

Calculate the pH of a buffer made from 3.2 grams of ammonium chloride and 29 mL of 18.1 M ammonia diluted to 50.0 mL with water.

$$\text{pH} = \text{p}K_{a,\text{acidic}} + \log\left(\frac{[\text{basic species}]}{[\text{acidic species}]}\right) \quad \left| \begin{array}{l} \text{Henderson-} \\ \text{Hasselbalch} \\ \text{Equation} \end{array} \right.$$



We need to find the concentrations of ammonia (the base) and ammonium ion (the acid)

$[\text{NH}_3]$ :  $m_1 V_1 = m_2 V_2$  Dilution equation  
 $(18.1 \text{ M})(29 \text{ mL}) = M_2 (50.0 \text{ mL})$   
 $10.498 \text{ M NH}_3 = M_2$

$[\text{NH}_4^+]$ :

$\text{NH}_4\text{Cl}$	$\text{N}: 1 \times 14.01$
	$\text{H}: 4 \times 1.008$
	$\text{Cl}: 1 \times 35.45$
	$53.492 \text{ g/mol}$

$$3.2 \text{ g NH}_4\text{Cl} \times \frac{\text{mol NH}_4\text{Cl}}{53.492 \text{ g NH}_4\text{Cl}} \times \frac{1}{0.0500 \text{ L}} = 1.196 \text{ M NH}_4^+$$

$\text{p}K_{a,\text{NH}_4^+}$      $K_{b,\text{NH}_3} = 1.8 \times 10^{-5}$  ( $\text{p}K_{b} = 4.745$ ), so  $\text{p}K_{a} = 9.255$

$4.745 + \text{p}K_{a} = 14$ ,  $\text{p}K_{a} = 9.255$

$$\text{pH} = 9.255 + \log\left(\frac{10.498}{1.196}\right) = 10.20$$

## 161 BUFFER SELECTION

- Buffer pH is controlled by the pKa of the acidic species in the buffer.

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

- Choose a buffer system so that the desired pH is within +/- 1 pH unit of the pKa
- You also need to ensure that the components of the buffer do not interact with your chemistry!

Some buffer "recipes" call for making the conjugate ion FROM the weak acid or base ... by adding a STRONG acid or base!



If you have more ammonia than nitric acid, you will end up with a solution containing a significant amount of both ammonia and ammonium ion ... a buffer!

## 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!

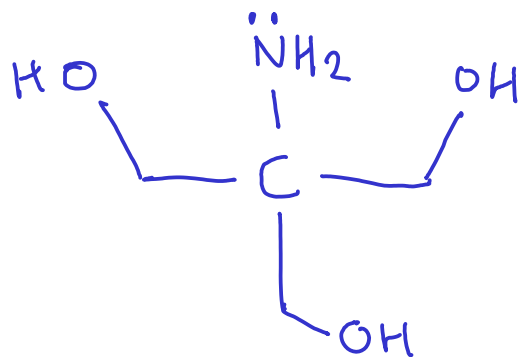
$$\text{pH} = \text{p}K_{a, \text{acidic}} + \log \left( \frac{[\text{basic species}]}{[\text{acidic species}]} \right)$$

Henderson-Hasselbalch Equation

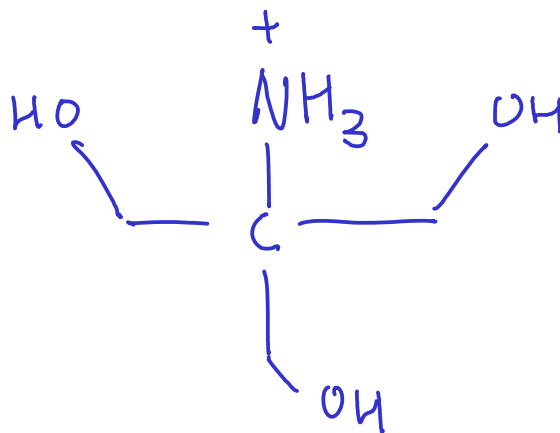
Ratio determines pH; the actual concentrations don't!

- So, if you make a buffer with 1.0M HA and 1.0M A<sup>-</sup>, it will have the same pH as a buffer with 2.0M HA and 2.0M A<sup>-</sup> .... 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



tris base



tris-HCl (conjugate acid of tris base)

$\text{Cl}^-$

$$pK_a = 8,06$$

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

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

$$[\text{tris}] = m_1 V_1 = m_2 V_2 \quad (0,10 \text{ M})(50 \text{ mL}) = m_2 (100 \text{ mL})$$

$$m_2 = [\text{tris}] = 0,050 \text{ M}$$

$$[\text{tris-HCl}] = m_1 V_1 = m_2 V_2 \quad (0,15 \text{ M})(50 \text{ mL}) = m_2 (100 \text{ mL})$$

$$m_2 = [\text{tris-HCl}] = 0,075 \text{ M}$$

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