

Molar Mass Calculations

A mole is a standard unit of measurement for amount of a substance. For an element, one mole is equivalent to the mass of that element as it is listed on the periodic table, and thus represents its molar mass. For a compound, one must take the sum of all elements in the compound with respect to the number of atoms in order to obtain the molar mass.

Example: CaCl_2

$$1 \text{ mol} = \text{mass of Ca} + 2 \cdot (\text{mass of Cl})$$

$$1 \text{ mol} = 40.078\text{g} + 2 \cdot (35.453\text{g})$$

$$1 \text{ mol of } \text{CaCl}_2 = 110.984 \text{ grams of } \text{CaCl}_2$$

Therefore, we say that CaCl_2 has a molar mass of 110.984 g/mol.

Practice

1. nickel (II) sulfate NiSO_4

$$154.76 \text{ g/mol}$$

2. magnesium nitride Mg_3N_2

$$100.93 \text{ g/mol}$$

3. phosphoric acid H_3PO_4

$$98.00 \text{ g/mol}$$

4. zinc acetate $\text{Zn}(\text{C}_2\text{H}_3\text{O}_2)_2$

$$183.51 \text{ g/mol}$$

5. dinitrogen pentoxide N_2O_5

$$108.02 \text{ g/mol}$$

6. copper (I) nitrite CuNO_2

$$109.56 \text{ g/mol}$$

7. strontium hydroxide Sr(OH)_2

$$121.64 \text{ g/mol}$$

8. potassium phosphate K_3PO_4

$$212.27 \text{ g/mol}$$

9. hydrochloric acid HCl

$$36.46 \text{ g/mol}$$

Mole Conversions Worksheet

There are three mole equalities:

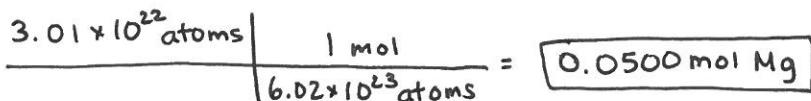
$$1 \text{ mol} = 6.02 \times 10^{23} \text{ particles}$$

$$1 \text{ mol} = \text{molar mass}$$

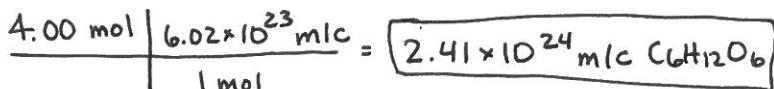
$$1 \text{ mol} = 22.4 \text{ L (for a gas at STP)}$$

Mole-Particle Conversions

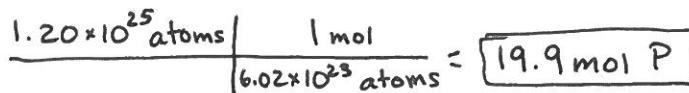
1. How many moles of magnesium is equivalent to 3.01×10^{22} atoms of magnesium?



2. How many molecules are there in 4.00 moles of glucose, $\text{C}_6\text{H}_{12}\text{O}_6$?



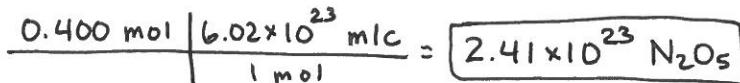
3. How many moles are in 1.20×10^{25} atoms of phosphorus?



4. How many atoms are in 0.750 moles of zinc?

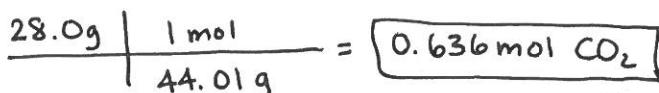


5. How many molecules are in 0.400 moles of dinitrogen pentoxide?

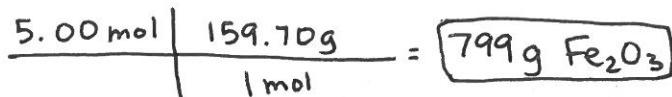


Mole-Mass Conversions

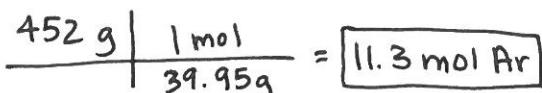
1. How many moles are in 28.0 grams of carbon dioxide?



2. What is the mass of 5.00 moles of iron (III) oxide?



3. Find the number of moles of argon that is equivalent to 452 grams.



- grams
4. Find the number of in 1.26×10^{-4} moles of acetic acid. $\text{HC}_2\text{H}_3\text{O}_2$

$$\frac{1.26 \times 10^{-4} \text{ mol}}{1 \text{ mol}} \left| \begin{array}{c} 60.06 \text{ g} \\ \hline 1 \text{ mol} \end{array} \right| = \boxed{7.57 \times 10^{-3} \text{ g } \text{HC}_2\text{H}_3\text{O}_2}$$

5. Find the mass equivalent to 2.6 moles of lithium bromide.

$$\frac{2.6 \text{ mol}}{1 \text{ mol}} \left| \begin{array}{c} 86.84 \text{ g} \\ \hline 1 \text{ mol} \end{array} \right| = \boxed{230 \text{ g LiBr}}$$

Mole-Volume Conversions

1. Determine the volume, in liters, occupied by 0.030 moles of a gas at STP.

$$\frac{0.030 \text{ mol}}{1 \text{ mol}} \left| \begin{array}{c} 22.4 \text{ L} \\ \hline 1 \text{ mol} \end{array} \right| = \boxed{0.67 \text{ L of gas}}$$

2. How many moles of gaseous argon atoms are present in 11.2 L at STP?

$$\frac{11.2 \text{ L}}{22.4 \text{ L}} \left| \begin{array}{c} 1 \text{ mol} \\ \hline \end{array} \right| = \boxed{0.500 \text{ mol Ar}}$$

3. What is the volume of 0.0520 moles of neon gas at STP?

$$\frac{0.0520 \text{ mol}}{1 \text{ mol}} \left| \begin{array}{c} 22.4 \text{ L} \\ \hline 1 \text{ mol} \end{array} \right| = \boxed{1.16 \text{ L Ne}}$$

4. What is the volume of 1.2 moles of water vapor at STP?

$$\frac{1.2 \text{ mol}}{1 \text{ mol}} \left| \begin{array}{c} 22.4 \text{ L} \\ \hline 1 \text{ mol} \end{array} \right| = \boxed{27 \text{ L H}_2\text{O}}$$

5. How many moles of gaseous sulfur trioxide are present in 2.50 L at STP?

$$\frac{2.50 \text{ L}}{22.4 \text{ L}} \left| \begin{array}{c} 1 \text{ mol} \\ \hline \end{array} \right| = \boxed{0.112 \text{ mol SO}_3}$$

Mixed Conversions

1. How many gaseous oxygen molecules are in 3.36 L at STP?

$$\frac{3.36 \text{ L}}{22.4 \text{ L}} \left| \begin{array}{c} 1 \text{ mol} \\ \hline \end{array} \right| \left| \begin{array}{c} 6.02 \times 10^{23} \text{ m/c} \\ \hline 1 \text{ mol} \end{array} \right| = \boxed{9.03 \times 10^{22} \text{ m/c O}_2}$$

2. Find the mass, in grams, of 2.00×10^{23} molecules of fluorine.

$$\frac{2.00 \times 10^{23} \text{ m/c}}{6.02 \times 10^{23} \text{ m/c}} \left| \begin{array}{c} 1 \text{ mol} \\ \hline \end{array} \right| \left| \begin{array}{c} 38.00 \text{ g} \\ \hline 1 \text{ mol} \end{array} \right| = \boxed{12.6 \text{ g F}_2}$$

3. Determine the volume, in liters, occupied by 14.0 grams of nitrogen gas at STP.

$$\frac{14.0 \text{ g}}{28.02 \text{ g}} \left| \begin{array}{c} 1 \text{ mol} \\ \hline 1 \text{ mol} \end{array} \right| \left| \begin{array}{c} 22.4 \text{ L} \\ \hline 1 \text{ mol} \end{array} \right| = \boxed{11.2 \text{ L N}_2}$$

Worksheet 11.2

4. Find the mass, in grams, of 1.00×10^{23} molecules of sodium iodide.

$$\frac{1.00 \times 10^{23} \text{ m/c}}{6.02 \times 10^{23} \text{ m/c}} \left| \begin{array}{c} 1 \text{ mol} \\ 1 \text{ mol} \end{array} \right| \frac{149.89 \text{ g}}{1 \text{ mol}} = 24.9 \text{ g NaI}$$

5. How many particles are present in a 1.43 gram sample of a molecular compound with a molar mass of 233 grams per mole?

$$\frac{1.43 \text{ g}}{233 \text{ g}} \left| \begin{array}{c} 1 \text{ mol} \\ 1 \text{ mol} \end{array} \right| \frac{6.02 \times 10^{23} \text{ particles}}{1 \text{ mol}} = 3.69 \times 10^{21} \text{ particles}$$

6. Aspartame is an artificial sweetener that is 160 times sweeter than sucrose (regular sugar) when dissolved in water. It is marketed by G.D. Searle as NutraSweet. The molecular formula of aspartame is C₁₄H₁₈N₂O₅.

- a. Calculate the molar mass of aspartame.

$$294.34 \text{ g/mol}$$

- b. How many moles of aspartame is equivalent to 10.0 grams?

$$\frac{10.0 \text{ g}}{294.34 \text{ g}} \left| \begin{array}{c} 1 \text{ mol} \\ 1 \text{ mol} \end{array} \right| = 0.0340 \text{ mol aspartame}$$

- c. What is the mass, in grams, of 1.56 moles of aspartame?

$$1.56 \text{ mol} \left| \begin{array}{c} 294.34 \text{ g} \\ 1 \text{ mol} \end{array} \right| = 459 \text{ g aspartame}$$

- d. How many molecules are present in 5.00 milligrams of aspartame?

$$\frac{5.00 \text{ mg}}{1000 \text{ mg}} \left| \begin{array}{c} 1 \text{ g} \\ 294.34 \text{ g} \end{array} \right| \left| \begin{array}{c} 1 \text{ mol} \\ 1 \text{ mol} \end{array} \right| \frac{6.02 \times 10^{23} \text{ m/c}}{1 \text{ mol}} = 1.02 \times 10^{19} \text{ m/c aspartame}$$

- e. How many atoms of nitrogen are in 1.20 grams of aspartame?

$$1.20 \text{ g} \left| \begin{array}{c} 1 \text{ mol} \\ 294.34 \text{ g} \end{array} \right| \left| \begin{array}{c} 6.02 \times 10^{23} \text{ m/c} \\ 1 \text{ mol} \end{array} \right| \frac{2 \text{ atoms N}}{1 \text{ m/c}} = 4.91 \times 10^{21} \text{ atoms N}$$

Percent Composition

Percent composition is used to determine how much each element contributes to the overall mass of the compound. This information is useful for determining formulas of unknown substances later on. Perform the following calculations, reporting the percent mass of each element.



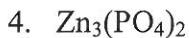
$$\text{Na: } \frac{22.99}{58.44} \times 100 = 39.3\% \quad \text{Cl: } 100 - 39.3 = 60.7\%$$



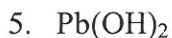
$$\text{S: } \frac{32.07}{64.07} \times 100 = 50.1\% \quad \text{O: } 100 - 50.1 = 49.9\%$$



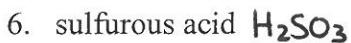
$$\text{Na: } \frac{22.99}{85.00} \times 100 = 27.0\% \quad \text{N: } \frac{14.01}{85.00} \times 100 = 16.5\% \quad \text{O: } 100 - 27.0 - 16.5 = 56.5\%$$



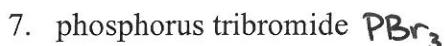
$$\text{Zn: } \frac{3(65.41)}{386.17} \times 100 = 50.8\% \quad \text{P: } \frac{2(30.97)}{386.17} \times 100 = 16.0\% \quad \text{O: } 100 - 50.8 - 16.0 = 33.2\%$$



$$\text{Pb: } \frac{207.19}{241.21} \times 100 = 85.9\% \quad \text{O: } \frac{32}{241.21} \times 100 = 13.3\% \quad \text{H: } 100 - 85.9 - 13.3 = 0.8\%$$



$$\text{H: } \frac{2.02}{82.09} \times 100 = 2.5\% \quad \text{S: } \frac{32.07}{82.09} \times 100 = 39.1\% \quad \text{O: } 100 - 2.5 - 39.1 = 58.4\%$$



$$\text{P: } \frac{30.97}{270.67} \times 100 = 11.4\% \quad \text{Br: } 100 - 11.4 = 88.6\%$$



$$\text{Ba: } \frac{137.33}{171.35} \times 100 = 80.1\% \quad \text{O: } \frac{32}{171.35} \times 100 = 18.7\% \quad \text{H: } 100 - 80.1 - 18.7 = 1.2\%$$



$$\text{Al: } \frac{26.98}{213.01} \times 100 = 12.7\% \quad \text{N: } \frac{3(14.01)}{213.01} \times 100 = 19.7\% \quad \text{O: } 100 - 12.7 - 19.7 = 67.6\%$$



$$\text{Fe: } \frac{55.85}{241.88} \times 100 = 23.1\% \quad \text{N: } \frac{3(14.01)}{241.88} \times 100 = 17.4\% \quad \text{O: } 100 - 23.1 - 17.4 = 59.5\%$$

Empirical & Molecular Formulas

Empirical formulas use percent composition information to determine the ratio of elements in a compound. Empirical formulas give the simplest formula possible, which does not necessarily represent the actual formula of a molecule. From experimentally collected data about the actual mass of the sample, the molecular formula can be determined using the empirical formula. See the example below for calculating empirical and molecular formulas from given data:

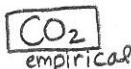
Given: 27.3% carbon + 72.7% oxygen; experimental mass = 88.02 g/mol

① Assume 100g of compound

② Convert to moles

③ Divide by smallest # of moles.

$$\frac{27.3 \text{ g C}}{12.01 \text{ g}} \frac{1 \text{ mol}}{1 \text{ mol}} = \frac{2.273 \text{ mol C}}{1} = 2.273$$



$$\frac{72.7 \text{ g O}}{16.00 \text{ g}} \frac{1 \text{ mol}}{1 \text{ mol}} = \frac{4.544 \text{ mol O}}{1} = 4.544$$

④ $\frac{\text{experimental}}{\text{empirical mass}} = \frac{88.02 \text{ g/mol}}{44.01 \text{ g/mol}} = 2$

⑤ multiply by factor from #4

$$2(\text{CO}_2) = \boxed{\text{C}_2\text{O}_4}$$

molecular

Determine the empirical formula:

1. 88.8% copper; 11.2% oxygen

$$\frac{88.8 \text{ g Cu}}{63.55 \text{ g}} \frac{1 \text{ mol}}{1 \text{ mol}} = 1.40 \text{ mol Cu}$$

$$\boxed{\text{Cu}_{1.40}\text{O}_{0.70}}$$

$$\frac{11.2 \text{ g O}}{16.00 \text{ g}} = 0.70 \text{ mol O}$$

2. 40.0% carbon; 6.7% hydrogen; 53.3% oxygen

$$\frac{40.0 \text{ g C}}{12.01 \text{ g}} \frac{1 \text{ mol}}{1 \text{ mol}} = 3.33 \text{ mol C}$$

$$\boxed{\text{C}_{3.33}\text{H}_{6.63}\text{O}_{3.34}}$$

$$\frac{6.7 \text{ g H}}{1.01 \text{ g}} \frac{1 \text{ mol}}{1 \text{ mol}} = 6.63 \text{ mol H}$$

$$\frac{53.5 \text{ g O}}{16.00 \text{ g}} \frac{1 \text{ mol}}{1 \text{ mol}} = 3.34 \text{ mol O}$$

3. 82.40% nitrogen; 17.60% hydrogen

$$\frac{82.40 \text{ g N}}{14.01 \text{ g}} \frac{1 \text{ mol}}{1 \text{ mol}} = 5.88 \text{ mol N}$$

$$\boxed{\text{N}_{5.88}\text{H}_{17.4}}$$

$$\frac{17.60 \text{ g H}}{1.01 \text{ g}} \frac{1 \text{ mol}}{1 \text{ mol}} = 17.4 \text{ mol H}$$

4. 10.04% carbon; 0.84% hydrogen; 89.12% chlorine

$$\frac{10.04 \text{ g C}}{12.01 \text{ g}} \frac{1 \text{ mol}}{1 \text{ mol}} = .836 \text{ mol C}$$

$$\frac{0.84 \text{ g H}}{1.01 \text{ g}} \frac{1 \text{ mol}}{1 \text{ mol}} = 0.832 \text{ mol H}$$

$$\boxed{\text{CHCl}_3}$$

5. 52.94% aluminum; 47.06% oxygen

$$\frac{52.94 \text{ g Al}}{26.98 \text{ g}} \frac{1 \text{ mol}}{1 \text{ mol}} = 1.96 \text{ mol Al}$$

$$\boxed{\text{Al}_{1.96}\text{O}_{2.94}} \quad 2(\text{AlO}_{1.5}) = \boxed{\text{Al}_2\text{O}_3}$$

$$\frac{47.06 \text{ g O}}{16.00 \text{ g}} \frac{1 \text{ mol}}{1 \text{ mol}} = 2.94 \text{ mol O}$$

6. 65.1% scandium; 34.9% oxygen

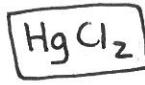
$$\frac{65.1 \text{ g Sc}}{49.96 \text{ g}} \frac{1 \text{ mol}}{1 \text{ mol}} = 1.45 \text{ mol Sc}$$

$$\boxed{\text{Sc}_{1.45}\text{O}_{2.18}} \quad 2(\text{ScO}_{1.5}) = \boxed{\text{Sc}_2\text{O}_3}$$

$$\frac{34.9 \text{ g O}}{16.00 \text{ g}} \frac{1 \text{ mol}}{1 \text{ mol}} = 2.18 \text{ mol O}$$

7. 74.17% mercury; 25.83% chlorine

$$\frac{74.17 \text{ g Hg}}{200.59 \text{ g}} \frac{1 \text{ mol}}{1 \text{ mol}} = .3698 \text{ mol Hg}$$



$$\frac{25.83 \text{ g Cl}}{35.45 \text{ g}} \frac{1 \text{ mol}}{1 \text{ mol}} = .729 \text{ mol Cl}$$

Determine the empirical and molecular formulas for each of the following.

1. 27.3% carbon; 72.7 oxygen; experimental molar mass is 132 g/mol.

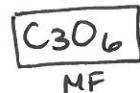
$$\frac{27.3 \text{ g}}{12.01 \text{ g}} \left| \begin{array}{l} \text{1 mol} \\ \hline \end{array} \right. = \frac{2.27 \text{ mol C}}{2.27}$$



$$\frac{132 \text{ g/mol}}{44.01 \text{ g/mol}} = 3$$



$$\frac{72.7 \text{ g}}{16.00 \text{ g}} \left| \begin{array}{l} \text{1 mol} \\ \hline \end{array} \right. = \frac{4.54 \text{ mol O}}{2.27}$$

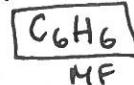
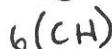


2. 92.3% carbon; 7.7% hydrogen; experimental molar mass is 78.1 g/mol.

$$\frac{92.3 \text{ g}}{12.01 \text{ g}} \left| \begin{array}{l} \text{1 mol} \\ \hline \end{array} \right. = 7.69 \text{ mol C}$$

$$\frac{78.1 \text{ g/mol}}{13.02 \text{ g/mol}} = 6$$

$$\frac{7.7 \text{ g}}{1.01 \text{ g}} \left| \begin{array}{l} \text{1 mol} \\ \hline \end{array} \right. = 7.62 \text{ mol H}$$



3. 94.12% sulfur; 5.88% hydrogen; experimental molar mass is 34.1 g/mol.

$$\frac{94.12 \text{ g}}{32.02 \text{ g}} \left| \begin{array}{l} \text{1 mol} \\ \hline \end{array} \right. = \frac{2.94 \text{ mol S}}{2.94}$$

$$\frac{34.1 \text{ g/mol}}{34.1 \text{ g/mol}} = 1$$

$$\frac{5.88 \text{ g}}{1.01 \text{ g}} \left| \begin{array}{l} \text{1 mol} \\ \hline \end{array} \right. = \frac{5.82 \text{ mol H}}{2.94}$$

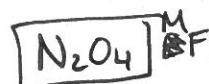
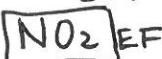


4. 30.43% nitrogen; 69.5% oxygen; experimental molar mass is 92.0 g/mol.

$$\frac{30.43 \text{ g N}}{14.01 \text{ g}} \left| \begin{array}{l} \text{1 mol} \\ \hline \end{array} \right. = \frac{2.17 \text{ mol N}}{2.17}$$

$$\frac{92.0 \text{ g/mol}}{46.01 \text{ g/mol}} = 2$$

$$\frac{69.5 \text{ g}}{16.00 \text{ g}} \left| \begin{array}{l} \text{1 mol} \\ \hline \end{array} \right. = \frac{4.34 \text{ mol O}}{2.17}$$



5. 73.8% carbon; 8.7% hydrogen; 17.5% nitrogen; experimental molar mass is 162.2 g/mol.

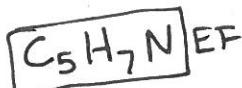
$$\frac{73.8 \text{ g C}}{12.01 \text{ g}} \left| \begin{array}{l} \text{1 mol} \\ \hline \end{array} \right. = \frac{6.14 \text{ mol C}}{1.25}$$

$$\frac{162.2 \text{ g/mol}}{81.1 \text{ g/mol}} = 2$$

$$\frac{8.7 \text{ g H}}{1.01 \text{ g}} \left| \begin{array}{l} \text{1 mol} \\ \hline \end{array} \right. = \frac{8.61 \text{ mol H}}{1.25}$$

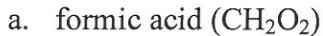


$$\frac{17.5 \text{ g}}{14.01 \text{ g}} \left| \begin{array}{l} \text{1 mol} \\ \hline \end{array} \right. = \frac{1.25 \text{ mol N}}{1.25}$$

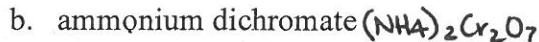


Practice with Ch. 11 Concepts – Honors Level

1. Determine the molar mass in each of the following quantities:

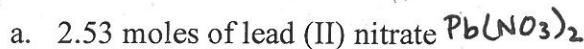


$$12.01 + 2(1.01) + 2(16.00) = \boxed{46.03 \text{ g/mol}}$$

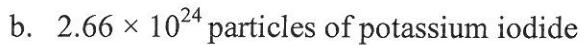


$$2(14.01) + 8(1.01) + 2(52.00) + 7(16.00) = \boxed{252.10 \text{ g/mol}}$$

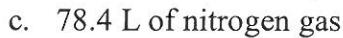
2. What is the mass, in grams, of each of the following quantities:



$$\frac{2.53 \text{ mol}}{1 \text{ mol}} \left| \begin{array}{c} 331.21 \text{ g} \\ \hline 1 \text{ mol} \end{array} \right| = \boxed{837 \text{ g } \text{Pb}(\text{NO}_3)_2}$$



$$\frac{2.66 \times 10^{24} \text{ particles}}{6.02 \times 10^{23} \text{ p}} \left| \begin{array}{c} 1 \text{ mol} \\ \hline 6.02 \times 10^{23} \text{ p} \end{array} \right| \left| \begin{array}{c} 166.00 \text{ g} \\ \hline 1 \text{ mol} \end{array} \right| = \boxed{733 \text{ g KI}}$$



$$\frac{78.4 \text{ L}}{22.4 \text{ L}} \left| \begin{array}{c} 1 \text{ mol} \\ \hline 1 \text{ mol} \end{array} \right| \left| \begin{array}{c} 28.02 \text{ g} \\ \hline 1 \text{ mol} \end{array} \right| = \boxed{98.07 \text{ g N}_2}$$

3. How many grams of copper (II) iodide are needed to get 8.00×10^{24} iodide ions?

$$\frac{8.00 \times 10^{24} \text{ I}^- \text{ ions}}{2 \text{ ions I}^-} \left| \begin{array}{c} 1 \text{ mol CuI}_2 \\ \hline 1 \text{ mol I}^- \end{array} \right| \left| \begin{array}{c} 1 \text{ mol} \\ \hline 6.02 \times 10^{23} \text{ mol} \end{array} \right| \left| \begin{array}{c} 317.35 \text{ g} \\ \hline 1 \text{ mol} \end{array} \right| = \boxed{2.11 \times 10^3 \text{ g CuI}_2}$$

4. If an element has 1.96×10^{22} atoms in 0.791 grams, what is its molar mass?

$$\frac{1.96 \times 10^{22} \text{ atoms}}{6.02 \times 10^{23} \text{ atoms}} = 0.0326 \text{ mol} \quad \frac{0.791 \text{ g}}{0.0326 \text{ mol}} = \boxed{24.3 \text{ g/mol}}$$

5. Find the density of gold, in g/cm^3 , when a cube measuring 0.787 inch by 3.15 inch by 1.57 inch contains 3.77×10^{24} atoms. (1 inch = 2.54 cm).

$$\frac{0.787 \text{ in}}{\text{in}} \left| \begin{array}{c} 2.54 \text{ cm} \\ \hline 1 \text{ in} \end{array} \right| = 2.00 \text{ cm} \quad (2.00)(8.00)(3.99) = 63.84 \text{ cm}^3$$

$$\frac{3.15 \text{ in}}{\text{in}} \left| \begin{array}{c} 2.54 \text{ cm} \\ \hline 1 \text{ in} \end{array} \right| = 8.00 \text{ cm} \quad \frac{1.57 \text{ in}}{\text{in}} \left| \begin{array}{c} 2.54 \text{ cm} \\ \hline 1 \text{ in} \end{array} \right| = 3.99 \text{ cm}$$

$$\frac{3.77 \times 10^{24} \text{ atoms}}{6.02 \times 10^{23} \text{ atoms}} \left| \begin{array}{c} 1 \text{ mol} \\ \hline 1 \text{ mol} \end{array} \right| \left| \begin{array}{c} 196.97 \text{ g} \\ \hline 1 \text{ mol} \end{array} \right| = 1230 \text{ g} \quad \frac{1230 \text{ g}}{63.84 \text{ cm}^3} = \boxed{19.3 \text{ g/cm}^3}$$

6. How many molecules of water would be present in a 15.0 mL sample, given that the density of water is 1.00 g/mL?

$$\frac{1.00 \text{ g}}{15.0 \text{ mL}} = \frac{x}{15.0 \text{ g}} \left| \begin{array}{c} 1 \text{ mol} \\ \hline 18.02 \text{ g} \end{array} \right| \left| \begin{array}{c} 6.02 \times 10^{23} \text{ mol} \\ \hline 1 \text{ mol} \end{array} \right| = \boxed{5.01 \times 10^{23} \text{ molecules H}_2\text{O}}$$

7. Determine the percent composition for each element in the following compounds:

a. manganese (II) oxide MnO

$$\text{Mn: } \frac{(54.94)}{70.94} \times 100 = 77.4\% \text{ Mn} \quad \text{O: } 100 - 77.4 = 22.6\% \text{ O}$$

b. propanol ($\text{C}_3\text{H}_8\text{O}$)

$$\text{C: } \frac{3(12.01)}{60.11} \times 100 = 59.9\% \text{ C} \quad \text{H: } \frac{8(1.01)}{60.11} \times 100 = 13.4\% \text{ H} \quad \text{O: } 100 - 59.9 - 13.4 = 26.7\% \text{ O}$$

c. calcium phosphate $\text{Ca}_3(\text{PO}_4)_2$

$$\text{Ca: } \frac{3(40.08)}{310.18} \times 100 = 38.8\% \quad \text{P: } \frac{2(30.97)}{310.18} \times 100 = 20.0\% \quad \text{O: } 100 - 38.8 - 20.0 = 81.2\% \text{ O}$$

8. Determine the empirical formula for a sample that has the following percent composition:

a. 5.93% hydrogen and 94.07% sulfur

$$\frac{5.93\text{ g}}{1.01\text{ g}} \mid \frac{1\text{ mol}}{1.01\text{ g}} = \frac{5.87\text{ mol}}{2.94} \quad \boxed{\text{H}_2\text{S}}$$

$$\frac{94.07\text{ g}}{32.01\text{ g}} \mid \frac{1\text{ mol}}{32.01\text{ g}} = \frac{2.94\text{ mol}}{2.94}$$

b. 80.68% mercury, 12.87% oxygen, and 6.45% sulfur

$$\frac{80.68\text{ g}}{200.59\text{ g}} \mid \frac{1\text{ mol}}{200.59\text{ g}} = .402\text{ mol Hg} \quad \frac{6.45\text{ g}}{32.01\text{ g}} \mid \frac{1\text{ mol}}{32.01\text{ g}} = .201\text{ mol S}$$

$$\frac{12.87\text{ g}}{16.00\text{ g}} \mid \frac{1\text{ mol}}{16.00\text{ g}} = .804\text{ mol O}$$

$$\boxed{\text{Hg}_2\text{SO}_4}$$

9. A 48.30 gram sample of a compound containing aluminum and iodine in a fixed ratio contains 3.20 grams of aluminum. What is the empirical formula of the compound?

$$\frac{3.20\text{ g}}{26.98\text{ g}} \mid \frac{1\text{ mol}}{1.01\text{ g}} = \frac{.119\text{ mol Al}}{.119} \quad \boxed{\text{AlI}_3}$$

$$48.30 - 3.20 = \frac{45.1\text{ g}}{126.90\text{ g}} \mid \frac{1\text{ mol}}{1.01\text{ g}} = \frac{.355\text{ mol I}}{.119}$$

10. Determine the empirical formula for caffeine, given that a sample contains 49.47 grams of carbon, 28.85 grams of nitrogen, 16.48 grams of oxygen, and 5.20 grams of hydrogen.

$$\frac{49.47\text{ g}}{12.01\text{ g}} \mid \frac{1\text{ mol}}{1.01\text{ g}} = \frac{4.12\text{ mol C}}{1.03} \quad \frac{16.48\text{ g}}{16.00\text{ g}} \mid \frac{1\text{ mol}}{1.01\text{ g}} = \frac{1.03\text{ mol O}}{1.03} \quad \boxed{\text{C}_4\text{N}_2\text{O}_5\text{H}_5}$$

$$\frac{28.85\text{ g}}{14.01\text{ g}} \mid \frac{1\text{ mol}}{1.01\text{ g}} = \frac{2.06\text{ mol N}}{1.03} \quad \frac{5.20\text{ g}}{1.01\text{ g}} \mid \frac{1\text{ mol}}{1.01\text{ g}} = \frac{5.15\text{ mol H}}{1.03}$$

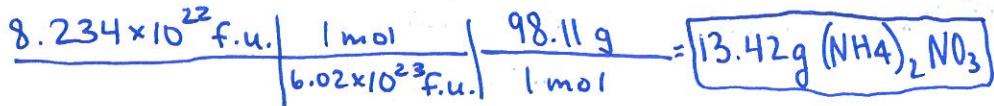
11. If the molar mass of caffeine is determined to be 194.19 g/mol, what is caffeine's molecular formula? (use information from #10 also)

$$\frac{194.19\text{ g/mol}}{97.11\text{ g/mol}} = 2 \quad 2(\text{C}_4\text{N}_2\text{O}_5\text{H}_5)$$

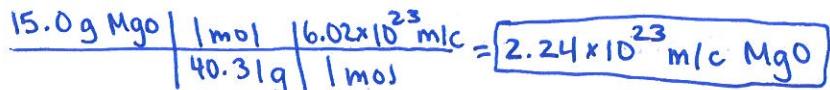
$$\boxed{\text{C}_8\text{N}_4\text{O}_2\text{H}_{10}}$$

Ch. 11 Practice II – Honors Level

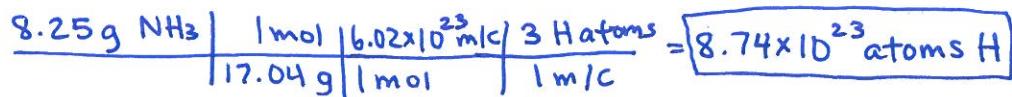
1. Determine the mass of 8.234×10^{22} formula units of ammonium nitrate. $(\text{NH}_4)_2\text{NO}_3$



2. How many molecules are in a 15.0 grams sample of magnesium oxide?



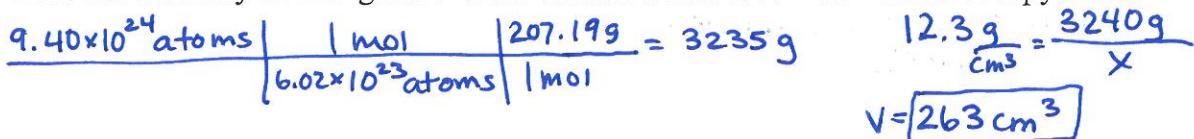
3. How many hydrogen atoms are in 8.25 grams of ammonia (NH_3)?



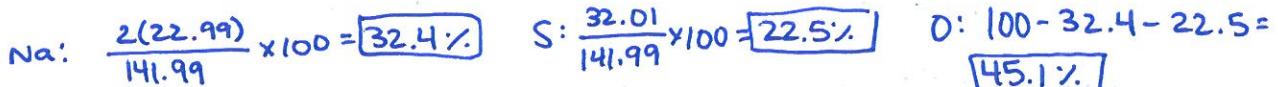
4. What is the molar mass of a compound that contains 2.25×10^{23} molecules in 14.45 grams of the compound?



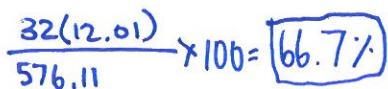
5. Lead has a density of 12.3 g/cm³. What volume would 9.40×10^{24} atoms occupy?



6. Find the percent composition for each element in sodium sulfate.



7. Determine the percent of carbon in copper phthalocyanine, Cu(C₈H₄N₂)₄.



8. Determine the mass of carbon found in 135 grams of silver oxalate.



$$\frac{2(12.01)}{303.76} \times 100 = 7.9\%$$

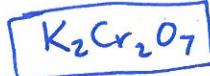
$$\frac{7.9}{100} = \frac{x}{135}$$

$$x = 10.7 \text{ g C}$$

9. Determine the empirical formula of a compound that contains 0.89 grams K, 1.18 grams Cr, and 1.27 grams O.

$$\frac{0.89 \text{ g K}}{39.10 \text{ g}} \left| \begin{array}{l} 1 \text{ mol} \\ \hline \end{array} \right. = 0.0228 \text{ mol K} \quad 2(\text{KCrO}_3 \cdot 5)$$

$$\frac{1.18 \text{ g Cr}}{52.00 \text{ g}} \left| \begin{array}{l} 1 \text{ mol} \\ \hline \end{array} \right. = 0.0227 \text{ mol Cr}$$



$$\frac{1.27 \text{ g O}}{16.00 \text{ g}} \left| \begin{array}{l} 1 \text{ mol} \\ \hline \end{array} \right. = 0.0794 \text{ mol O}$$

10. Use the following data to determine the empirical formula of the compound:

Mass of crucible (g)	15.00
Mass of crucible + iron (g)	24.68
Mass of crucible + iron oxide compound (g)	28.85

$$24.68 - 15.00 = 9.68 \text{ g iron}$$

$$28.85 - 24.68 = 4.17 \text{ g oxygen}$$

$$\frac{9.68 \text{ g Fe}}{55.85 \text{ g}} \left| \begin{array}{l} 1 \text{ mol} \\ \hline \end{array} \right. = .173 \text{ mol Fe} \quad 2(\text{FeO}_{1.5}) \quad \boxed{\text{Fe}_2\text{O}_3}$$

$$\frac{4.17 \text{ g O}}{16.00 \text{ g}} \left| \begin{array}{l} 1 \text{ mol} \\ \hline \end{array} \right. = .261 \text{ mol O}$$

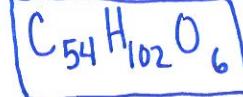
11. A fat is composed of 76.5% carbon, 11.3% oxygen, and 12.2% hydrogen. What is its molecular formula if it has a molar mass of 847 g/mol?

$$\frac{76.5 \text{ g C}}{12.01 \text{ g}} \left| \begin{array}{l} 1 \text{ mol} \\ \hline \end{array} \right. = 6.37 \text{ mol C} \quad \text{C}_9\text{H}_{17}\text{O}$$

$$\frac{11.3 \text{ g O}}{16.00 \text{ g}} \left| \begin{array}{l} 1 \text{ mol} \\ \hline \end{array} \right. = .706 \text{ mol O}$$

$$\frac{12.2 \text{ g H}}{1.01 \text{ g}} \left| \begin{array}{l} 1 \text{ mol} \\ \hline \end{array} \right. = 12.08 \text{ mol H}$$

$$\frac{847 \text{ g/mol}}{141.26 \text{ g/mol}} = 6 (\text{C}_9\text{H}_{17}\text{O})$$



Hydrate Problems

1. A 50.00 gram sample of hydrated manganese (II) chloride yields 31.75 grams of the anhydrous compound after heating. Determine the chemical formula and the name of the hydrate.

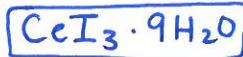
$$50.00\text{ g} - 31.75\text{ g} = 18.25\text{ g H}_2\text{O} \quad \frac{1\text{ mol}}{18.02\text{ g}} = \frac{1.01\text{ mol H}_2\text{O}}{0.252\text{ mol}}$$



$$\frac{31.75\text{ g}}{125.84\text{ g}} = \frac{1\text{ mol}}{0.252\text{ mol}}$$

2. Cerium (III) iodide occurs as a hydrate with a composition of 76.3% of cerium (III) iodide and 23.7% water. Determine the chemical formula of the hydrate.

$$\frac{76.3\text{ g}}{520.82\text{ g}} = .146\text{ mol}$$



$$\frac{23.7\text{ g}}{18.02\text{ g}} = 1.32\text{ mol}$$

3. A hydrate with a molar mass of 174.37 grams has 31.0% water. Find the formula of this hydrate containing 13.9% magnesium, 4.1% hydrogen, 17.8% phosphorus, and 64.2% oxygen.

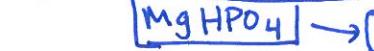
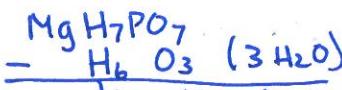
$$\frac{13.9\text{ g}}{24.31\text{ g}} = .57\text{ mol Mg} \quad \frac{64.2\text{ g}}{16.00\text{ g}} = 4.02\text{ mol O}$$

$$69\% \text{ anhydrite} \times 174.37\text{ g/mol} = 120.32\text{ g/mol}$$

M Manhydrate

$$\frac{4.1\text{ g}}{1.01\text{ g}} = 4.06\text{ mol H}$$

$$\frac{17.8\text{ g}}{30.97\text{ g}} = .57\text{ mol P}$$



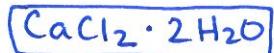
$$\frac{31.0\text{ g H}_2\text{O}}{18.02\text{ g}} = 1.720\text{ mol H}_2\text{O}$$

$$\frac{69.0\text{ g}}{120.32\text{ g}} = .570\text{ mol anh.}$$

$$\begin{array}{c} \text{MgHPO}_4 \cdot 3\text{H}_2\text{O} \\ | \\ 3\text{H}_2\text{O} \end{array}$$

4. A calcium chloride hydrate has a mass of 4.72 grams. After heating, the mass of the anhydrite is found to be 3.56 grams. Determine the formula of the hydrate.

$$4.72 - 3.56 = 1.16\text{ g H}_2\text{O} \quad \frac{1\text{ mol}}{18.02\text{ g}} = 0.064\text{ mol}$$

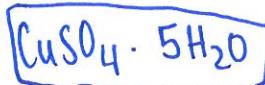


$$\frac{3.56\text{ g CaCl}_2}{111.16\text{ g}} = 0.032\text{ mol}$$

5. A sample of copper (II) sulfate hydrate has a mass of 3.97 grams. After heating, the anhydrite has a mass of 2.54 grams. Determine the formula of the hydrate.

$$3.97 - 2.54 = 1.43\text{ g H}_2\text{O} \quad \frac{1\text{ mol}}{18.02\text{ g}} = 0.0794\text{ mol}$$

$$\frac{2.54\text{ g CuSO}_4}{159.56\text{ g}} = 0.0159\text{ mol}$$



6. A sample of sodium carbonate hydrate has a mass of 8.85 grams. It loses 1.28 grams when heated. Find the formula of the hydrate.

$$\frac{8.85 - 1.28}{1.28} = \frac{7.57 \text{ g anhydrous}}{18.02 \text{ g}} \times \frac{1 \text{ mol}}{105.99 \text{ g}} = 0.0714 \text{ mol}$$

$$\frac{1.28 \text{ g}}{18.02 \text{ g}} \times \frac{1 \text{ mol}}{105.99 \text{ g}} = 0.0710 \text{ mol}$$



7. A hydrate is determined to be 45.43% water and 54.57% CoCl_2 . Find the chemical formula and the name of this hydrate.

$$\frac{45.43 \text{ g}}{100 \text{ g}} \times \frac{1 \text{ mol}}{18.02 \text{ g}} = 2.52 \text{ mol H}_2\text{O}$$

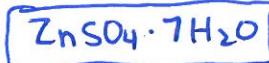
$$\frac{54.57 \text{ g}}{100 \text{ g}} \times \frac{1 \text{ mol}}{129.83 \text{ g}} = .420 \text{ mol}$$



8. A 17.44 gram sample of a zinc sulfate hydrate is heated in a crucible to drive off all water of hydration. After heating, the anhydrous compound has a mass of 9.79 grams. Determine the formula of this hydrate.

$$\frac{17.44 - 9.79}{9.79} = \frac{7.65 \text{ g}}{161.42 \text{ g}} \times \frac{1 \text{ mol}}{18.02 \text{ g}} = .425 \text{ mol H}_2\text{O}$$

$$\frac{9.79 \text{ g}}{161.42 \text{ g}} \times \frac{1 \text{ mol}}{18.02 \text{ g}} = .0606 \text{ mol}$$

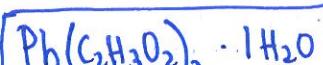


9. Use the following data to find the formula of a lead (II) acetate hydrate.

Mass of crucible (g)	20.00
Mass of crucible + hydrate (g)	21.04
Mass of crucible + anhydrite (g)	20.99

$$21.04 - 20.99 = \frac{0.05 \text{ g anhydrite}}{325.29 \text{ g}} \times \frac{1 \text{ mol}}{18.02 \text{ g}} = \frac{0.05 \text{ g} \times 10^{-3} \text{ mol}}{325.29 \text{ g}} = 2.77 \times 10^{-5} \text{ mol}$$

$$\frac{0.99 \text{ g anhydrite}}{325.29 \text{ g}} \times \frac{1 \text{ mol}}{18.02 \text{ g}} = 3.04 \times 10^{-3} \text{ mol}$$



Chapter 11 Review

1. Calculate the number of moles of FeSO_4 in 25.6 grams of this compound.

$$\frac{25.6 \text{ g FeSO}_4}{151.86 \text{ g}} \left| \begin{array}{c} |1\text{ mol}| \\ |1\text{ mol}| \end{array} \right| = [0.169 \text{ mol}]$$

2. Calculate the number of liters of NO_2 gas at STP in 79.7 grams of this compound.

$$\frac{79.7 \text{ g}}{46.01 \text{ g}} \left| \begin{array}{c} |1\text{ mol}| \\ |1\text{ mol}| \end{array} \right| \left| \begin{array}{c} |22.4 \text{ L}| \\ |1\text{ mol}| \end{array} \right| = [38.8 \text{ L}]$$

3. How many chloride ions are there in 2.19 grams of aluminum chloride?

$$\frac{2.19 \text{ g}}{133.33 \text{ g}} \left| \begin{array}{c} |1\text{ mol}| \\ |1\text{ mol}| \end{array} \right| \left| \begin{array}{c} |6.02 \times 10^{23} \text{ mol}^{-1}| \\ |1\text{ mol}| \end{array} \right| \left| \begin{array}{c} |3 \text{ ions Cl}^-| \\ |1\text{ mol}| \end{array} \right| = [2.70 \times 10^{22} \text{ ions Cl}^-]$$

4. How many atoms of carbon are in 19.1 grams of glucose?

$$\frac{19.1 \text{ g}}{180.18 \text{ g}} \left| \begin{array}{c} |1\text{ mol}| \\ |1\text{ mol}| \end{array} \right| \left| \begin{array}{c} |6.02 \times 10^{23} \text{ mol}^{-1}| \\ |1\text{ mol}| \end{array} \right| \left| \begin{array}{c} |6 \text{ atoms C}| \\ |1\text{ mol}| \end{array} \right| = [3.83 \times 10^{23} \text{ atoms C}]$$

5. Determine the following for 45.2 grams of the compound $\text{PtCl}_2(\text{NH}_3)_2$:

- a. number of moles of nitrogen
- b. number of atoms of nitrogen
- c. mass percent (percent composition) of nitrogen
- d. actual mass of nitrogen in the compound

(a) $\frac{45.2 \text{ g}}{300.06 \text{ g}} \left| \begin{array}{c} |1\text{ mol}| \\ |1\text{ mol}| \end{array} \right| \left| \begin{array}{c} |1\text{ mol N}| \\ |1\text{ mol compd}| \end{array} \right| = [.301 \text{ mol N}]$

(b) $.301 \text{ mol} \left| \begin{array}{c} |6.02 \times 10^{23} \text{ atoms}| \\ |1\text{ mol}| \end{array} \right| = [1.81 \times 10^{23} \text{ atoms N}]$

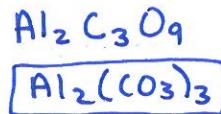
(c) $\frac{2(14.01)}{300.06} \times 100 = [9.34\%]$

(d) $\frac{9.34}{100} = \frac{x}{45.2} \quad [4.22 \text{ g N}]$

6. Determine the empirical formula of a compound composed of 23.1% aluminum, 15.4% carbon, and 61.5% oxygen.

$$\frac{23.1 \text{ g Al}}{26.98 \text{ g}} \left| \begin{array}{c} |1\text{ mol}| \\ |1\text{ mol}| \end{array} \right| = .856 \text{ mol Al} \quad 2(\text{AlC}_{1.5}\text{O}_{4.5})$$

$$\frac{15.4 \text{ g C}}{12.01 \text{ g}} \left| \begin{array}{c} |1\text{ mol}| \\ |1\text{ mol}| \end{array} \right| = 1.26 \text{ mol C}$$



$$\frac{61.5 \text{ g O}}{16.00 \text{ g}} \left| \begin{array}{c} |1\text{ mol}| \\ |1\text{ mol}| \end{array} \right| = 3.84 \text{ mol O}$$

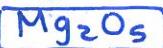
7. To find the experimental empirical formula of a compound, a student heats a piece of magnesium and collects the following data. Determine the empirical formula for oxide of magnesium that is produced.

Length of magnesium strip	35.00 cm
Mass of empty crucible	20.74 g
Mass of 2.00 meters of magnesium	1.44 g
Mass of crucible + oxide of magnesium product	21.17 g

$$\frac{1.44 \text{ g}}{2.00 \text{ m}} = \frac{x}{.035 \text{ m}} \quad x = 0.252 \text{ g Mg} \quad \left| \begin{array}{l} 1 \text{ mol} \\ 24.305 \text{ g} \end{array} \right. = 0.0104 \text{ mol}$$

$$21.17 - 20.74 \text{ g} = .43 \text{ g MgO}$$

$$\frac{.43}{.0104} = \frac{.420}{\left| \begin{array}{l} 1 \text{ mol} \\ 16.00 \text{ g} \end{array} \right.} = .0253 \text{ mol}$$

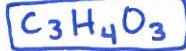


8. Vitamin C is composed of 40.92% carbon, 4.580% hydrogen, and 54.50% oxygen. Its molar mass is 176.0 g/mol. Calculate its molecular formula.

$$\frac{40.92 \text{ g C}}{12.01 \text{ g}} \left| \begin{array}{l} 1 \text{ mol} \\ \hline \end{array} \right. = 3.41 \text{ mol C}$$



$$\frac{4.580 \text{ g H}}{1.01 \text{ g}} \left| \begin{array}{l} 1 \text{ mol} \\ \hline \end{array} \right. = 4.53 \text{ mol H}$$



$$\frac{54.50 \text{ g O}}{16.00 \text{ g}} \left| \begin{array}{l} 1 \text{ mol} \\ \hline \end{array} \right. = 3.41 \text{ mol O}$$

9. A compound is found to contain 36.84% and 63.16% oxygen. If the molar mass is 152.0 g/mol, determine the molecular formula of the compound.

$$\frac{36.84 \text{ g N}}{14.01 \text{ g}} \left| \begin{array}{l} 1 \text{ mol} \\ \hline \end{array} \right. = 2.63 \text{ mol N}$$



$$\frac{63.16 \text{ g O}}{16.00 \text{ g}} \left| \begin{array}{l} 1 \text{ mol} \\ \hline \end{array} \right. = 3.95 \text{ mol O}$$

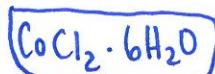
$$\frac{152.0 \text{ g/mol}}{76.02 \text{ g/mol}} = 2$$



10. 1.000 gram of the hydrate of cobalt (II) chloride was heated until all of the water was driven off. After heating, 0.546 grams of the anhydride remained. What is the formula of this hydrate?

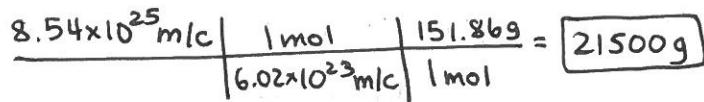
$$1.000 \text{ g} - 0.546 = .454 \text{ g H}_2\text{O} \left| \begin{array}{l} 1 \text{ mol} \\ 18.02 \text{ g} \end{array} \right. = .025 \text{ mol}$$

$$\frac{.546 \text{ g CoCl}_2}{129.83 \text{ g}} \left| \begin{array}{l} 1 \text{ mol} \\ \hline \end{array} \right. = .00421 \text{ mol}$$

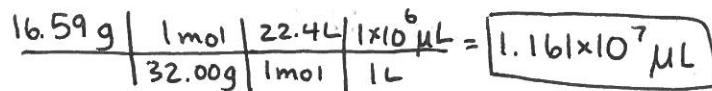


Chapter 11 Review II

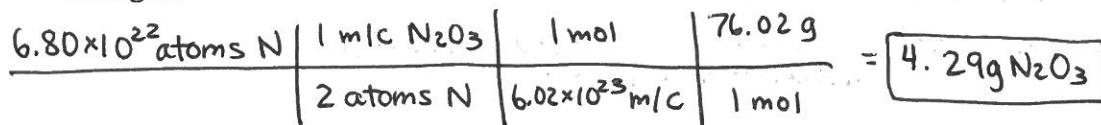
1. Calculate the number of grams of FeSO_4 in 8.54×10^{25} molecules of this compound.



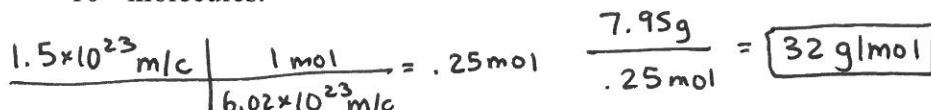
2. Calculate the number of microliters of O_2 gas at STP in 16.59 grams of this compound.



3. How many grams of dinitrogen trioxide are present when you have 6.80×10^{22} atoms of nitrogen?

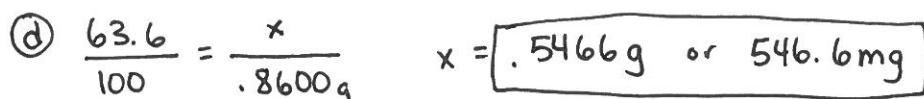
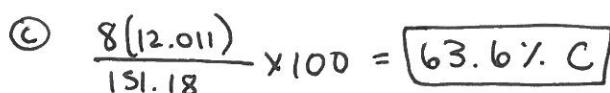
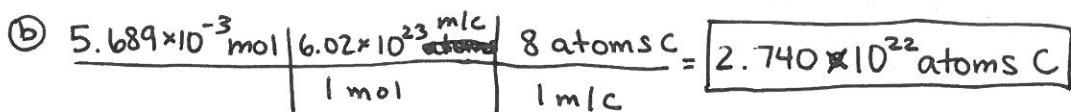
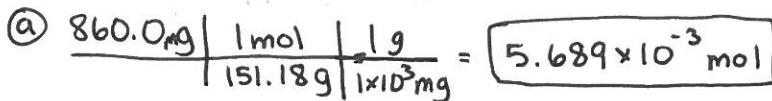


4. Determine the molar mass of a substance with a mass of 7.95 grams that contains 1.5×10^{23} molecules.



5. Determine the following for 860.0 milligrams of acetaminophen ($\text{C}_8\text{H}_9\text{NO}_2$)

- a. number of moles of acetaminophen
- b. number of atoms of carbon
- c. mass percent (percent composition) of carbon
- d. actual mass of carbon in the compound



6. Caffeine has the following percent composition: carbon 49.48%, hydrogen 5.19%, oxygen 16.48% and nitrogen 28.85%. Its molecular weight is 194.19 g/mol. What is its molecular formula?

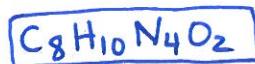
$$\frac{49.48 \text{ g C}}{12.01 \text{ g}} \times \frac{1 \text{ mol}}{1.03} = \frac{4.12 \text{ mol C}}{1.03}$$



$$\frac{5.19 \text{ g H}}{1.01 \text{ g}} \times \frac{1 \text{ mol}}{1.03} = \frac{5.14 \text{ mol H}}{1.03}$$

$$\frac{194.19 \text{ g/mol}}{97.11 \text{ g/mol}} = 2$$

$$\frac{28.85 \text{ g N}}{14.01 \text{ g}} \times \frac{1 \text{ mol}}{1.03} = \frac{2.06 \text{ mol N}}{1.03}$$



$$\frac{16.48 \text{ g O}}{16.00 \text{ g}} \times \frac{1 \text{ mol}}{1.03} = \frac{1.03 \text{ mol O}}{1.03}$$

7. Given that a sample of lead contains 1.20×10^{24} atoms and during water displacement, the water level increases from 45.0 mL to 81.4 mL, determine the density of lead.

$$81.4 - 45.0 = 36.4 \text{ mL} = \text{volume}$$

$$\frac{1.20 \times 10^{24} \text{ atoms}}{6.02 \times 10^{23} \text{ atoms}} \times \frac{1 \text{ mol}}{1 \text{ mol}} \times \frac{207.19 \text{ g}}{1 \text{ mol}} = 413.0 \text{ g}$$

$$\frac{413.0 \text{ g}}{36.4 \text{ mL}} = \boxed{11.3 \text{ g/cm}^3}$$

8. Given the lab data below, determine the number of waters of hydration present.

Mass of crucible: 18.050 grams

Mass of crucible + magnesium carbonate hydrate: 33.720 grams

Mass of crucible + magnesium carbonate anhydrite: 25.63 grams

$$33.720 - 25.63 = \frac{8.09 \text{ g H}_2\text{O}}{18.02 \text{ g}} \times \frac{1 \text{ mol}}{1 \text{ mol}} = \frac{.449 \text{ mol H}_2\text{O}}{.0899}$$

$$25.63 - 18.050 = \frac{7.58 \text{ g MgCO}_3}{84.32 \text{ g}} \times \frac{1 \text{ mol}}{1 \text{ mol}} = \frac{.0899 \text{ mol MgCO}_3}{.0899}$$

