

## *Chem 151 Drilling Question Set #2*

### **Mole and Molarity**

1. a. What is the molar mass of copper(II) nitrate?

$$\text{molar mass of Cu(NO}_3)_2 = 63.5 + 2(14.0 + 3 \times 16.0) = 187.5 \text{ g/mole}$$

- b. What is the mass, in grams, of 0.120 moles of copper(II) nitrate?

$$\text{Mass} = (0.120 \text{ mole})(187.5 \text{ g/mol}) = 22.5 \text{ mole}$$

- c. How many nitrogen atoms are present in 1.25 mg of copper(II) nitrate? (1 mg =  $1 \times 10^{-3}$  g)

$$(1.25 \times 10^{-3} \text{ g}) \left( \frac{1 \text{ mol Cu(NO}_3)_2}{187.5 \text{ g}} \right) \left( \frac{2 \text{ mol N atoms}}{1 \text{ mol Cu(NO}_3)_2} \right) = 1.33 \times 10^{-5} \text{ moles N atoms}$$

*Note that "mole" is a unit of number.*

2. Calculate the mass in grams of each of the following:

- a. 0.0287 mole of aspirin,  $\text{C}_9\text{H}_8\text{O}_4$

$$\text{Molar mass of aspirin} = 180 \text{ g/mol}$$

$$\text{Mass of aspirin} = (0.287 \text{ mol})(180 \text{ g/mol}) = 5.17 \text{ g}$$

- b.  $1.75 \times 10^{21}$  molecules of ozone

$$\text{molar mass of O}_3 = 3 \times 16.0 \text{ g/mol} = 48.0 \text{ g}$$

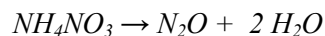
$$\text{Mass} = (1.75 \times 10^{21} \text{ molecules}) \left( \frac{1 \text{ mole}}{6.022 \times 10^{23} \text{ molecules}} \right) \left( \frac{48.0 \text{ g}}{1 \text{ mol}} \right) = 0.139 \text{ g}$$

- c. two molecules of cholesterol,  $\text{C}_{27}\text{H}_{46}\text{O}$

$$\text{molar mass of cholesterol} = 386 \text{ g/mol}$$

$$\text{Mass} = (2 \text{ molecules}) \left( \frac{1 \text{ mole}}{6.022 \times 10^{23} \text{ molecules}} \right) \left( \frac{386 \text{ g}}{1 \text{ mol}} \right) = 1.28 \times 10^{-21} \text{ g}$$

3. Laughing gas,  $N_2O$ , can be made by the thermal decomposition of ammonium nitrate according to the below equation. How many grams of laughing gas can be obtained from 18.5 g of ammonium nitrate?



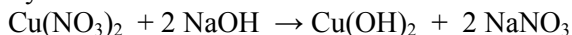
$$(18.5 \text{ g } NH_4NO_3) \left( \frac{1 \text{ mol } NH_4NO_3}{80.0 \text{ g}} \right) \left( \frac{1 \text{ mol } N_2O}{1 \text{ mol } NH_4NO_3} \right) \left( \frac{44.0 \text{ g}}{1 \text{ mol } N_2O} \right) = 10.2 \text{ g}$$

*Do not just focus on the cancellation of units in your calculation. Instead, you should focus on the method. First, the amount of reactant is converted to mole. Then, the mole of reactant is related to the mole of product using the reaction equation. Last the mole of product is converted to the mass of product.*

4. A copper wire weighed 3.25 g is dissolved in nitric acid according to the reaction equation:



Sodium hydroxide is added to the above solution containing copper(II) nitrate and  $\text{Cu}(\text{OH})_2$  is formed.



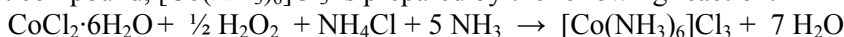
The solution is then heated and  $\text{Cu}(\text{OH})_2$  decomposes to  $\text{CuO}$ .  $\text{Cu}(\text{OH})_2 \rightarrow \text{CuO} + \text{H}_2\text{O}$

How many grams of  $\text{CuO}$  can be produced by this process?

*Since the above reactions has no copper by-products, all copper metal is converted to  $\text{CuO}$ . Since one Cu makes one  $\text{CuO}$ , the mass of  $\text{CuO}$  produced by this process is:*

$$\text{Mass of CuO} = (3.25 \text{ g Cu}) \left( \frac{1 \text{ mol Cu}}{63.55 \text{ g}} \right) \left( \frac{1 \text{ mol CuO}}{1 \text{ mol Cu}} \right) \left( \frac{79.55 \text{ g}}{1 \text{ mol CuO}} \right) = 4.07 \text{ g CuO}$$

5. A cobalt compound,  $[\text{Co}(\text{NH}_3)_6]\text{Cl}_3$  is prepared by the following reaction:



How many grams of  $[\text{Co}(\text{NH}_3)_6]\text{Cl}_3$  can be produced from 0.123 mole of  $\text{NH}_3$  if the quantity of the other reactants are more than sufficient and the reaction goes to completeness?

*Molar mass of  $[\text{Co}(\text{NH}_3)_6]\text{Cl}_3 = 267.5 \text{ g/mol}$*

*First, find out how many moles of  $[\text{Co}(\text{NH}_3)_6]\text{Cl}_3$  can be produced from 0.123 mole  $\text{NH}_3$ .*

$$\text{Moles of } [\text{Co}(\text{NH}_3)_6]\text{Cl}_3 = (0.123 \text{ mol NH}_3) \left( \frac{1 \text{ mol } [\text{Co}(\text{NH}_3)_6]\text{Cl}_3}{5 \text{ mol NH}_3} \right) = 0.0246 \text{ mol}$$

$$\text{Mass of } [\text{Co}(\text{NH}_3)_6]\text{Cl}_3 = (0.0246 \text{ mol})(267.5 \text{ g/mol}) = 6.58 \text{ g}$$

6. Glucose can be converted into ethanol and carbon dioxide by fermentation. If the reaction goes to completion, how many liters of ethanol can be produced from 500. g of glucose? Density of ethanol is 0.789 g/mL.



*Molar mass of glucose = 180 g/mol Molar mass of ethanol = 46.0 g/mol*

*Mass of ethanol can be produced*

$$= (500 \text{ g glucose}) \left( \frac{1 \text{ mol glucose}}{180 \text{ g}} \right) \left( \frac{2 \text{ mol ethanol}}{1 \text{ mol glucose}} \right) \left( \frac{46.0 \text{ g}}{1 \text{ mol ethanol}} \right) = 256 \text{ g ethanol}$$

$$\text{Volume of ethanol} = (256 \text{ g}) \left( \frac{1 \times 10^{-3} \text{ L}}{0.789 \text{ g}} \right) = 0.324 \text{ L}$$

7. The molecular formula of the iron core in ferritin is  $[\text{FeO}(\text{OH})]_8\text{FeO}(\text{H}_2\text{PO}_4)$ . What is the molarity of  $\text{Fe}^{3+}$  ions if 1.00 mg of the iron core dissolves in acid to make 500. mL of solution?

*Molar mass of  $[\text{FeO}(\text{OH})]_8\text{FeO}(\text{H}_2\text{PO}_4) = 879.7 \text{ g/mol}$*

$$\text{Moles of } [\text{FeO}(\text{OH})]_8\text{FeO}(\text{H}_2\text{PO}_4) = (1.00 \text{ mg}) \left( \frac{1 \text{ g}}{1000 \text{ mg}} \right) \left( \frac{1 \text{ mol}}{879.7 \text{ g}} \right) = 1.14 \times 10^{-6} \text{ mol}$$

$$\text{Moles of } \text{Fe}^{3+} = (1.14 \times 10^{-6} \text{ mol}) \left( \frac{9 \text{ mol } \text{Fe}^{3+}}{1 \text{ mol } [\text{FeO}(\text{OH})]_8\text{FeO}(\text{H}_2\text{PO}_4)} \right) = 1.02 \times 10^{-5} \text{ mol } \text{Fe}^{3+}$$

$$\text{Molarity of } \text{Fe}^{3+} = \frac{1.02 \times 10^{-5} \text{ mol}}{0.500 \text{ L}} = 2.05 \times 10^{-5} \text{ mol/L} = 2.05 \times 10^{-5} \text{ M}$$

8. A solution containing  $\text{NaCl}$ ,  $\text{MgCl}_2$ , and  $\text{AlCl}_3$ . Given that the molarity of chloride ions in the solution is 1.00 M, the molarity of  $\text{Na}^+$  and  $\text{Mg}^{2+}$  is 0.200 M and 0.300 M respectively. What is the molarity of  $\text{Al}^{3+}$  in the solution?

*Let the total volume of solution be 1.00 L.*

*Since the total charges of the cations and the charges of the anions must be balanced, we have the following relationship:*

$$\text{Moles of } \text{Cl}^- \text{ ions} = 1(\text{moles of } \text{Na}^+) + 2(\text{moles of } \text{Mg}^{2+}) + 3(\text{moles of } \text{Al}^{3+})$$

*Thus, we have:  $1.00 \text{ mol} = 0.200 \text{ mol} + 0.600 \text{ mol} + 3(\text{moles of } \text{Al}^{3+})$*

$$\text{Moles of } \text{Al}^{3+} = \frac{1}{3}(1.00 - 0.200 - 0.600) = 0.067 \text{ mol}$$

*Since the volume is assumed as 1.00 L, Molarity of  $\text{Al}^{3+} = 0.067 \text{ mol/L} = 0.067 \text{ M}$*

9. 40.00 mL of 0.300 M sodium sulfate,  $\text{Na}_2\text{SO}_4$ , solution is added to 70.00 mL of 0.200 M aluminum sulfate,  $\text{Al}_2(\text{SO}_4)_3$ , solution. (No reaction occurs.)

- a) Calculate the molarity of the sulfate ions in the resultant solution.

$$\text{Moles of } \text{SO}_4^{2-} = (0.300 \text{ mol/L})(0.0400 \text{ L}) + 3(0.200 \text{ mol/L})(0.0700 \text{ L}) = 0.0540 \text{ mol}$$

$$\text{Volume of resultant solution} = 40.00 \text{ mL} + 70.00 \text{ mL} = 110.00 \text{ mL} = 0.11000 \text{ L}$$

$$\text{Molarity of } \text{SO}_4^{2-} \text{ in solution} = \frac{0.0540 \text{ mol}}{0.11000 \text{ L}} = 0.491 \text{ M}$$

- b) Calculate the number of moles of barium ions needed in order to precipitate all the sulfate ions from the above solution.

*The formula of barium sulfate is  $\text{BaSO}_4$ . Thus, one  $\text{Ba}^{2+}$  needs one  $\text{SO}_4^{2-}$ .*

$$\text{Moles of } \text{Ba}^{2+} \text{ needed} = \text{moles of } \text{SO}_4^{2-} \text{ in solution} = 0.0540 \text{ mol}$$

- c) If the source of barium ions is a 0.150M barium chloride solution, calculate the volume of the barium chloride solution needed in order to precipitate all the sulfate ions from the solution.

*Volume of  $\text{BaCl}_2$  solution needed*

$$= (0.0540 \text{ mol Ba}^{2+} \text{ needed}) \left( \frac{1 \text{ mol BaCl}_2}{1 \text{ mol Ba}^{2+}} \right) \left( \frac{1 \text{ L solution}}{0.150 \text{ mol BaCl}_2} \right) = 0.360 \text{ L}$$

10. 255 mL of 0.100 M nitric acid is added to 745 mL of 0.200 M nitric acid. What is the molarity of nitric acid in the resultant solution?

$$\text{Total moles of HNO}_3 = (0.100 \text{ mol/L})(0.255 \text{ L}) + (0.200 \text{ mol/L})(0.745 \text{ L}) = 0.1745 \text{ mol}$$

$$\text{Total volume of solution} = 255 \text{ mL} + 745 \text{ mL} = 1000. \text{ mL} = 1.000 \text{ L}$$

$$\text{Molarity of HNO}_3 = \frac{0.1745 \text{ mole}}{1.000 \text{ L}} = 0.1745 \text{ M}$$

11. The concentrated hydrochloric acid in the stockroom is 36.0% HCl by mass. Given that the density of this acid is 1.18 g/mL.

- a) How many gram of HCl in 800. mL of this acid?

$$\text{Grams of HCl} = (800. \text{ mL solution})\left(\frac{1.18 \text{ g}}{1 \text{ mL}}\right)(36.0\%) = 340. \text{ g}$$

- b) What is the molarity of this acid?

$$\text{Moles of HCl in acid} = (340. \text{ g})\left(\frac{1 \text{ mol}}{36.46 \text{ g}}\right) = 9.33 \text{ mol}$$

$$\text{Molarity of HCl in solution} = \frac{9.33 \text{ mol}}{0.8000 \text{ L}} = 11.6 \text{ M}$$

12. If 70.0 mL of the hydrochloric acid in Question 10 is diluted to 100. mL, what is the molarity of HCl in the diluted solution?

*When a solution is diluted by adding water, the number of moles of solute in the solution remains unchanged. Therefore we have the equation:*

$$M_{\text{dilute}} V_{\text{dilute}} = M_{\text{concentrated}} V_{\text{concentrated}}$$

*Applying the equation to the problem,  $M_{\text{conc}} = 11.6 \text{ M}$*

$$V_{\text{conc}} = 70.0 \text{ mL}$$

$$M_{\text{dil}} = ?$$

$$V_{\text{dil}} = 100. \text{ mL}$$

$$M_{\text{dil}} = \frac{(11.6 \text{ M})(70.0 \text{ mL})}{100. \text{ mL}} = 8.16 \text{ M}$$