

Solutions to the Titrations Practice Worksheet

For questions 1 and 2, the units for your final answer should be "M", or "molar", because you're trying to find the molarity of the acid or base solution. To solve these problems, use $M_1V_1 = M_2V_2$.

- 1) 0.043 M HCl
- 2) 0.0036 M NaOH

For problem 3, you need to divide your final answer by two, because H_2SO_4 is a diprotic acid, meaning that there are two acidic hydrogens that need to be neutralized during the titration. As a result, it takes twice as much base to neutralize it, making the concentration of the acid appear twice as large as it really is.

- 3) 0.1 M H_2SO_4

4) You cannot do a titration without knowing the molarity of at least one of the substances, because you'd then be solving one equation with two unknowns (the unknowns being M_1 and M_2).

5) Endpoint: When you actually stop doing the titration (usually, this is determined by a color change in an indicator or an indication of $pH=7.0$ on an electronic pH probe)

Equivalence point: When the solution is exactly neutralized. It's important to keep in mind that the equivalence point and the endpoint are not exactly the same because indicators don't change color at exactly 7.0000 pH and pH probes aren't infinitely accurate. Generally, you can measure the effectiveness of a titration by the closeness of the endpoint to the equivalence point.



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

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)

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

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

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

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