

TYPE 2 (base buffer)

(both a weak BASE and the SALT of that weak base are put in solution)

For a weak base (B) in water:

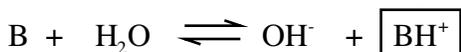
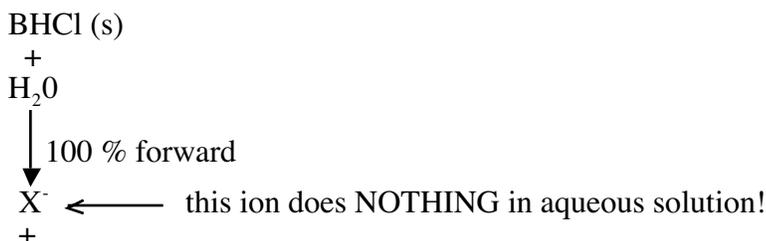


For a salt, BHX, of the weak base (B) in water:



NOTICE: There are now TWO sources for the ion BH^+ . When there is more than one source for an ion in solution, this is called the COMMON ION EFFECT. Please be VERY aware that the main source for the ion, BH^+ , is from the salt and not from the base. This is the key to understanding TYPE 2 problems.

Here, the two reactions are shown intersecting at the point of the common ion.



The following equilibrium condition still applies: $K_b = \frac{[OH^-][BH^+]}{[B]}$

now let C_B be the analytical concentration (label) of the weak base B, and C_{BH^+} be the label concentration of the salt of the weak base BHX.

Our Equilibrium "Chart"

	[B]	[OH ⁻]	[BH ⁺]
initial	C_B	0	C_{BH^+}
change	-x	+x	+x
final	$C_B - x$	x	$C_{BH^+} + x$

$$K_b = \frac{x(C_{BH^+} + x)}{(C_B - x)} \quad \text{Equation 2.1}$$

$$K_b C_B - K_b(x) = C_{BH^+}(x) + x^2$$

$$0 = x^2 + (K_b + C_{BH^+})(x) - K_b C_B$$

$$\boxed{[OH^-] = x = \frac{-(K_b + C_{BH^+}) + \sqrt{(K_b + C_{BH^+})^2 + 4K_b C_B}}{2}}$$

This equation is the exact solution for calculating the [OH] for any solution made by mixing both a concentration of weak base AND a concentration of the salt of the weak base in water.

NOTE: you should (almost) never ever have to use this equation. Why? continue...

TYPE 2 (base) continued

If K_b is small enough (less than 10^{-4}) and there are REASONABLE concentrations for B and BH^+ , the following assumption can and should be made:

$$C_B - x \cong C_B \quad \text{all this is saying is that } x \text{ is so small compared to } C_B, \text{ it doesn't change it}$$

$$C_{BH^+} + x \cong C_{BH^+} \quad \text{this is saying the same thing except compared to } C_{BH^+}, \text{ it doesn't change it}$$

So that Equation 2.1 shown above becomes

$$K_b = \frac{x(C_{BH^+})}{(C_B)}$$

and solving:

$$[OH^-] = x = K_b \left(\frac{C_B}{C_{BH^+}} \right) \quad \text{Equation 2.2}$$

This equation is the approximate solution for calculating the $[OH^-]$ for any solution made by mixing both a concentration of weak base (B) AND a concentration of the salt of that weak base (BHX) in water. As long as any REASONABLE concentration of B and BHX are used, this "approximation" is really almost EXACT!!

Notice that it is Equation 2.2 shown above that is the basis for the Henderson-Hasselbalch Equation. The $-\log$ function is taken on both sides of Equation 2.2 to give the Henderson-Hasselbalch Equation for acids:

$$pOH = pK_b + \log \left(\frac{C_{BH^+}}{C_B} \right) \quad \text{Henderson-Hasselbalch Equation for Bases}$$

Because the BH^+ usually comes from a salt, sometimes the H-H Equation is written in the following form:

$$pOH = pK_b + \log \left(\frac{[SALT]}{[BASE]} \right) \quad \text{Henderson-Hasselbalch Equation for Bases}$$