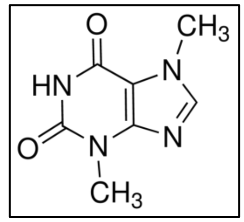
Moles and Stoichiometry Problems

1. Theobromine is an organic compound that gives dark chocolate a somewhat bitter taste. The structure of theobromine is shown at the right.



   a. If the molecular formula of theobromine is $C_7H_8N_4O_2$, what is the molar mass of theobromine?

   7(12.011) + 8(1.008) + 4(14.007) + 2(15.999) = 180.167

   Thus, the molar mass is 180.167 g/mol

   Given that you should use at least two decimal places for molar masses from the periodic table, your molar mass should have at least two decimal places.

   b. Theobromine is also toxic to dogs at doses greater than 300 mg of theobromine per kg body weight. That means my 10.0 kg dog would experience toxic effects if she eats 3.0 grams of theobromine. How many moles are present in 3.0 grams of theobromine?

   $$3.0\ \text{\sout{g theobromine}} \times \frac{1\ \text{mol theobromine}}{180.167\ \text{\sout{g theobromine}}} = 0.01665\ \text{mol theobromine}$$

   To determine the number of significant figures, we count the number of sig figs in the data used in the calculation. 3.0 g has 2 sig figs and 180.167 g/mol has 6 sig figs. Since we are multiplying, the number of sig figs in the result must be the smallest of 2 or 6. Therefore, our result should have 2 sig figs: **0.017 mol theobromine.**

2. Calculate the mass in grams of a single water molecule

   1 mol H_2O= 18.02 g

   $$\frac{18.02\ \text{g}\ H_2O}{1\ \text{mol}H_2O} \times \frac{1\ \text{mol}}{6.02 \times 10^{23}\ \text{atoms}} = \frac{\mathbf{2.99 \times 10^{-23}\ g}}{\mathbf{1\ atom}}$$

   Therefore, 1 atom has a mass of 2.99 x 10^{-23} g

3. A cube of gold (Au) that is 1.00 cm on each side has a mass of 19.32 g. How many gold atoms does this cube contain?

   This problem requires us to convert the mass of gold to a number of atoms. Avogadro's number will allow the conversion of *moles* to atoms, but first we need to convert to moles using the molar mass.

$$19.32\ \cancel{g\ Au} \times \frac{1\ mol\ Au}{196.967\ \cancel{g\ Au}} = 0.09809\ moles\ Au$$

Now Avogadro's number will allow us to calculate atoms:

$$0.09809\ \cancel{mol\ Au} \times \frac{6.022 \times 10^{23}\ atoms\ Au}{1\ \cancel{mol\ Au}} = 5.91 \times 10^{22}\ atoms\ Au$$

4. Consider the rusting of iron in air. In this reaction iron metal reacts with oxygen from the atmosphere to produce iron (III) oxide. How many grams of iron must react with an excess of oxygen to produce 33.3 grams of iron (III) oxide?

$$4Fe(s) + 3O_2(g) \rightarrow 2Fe_2O_3(s)$$

$$33.3\ \cancel{g\ Fe_2O_3} \times \frac{1\ \cancel{mol\ Fe_2O_3}}{159.7\ \cancel{g\ Fe_2O_3}} \times \frac{4\ \cancel{mol\ Fe}}{2\ \cancel{mol\ Fe_2O_3}} \times \frac{55.85\ g\ Fe}{1\ \cancel{mol\ Fe}} = 23.3\ g\ Fe$$

5. Ammonia (NH_3) reacts with oxygen (O_2) to produce nitric acid (HNO_3) and water. Write the balanced reaction for this process and calculate the mass of nitric acid that can be produced from 549 g of ammonia.

$$NH_3 + 2O_2 \rightarrow HNO_3 + H_2O$$

$$549\ \cancel{g\ NH_3} \times \frac{1\ mol\ NH_3}{17.03\ \cancel{g\ NH_3}} = 32.23\ mol\ NH_3$$

$$32.23\ \cancel{mol\ NH_3} \times \frac{1\ mol\ HNO_3}{1\ \cancel{mol\ NH_3}} = 32.23\ mol\ HNO_3$$

$$32.23\ \cancel{mol\ HNO_3} \times \frac{63.01\ g\ HNO_3}{1\ \cancel{mol\ HNO_3}} = 2031\ g\ HNO_3 = \mathbf{2030\ g\ HNO_3\ (3\ sig\ figs)}$$

6. Toluene (C_7H_8) can react with nitric acid (HNO_3) to produce the explosive TNT ($C_7H_5N_3O_6$) and water by the balanced reaction below. What mass of TNT can be made from 829 g of toluene?

$$C_7H_8 + HNO_3 \rightarrow C_7H_5N_3O_6 + H_2O \quad \text{(not balanced)}$$

First balance the reaction:

$$C_7H_8 + 3\ HNO_3 \rightarrow C_7H_5N_3O_6 + 3\ H_2O$$

Now tackle the stoichiometry

$$829 \text{ \sout{g C_7H_8}} \times \frac{1 \text{ mol } C_7H_8}{92.14 \text{ \sout{g C_7H_8}}} = 8.997 \text{ mol } C_7H_8$$

$$8.997 \text{ \sout{mol C_7H_8}} \times \frac{1 \text{ mol } C_7H_5N_3O_6}{1 \text{ \sout{mol C_7H_8}}} = 8.997 \text{ mol } C_7H_5N_3O_6$$

$$8.997 \text{ \sout{mol C_7H_5N_3O_6}} \times \frac{227.13 \text{ g } C_7H_5N_3O_6}{1 \text{ \sout{mol C_7H_5N_3O_6}}} = \mathbf{2040 \text{ g } C_7H_5N_3O_6 \text{ (3 sig figs)}}$$

7. For the reaction below, what is the limiting reactant when 0.253 g aluminum reacts with 0.482 g Cl_2? You must justify your answer with a calculation.

$$2Al(s) + 3Cl_2(g) \rightarrow 2AlCl_3(s)$$

If aluminum is the L.R. how much $AlCl_3$ could be made?.

$$0.253 \text{ \sout{g Al}} \times \frac{1 \text{ \sout{mol Al}}}{26.98 \text{ \sout{g Al}}} \times \frac{2 \text{ \sout{mol AlCl_3}}}{2 \text{ \sout{mol Al}}} = \mathbf{0.00938 \text{ mol } AlCl_3} \times \frac{133.34 \text{ g } AlCl_3}{1 \text{ mol } AlCl_3} = \mathbf{1.25g\ AlCl_3}$$

If chlorine is the L.R. how much $AlCl_3$ could be made?

$$0.482 \text{ \sout{g Cl_2}} \times \frac{1 \text{ \sout{mol Cl_2}}}{70.90 \text{ \sout{g Cl_2}}} \times \frac{2 \text{ \sout{mol AlCl_3}}}{3 \text{ \sout{mol Cl_2}}} = \mathbf{0.00453 \text{ mol } AlCl_3} \times \frac{133.34 \text{ g } AlCl_3}{1 \text{ mol } AlCl_3} = \mathbf{0.604g\ AlCl_3}$$

Since less $AlCl_3$ is produced when all of the Cl_2 is consumed, Cl_2 must be the limiting reagent.