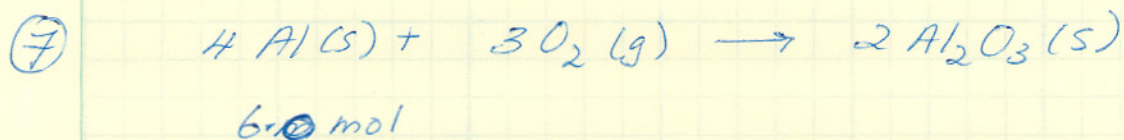
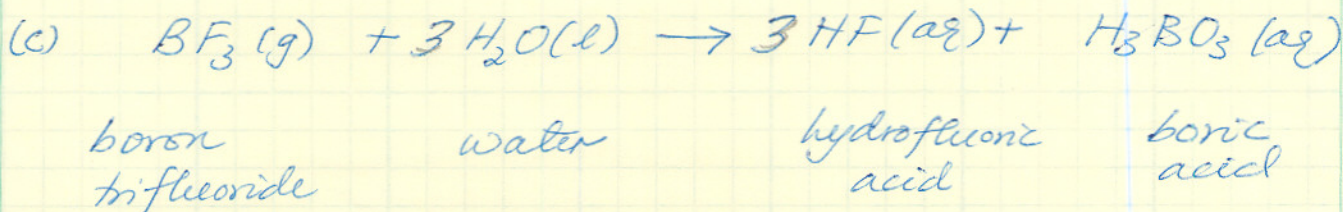
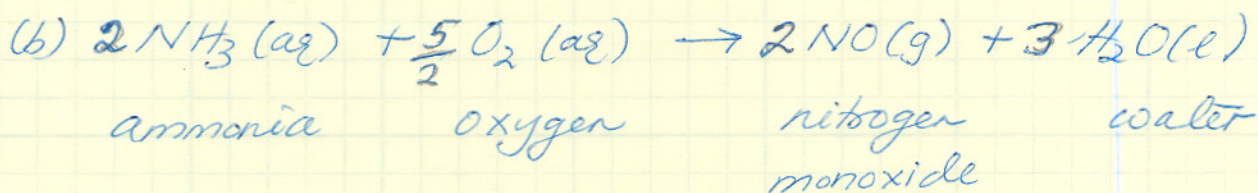
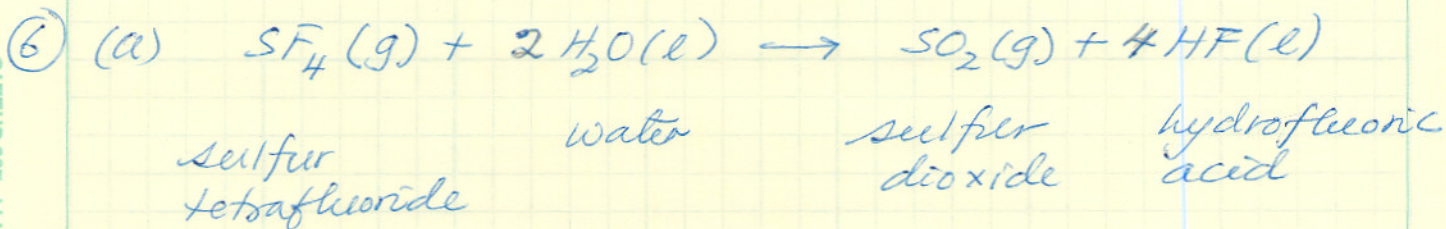


INTRODUCTION TO NATURAL SCIENCE  
CHEMISTRY HOMEWORK - FALL 2006 - WEEK 5

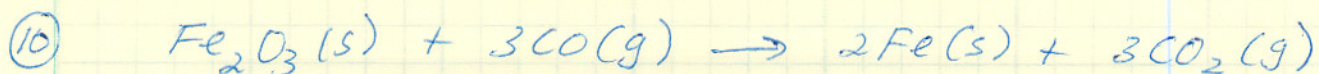
Chapter 4



$$\text{Moles of } O_2 \text{ needed} = 6.0 \text{ mol Al} \times \frac{3 \text{ mol } O_2}{4 \text{ mol Al}} = \underline{\underline{4.5 \text{ mol}}}$$

$$\text{moles of } Al_2O_3 \text{ produced} = 6.0 \text{ mol Al} \times \frac{2 \text{ mol } Al_2O_3}{4 \text{ mol Al}} = \underline{\underline{3.0 \text{ mol}}}$$

$$3.0 \text{ mol } Al_2O_3 \times \frac{102 \text{ g}}{1 \text{ mol } Al_2O_3} = 306 \text{ g} = \underline{\underline{310 \text{ g}}}$$



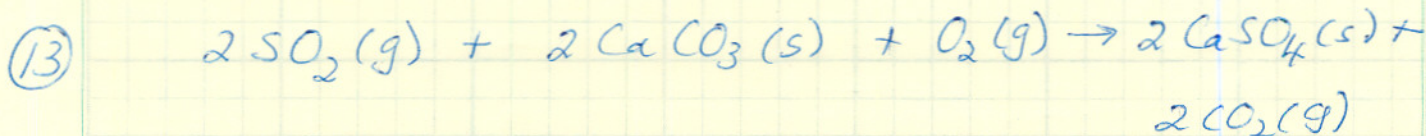
$$(a) \quad 454 \text{ g Fe}_2\text{O}_3 \times \frac{1 \text{ mol}}{160 \text{ g Fe}_2\text{O}_3} = 2.8375 \text{ mol}$$

$$\begin{aligned} \text{moles of Fe produced} &= 2.8375 \text{ mol Fe}_2\text{O}_3 \times \frac{2 \text{ mol Fe}}{1 \text{ mol Fe}_2\text{O}_3} \\ &= 5.675 \text{ mol Fe} \end{aligned}$$

$$\begin{aligned} &= 5.675 \text{ mol Fe} \times \frac{55.85 \text{ g}}{1 \text{ mol Fe}} \\ &= 316.95 \text{ g} \\ &= \underline{\underline{317 \text{ g}}} \end{aligned}$$

$$(b) \quad \text{moles of CO required} = 2.8375 \text{ mol Fe}_2\text{O}_3 \times \frac{3 \text{ mol CO}}{1 \text{ mol Fe}_2\text{O}_3} = 8.5125 \text{ mol}$$

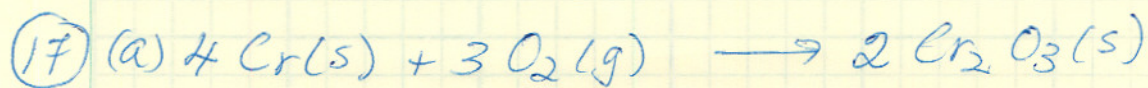
$$\begin{aligned} 8.5125 \text{ mol CO} \times \frac{28 \text{ g}}{1 \text{ mol CO}} &= 238.35 \text{ g CO} \\ &= \underline{\underline{238 \text{ g CO}}} \end{aligned}$$



$$(a) \quad 155 \text{ g SO}_2 \times \frac{1 \text{ mol}}{64 \text{ g}} = \frac{155}{64} \text{ mol SO}_2 = 2.4219 \text{ mol SO}_2$$

$$\begin{aligned}
 \text{Moles of CaCO}_3 \text{ required for the reaction} \} &= \frac{2.4219}{\text{SO}_2} \text{ mol} \times \boxed{\frac{2 \text{ mol CaCO}_3}{2 \text{ mol SO}_2}} \\
 &= 2.4219 \text{ mol CaCO}_3 \\
 &= 2.4219 \text{ mol} \times \boxed{\frac{100 \text{ g}}{1 \text{ mol CaCO}_3}} \\
 &= 242.19 \text{ g CaCO}_3 \\
 &= \underline{\underline{242 \text{ g}}}
 \end{aligned}$$

$$\begin{aligned}
 \text{(b) Moles of CaSO}_4 \text{ formed} &= 2.4219 \text{ mol} \times \boxed{\frac{2 \text{ mol CaSO}_4}{2 \text{ mol SO}_2}} \\
 &= 2.4219 \text{ mol CaSO}_4 \\
 &= 2.4219 \text{ mol} \times \boxed{\frac{136 \text{ g CaSO}_4}{1 \text{ mol}}} \\
 &= 329.38 \text{ g} = \underline{\underline{329 \text{ g}}}
 \end{aligned}$$



$$\text{(b) } 0.175 \text{ g Cr} \times \boxed{\frac{1 \text{ mol Cr}}{51.996 \text{ g}}} = 3.366 \times 10^{-3} \text{ mol Cr}$$

$$\begin{aligned}
 \text{moles of Cr}_2\text{O}_3 \text{ produced} \} &= 3.366 \times 10^{-3} \text{ mol Cr} \times \boxed{\frac{2 \text{ mol Cr}_2\text{O}_3}{4 \text{ mol Cr}}} \\
 &= 1.683 \times 10^{-3} \text{ mol Cr}_2\text{O}_3 \\
 &= 1.683 \times 10^{-3} \text{ mol Cr}_2\text{O}_3 \times \boxed{\frac{151.992 \text{ g}}{1 \text{ mol Cr}_2\text{O}_3}}
 \end{aligned}$$

$$= 2.1558 \text{ g} \times 10^{-1} \text{ Cr}_2\text{O}_3$$

$$= \underline{\underline{215.58 \text{ g} \quad 0.256 \text{ g}}}$$

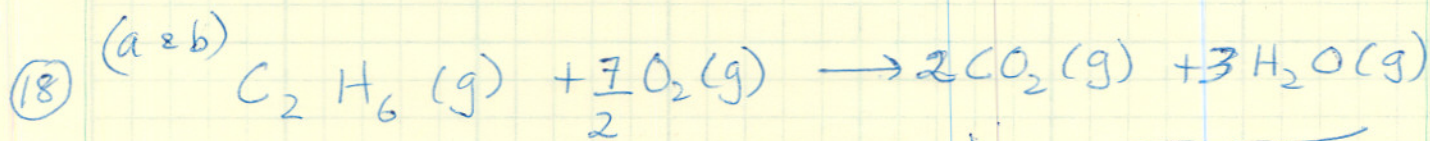
$$(c) \text{ moles of } O_2 \text{ required} = 3.366 \times 10^{-3} \text{ mol Cr} \times \frac{3 \text{ moles } O_2}{4 \text{ mol Cr}}$$

$$= 2.5245 \times 10^{-3} \text{ mol } O_2$$

$$= 2.5245 \times 10^{-3} \text{ mol } O_2 \times \frac{32 \text{ g}}{1 \text{ mol } O_2}$$

$$= 8.078 \times 10^{-2} \text{ g } O_2$$

$$= \underline{\underline{8.08 \times 10^{-2} \text{ g } O_2}}$$



products

$$(e) \quad 13.6 \text{ g } C_2H_6 \times \frac{1 \text{ mol}}{30 \text{ g}} = 0.4533 \text{ mol } C_2H_6$$

$$\text{moles of } O_2 \text{ required} = 0.4533 \text{ mol } C_2H_6 \times \frac{7/2 \text{ mol } O_2}{1 \text{ mol } C_2H_6}$$

$$= 1.5867 \text{ mol } O_2$$

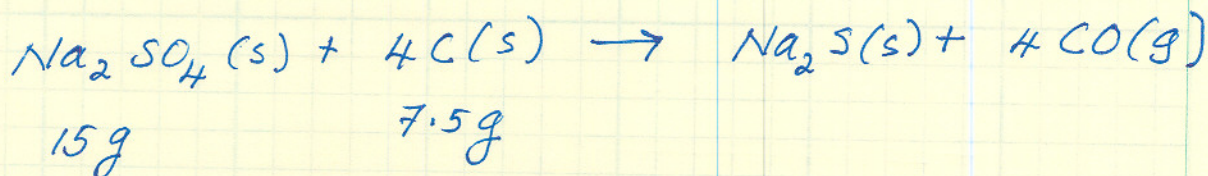
$$= 1.5867 \text{ mol } O_2 \times \frac{32 \text{ g}}{1 \text{ mol } O_2}$$

$$\begin{aligned}
 \text{(d) Total mass of reactants} &= \text{mass of } C_2H_6 + \\
 &\quad \text{mass of } O_2 \\
 &= 13.6g + 50.8g \\
 &= 64.4g
 \end{aligned}$$

Due to law of conservation of mass

$$\begin{aligned}
 \text{total mass of products} &= \text{total mass of reactants} \\
 &= \underline{\underline{64.4g}}
 \end{aligned}$$

19



$$\text{moles of } Na_2SO_4 = 15g \times \frac{\text{mol}}{142g} = 0.1056 \text{ mol}$$

$$\text{moles of } C = 7.5g \times \frac{\text{mol}}{12g} = 0.625 \text{ mol}$$

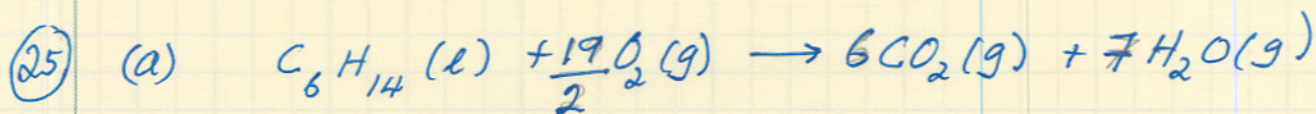
$$\begin{aligned}
 \left. \begin{array}{l} \text{\# of moles of } C \text{ needed} \\ \text{to react with } 0.1261 \text{ mol of} \\ Na_2SO_4 \end{array} \right\} &= \frac{0.1056}{\cancel{0.1261}} \text{ mol } Na_2SO_4 \times \frac{4 \text{ mol } C}{1 \text{ mol } Na_2SO_4} \\
 &= \frac{0.4225}{\cancel{0.5044}} \text{ mol } C
 \end{aligned}$$

We have excess C.  $Na_2SO_4$  is the limiting reactant.

$$\begin{aligned}
 \text{mol. of } Na_2S \text{ produced} &= \frac{0.1056}{\cancel{0.1261}} \text{ mol } Na_2SO_4 \times \frac{1 \text{ mol } Na_2S}{1 \text{ mol } Na_2SO_4} \\
 &= \frac{0.1056}{\cancel{0.1261}} \text{ mol } Na_2S
 \end{aligned}$$

$$\text{molar mass of Na}_2\text{S} = (2 \times 23 + 32) \text{ g/mol} = 78 \text{ g/mol}$$

$$\begin{aligned} 0.1056 \\ \cancel{0.125} \text{ mol Na}_2\text{S} \times \boxed{\frac{78 \text{ g}}{1 \text{ mol}}} &= \frac{8.236}{\cancel{9.8333}} \text{ g Na}_2\text{S} \\ &= \underline{\underline{8.2 \text{ g Na}_2\text{S}}} \end{aligned}$$



$$\begin{array}{r} 12 \\ 7 \\ \hline 19 \end{array}$$

(b)  $215 \text{ g} \quad 215 \text{ g}$

$$\begin{aligned} \text{moles of C}_6\text{H}_{14} &= 215 \text{ g} \times \boxed{\frac{\text{mol}}{86 \text{ g}}} = 2.50 \text{ mol} \\ \text{moles of O}_2 &= 215 \text{ g} \times \boxed{\frac{\text{mol}}{32 \text{ g}}} = 6.719 \text{ mol} \end{aligned}$$

$$\begin{aligned} \left. \begin{array}{l} \text{mol of O}_2 \text{ needed to react with} \\ 2.50 \text{ mol of C}_6\text{H}_{14} \end{array} \right\} &= 2.50 \text{ mol C}_6\text{H}_{14} \times \boxed{\frac{\left(\frac{19}{2}\right) \text{ mol O}_2}{1 \text{ mol C}_6\text{H}_{14}}} \\ &= 23.75 \text{ mol O}_2 \end{aligned}$$

since we only have 6.719 mol O<sub>2</sub>, O<sub>2</sub> is the limiting reactant.

$$\begin{aligned} \text{moles of CO}_2 \text{ produced} &= 6.719 \text{ mol O}_2 \times \boxed{\frac{6 \text{ mol CO}_2}{\frac{19}{2} \text{ mol O}_2}} \\ &= 4.2436 \text{ mol CO}_2 \end{aligned}$$

$$= 186.718 \text{ g CO}_2$$

$$= \underline{\underline{187 \text{ g CO}_2}}$$

$$\text{mass of H}_2\text{O produced} = 6.719 \text{ mol O}_2 \times \boxed{\frac{7 \text{ mol H}_2\text{O}}{\frac{19 \text{ mol O}_2}{2}}}$$

$$= 4.9508 \text{ mol H}_2\text{O}$$

$$= 4.9508 \text{ mol H}_2\text{O} \times \boxed{\frac{18 \text{ g}}{1 \text{ mol H}_2\text{O}}}$$

$$= 89.115 \text{ g H}_2\text{O} = \underline{\underline{89.1 \text{ g H}_2\text{O}}}$$

(c) hexane is the excess agent.

$$\text{moles of hexane that reacts} = 6.719 \text{ mol O}_2 \times \boxed{\frac{1 \text{ mol C}_6\text{H}_{14}}{\frac{19 \text{ mol O}_2}{2}}}$$

$$= 0.70726 \text{ mol C}_6\text{H}_{14}$$

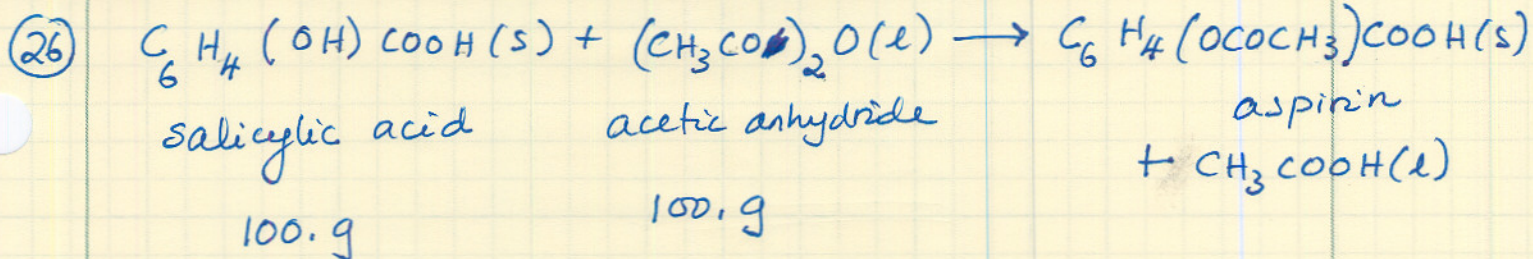
$$= 0.70726 \text{ mol C}_6\text{H}_{14} \times \boxed{\frac{86 \text{ g}}{1 \text{ mol C}_6\text{H}_{14}}}$$

$$= 60.824 \text{ g}$$

$$\text{mass of hexane in excess} = 215 \text{ g} - 60.824 \text{ g}$$

$$= 154.176 \text{ g}$$

$$= \underline{\underline{154 \text{ g}}}$$



molar mass of salicylic acid = 138.095 g/mol

molar mass of acetic anhydride = 102.062 g/mol

$$\text{moles of salicylic acid} = 100. \text{g} \times \frac{\text{mol}}{138.095 \text{g}} = 0.7241 \text{mol}$$

$$\text{moles of acetic anhydride} = 100. \text{g} \times \frac{\text{mol}}{102.062 \text{g}} = 0.9798 \text{mol}$$

$$\left. \begin{array}{l} \text{moles of acetic anhydride needed} \\ \text{to react with } 0.7241 \text{mol salicylic acid} \end{array} \right\} = 0.7241 \text{mol salicylic acid} \times \frac{1 \text{mol acetic anhydride}}{1 \text{mol salicylic acid}}$$

$$= 0.7241 \text{mol acetic anhydride}$$

∴ We have excess of acetic anhydride. Salicylic acid is the limiting reagent.

$$\text{moles of aspirin produced} = 0.7241 \text{mol salicylic acid} \times \frac{1 \text{mol aspirin}}{1 \text{mol salicylic acid}}$$

$$= 0.7241 \text{mol aspirin}$$

molar mass of aspirin = 180.123 g/mol

$$\begin{aligned}
 0.7241 \text{mol aspirin} \times \frac{180.123 \text{g}}{1 \text{mol aspirin}} &= 130.427 \text{g} \\
 &= \underline{\underline{130. \text{g}}}
 \end{aligned}$$



(28)

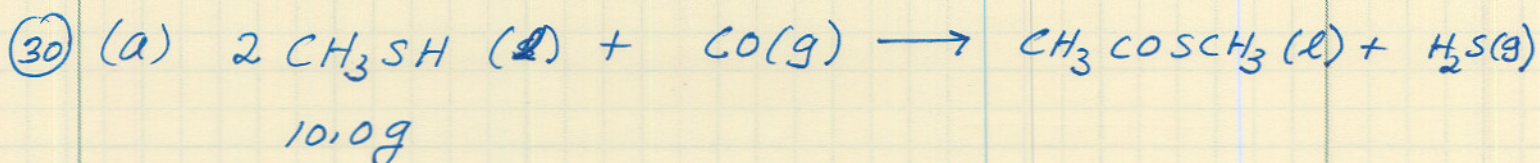
$$\text{Theoretical yield} = 68.0 \text{ g}$$

$$\text{actual yield} = 16.3 \text{ g}$$

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

$$= \frac{16.3 \text{ g}}{68.0 \text{ g}} \times 100\% = 23.97\%$$

$$= \underline{\underline{24.0\%}}$$



$$\text{molar mass of CH}_3\text{SH} = 48.109 \text{ g/mol}$$

$$10.0 \text{ g CH}_3\text{SH} \times \frac{1 \text{ mol}}{48.109 \text{ g}} = 0.2079 \text{ mol CH}_3\text{SH}$$

$$\text{moles of CH}_3\text{COSCH}_3 \text{ produced} = 0.2079 \text{ mol CH}_3\text{SH} \times \frac{1 \text{ mol CH}_3\text{COSCH}_3}{2 \text{ mol CH}_3\text{SH}}$$

$$= 0.10393 \text{ mol CH}_3\text{COSCH}_3$$

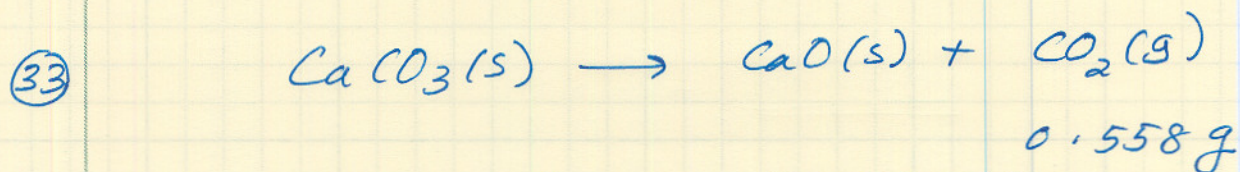
$$\text{molar mass of CH}_3\text{COSCH}_3 = 90.137 \text{ g/mol}$$

$$0.10393 \text{ mol CH}_3\text{COSCH}_3 \times \frac{90.137 \text{ g}}{1 \text{ mol CH}_3\text{COSCH}_3} = 9.36799 \text{ g CH}_3\text{COSCH}_3$$

$$\text{theoretical yield} = \underline{\underline{9.37 \text{ g CH}_3\text{COSCH}_3}}$$

(b)  $\text{actual yield} = 8.65 \text{ g}$

$$\begin{aligned}
 \text{Percentage yield} &= \frac{\text{actual yield}}{\text{theoretical yield}} \times 100\% \\
 &= \frac{8.659}{9.36799} \times 100\% = 92.336\% \\
 &= \underline{\underline{92.3\%}}
 \end{aligned}$$



molar mass of  $\text{CO}_2 = 43.9919 \text{ g/mol}$

moles of  $\text{CO}_2$  recovered =  $0.558 \text{ g} \times \frac{1 \text{ mol}}{43.9919 \text{ g}}$

$$= 1.268 \times 10^{-2} \text{ mol CO}_2$$

moles of  $\text{CaCO}_3$  reacted =  $1.268 \times 10^{-2} \text{ mol CO}_2 \times \frac{1 \text{ mol CaCO}_3}{1 \text{ mol CO}_2}$

$$= 1.268 \times 10^{-2} \text{ mol CaCO}_3$$

Molar mass of  $\text{CaCO}_3 = 100.059 \text{ g/mol}$

$$1.268 \times 10^{-2} \text{ mol CaCO}_3 \times \frac{100.059 \text{ g}}{1 \text{ mol CaCO}_3} = 1.2687 \text{ g CaCO}_3$$

Mass percent of  $\text{CaCO}_3$  in the limestone sample } =  $\frac{1.2687 \text{ g}}{1.506 \text{ g}} \times 100\%$

$$= 84.246\%$$

$$= \underline{\underline{84.2\%}}$$