

- 1) It takes 83 mL of a 0.45 M NaOH solution to neutralize 235 mL of an HCl solution. What is the concentration of the HCl solution?
  
- 2) You are titrating an acid into a base to determine the concentration of the base. The endpoint of the neutralization is reached but the stopcock on the buret sticks slightly and allows a few more drops of acid to fall into the solution. How will this affect your calculations for the concentration of the base?
  
- 3) It takes 38 mL of 0.75 M NaOH solution to completely neutralize 155 mL of a sulfuric acid solution ( $\text{H}_2\text{SO}_4$ ). What is the concentration of the  $\text{H}_2\text{SO}_4$  solution?
  
- 4) A few small drops of water are left in a buret that is then used to titrate a base into an acid solution to determine the concentration of the acid. Will this small amount of water have any effect on the determined value for the concentration of the acid? If so, how is it affected?
  
- 5) It takes 12.5 mL of a 0.30 M HCl solution to neutralize 285 mL of NaOH solution. What is the concentration of the NaOH solution?
  
- 6) Lulu Labwrecker carefully pipets 25.0 mL of 0.525 M NaOH into a test tube. She places the test tube into a small beaker to keep it from spilling and then pipets 75.0 mL of 0.355 M HCl into another test tube. When Lulu reaches to put this test tube of acid into the beaker along with test tube of base she accidentally knocks the test tubes together hard enough to break them and their respective contents combine in the bottom of the beaker. Is the solution formed from the contents of the two test tubes acidic or basic? What is the pH of the resulting solution?



$$\frac{M_A V_A}{n_A} = \frac{M_B V_B}{n_B}$$

**Note:**  $n_A$  and  $n_B$  are the # of moles from the balanced equation.

$$M_A = \frac{M_B V_B n_A}{V_A n_B} = \frac{(0.45 \text{ M})(83 \text{ mL})(1)}{(235 \text{ mL})(1)} = 0.16 \text{ M HCl}$$

- 2) Those extra few drops of acid will cause the calculation for the concentration of the base to be too high. This is because it will seem that it took more acid to neutralize the base than it really did and so it will appear that the base is of stronger concentration than it really was.



$$\frac{M_A V_A}{n_A} = \frac{M_B V_B}{n_B}$$

$$M_A = \frac{M_B V_B n_A}{V_A n_B} = \frac{(0.75 \text{ M})(38 \text{ mL})(1)}{(155 \text{ mL})(2)} = 0.092 \text{ M H}_2\text{SO}_4$$

- 4) Yes, even this small amount of water will cause an error because the drops of water add to the volume of base, actually diluting it slightly. This means it will take a tad more base solution to neutralize the acid, making it seem as if the acidic solution was of stronger concentration than it actually was.



$$\frac{M_A V_A}{n_A} = \frac{M_B V_B}{n_B}$$

$$M_B = \frac{M_A V_A n_B}{V_B n_A} = \frac{(0.30 \text{ M})(12.5 \text{ mL})(1)}{(285 \text{ mL})(1)} = 0.013 \text{ M NaOH}$$



$$0.0750 \text{ L NaOH} \times \frac{0.355 \text{ mole NaOH}}{1 \text{ L NaOH}} = 0.026625 \text{ mole NaOH}$$

$$0.0250 \text{ L HCl} \times \frac{0.525 \text{ mole HCl}}{1 \text{ L HCl}} = 0.013125 \text{ mole HCl}$$

Because the ration of moles of NaOH to HCl is 1:1 we can subtract the number of moles of each to find the unreacted part and since this is the number of moles not neutralized, we can get the molarity of the final solution in the usual way. (0.0250 L acid + 0.0750 L base = 0.100 L soln)

$$0.026625 \text{ mole base} - 0.013125 \text{ mole acid} = 0.013500 \text{ moles NaOH excess}$$

$$\frac{0.0135 \text{ mol NaOH}}{(0.100 \text{ L soln})} = 0.135 \text{ M NaOH}$$

$$\text{pOH} = -\log(0.135) = 0.870$$

$$\text{pH} = 14.000 - \text{pOH} = 14.000 - 0.870 = 13.130$$