3.7 What is the mass in grams of 13.2 amu?

The unit factor required is $\left(\dfrac{6.022 \times 10^{23}\ \text{amu}}{1\ \text{g}}\right)$

$$\textbf{? g} = 13.2\ \cancel{\text{amu}} \times \frac{1\ \text{g}}{6.022 \times 10^{23}\ \cancel{\text{amu}}} = \mathbf{2.19 \times 10^{-23}\ g}$$

3.8 How many amu are there in 8.4 g?

The unit factor required is $\left(\dfrac{6.022 \times 10^{23}\ \text{amu}}{1\ \text{g}}\right)$

$$\textbf{? amu} = 8.4\ \cancel{\text{g}} \times \frac{6.022 \times 10^{23}\ \text{amu}}{1\ \cancel{\text{g}}} = \mathbf{5.1 \times 10^{24}\ amu}$$

3.26 How many molecules of ethane (C_2H_6) are present in 0.334 g of C_2H_6?

Strategy: We are given grams of ethane and asked to solve for molecules of ethane. We cannot convert directly from grams ethane to molecules of ethane. What unit do we need to obtain first before we can convert to molecules? How should Avogadro's number be used here?

Solution: To calculate number of ethane molecules, we first must convert grams of ethane to moles of ethane. We use the molar mass of ethane as a conversion factor. Once moles of ethane are obtained, we can use Avogadro's number to convert from moles of ethane to molecules of ethane.

$$\text{molar mass of } C_2H_6 = 2(12.01\text{ g}) + 6(1.008\text{ g}) = 30.068\text{ g}$$

The conversion factor needed is

$$\frac{1\text{ mol } C_2H_6}{30.068\text{ g } C_2H_6}$$

Avogadro's number is the key to the second conversion. We have

$$1\text{ mol} = 6.022 \times 10^{23}\text{ particles (molecules)}$$

From this equality, we can write the conversion factor:

$$\frac{6.022 \times 10^{23}\text{ ethane molecules}}{1\text{ mol ethane}}$$

Let's complete the two conversions in one step.

$$\text{grams of ethane} \rightarrow \text{moles of ethane} \rightarrow \text{number of ethane molecules}$$

$$\textbf{? molecules of } \mathbf{C_2H_6} = 0.334\cancel{\text{ g}}\ C_2H_6 \times \frac{1\text{ m}\cancel{\text{o}}\text{l } C_2H_6}{30.068\cancel{\text{ g}}\ C_2H_6} \times \frac{6.022 \times 10^{23}\ C_2H_6\text{ molecules}}{1\text{ m}\cancel{\text{o}}\text{l } C_2H_6}$$

$$= \mathbf{6.69 \times 10^{21}\ C_2H_6\ molecules}$$

Check: Should 0.334 g of ethane contain fewer than Avogadro's number of molecules? What mass of ethane would contain Avogadro's number of molecules?

3.27 Calculate the number of C, H, and O atoms in 1.50 g of glucose ($C_6H_{12}O_6$), a sugar.

$$1.50\cancel{\text{ g}}\text{ glucose} \times \frac{1\text{ m}\cancel{\text{o}}\text{l glucose}}{180.2\cancel{\text{ g}}\text{ glucose}} \times \frac{6.022 \times 10^{23}\text{ mol}\cancel{\text{e}}\text{cules glucose}}{1\text{ m}\cancel{\text{o}}\text{l glucose}} \times \frac{6\text{ C atoms}}{1\text{ mol}\cancel{\text{e}}\text{cule glucose}}$$

$$= \mathbf{3.01 \times 10^{22}\ C\ atoms}$$

The ratio of O atoms to C atoms in glucose is 1:1. Therefore, there are the same number of O atoms in glucose as C atoms, so the number of O atoms = $\mathbf{3.01 \times 10^{22}}$ **O atoms**.

The ratio of H atoms to C atoms in glucose is 2:1. Therefore, there are twice as many H atoms in glucose as C atoms, so the number of H atoms = $2(3.01 \times 10^{22}\text{ atoms})$ = $\mathbf{6.02 \times 10^{22}}$ **H atoms**.

3.40 For many years chloroform ($CHCl_3$) was used as an inhalation anesthetic in spite of the fact that it is also a toxic substance that may cause severe liver, kidney, and heart damage. Calculate the percent composition by mass of this compound.

Strategy: Recall the procedure for calculating a percentage. Assume that we have 1 mole of $CHCl_3$. The percent by mass of each element (C, H, and Cl) is given by the mass of that element in 1 mole of $CHCl_3$ divided by the molar mass of $CHCl_3$, then multiplied by 100 to convert from a fractional number to a percentage.

Solution: The molar mass of $CHCl_3$ = 12.01 g/mol + 1.008 g/mol + 3(35.45 g/mol) = 119.4 g/mol. The percent by mass of each of the elements in $CHCl_3$ is calculated as follows:

$$\%\text{C} = \frac{12.01\text{ g/mol}}{119.4\text{ g/mol}} \times 100\% = \mathbf{10.06\%}$$

$$\%\text{H} = \frac{1.008\text{ g/mol}}{119.4\text{ g/mol}} \times 100\% = \mathbf{0.8442\%}$$

$$\%\text{Cl} = \frac{3(35.45)\text{ g/mol}}{119.4\text{ g/mol}} \times 100\% = \mathbf{89.07\%}$$

Check: Do the percentages add to 100%? The sum of the percentages is (10.06% + 0.8442% + 89.07%) = 99.97%. The small discrepancy from 100% is due to the way we rounded off.

3.44 Peroxyacylnitrate (PAN) is one of the components of smog. It is a compound of C, H, N, and O. Determine the percent composition of oxygen and the empirical formula from the following percent composition by mass: 19.8 percent C, 2.50 percent H, 11.6 percent N. What is its molecular formula given that its molar mass is about 120 g?

METHOD 1:

Step 1: Assume you have exactly 100 g of substance. 100 g is a convenient amount, because all the percentages sum to 100%. The percentage of oxygen is found by difference:

$$100\% - (19.8\% + 2.50\% + 11.6\%) = 66.1\%$$

In 100 g of PAN there will be 19.8 g C, 2.50 g H, 11.6 g N, and 66.1 g O.

Step 2: Calculate the number of moles of each element in the compound. Remember, an *empirical formula* tells us which elements are present and the simplest whole-number ratio of their atoms. This ratio is also a mole ratio. Use the molar masses of these elements as conversion factors to convert to moles.

$$n_{\text{C}} = 19.8\cancel{\text{g}}\text{ C} \times \frac{1\text{ mol C}}{12.01\cancel{\text{g}}\text{ C}} = 1.649\text{ mol C}$$

$$n_{\text{H}} = 2.50\cancel{\text{g}}\text{ H} \times \frac{1\text{ mol H}}{1.008\cancel{\text{g}}\text{ H}} = 2.480\text{ mol H}$$

$$n_{\text{N}} = 11.6\cancel{\text{g}}\text{ N} \times \frac{1\text{ mol N}}{14.01\cancel{\text{g}}\text{ N}} = 0.8280\text{ mol N}$$

$$n_{\text{O}} = 66.1\cancel{\text{g}}\text{ O} \times \frac{1\text{ mol O}}{16.00\cancel{\text{g}}\text{ O}} = 4.131\text{ mol O}$$

Step 3: Try to convert to whole numbers by dividing all the subscripts by the smallest subscript. The formula is $C_{1.649}H_{2.480}N_{0.8280}O_{4.131}$. Dividing the subscripts by 0.8280 gives the empirical formula, $\mathbf{C_2H_3NO_5}$.

To determine the molecular formula, remember that the molar mass/empirical mass will be an integer greater than or equal to one.

$$\frac{\text{molar mass}}{\text{empirical molar mass}} \geq 1 \text{ (integer values)}$$

In this case,

$$\frac{\text{molar mass}}{\text{empirical molar mass}} = \frac{120 \text{ g}}{121.05 \text{ g}} \approx 1$$

Hence, the molecular formula and the empirical formula are the same, $\mathbf{C_2H_3NO_5}$.

METHOD 2:

Step 1: Multiply the mass % (converted to a decimal) of each element by the molar mass to convert to grams of each element. Then, use the molar mass to convert to moles of each element.

$$n_C = (0.198) \times (120 \cancel{\text{g}}) \times \frac{1 \text{ mol C}}{12.01 \cancel{\text{g}} \text{ C}} = 1.98 \text{ mol C} \approx \mathbf{2 \text{ mol C}}$$

$$n_H = (0.0250) \times (120 \cancel{\text{g}}) \times \frac{1 \text{ mol H}}{1.008 \cancel{\text{g}} \text{ H}} = 2.98 \text{ mol H} \approx \mathbf{3 \text{ mol H}}$$

$$n_N = (0.116) \times (120 \cancel{\text{g}}) \times \frac{1 \text{ mol N}}{14.01 \cancel{\text{g}} \text{ N}} = 0.994 \text{ mol N} \approx \mathbf{1 \text{ mol N}}$$

$$n_O = (0.661) \times (120 \cancel{\text{g}}) \times \frac{1 \text{ mol O}}{16.00 \cancel{\text{g}} \text{ O}} = 4.96 \text{ mol O} \approx \mathbf{5 \text{ mol O}}$$

Step 2: Since we used the molar mass to calculate the moles of each element present in the compound, this method directly gives the molecular formula. The formula is $\mathbf{C_2H_3NO_5}$.

Step 3: Try to reduce the molecular formula to a simpler whole number ratio to determine the empirical formula. The formula is already in its simplest whole number ratio. The molecular and empirical formulas are the same. The empirical formula is $\mathbf{C_2H_3NO_5}$.

3.45 The formula for rust can be represented by Fe_2O_3. How many moles of Fe are present in 24.6 g of the compound?

$$24.6 \cancel{\text{g}} \text{ Fe}_2\text{O}_3 \times \frac{1 \cancel{\text{mol}} \text{ Fe}_2\text{O}_3}{159.7 \cancel{\text{g}} \text{ Fe}_2\text{O}_3} \times \frac{2 \text{ mol Fe}}{1 \cancel{\text{mol}} \text{ Fe}_2\text{O}_3} = \mathbf{0.308 \text{ mol Fe}}$$

3.47 Calculate the mass in grams of iodine (I_2) that will react completely with 20.4 g of aluminum (Al) to form aluminum iodide (AlI_3).

The balanced equation is: $2\text{Al}(s) + 3\text{I}_2(s) \longrightarrow 2\text{AlI}_3(s)$

Using unit factors, we convert: g of Al $\rightarrow$ mol of Al $\rightarrow$ mol of I_2 $\rightarrow$ g of I_2

$$20.4 \cancel{\text{g}} \text{ Al} \times \frac{1 \cancel{\text{mol}} \text{ Al}}{26.98 \cancel{\text{g}} \text{ Al}} \times \frac{3 \cancel{\text{mol}} \text{ I}_2}{2 \cancel{\text{mol}} \text{ Al}} \times \frac{253.8 \text{ g I}_2}{1 \cancel{\text{mol}} \text{ I}_2} = \mathbf{288 \text{ g I}_2}$$

3.48 Tin(II) fluoride (SnF_2) is often added to toothpaste as an ingredient to prevent tooth decay. What is the mass of F in grams in 24.6 g of the compound?

Strategy: Tin(II) fluoride is composed of Sn and F. The mass due to F is based on its percentage by mass in the compound. How do we calculate mass percent of an element?

Solution: First, we must find the mass % of fluorine in SnF_2. Then, we convert this percentage to a fraction and multiply by the mass of the compound (24.6 g), to find the mass of fluorine in 24.6 g of SnF_2.

The percent by mass of fluorine in tin(II) fluoride, is calculated as follows:

$$\text{mass \% F} = \frac{\text{mass of F in 1 mol SnF}_2}{\text{molar mass of SnF}_2} \times 100\%$$

$$= \frac{2(19.00\text{ g})}{156.7\text{ g}} \times 100\% = 24.25\%\text{ F}$$

Converting this percentage to a fraction, we obtain 24.25/100 = 0.2425.

Next, multiply the fraction by the total mass of the compound.

$$\textbf{? g F in 24.6 g SnF}_\mathbf{2} = (0.2425)(24.6\text{ g}) = \textbf{5.97 g F}$$

Check: As a ball-park estimate, note that the mass percent of F is roughly 25 percent, so that a quarter of the mass should be F. One quarter of approximately 24 g is 6 g, which is close to the answer.

Note: This problem could have been worked in a manner similar to Problem 3.46. You could complete the following conversions:

g of SnF_2 $\rightarrow$ mol of SnF_2 $\rightarrow$ mol of F $\rightarrow$ g of F

3.54 Monosodium glutamate (MSG), a food-flavor enhancer, has been blamed for "Chinese restaurant syndrome," the symptoms of which are headaches and chest pains. MSG has the following composition by mass: 35.51 percent C, 4.77 percent H, 37.85 percent O, 8.29 percent N, and 13.60 percent Na. What is its molecular formula if its molar mass is about 169 g?

METHOD 1:

Step 1: Assume you have exactly 100 g of substance. 100 g is a convenient amount, because all the percentages sum to 100%. In 100 g of MSG there will be 35.51 g C, 4.77 g H, 37.85 g O, 8.29 g N, and 13.60 g Na.

Step 2: Calculate the number of moles of each element in the compound. Remember, an *empirical formula* tells us which elements are present and the simplest whole-number ratio of their atoms. This ratio is also a mole ratio. Let n_C, n_H, n_O, n_N, and n_{Na} be the number of moles of elements present. Use the molar masses of these elements as conversion factors to convert to moles.

$$n_C = 35.51\,\cancel{g}\,C \times \frac{1\text{ mol C}}{12.01\,\cancel{g}\,C} = 2.9567\text{ mol C}$$

$$n_H = 4.77\,\cancel{g}\,H \times \frac{1\text{ mol H}}{1.008\,\cancel{g}\,H} = 4.732\text{ mol H}$$

$$n_O = 37.85\,\cancel{g}\,O \times \frac{1\text{ mol O}}{16.00\,\cancel{g}\,O} = 2.3656\text{ mol O}$$

$$n_N = 8.29\,\cancel{g}\,N \times \frac{1\text{ mol N}}{14.01\,\cancel{g}\,N} = 0.5917\text{ mol N}$$

$$n_{Na} = 13.60\,\cancel{g}\,Na \times \frac{1\text{ mol Na}}{22.99\,\cancel{g}\,Na} = 0.59156\text{ mol Na}$$

Thus, we arrive at the formula $C_{2.9567}H_{4.732}O_{2.3656}N_{0.5917}Na_{0.59156}$, which gives the identity and the ratios of atoms present. However, chemical formulas are written with whole numbers.

Step 3: Try to convert to whole numbers by dividing all the subscripts by the smallest subscript.

$$\textbf{C:}\ \frac{2.9567}{0.59156} = 4.9981 \approx 5 \qquad \textbf{H:}\ \frac{4.732}{0.59156} = 7.999 \approx 8 \qquad \textbf{O:}\ \frac{2.3656}{0.59156} = 3.9989 \approx 4$$

$$\textbf{N:}\ \frac{0.5917}{0.59156} = 1.000 \qquad \textbf{Na:}\ \frac{0.59156}{0.59156} = 1$$

This gives us the empirical formula for MSG, $C_5H_8O_4NNa$.

To determine the molecular formula, remember that the molar mass/empirical mass will be an integer greater than or equal to one.

$$\frac{\text{molar mass}}{\text{empirical molar mass}} \geq 1\ \text{(integer values)}$$

In this case,

$$\frac{\text{molar mass}}{\text{empirical molar mass}} = \frac{169\text{ g}}{169.11\text{ g}} \approx 1$$

Hence, the molecular formula and the empirical formula are the same, $\mathbf{C_5H_8O_4NNa}$. It should come as no surprise that the empirical and molecular formulas are the same since MSG stands for *monosodium*glutamate.

METHOD 2:

Step 1: Multiply the mass % (converted to a decimal) of each element by the molar mass to convert to grams of each element. Then, use the molar mass to convert to moles of each element.

$$n_{\text{C}} = (0.3551) \times (169\cancel{\text{ g}}) \times \frac{1\text{ mol C}}{12.01\cancel{\text{ g}}\text{ C}} = 5.00\text{ mol C}$$

$$n_{\text{H}} = (0.0477) \times (169\cancel{\text{ g}}) \times \frac{1\text{ mol H}}{1.008\cancel{\text{ g}}\text{ H}} = 8.00\text{ mol H}$$

$$n_{\text{O}} = (0.3785) \times (169\cancel{\text{ g}}) \times \frac{1\text{ mol O}}{16.00\cancel{\text{ g}}\text{ O}} = 4.00\text{ mol O}$$

$$n_{\text{N}} = (0.0829) \times (169\cancel{\text{ g}}) \times \frac{1\text{ mol N}}{14.01\cancel{\text{ g}}\text{ N}} = 1.00\text{ mol N}$$

$$n_{\text{Na}} = (0.1360) \times (169\cancel{\text{ g}}) \times \frac{1\text{ mol Na}}{22.99\cancel{\text{ g}}\text{ Na}} = 1.00\text{ mol Na}$$

Step 2: Since we used the molar mass to calculate the moles of each element present in the compound, this method directly gives the molecular formula. The formula is $\mathbf{C_5H_8O_4NNa}$.

3.59 Balance the following equations using the method outlined in Section 3.7:

The balanced equations are as follows:

(a) $2C + O_2 \rightarrow 2CO$

(b) $2CO + O_2 \rightarrow 2CO_2$

(c) $H_2 + Br_2 \rightarrow 2HBr$

(d) $2K + 2H_2O \rightarrow 2KOH + H_2$

(e) $2Mg + O_2 \rightarrow 2MgO$

(f) $2O_3 \rightarrow 3O_2$

(g) $2H_2O_2 \rightarrow 2H_2O + O_2$

(h) $N_2 + 3H_2 \rightarrow 2NH_3$

(i) $Zn + 2AgCl \rightarrow ZnCl_2 + 2Ag$

(j) $S_8 + 8O_2 \rightarrow 8SO_2$

(k) $2NaOH + H_2SO_4 \rightarrow Na_2SO_4 + 2H_2O$

(l) $Cl_2 + 2NaI \rightarrow 2NaCl + I_2$

(m) $3KOH + H_3PO_4 \rightarrow K_3PO_4 + 3H_2O$

(n) $CH_4 + 4Br_2 \rightarrow CBr_4 + 4HBr$

3.63 Which of the following equations best represents the reaction shown in the diagram?

(Check page 112 on change book)

On the reactants side there are 8 A atoms and 4 B atoms. On the products side, there are 4 C atoms and 4 D atoms. Writing an equation,

$$8A + 4B \rightarrow 4C + 4D$$

Chemical equations are typically written with the smallest set of whole number coefficients. Dividing the equation by four gives,

$$\mathbf{2A + B \rightarrow C + D}$$

The correct answer is choice **(c)**.

3.64 Which of the following equations best represents the reaction shown in the diagram?

(Check page 112 on change book)

On the reactants side there are 6 A atoms and 4 B atoms. On the products side, there are 4 C atoms and 2 D atoms. Writing an equation,

$$6A + 4B \rightarrow 4C + 2D$$

Chemical equations are typically written with the smallest set of whole number coefficients. Dividing the equation by two gives,

$$\mathbf{3A + 2B \rightarrow 2C + D}$$

The correct answer is choice **(d)**.

3.60 Balance the following equations using the method outlined in Section 3.7:

The balanced equations are as follows:

(a) $2N_2O_5 \rightarrow 2N_2O_4 + O_2$

(b) $2KNO_3 \rightarrow 2KNO_2 + O_2$

(c) $NH_4NO_3 \rightarrow N_2O + 2H_2O$

(d) $NH_4NO_2 \rightarrow N_2 + 2H_2O$

(e) $2NaHCO_3 \rightarrow Na_2CO_3 + H_2O + CO_2$

(f) $P_4O_{10} + 6H_2O \rightarrow 4H_3PO_4$

(g) $2HCl + CaCO_3 \rightarrow CaCl_2 + H_2O + CO_2$

(h) $2Al + 3H_2SO_4 \rightarrow Al_2(SO_4)_3 + 3H_2$

(i) $CO_2 + 2KOH \rightarrow K_2CO_3 + H_2O$

(j) $CH_4 + 2O_2 \rightarrow CO_2 + 2H_2O$

(k) $Be_2C + 4H_2O \rightarrow 2Be(OH)_2 + CH_4$

(l) $3Cu + 8HNO_3 \rightarrow 3Cu(NO_3)_2 + 2NO + 4H_2O$

(m) $S + 6HNO_3 \rightarrow H_2SO_4 + 6NO_2 + 2H_2O$

(n) $2NH_3 + 3CuO \rightarrow 3Cu + N_2 + 3H_2O$

3.83 Nitric oxide (NO) reacts with oxygen gas to form nitrogen dioxide (NO_2), a dark-brown gas: $2NO_{(g)} + O_{2(g)} \rightarrow 2NO_{2\,(g)}$ In one experiment 0.886 mole of NO is mixed with 0.503 mole of O_2. Calculate which of the two reactants is the limiting reagent. Calculate also the number of moles of NO_2 produced.

This is a limiting reagent problem. Let's calculate the moles of NO_2 produced assuming complete reaction for each reactant.

$2NO(g) + O_2(g) \rightarrow 2NO_2(g)$

$$0.886\ \cancel{\text{mol}}\ \text{NO} \times \frac{2\ \text{mol NO}_2}{2\ \cancel{\text{mol}}\ \text{NO}} = 0.886\ \text{mol NO}_2$$

$$0.503\ \cancel{\text{mol}}\ \text{O}_2 \times \frac{2\ \text{mol NO}_2}{1\ \cancel{\text{mol}}\ \text{O}_2} = 1.01\ \text{mol NO}_2$$

NO is the **limiting reagent**; it limits the amount of product produced. The amount of product produced is **0.886 mole NO_2**.

3.85 Propane (C_3H_8) is a component of natural gas and is used in domestic cooking and heating. (a) Balance the following equation representing the combustion of propane in air: $C_3H_8 + O_2 \rightarrow CO_2 + H_2O$ (b) How many grams of carbon dioxide can be produced by burning 3.65 moles of propane? Assume that oxygen is the excess reagent in this reaction.

(a) The balanced equation is: $C_3H_8(g) + 5O_2(g) \longrightarrow 3CO_2(g) + 4H_2O(l)$

(b) The balanced equation shows a mole ratio of 3 moles CO_2 : 1 mole C_3H_8. The mass of CO_2 produced is:

$$3.65\ \cancel{\text{mol}}\ \text{C}_3\text{H}_8 \times \frac{3\ \cancel{\text{mol}}\ \text{CO}_2}{1\ \cancel{\text{mol}}\ \text{C}_3\text{H}_8} \times \frac{44.01\ \text{g CO}_2}{1\ \cancel{\text{mol}}\ \text{CO}_2} = \mathbf{482\ g\ CO_2}$$

3.86 Consider the reaction: MnO_2 + 4HCl -> $MnCl_2$ + Cl_2 + $2H_2O$
If 0.86 mole of MnO_2 and 48.2 g of HCl react, which reagent will be used up first? How many grams of Cl_2 will be produced?

This is a limiting reagent problem. Let's calculate the moles of Cl_2 produced assuming complete reaction for each reactant.

$$0.86\ \cancel{\text{mol}}\ \text{MnO}_2 \times \frac{1\ \text{mol Cl}_2}{1\ \cancel{\text{mol}}\ \text{MnO}_2} = 0.86\ \text{mol Cl}_2$$

$$48.2\ \cancel{\text{g}}\ \text{HCl} \times \frac{1\ \cancel{\text{mol}}\ \text{HCl}}{36.458\ \cancel{\text{g}}\ \text{HCl}} \times \frac{1\ \text{mol Cl}_2}{4\ \cancel{\text{mol}}\ \text{HCl}} = 0.3305\ \text{mol Cl}_2$$

HCl is the limiting reagent; it limits the amount of product produced. It will be used up first. The amount of product produced is 0.3305 mole Cl_2. Let's convert this to grams.

$$\mathbf{?\ g\ Cl_2} = 0.3305\ \cancel{\text{mol}}\ \text{Cl}_2 \times \frac{70.90\ \text{g Cl}_2}{1\ \cancel{\text{mol}}\ \text{Cl}_2} = \mathbf{23.4\ g\ Cl_2}$$

3.91 Titanium(IV) oxide (TiO_2) is a white substance produced by the action of sulfuric acid on the mineral ilmenite ($FeTiO_3$):
$FeTiO_3$ + H_2SO_4 -> TiO_2 + $FeSO_4$ + H_2O
Its opaque and nontoxic properties make it suitable as a pigment in plastics and paints. In one process, 8.00 3 10 3 kg of $FeTiO_3$ yielded 3.67 3 10 3 kg of TiO_2. What is the percent yield of the reaction?

The balanced equation shows a mole ratio of 1 mole TiO_2 : 1 mole $FeTiO_3$. The molar mass of $FeTiO_3$ is 151.73 g/mol, and the molar mass of TiO_2 is 79.88 g/mol. The theoretical yield of TiO_2 is:

$$8.00 \times 10^6\ \cancel{\text{g}}\ \text{FeTiO}_3 \times \frac{1\ \cancel{\text{mol}}\ \text{FeTiO}_3}{151.73\ \cancel{\text{g}}\ \text{FeTiO}_3} \times \frac{1\ \cancel{\text{mol}}\ \text{TiO}_2}{1\ \cancel{\text{mol}}\ \text{FeTiO}_3} \times \frac{79.88\ \cancel{\text{g}}\ \text{TiO}_2}{1\ \cancel{\text{mol}}\ \text{TiO}_2} \times \frac{1\ \text{kg}}{1000\ \cancel{\text{g}}}$$
$$= \mathbf{4.21 \times 10^3\ kg\ TiO_2}$$

The actual yield is given in the problem (3.67×10^3 kg TiO_2).

$$\%\ \text{yield} = \frac{\text{actual yield}}{\text{theoretical yield}} \times 100\% = \frac{3.67 \times 10^3\ \text{kg}}{4.21 \times 10^3\ \text{kg}} \times 100\% = \mathbf{87.2\%}$$

f