

Examples: One strong and one weak

What is the pH when 50.0 mL of 0.25 M NaOH are added to 40.0 mL of 0.20 M HF?

Examples: One strong and one weak

What is the pH when 50.0 mL of 0.25 M NaOH are added to 40.0 mL of 0.20 M HF?

Step 1. Calculate dilutions. First add the volumes

$$\text{Total volume} = 50 \text{ mL} + 40 \text{ mL} = 90 \text{ mL}$$

Calculate concentrations in the solution

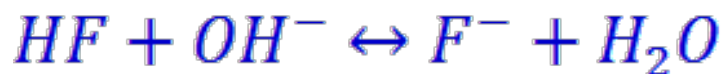
$$[HF] = [0.20] \left(\frac{40}{90} \right) = 0.0888 \text{ M}$$

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

Examples: One strong and one weak

What is the pH when 50.0 mL of 0.25 M NaOH are added to 40.0 mL of 0.20 M HF?

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



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 $[OH^-] = 0.1390 - 0.0888 = 0.511 \text{ M}$

Examples: One strong and one weak

What is the pH when 50.0 mL of 0.25 M NaOH are added to 40.0 mL of 0.20 M HF?

Step 5. Calculate pOH.

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

Therefore, $pH = 14 - pOH = 12.71$

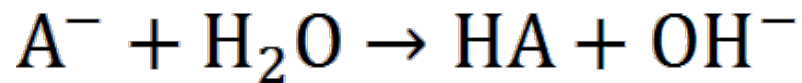
Strong base exceeds weak acid

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:

$$[\text{OH}^-]_0 > [\text{HA}]_0$$

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

If you need to use K_b then use:



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

