

MATTER & MINERALS

CHEMISTRY HOMEWORK - FALL - WEEK 4

Chapter 3

(59) (a) CH_2O molar mass = $(12.01 + 2(1.008) + 16.00) \text{ g/mol}$
 $= 30.026 \text{ g/mol}$

mass of C in 1 mol of $\text{CH}_2\text{O} = 12.01 \text{ g/mol}$

mass % of C = $\frac{12.01 \text{ g/mol} \times 100\%}{30.026 \text{ g/mol}} = \underline{\underline{39.99\%}}$

mass of H in 1 mol of $\text{CH}_2\text{O} = 2(1.008) \text{ g/mol}$
 $= 2.016 \text{ g/mol}$

mass % of H = $\frac{2.016 \text{ g/mol} \times 100\%}{30.026 \text{ g/mol}} = \underline{\underline{6.71\%}}$

mass of O in 1 mol of $\text{CH}_2\text{O} = 16.00 \text{ g/mol}$

mass % of O = $\frac{16.00 \text{ g/mol} \times 100\%}{30.026 \text{ g/mol}}$
 $= \underline{\underline{53.29\%}}$

(b) molar mass of $\text{C}_6\text{H}_{12}\text{O}_6 = [6(12.01) + 12(1.008) + 6(16.00)] \text{ g/mol}$
 $= 180.156 \text{ g/mol}$

mass of C in 1 mol of glucose = $6 \times 12.01 \text{ g/mol} = 72.06 \text{ g/mol}$
mass % of C = $\frac{72.06 \text{ g/mol}}{180.156 \text{ g/mol}} \times 100\%$

2/

$$= \underline{\underline{39.99\%}}$$

$$\begin{aligned} \text{mass of H in 1 mol of glucose} &= (12 \times 1.008) \text{ g/mol} \\ &= 12.096 \text{ g/mol} \end{aligned}$$

$$\begin{aligned} \text{mass \% of H} &= \frac{12.096 \text{ g/mol}}{180.156 \text{ g/mol}} \times 100\% \\ &= \underline{\underline{6.71\%}} \end{aligned}$$

$$\begin{aligned} \text{mass of O in 1 mol of glucose} &= (6 \times 16.00) \text{ g/mol} \\ &= 96.00 \text{ g/mol} \end{aligned}$$

$$\begin{aligned} \text{mass \% of O} &= \frac{96.00 \text{ g/mol}}{180.156 \text{ g/mol}} \times 100\% \\ &= \underline{\underline{53.29\%}} \end{aligned}$$

$$\begin{aligned} \text{(c) molar mass of } \text{HC}_2\text{H}_3\text{O}_2 &= [4(1.008) + 2(12.01) + \\ &\quad 2(16.00)] \text{ g/mol} \\ &= 60.052 \text{ g/mol} \end{aligned}$$

$$\begin{aligned} \text{mass of C in 1 mol of } \text{HC}_2\text{H}_3\text{O}_2 &= 2(12.01) \text{ g/mol} \\ &= 24.02 \text{ g/mol} \end{aligned}$$

$$\begin{aligned} \text{mass \% of C} &= \frac{24.02 \text{ g/mol}}{60.052 \text{ g/mol}} \times 100\% \\ &= \underline{\underline{39.99\%}} \end{aligned}$$

$$\begin{aligned} \text{mass of H in 1 mol of } \text{HC}_2\text{H}_3\text{O}_2 &= 4(1.008) \text{ g/mol} \\ &= 4.032 \text{ g/mol} \end{aligned}$$

$$\begin{aligned} \text{mass \% of H} &= \frac{4.032 \text{ g/mol}}{60.052 \text{ g/mol}} \times 100\% \end{aligned}$$

$$= \underline{\underline{6.71\%}}$$

$$\begin{aligned} \text{mass of O in 1 mol of } \text{HC}_2\text{H}_3\text{O}_2 &= 2(16.00) \text{ g/mol} \\ &= 32.00 \text{ g/mol} \end{aligned}$$

$$\begin{aligned} \text{mass \% of O} &= \frac{32.00 \text{ g/mol} \times 100}{60.052 \text{ g/mol}} \end{aligned}$$

$$= \underline{\underline{53.29\%}}$$

$$(63) \quad \text{Percentage of Ti} = 59.9$$

$$\therefore \text{Percentage of O} = 100 - 59.9 = 40.1$$

<u>Ti</u>	<u>O</u>
59.9 g	40.1 g
$\frac{59.99}{47.88 \text{ g/mol}}$	$\frac{40.19}{16.00 \text{ g/mol}}$
= 1.25 mol	= 2.5 mol

molar ratios

$$\frac{1.25 \text{ mol}}{1.25 \text{ mol}} = 1 \quad \frac{2.5 \text{ mol}}{1.25 \text{ mol}} = 2$$

$$1 : 2$$

$$\therefore \text{Empirical formula} = \underline{\underline{\text{TiO}_2}}$$

(65) Compound ①

$$\begin{aligned} \text{mass of compound} &= 0.6498 \text{ g} \\ \text{mass of residue} = \text{mass of Hg} &= 0.6018 \text{ g} \\ \therefore \text{mass of O}_2 \text{ gas} &= (0.6498 - 0.6018) \text{ g} \\ &= 0.0480 \text{ g} \end{aligned}$$

$$\begin{aligned} \text{moles of Hg} &= \frac{0.6018 \text{ g}}{200.6 \text{ g/mol}} \\ &= 3.000 \times 10^{-3} \text{ mol} \end{aligned}$$

$$\begin{aligned} \text{moles of O}_2 &= \frac{0.0480 \text{ g}}{2(16.00) \text{ g/mol}} \\ &= 1.5 \times 10^{-3} \text{ mol} \end{aligned}$$

$$\begin{aligned} \therefore \text{moles of O} &= 2(1.5 \times 10^{-3}) \text{ mol} \\ &= 3.000 \times 10^{-3} \text{ mol} \end{aligned}$$

$$\begin{array}{ccc} \text{molar ratios} & \frac{\text{Hg}}{3.000 \times 10^{-3} \text{ mol}} & \frac{\text{O}}{3.000 \times 10^{-3} \text{ mol}} \\ & \frac{3.000 \times 10^{-3} \text{ mol}}{3.000 \times 10^{-3} \text{ mol}} & \frac{3.000 \times 10^{-3} \text{ mol}}{3.000 \times 10^{-3} \text{ mol}} \\ & & 1 : 1 \end{array}$$

$$\therefore \text{empirical formula} = \underline{\underline{\text{HgO}}}$$

Compound ②

$$\begin{aligned} \text{mass of compound} &= 0.4172 \text{ g} \\ \text{mass loss} = \text{mass of O}_2 \text{ gas} &= 0.016 \text{ g} \\ \therefore \text{mass of mercury} &= (0.4172 - 0.016) \text{ g} \\ &= 0.4012 \text{ g} \end{aligned}$$

$$\begin{aligned} \text{moles of mercury} &= \frac{0.4012 \text{ g}}{200.6 \text{ g/mol}} \\ &= 2.000 \times 10^{-3} \text{ mol} \end{aligned}$$

$$\begin{aligned} \text{moles of } O_2 &= \frac{0.016 \text{ g}}{2(16.00 \text{ g/mol})} \\ &= 5.000 \times 10^{-4} \text{ mol} \end{aligned}$$

$$\begin{aligned} \text{moles of } O &= 2(5.000 \times 10^{-4}) \text{ mol} \\ &= 1.000 \times 10^{-3} \text{ mol} \end{aligned}$$

	<u>Hg</u>	<u>O</u>
molar ratios	$\frac{2.000 \times 10^{-3} \text{ mol}}{1.000 \times 10^{-3} \text{ mol}}$	$\frac{1.000 \times 10^{-3} \text{ mol}}{1.000 \times 10^{-3} \text{ mol}}$
	= 2	= 1

$$\text{empirical formula} = \underline{\underline{Hg_2O}}$$

(74) mass of compound heated = 10.68 mg

mass of CO_2 produced = 16.01 mg

$$\begin{aligned} \text{moles of } CO_2 \text{ produced} &= 16.01 \text{ mg} \times \left(\frac{\text{g}}{10^3 \text{ mg}} \right) \times \left(\frac{\text{mol}}{44.01 \text{ g}} \right) \\ &= 3.638 \times 10^{-4} \text{ mol} \end{aligned}$$

$$\begin{aligned} \text{moles of } C &= 3.638 \times 10^{-4} \text{ mol } CO_2 \times \frac{1 \text{ mol } C}{1 \text{ mol } CO_2} \\ &= 3.638 \times 10^{-4} \text{ mol } C \end{aligned}$$

6/

$$\text{Mass of H}_2\text{O produced} = 4.37 \text{ mg}$$

$$\begin{aligned} \text{moles of H}_2\text{O} &= 4.37 \text{ mg} \times \frac{\text{g}}{10^3 \text{ mg}} \times \frac{1 \text{ mol}}{18.016 \text{ g}} \\ &= 2.426 \times 10^{-4} \text{ mol} \end{aligned}$$

$$\begin{aligned} \text{mol of H} &= 2.426 \times 10^{-4} \text{ mol H}_2\text{O} \times \frac{2 \text{ mol H}}{1 \text{ mol H}_2\text{O}} \\ &= 4.852 \times 10^{-4} \text{ mol} \end{aligned}$$

$$\begin{aligned} \text{mass of C} &= 3.638 \times 10^{-4} \text{ mol} \times \left(\frac{12.01 \text{ g}}{1 \text{ mol}} \right) \\ &= 4.369 \times 10^{-3} \text{ g} \end{aligned}$$

$$\begin{aligned} \text{mass of H} &= 4.852 \times 10^{-4} \text{ mol} \times \left(\frac{1.008 \text{ g}}{1 \text{ mol H}} \right) \\ &= 4.890 \times 10^{-4} \text{ g} \end{aligned}$$

$$\begin{aligned} \therefore \text{mass of O} & \left. \begin{array}{l} \text{in the compound} \end{array} \right\} = 10.68 \times 10^{-3} \text{ g} - \left[4.369 \times 10^{-3} \text{ g} + 4.890 \times 10^{-4} \text{ g} \right] \\ &= 5.822 \times 10^{-3} \text{ g} \end{aligned}$$

$$\begin{aligned} \text{moles of O} &= 5.822 \times 10^{-3} \text{ g} \times \frac{\text{mol}}{16.00 \text{ g}} \\ &= 3.639 \times 10^{-4} \text{ mol} \end{aligned}$$

7.

	C	H	O
molar ratios	$\frac{3.638 \times 10^{-4} \text{ mol}}{3.638 \times 10^{-4} \text{ mol}}$	$\frac{4.852 \times 10^{-4} \text{ mol}}{3.638 \times 10^{-4} \text{ mol}}$	$\frac{3.639 \times 10^{-4} \text{ mol}}{3.638 \times 10^{-4} \text{ mol}}$
	1	1.33	1
	3	4	3

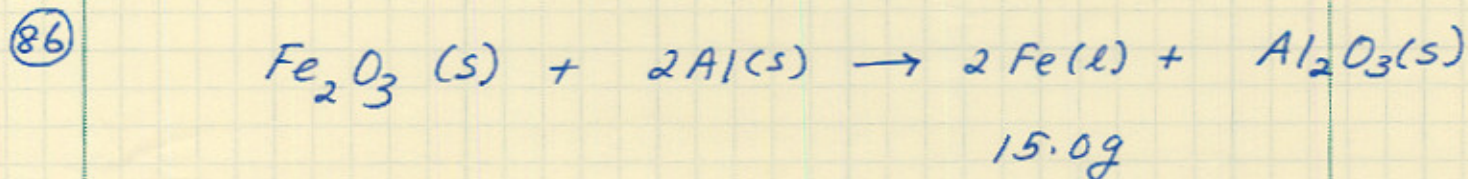
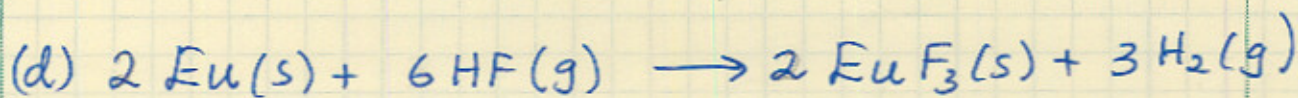
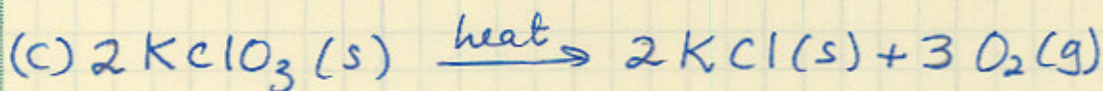
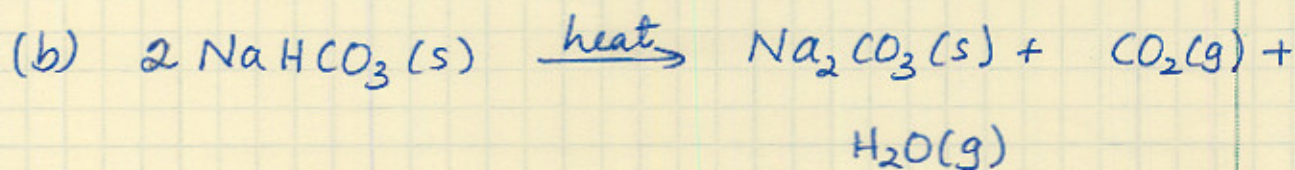
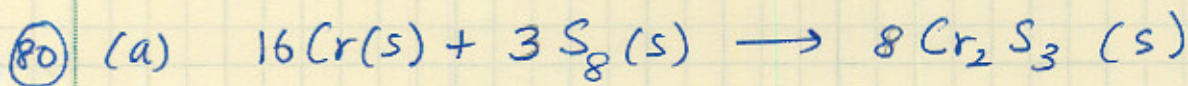
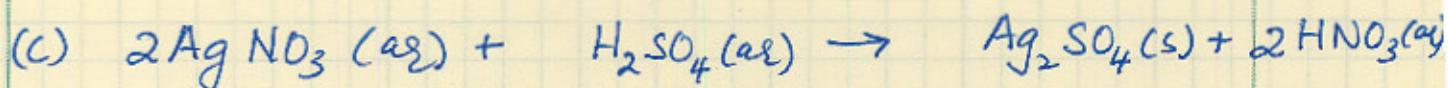
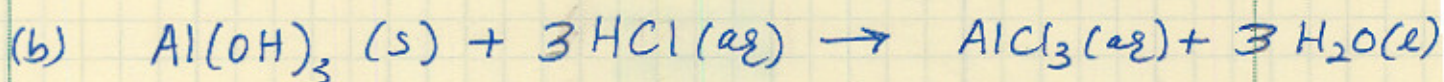
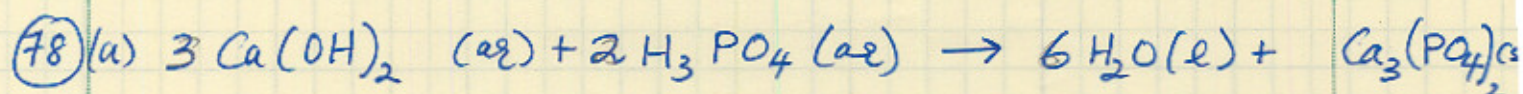
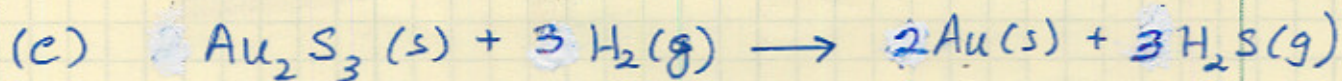
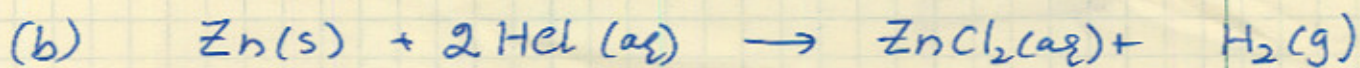
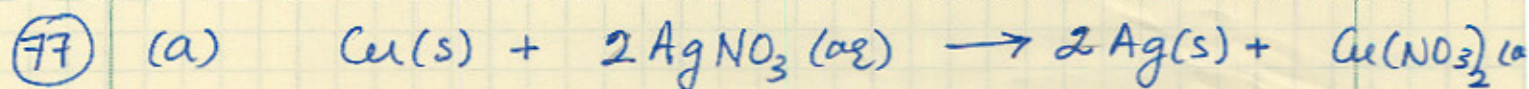
∴ empirical formula = $C_3H_4O_3$

empirical formula mass = $[3(12.01) + 4(1.008) + 3(16.00)]$
 g/mol
 $= 88.062 \text{ g}$

molar mass = 176.1 g/mol

of empirical formula units }
 in one molecular formula } = $\frac{176.1 \text{ g/mol}}{88.062 \text{ g}}$
 $= 2/\text{mol}$

∴ Molecular formula = $C_6H_8O_6$



$$15.0\text{g Fe} \times \frac{1\text{mol Fe}}{55.85\text{g Fe}} = 0.269\text{mol Fe}$$

$$0.269\text{mol Fe} \times \frac{1\text{mol Fe}_2\text{O}_3}{2\text{mol Fe}} = 0.134\text{mol Fe}_2\text{O}_3$$

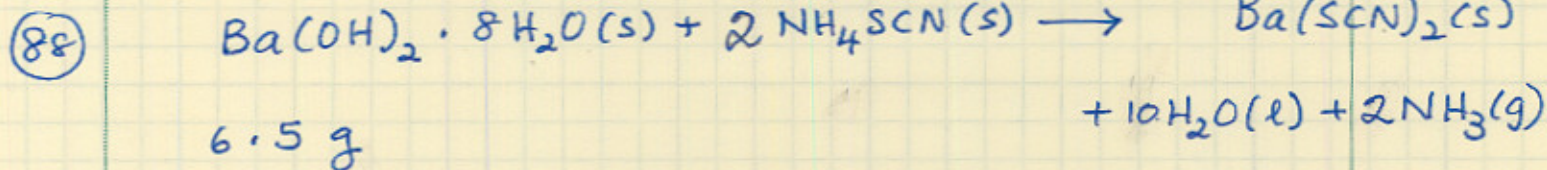
$$0.134 \text{ mol Fe}_2\text{O}_3 \times \frac{[2(55.85) + 3(16.00)] \text{ g}}{1 \text{ mol Fe}_2\text{O}_3}$$

$$= 21.39 \text{ g Fe}_2\text{O}_3 = \underline{\underline{21.4 \text{ g Fe}_2\text{O}_3}}$$

$$0.269 \text{ mol Fe} \times \frac{1 \text{ mol Al}_2\text{O}_3}{2 \text{ mol Fe}} = 0.1345 \text{ mol Al}_2\text{O}_3$$

$$0.1345 \text{ mol Al}_2\text{O}_3 \times \frac{[2(26.98) + 3(16.00)] \text{ g}}{1 \text{ mol Al}_2\text{O}_3} = 13.71 \text{ g}$$

$$= \underline{\underline{13.7 \text{ g Al}_2\text{O}_3}}$$



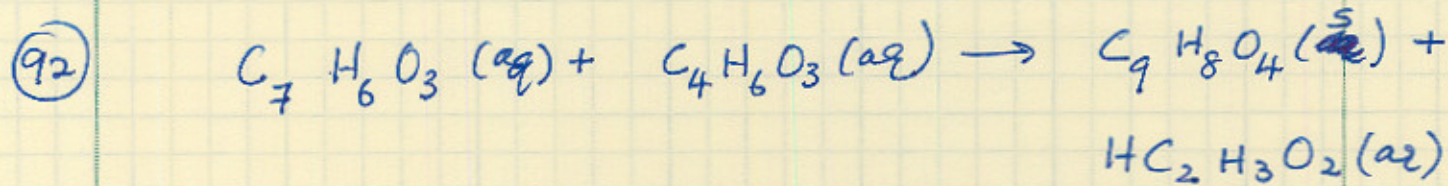
$$\begin{aligned} \text{Molar mass of Ba(OH)}_2 \cdot 8\text{H}_2\text{O} &= [137.3 + 10(16.00) + 18(1.008)] \text{ g/mol} \\ &= 315.444 \text{ g/mol} \end{aligned}$$

$$\begin{aligned} \text{number of moles of Ba(OH)}_2 \cdot 8\text{H}_2\text{O} &= 6.5 \text{ g} \times \frac{1 \text{ mol}}{315.444 \text{ g}} \\ &= 2.061 \times 10^{-2} \text{ mol} \end{aligned}$$

$$2.061 \times 10^{-2} \text{ mol Ba(OH)}_2 \cdot 8\text{H}_2\text{O} \times \frac{2 \text{ mol NH}_4\text{SCN}}{1 \text{ mol Ba(OH)}_2 \cdot 8\text{H}_2\text{O}} = 4.121 \times 10^{-2} \text{ mol NH}_4\text{SCN}$$

$$\begin{aligned} \text{Molar mass of } \text{NH}_4\text{SCN} &= [2(14.01) + 4(1.008) + 32.07 \\ &\quad + 12.01] \text{ g/mol} \\ &= 76.132 \text{ g/mol} \end{aligned}$$

$$\begin{aligned} 4.121 \times 10^{-2} \text{ mol } \text{NH}_4\text{SCN} \times \frac{76.132 \text{ g}}{\text{mol}} &= 3.137 \text{ g} \\ &= \underline{\underline{3.1 \text{ g}}} \end{aligned}$$



$$\begin{aligned} \text{(a) Molar mass of salicylic acid} &= [7(12.01) + 6(1.008) + \\ &\quad 3(16.00)] \text{ g/mol} \\ &= 138.118 \text{ g/mol} \end{aligned}$$

$$\begin{aligned} \# \text{ of moles of salicylic acid} &= 1.00 \times 10^2 \text{ g} \times \left(\frac{\text{mol}}{138.118 \text{ g}} \right) \\ &= 0.724 \text{ mol} \end{aligned}$$

$$0.724 \text{ mol salicylic acid} \times \left(\frac{1 \text{ mol acetic anhydride}}{1 \text{ mol salicylic acid}} \right) = 0.724 \text{ mol acetic anhydride}$$

$$\begin{aligned} \text{Molar mass of acetic anhydride} &= [4(12.01) + 6(1.008) + 3(16.00)] \text{ g/mol} \\ &= 102.088 \text{ g/mol} \end{aligned}$$

$$\begin{aligned} \text{Mass of acetic anhydride} &= 0.724 \text{ mol} \times \left(\frac{102.088 \text{ g}}{\text{mol}} \right) \\ &= \underline{\underline{73.9 \text{ g}}} \end{aligned}$$

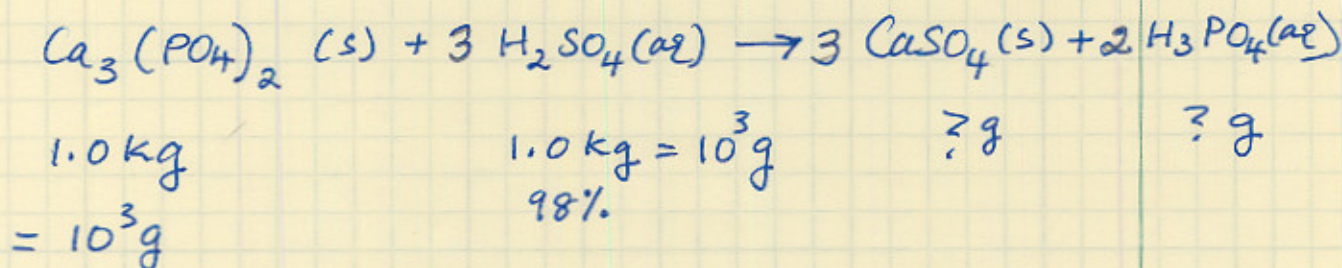
$$(b) \quad 0.724 \text{ mol salicylic acid} \times \left(\frac{1 \text{ mol aspirin}}{1 \text{ mol salicylic acid}} \right)$$

$$= 0.724 \text{ mol aspirin}$$

$$\begin{aligned} \text{Molar mass of aspirin} &= [9(12.01) + 8(1.008) + 4(16.00)] \text{ g/mol} \\ &= 180.154 \text{ g/mol} \end{aligned}$$

$$\begin{aligned} 0.724 \text{ mol aspirin} \times \left(\frac{180.154}{1 \text{ mol}} \right) &= 130.43 \text{ g} \\ &= \underline{\underline{130 \text{ g aspirin}}} \end{aligned}$$

(97)



$$\begin{aligned} \text{Molar mass of Ca}_3(\text{PO}_4)_2 &= [3(40.08) + 2(30.97) + 8(16.00)] \text{ g/mol} \\ &= 310.18 \text{ g/mol} \end{aligned}$$

$$\begin{aligned} \text{Molar mass of H}_2\text{SO}_4 &= [2(1.008) + 32.07 + 4(16.00)] \text{ g/mol} \\ &= 98.086 \text{ g/mol} \end{aligned}$$

12/

$$\begin{aligned} \# \text{ of moles of } \text{Ca}_3(\text{PO}_4)_2 &= 10^3 \text{ g} \times \frac{\text{mol}}{310.18 \text{ g}} \\ &= 3.223 \text{ mol} \end{aligned}$$

$$\# \text{ of grams of } \text{H}_2\text{SO}_4 = 10^3 \text{ g} \times \frac{98}{100} = 980 \text{ g}$$

$$\# \text{ of moles of } \text{H}_2\text{SO}_4 = 980 \text{ g} \times \frac{\text{mol}}{98.086 \text{ g}} = 9.99 \text{ mol}$$

$$\begin{aligned} \# \text{ of moles of } \text{H}_2\text{SO}_4 \text{ required} \\ \text{to react with } 3.223 \text{ mol} \\ \text{of } \text{Ca}_3(\text{PO}_4)_2 \end{aligned} \left. \vphantom{\begin{aligned} \# \text{ of moles of } \text{H}_2\text{SO}_4 \text{ required} \\ \text{to react with } 3.223 \text{ mol} \\ \text{of } \text{Ca}_3(\text{PO}_4)_2 \end{aligned}} \right\} = 3.223 \text{ mol } \text{Ca}_3(\text{PO}_4)_2 \times \left(\frac{3 \text{ mol } \text{H}_2\text{SO}_4}{1 \text{ mol } \text{Ca}_3\text{PO}_4} \right) \\ = 9.672 \text{ mol } \text{H}_2\text{SO}_4$$

Since there are 9.99 mol H_2SO_4 , there is excess H_2SO_4 . Therefore the limiting reagent is $\text{Ca}_3(\text{PO}_4)_2$.

$$3.223 \text{ mol } \text{Ca}_3(\text{PO}_4)_2 \times \left(\frac{3 \text{ mol } \text{CaSO}_4}{1 \text{ mol } \text{Ca}_3(\text{PO}_4)_2} \right) = 9.669 \text{ mol } \text{CaSO}_4$$

$$3.223 \text{ mol } \text{Ca}_3(\text{PO}_4)_2 \times \left(\frac{2 \text{ mol } \text{H}_3\text{PO}_4}{1 \text{ mol } \text{Ca}_3(\text{PO}_4)_2} \right) = 6.446 \text{ mol } \text{H}_3\text{PO}_4$$

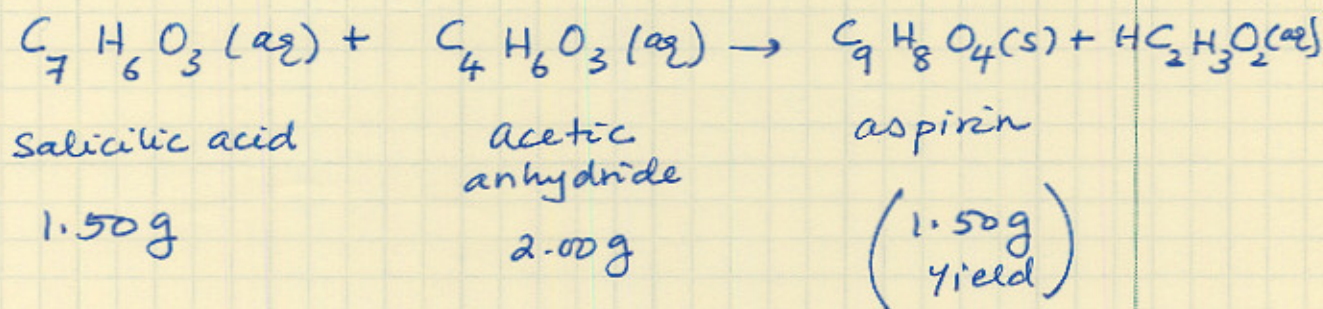
$$\begin{aligned} \text{Molar mass of } \text{CaSO}_4 &= [40.08 + 32.07 + 4(16.00)] \text{ g/mol} \\ &= 136.15 \text{ g/mol} \end{aligned}$$

$$\begin{aligned} \text{Molar mass of } \text{H}_3\text{PO}_4 &= [3(1.008) + 30.97 + 4(16.00)] \text{ g/mol} \\ &= 97.994 \text{ g/mol} \end{aligned}$$

$$\begin{aligned} 9.669 \text{ mol CaSO}_4 \times \left(\frac{136.15 \text{ g}}{1 \text{ mol CaSO}_4} \right) &= 1316.43 \text{ g CaSO}_4 \\ &= \underline{\underline{1.3 \text{ kg CaSO}_4}} \end{aligned}$$

$$\begin{aligned} 6.446 \text{ mol H}_3\text{PO}_4 \times \left(\frac{97.994 \text{ g}}{1 \text{ mol H}_3\text{PO}_4} \right) &= 631.67 \text{ g H}_3\text{PO}_4 \\ &= \underline{\underline{6.3 \times 10^2 \text{ g H}_3\text{PO}_4}} \end{aligned}$$

(100)



$$\begin{aligned} \# \text{ of moles of salicylic acid} &= 1.50 \text{ g} \times \left(\frac{\text{mol}}{138.118 \text{ g}} \right) \\ &= 1.086 \times 10^{-2} \text{ mol salicylic acid} \end{aligned}$$

$$\begin{aligned} \# \text{ of moles of acetic anhydride} &= 2.00 \text{ g} \times \left(\frac{\text{mol}}{102.088 \text{ g}} \right) \\ &= 1.959 \times 10^{-2} \text{ mol acetic anhydride} \end{aligned}$$

$$\begin{aligned} \left. \begin{array}{l} \# \text{ of moles of salicylic acid} \\ \text{needed to react with } 1.959 \times 10^{-2} \\ \text{mol acetic anhydride} \end{array} \right\} &= 1.959 \times 10^{-2} \text{ mol} \times \left(\frac{1 \text{ mol salicylic acid}}{1 \text{ mol acetic anhydride}} \right) \\ &= 1.959 \times 10^{-2} \text{ mol salicylic acid} \end{aligned}$$

There is only 1.086×10^{-2} mol salicylic acid.
Therefore salicylic acid is the limiting reagent.

$$\left. \begin{array}{l} \text{Amount of aspirin that can} \\ \text{be made from } 1.086 \times 10^{-2} \text{ mol} \\ \text{of salicylic acid} \end{array} \right\} = 1.086 \times 10^{-2} \text{ mol salicylic acid} \times \left(\frac{1 \text{ mol aspirin}}{1 \text{ mol salicylic acid}} \right)$$

$$= 1.086 \times 10^{-2} \text{ mol aspirin}$$

$$= 1.086 \times 10^{-2} \text{ mol aspirin} \times \left(\frac{180.15 \text{ g}}{1 \text{ mol aspirin}} \right)$$

$$\text{theoretical yield} = \underline{1.96 \text{ g aspirin}}$$

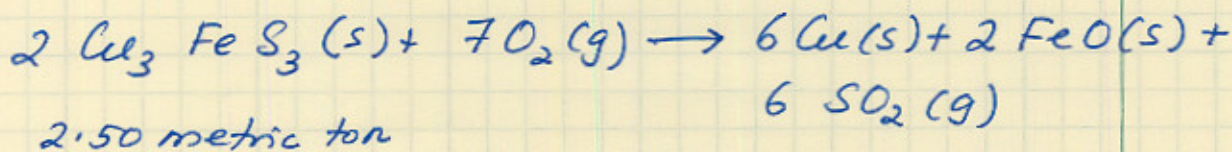
$$\text{Percentage yields} = \frac{\text{actual yield}}{\text{theoretical yield}} \times 100\%$$

$$= \frac{1.50 \text{ g}}{1.96 \text{ g}} \times 100\%$$

$$= 76.67\%$$

$$= \underline{\underline{76.7\%}}$$

(101)



2.50 metric ton

$$= 2.50 \times 1000 \text{ kg}$$

$$= 2.50 \times 10^6 \text{ g}$$

$$\text{Molar mass of } \left. \begin{array}{l} \text{Cu}_3 \text{ Fe S}_3 \end{array} \right\} = [3(63.55) + 55.85 + 3(32.07)] \text{ g/mol}$$

$$= 342.71 \text{ g/mol}$$

$$2.50 \times 10^6 \text{ g Cu}_3\text{FeS}_3 \times \left(\frac{\text{mol}}{342.71 \text{ g}} \right) = 7294.797 \text{ mol}$$

$$\left(\begin{array}{l} 7294.797 \text{ mol} \\ \text{Cu}_3\text{FeS}_3 \end{array} \right) \times \left(\frac{6 \text{ mol Cu}}{2 \text{ mol Cu}_3\text{FeS}_3} \right) = \begin{array}{l} 21884.39 \text{ mol Cu} \\ 10942.196 \text{ mol Cu} \end{array}$$

$$\text{theoretical yield} = 2.18842 \times 10^4 \text{ mol Cu}$$

$$= 2.1884 \times 10^4 \text{ mol} \times \left(\frac{63.55 \text{ g Cu}}{1 \text{ mol Cu}} \right)$$

$$= 1.39085 \times 10^6 \text{ g Cu}$$

= theoretical yield.

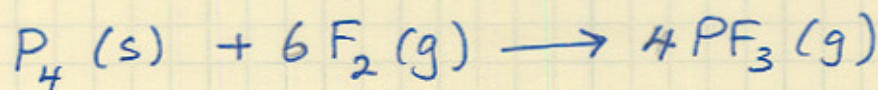
Actual yield = 86.3% of theoretical yield

$$= \frac{86.3}{100} \times 1.39085 \times 10^6 \text{ g Cu}$$

$$= 1.2002 \times 10^6 \text{ g Cu}$$

$$= 1.20 \times 10^6 \text{ g Cu} = \underline{\underline{1.20 \text{ metric tons of Cu}}}$$

(102)



120. g

78.1%
yield.

Actual yield = 120. g PF₃

Percentage yield = 78.1%.

$$= 153.6 \text{ g PF}_3$$

$$\begin{aligned} \text{Molar mass of PF}_3 &= [30.97 + 3(19.00)] \text{ g/mol} \\ &= 87.97 \text{ g/mol} \end{aligned}$$

$$\begin{aligned} \# \text{ of moles of PF}_3 &= 153.6 \text{ g} \times \left(\frac{\text{mol}}{87.97 \text{ g}} \right) \\ &= 1.746 \text{ mol PF}_3. \end{aligned}$$

$$1.746 \text{ mol PF}_3 \times \frac{6 \text{ mol F}_2}{4 \text{ mol PF}_3} = 2.619 \text{ mol F}_2$$

$$\begin{aligned} \text{molar mass of F}_2 &= 2(19.00) \text{ g/mol} \\ &= 38.00 \text{ g/mol} \end{aligned}$$

$$\begin{aligned} 2.619 \text{ mol F}_2 \times \left(\frac{38.00 \text{ g F}_2}{1 \text{ mol F}_2} \right) &= 99.525 \text{ g F}_2 \\ &= \underline{\underline{99.5 \text{ g F}_2}} \end{aligned}$$