\ m}\right) = \boxed{7.81}$$

The original pH was 7.88, so the pH dropped by 0.07 pH units.

Compare this 0.07 unit pH change with adding 5.0 mL of 0.10 M HCl to 100. mL of pure water.

(We're just diluting the acid...)

$$M_1 V_1 = M_2 V_2$$

$$(0.10 \, m)(5.0 \, mL) = M_2 (105 \, mL)$$

$$0.0047619 \, m = M_2$$

Since this is a strong acid, hydronium ion concentration equals nominal acid concentration:

$$[H_3O^+] = 0.0047619 \, m$$

$$\boxed{pH = 2.32}$$

... which is a change of 4.68 pH units from water's original pH of 7.00!

# 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!

RED                                    BLUE

$$HA + H_2O \rightleftharpoons H_3O^+ + A^-$$
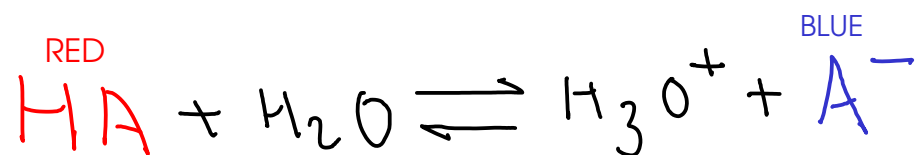
The indicator must be present in very low concentrations -
so that the indicator's equilibrium DOES NOT CONTROL
the pH of the solution!

$$HA + H_2O \rightleftharpoons H_3O^+ + A^-$$
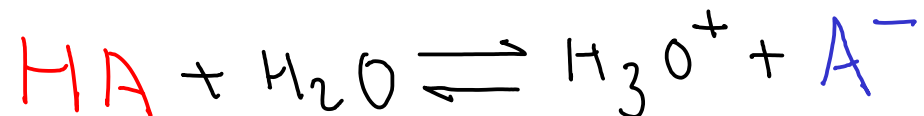
Look at the Henderson-Hasselbalch equation - we want to know how much of the red form and how much of the blue form are present!

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

When does the color of the indicator change?

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

- What color is the solution?

$$[HA] >> [A^-]$$ ... and the solution is RED.

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

- What color is the solution?

$$[A^-] >> [HA]$$ ... and the solution is BLUE

- 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.