

Basic buffer solutions

We are talking here about a mixture of a weak base and one of its salts - for example, a solution containing the weak base, ammonia,  $\text{NH}_3$  and the salt, ammonium chloride,  $\text{NH}_4\text{Cl}$ .

**Example: Calculate the pH of a solution containing  $0.100 \text{ mol L}^{-1} \text{NH}_3$  and  $0.0500 \text{ mol L}^{-1} \text{NH}_4\text{Cl}$ ?**

$$K_a(\text{NH}_4^+) = 5.62 \times 10^{-10}$$

You may have learnt the (Henderson-Hasselbalch) buffer equation in class:

$$\text{pH} = \text{p}K_a + \log \frac{[\text{base}]}{[\text{acid}]}$$

But this equation is **NOT** on the supplied resource sheet so you either must rely on remembering it OR use the equation that is supplied....

That looks a bit unlikely as it is a  $K_a$  expression..... and weak acid and all.... 😊 Don't panic!

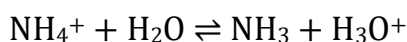
$$K_a = \frac{[\text{H}_3\text{O}^+][\text{A}^-]}{[\text{HA}]}$$

The way of doing this basic buffer calculation is to *re-think it* from the point of view of the ammonium ion rather than of the ammonia solution.

Once you have taken this different view-point, you can easily use the equation supplied by NZQA.

$$K_a = \frac{[\text{H}_3\text{O}^+][\text{A}^-]}{[\text{HA}]}$$

The buffer will contain lots of unreacted ammonia molecules and lots of ammonium ions from the salt,  $\text{NH}_4\text{Cl}$ . These ammonium ions are weakly acidic.



Write the  $K_a$  expression for the ammonium ion.

$$K_a = \frac{[\text{NH}_3][\text{H}_3\text{O}^+]}{[\text{NH}_4^+]}$$

We assume  $[\text{NH}_3]$  is the same concentration as the original ammonia solution and that  $[\text{NH}_4^+]$  is the same as the concentration of the ammonium chloride.

Put the values into the  $K_a$  expression and calculate  $[\text{H}_3\text{O}^+]$  and then the pH.

$$5.62 \times 10^{-10} = \frac{0.100 \times [\text{H}_3\text{O}^+]}{0.0500}$$

$$[\text{H}_3\text{O}^+] = (5.62 \times 10^{-10} \times 0.0500) \div 0.100 = 2.81 \times 10^{-10}$$

$$\text{pH} = -\log [\text{H}_3\text{O}^+] = -\log 2.81 \times 10^{-10} = 9.55$$

*If you are looking for a way to calculate buffer composition, you can reverse the equation. Using known pH to calculate  $\text{H}_3\text{O}^+$  and the known  $K_a$ , you can calculate the ratio of concentrations of the acid and conjugate base, necessary to prepare the buffer.*

Alternatively..... using the buffer equation that you have carefully memorised....

$$\text{pH} = \text{pK}_a + \log \frac{[\text{base}]}{[\text{acid}]}$$

$$\text{pH} = -\log 5.62 \times 10^{-10} + \log (0.100 / 0.0500)$$

$$\text{pH} = 9.25 + \log (0.100 / 0.0500)$$

$$\text{pH} = 9.25 + \log (2)$$

$$\text{pH} = 9.55$$

*Again, if you are looking for a way to calculate buffer composition, you can reverse the equation. Using known **pH** and known **pK<sub>a</sub>** you can calculate the ratio of concentrations of the acid and conjugate base, necessary to prepare the buffer.*