Examples: One strong and one weak

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Step 1. Calculate dilutions. First add the volumes

Total volume = 50 mL + 40 mL = 90 mL

Calculate concentrations in the solution

$$[HF] = [0.20]\left(\frac{40}{90}\right) = 0.0888\ M$$

$$[NaOH] = [0.25]\left(\frac{50}{90}\right) = 0.139\ M$$

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Step 2. Write a balanced chemical reaction and determine the form of the equilibrium constant.

$$HF + OH^{-} \leftrightarrow F^{-} + H_2O$$

Step 3. Although the equilibrium constant is given by:

$$\frac{1}{K_b} = \frac{[F^{-}]}{[HF][OH^{-}]}$$

we can assume that the OH^{-} reacts complete with HF since it is a strong base. In this case the limiting reagent is HF so We have excess $[OH^{-}]$ = 0.1390 – 0.0888 = 0.511 M

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Step 5. Calculate pOH.

$$pOH = -\log_{10}(0.0511) = 1.29$$

Therefore, pH = 14 – pOH = 12.71

Strong base exceeds weak acid

The key point of the previous problem is that we are no longer in the buffer range. We cannot use H-H in this case. Since:

$$[OH^-]_0 > [HA]_0$$

While K_b still applies it is often unnecessary since $[OH^-]$ is in excess.

If you need to use K_b then use:

$$A^- + H_2O \rightarrow HA + OH^-$$

$$K_b = \frac{[HA][OH^-]}{[A^-]}$$