

16. One mole of NaOH dissolved in 1.00 L of solution will produce 1.00 *M* NaOH. First, weigh out 40.00 g of NaOH (1.000 mol). Next, add some water to a 1-L volumetric flask (an instrument that is precise to 1.000 L). Dissolve the NaOH in the flask, add some more water, mix, add more water, mix, etc. until water has been added to 1.000-L mark of the volumetric flask. The result is 1.000 L of a 1.000 *M* NaOH solution. Because we know the volume to four significant figures as well as the mass, the molarity will be known to four significant figures. This is good practice, if you need a three-significant-figure molarity, your measurements should be taken to four significant figures.

When you need to dilute a more concentrated solution with water to prepare a solution, again make all measurements to four significant figures to ensure three significant figures in the molarity. Here, we need to cut the molarity in half from 2.00 *M* to 1.00 *M*. We would start with 1 mole of NaOH from the concentrated solution. This would be 500.0 mL of 2.00 *M* NaOH. Add this to a 1-L volumetric flask with addition of more water and mixing until the 1.000-L mark is reached. The resulting solution would be 1.00 *M*.

20.

The acid is a diprotic acid (H_2A), meaning that it has two H^+ ions in the formula to donate to a base. The reaction is $\text{H}_2\text{A}(\text{aq}) + 2 \text{NaOH}(\text{aq}) \rightarrow 2 \text{H}_2\text{O}(\text{l}) + \text{Na}_2\text{A}(\text{aq})$, where A^{2-} is what is left over from the acid formula when the two protons (H^+ ions) are reacted.

$$(36) \quad a. \quad 1.00 \text{ L solution} \times \frac{0.50 \text{ mol H}_2\text{SO}_4}{\text{L}} = 0.50 \text{ mol H}_2\text{SO}_4$$

$$0.50 \text{ mol H}_2\text{SO}_4 \times \frac{1 \text{ L}}{18 \text{ mol H}_2\text{SO}_4} = 2.8 \times 10^{-2} \text{ L conc. H}_2\text{SO}_4 \text{ or } 28 \text{ mL}$$

Dilute 28 mL of concentrated H_2SO_4 to a total volume of 1.00 L with water. The resulting 1.00 L of solution will be a 0.50 M H_2SO_4 solution.

b. We will need 0.50 mol HCl.

$$0.50 \text{ mol HCl} \times \frac{1 \text{ L}}{12 \text{ mol HCl}} = 4.2 \times 10^{-2} \text{ L} = 42 \text{ mL}$$

Dilute 42 mL of concentrated HCl to a final volume of 1.00 L.

c. We need 0.50 mol NiCl_2 .

$$0.50 \text{ mol NiCl}_2 \times \frac{1 \text{ mol NiCl}_2 \cdot 6\text{H}_2\text{O}}{\text{mol NiCl}_2} \times \frac{237.69 \text{ g NiCl}_2 \cdot 6\text{H}_2\text{O}}{\text{mol NiCl}_2 \cdot 6\text{H}_2\text{O}} = 118.8 \text{ g NiCl}_2 \cdot 6\text{H}_2\text{O} \approx 120 \text{ g}$$

Dissolve 120 g $\text{NiCl}_2 \cdot 6\text{H}_2\text{O}$ in water, and add water until the total volume of the solution is 1.00 L.

$$d. \quad 1.00 \text{ L} \times \frac{0.50 \text{ mol HNO}_3}{\text{L}} = 0.50 \text{ mol HNO}_3$$

$$0.50 \text{ mol HNO}_3 \times \frac{1 \text{ L}}{16 \text{ mol HNO}_3} = 0.031 \text{ L} = 31 \text{ mL}$$

Dissolve 31 mL of concentrated reagent in water. Dilute to a total volume of 1.00 L.

e. We need 0.50 mol Na_2CO_3 .

$$0.50 \text{ mol Na}_2\text{CO}_3 \times \frac{105.99 \text{ g Na}_2\text{CO}_3}{\text{mol}} = 53 \text{ g Na}_2\text{CO}_3$$

Dissolve 53 g Na_2CO_3 in water, dilute to 1.00 L.

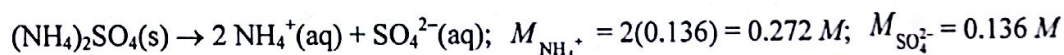
$$(37) \quad 10.8 \text{ g (NH}_4\text{)}_2\text{SO}_4 \times \frac{1 \text{ mol}}{132.15 \text{ g}} = 8.17 \times 10^{-2} \text{ mol (NH}_4\text{)}_2\text{SO}_4$$

$$\text{Molarity} = \frac{8.17 \times 10^{-2} \text{ mol}}{100.0 \text{ mL}} \times \frac{1000 \text{ mL}}{\text{L}} = 0.817 \text{ M (NH}_4\text{)}_2\text{SO}_4$$

Moles of $(\text{NH}_4)_2\text{SO}_4$ in final solution:

$$10.00 \times 10^{-3} \text{ L} \times \frac{0.817 \text{ mol}}{\text{L}} = 8.17 \times 10^{-3} \text{ mol}$$

$$\text{Molarity of final solution} = \frac{8.17 \times 10^{-3} \text{ mol}}{(10.00 + 50.00) \text{ mL}} \times \frac{1000 \text{ mL}}{\text{L}} = 0.136 \text{ M } (\text{NH}_4)_2\text{SO}_4$$



38. Molarity = $\frac{\text{total mol HNO}_3}{\text{total volume}}$; total volume = 0.05000 L + 0.10000 L = 0.15000 L

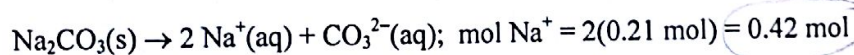
$$\text{Total mol HNO}_3 = 0.05000 \text{ L} \times \frac{0.100 \text{ mol HNO}_3}{\text{L}} + 0.10000 \text{ L} \times \frac{0.200 \text{ mol HNO}_3}{\text{L}}$$

$$\text{Total mol HNO}_3 = 5.00 \times 10^{-3} \text{ mol} + 2.00 \times 10^{-2} \text{ mol} = 2.50 \times 10^{-2} \text{ mol HNO}_3$$

$$\text{Molarity} = \frac{2.50 \times 10^{-2} \text{ mol HNO}_3}{0.15000 \text{ L}} = 0.167 \text{ M HNO}_3$$

As expected, the molarity of HNO_3 is between 0.100 M and 0.200 M.

39. Mol $\text{Na}_2\text{CO}_3 = 0.0700 \text{ L} \times \frac{3.0 \text{ mol Na}_2\text{CO}_3}{\text{L}} = 0.21 \text{ mol Na}_2\text{CO}_3$



$$\text{Mol NaHCO}_3 = 0.0300 \text{ L} \times \frac{1.0 \text{ mol NaHCO}_3}{\text{L}} = 0.030 \text{ mol NaHCO}_3$$



$$M_{\text{Na}^+} = \frac{\text{total mol Na}^+}{\text{total volume}} = \frac{0.42 \text{ mol} + 0.030 \text{ mol}}{0.0700 \text{ L} + 0.0300 \text{ L}} = \frac{0.45 \text{ mol}}{0.1000 \text{ L}} = 4.5 \text{ M Na}^+$$

40. Mol $\text{CoCl}_2 = 0.0500 \text{ L} \times \frac{0.250 \text{ mol CoCl}_2}{\text{L}} = 0.0125 \text{ mol}$

$$\text{Mol NiCl}_2 = 0.0250 \text{ L} \times \frac{0.350 \text{ mol NiCl}_2}{\text{L}} = 0.00875 \text{ mol}$$

Both CoCl_2 and NiCl_2 are soluble chloride salts by the solubility rules. A 0.0125-mol aqueous sample of CoCl_2 is actually 0.0125 mol Co^{2+} and $2(0.0125 \text{ mol}) = 0.0250 \text{ mol Cl}^-$. A 0.00875-mol aqueous sample of NiCl_2 is actually 0.00875 mol Ni^{2+} and $2(0.00875) = 0.0175 \text{ mol Cl}^-$. The total volume of solution that these ions are in is $0.0500 \text{ L} + 0.0250 \text{ L} = 0.0750 \text{ L}$.

$$M_{\text{Co}^{2+}} = \frac{0.0125 \text{ mol Co}^{2+}}{0.0750 \text{ L}} = 0.167 \text{ M}; M_{\text{Ni}^{2+}} = \frac{0.00875 \text{ mol Ni}^{2+}}{0.0750 \text{ L}} = 0.117 \text{ M}$$

$$M_{\text{Cl}^-} = \frac{0.0250 \text{ mol Cl}^- + 0.0175 \text{ mol Cl}^-}{0.0750 \text{ L}} = 0.567 \text{ M}$$

$$(41) \text{ Stock solution} = \frac{10.0 \text{ mg}}{500.0 \text{ mL}} = \frac{10.0 \times 10^{-3} \text{ g}}{500.0 \text{ mL}} = \frac{2.00 \times 10^{-5} \text{ g steroid}}{\text{mL}}$$

$$100.0 \times 10^{-6} \text{ L stock} \times \frac{1000 \text{ mL}}{\text{L}} \times \frac{2.00 \times 10^{-5} \text{ g steroid}}{\text{mL}} = 2.00 \times 10^{-6} \text{ g steroid}$$

This is diluted to a final volume of 100.0 mL.

$$\frac{2.00 \times 10^{-6} \text{ g steroid}}{100.0 \text{ mL}} \times \frac{1000 \text{ mL}}{\text{L}} \times \frac{1 \text{ mol steroid}}{336.43 \text{ g steroid}} = 5.94 \times 10^{-8} \text{ M steroid}$$

$$(42) \text{ Stock solution:}$$

$$1.584 \text{ g Mn}^{2+} \times \frac{1 \text{ mol Mn}^{2+}}{54.94 \text{ g Mn}^{2+}} = 2.883 \times 10^{-2} \text{ mol Mn}^{2+}$$

$$\text{Molarity} = \frac{2.833 \times 10^{-2} \text{ mol Mn}^{2+}}{1.000 \text{ L}} = 2.883 \times 10^{-2} \text{ M}$$

Solution A:

$$50.00 \text{ mL} \times \frac{1 \text{ L}}{1000 \text{ mL}} \times \frac{2.833 \times 10^{-2} \text{ mol}}{\text{L}} = 1.442 \times 10^{-3} \text{ mol Mn}^{2+}$$

$$\text{Molarity} = \frac{1.442 \times 10^{-3} \text{ mol}}{1000 \text{ mL}} \times \frac{1000 \text{ mL}}{1 \text{ L}} = 1.442 \times 10^{-3} \text{ M}$$

Solution B:

$$10.0 \text{ mL} \times \frac{1 \text{ L}}{1000 \text{ mL}} \times \frac{1.442 \times 10^{-3} \text{ mol}}{\text{L}} = 1.442 \times 10^{-5} \text{ mol Mn}^{2+}$$

$$\text{Molarity} = \frac{1.442 \times 10^{-5} \text{ mol}}{0.2500 \text{ L}} = 5.768 \times 10^{-5} \text{ M}$$

Solution C:

$$10.00 \times 10^{-3} \text{ L} \times \frac{5.768 \times 10^{-5} \text{ mol}}{\text{L}} = 5.768 \times 10^{-7} \text{ mol Mn}^{2+}$$

$$\text{Molarity} = \frac{5.768 \times 10^{-7} \text{ mol}}{0.5000 \text{ L}} = 1.154 \times 10^{-6} \text{ M}$$