TYPE 2 (base buffer)
(both B and BH⁺ are put into solution)

For a weak base (B) in water:

B + H_2O $\rightleftharpoons$ OH^- + BH^+
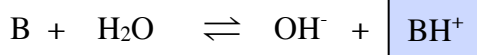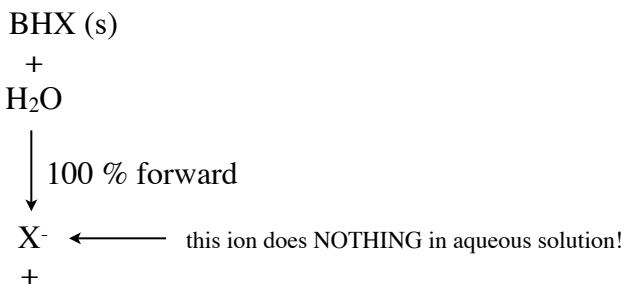
For a salt, BHX, of the weak base (B) in water:

BHX (s) + H_2O $\rightleftharpoons$ BH^+ + X^- (100% ionization!!!, and X^- is some anion of a strong acid - it does nothing)

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 <u>main</u> source for the ion, BH^+, is from the <u>salt</u> 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.

BHX (s)
+
H_2O

100 % forward

X^- ⟵ this ion does NOTHING in aqueous solution!
+

B + H_2O $\rightleftharpoons$ OH^- + $\boxed{BH^+}$

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

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 "ICE" Table

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

$$K_b = \frac{x\,(C_{BH+} + x)}{(C_B - x)} \qquad \textbf{Equation 2.1}$$

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

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

solve with quadratic formula...

TYPE 2 (base buffer) continued

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

This equation is the <u>exact</u> 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) **<u>never</u> ever have to use this equation. Why? continue...**

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:

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

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

So that Equation 2.1 shown above becomes

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

and solving:

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

This equation is the <u>approximate</u> 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:

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

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

$$\text{pOH} = \text{p}K_b + \log\left(\frac{[\text{Salt}]}{[\text{Base}]}\right) \qquad \textbf{Henderson-Hasselbalch Equation for Bases}$$