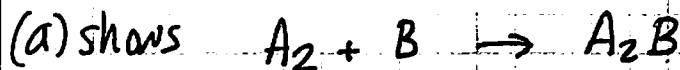
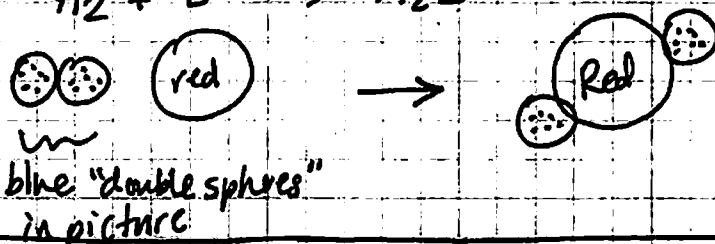


① The reaction shown in the diagram could match (a).

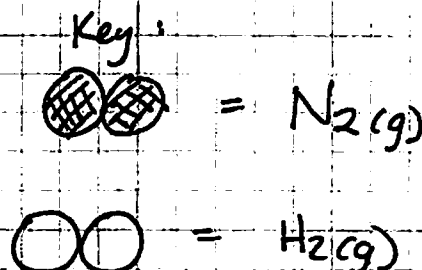
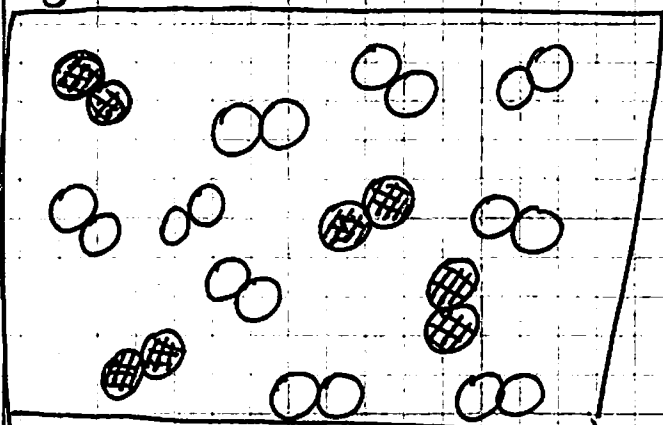


which would look like this



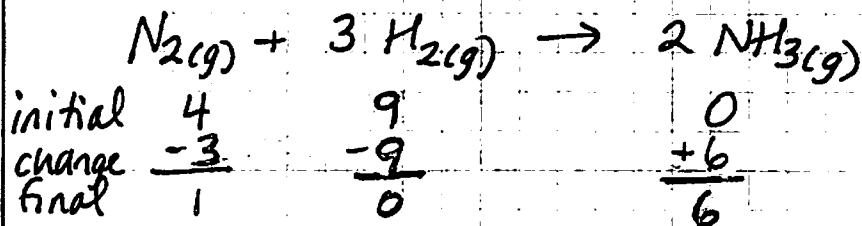
#2 is on the next page. D'oh!

⑦ given picture :



Q: draw box after rxn complete as much as possible, explain.

The balanced rxn to form  $NH_3$  would be :



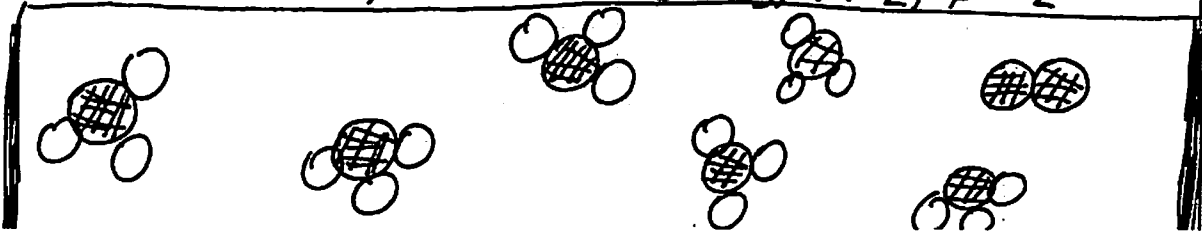
We know  $H_2$  was the limiting reactant

Since the mole ratio is 3 mole  $H_2$  / 1 mole  $N_2$ :

4 molecules  $N_2$  could react with 12 molecules  $H_2$ , but only 9  $H_2$  are present initially.

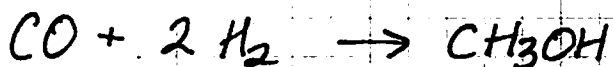
If  $H_2$  is used up, the 9  $H_2$  molecules that react will only need 3  $N_2$  to react with acc to mole ratio. if 3  $N_2$  react, 6  $NH_3$  will form due to 2:1 mole ratio between  $NH_3$  and  $N_2$ .

So, after rxn, will have 6  $NH_3$ , 1  $N_2$ , 0  $H_2$  :



(2) The given picture shows 8 molecules of  $H_2$ .

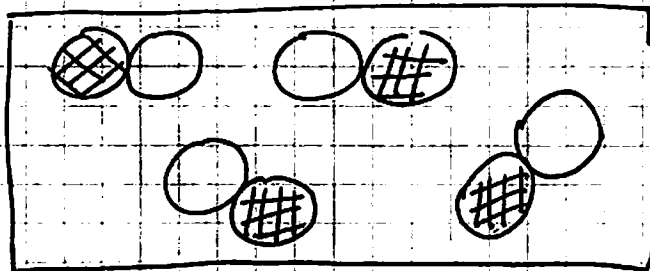
Rxn of CO with  $H_2$  to form  $CH_3OH$  would be:



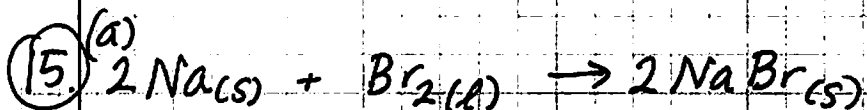
Acc to balanced rxn, we only need 1 molecule CO per 2 molecules  $H_2$ , or 4 molecules CO per 8 molecules  $H_2$ .

For CO molecules..

Not very good scale... need more spacing between molecules.



(#7) is on the previous page...



I determined the formula by balancing ion charges...

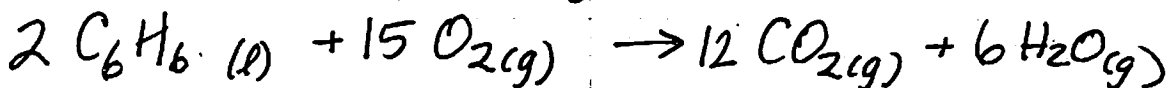
$Na^{+1}$  with  $Br^{-1}$

$NaBr$  must be solid at room temperature because it is ionic (metal ( $Na$ ) bonded w/ nonmetal ( $Br$ )). Ionic compounds have very strong cohesive forces so tend to have very high melting points.

(acc to wikipedia,  $NaBr$  has a melting point of  $747^\circ C$ , well above room temp of  $\approx 20^\circ C$ . but don't use that as your explanation for how you know.)

(b) When a hydrocarbon burns, the other reactant, besides hydrocarbon, is oxygen gas;  $O_2(g)$

Rxn for benzene Combusting:



note: these are covalent so do not have very high mp/bp.

24. (a) Calculate % C in acetylene;  $C_2H_2$

$$\% C = \frac{2(12.011)}{2(12.011) + 2(1.0079)} \times 100 = \boxed{92.258 \% C \text{ by mass}}$$

(b) calculate % H in ascorbic acid (vitamin C);  $HC_6H_7O_6$   
(or  $C_6H_8O_6$ !)

$$\% H = \frac{8(1.0079)}{8(1.0079) + 6(12.011) + 6(15.9994)} \times 100 = \boxed{4.5781 \% H \text{ by mass}}$$

35. (a) convert  $2.50 \times 10^{-3}$  mole ammonium phosphate to grams

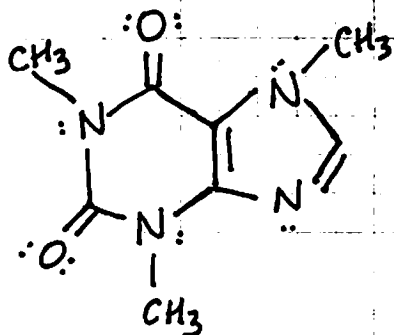
$$(2.50 \times 10^{-3} \text{ mole } (NH_4)_3PO_4) \left( \frac{149.08626g}{1 \text{ mole}} \right) = \boxed{0.373 \text{ grams}}$$

(b) <sup>find</sup> moles chloride ions in 0.2550 grams Aluminum chloride

$$(0.2550 \text{ g } AlCl_3) \left( \frac{1 \text{ mole } AlCl_3}{133.3405g} \right) \left( \frac{3 \text{ moles } Cl^-}{1 \text{ mole } AlCl_3} \right) = \boxed{0.005737 \text{ moles}}$$

(c) convert  $7.70 \times 10^{20}$  caffeine molecules to grams

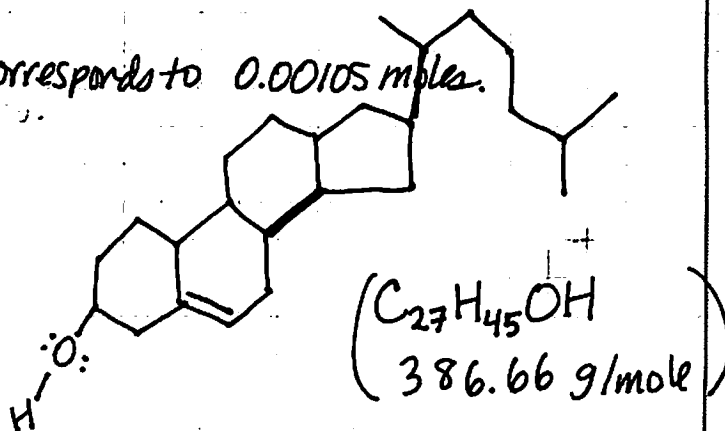
$$(7.70 \times 10^{20} \text{ molecules of } C_8H_{10}N_4O_2) \left( \frac{1 \text{ mole}}{6.02 \times 10^{23} \text{ molecules}} \right) \left( \frac{194.1926g}{\text{mole}} \right) = \boxed{0.248g}$$



(d) 0.406 g cholesterol corresponds to 0.00105 moles.

So, the molar mass is:

$$\frac{0.406g}{0.00105 \text{ moles}} = \boxed{387 \frac{g}{\text{mole}}}$$



37) 5.00 mg "allicin" → # of sulfur atoms

$$\left( \frac{5.00 \times 10^{-3} \text{ g}}{\text{C}_6\text{H}_{10}\text{OS}_2} \right) \left( \frac{1 \text{ mole}}{162.2724 \text{ g}} \right) \left( \frac{6.02 \times 10^{23} \text{ molecules}}{1 \text{ mole}} \right) \left( \frac{2 \text{ sulfur atoms}}{1 \text{ C}_6\text{H}_{10}\text{OS}_2 \text{ molecule}} \right)$$

$$= \boxed{3.71 \times 10^{19} \text{ sulfur atoms}}$$

42) 25 μg THC (C<sub>21</sub>H<sub>30</sub>O<sub>2</sub>) → moles

$$(25 \mu\text{g}) \left( \frac{1 \text{ g}}{10^6 \mu\text{g}} \right) \left( \frac{1 \text{ mole}}{314.4668 \text{ g}} \right) = 7.94996 \times 10^{-8} \text{ moles}$$

$$\boxed{7.9 \times 10^{-8} \text{ moles}}$$

$$(7.94996 \times 10^{-8} \text{ moles}) \left( \frac{6.02 \times 10^{23} \text{ molecules}}{1 \text{ mole}} \right) = \boxed{4.8 \times 10^{16} \text{ molecules}}$$

43) Find empirical formula:

(a) Cpd contains 0.0130 mole C, 0.0390 mole H, 0.0065 mole O

$$\frac{0.0130}{0.0065} = 2.00 \text{ mole C per mole O}$$

$$\frac{0.0390 \text{ mole H}}{0.0065 \text{ mole O}} = 6.00 \text{ mole H per mole O}$$

$$\left. \begin{array}{l} \\ \end{array} \right\} \boxed{\text{C}_2\text{H}_6\text{O}} \text{ (a)}$$

(b) cpd contains 11.66 g Fe and 5.01 g Oxygen

$$(11.66 \text{ g Fe}) \left( \frac{1 \text{ mole}}{55.847 \text{ g}} \right) = 0.20878 \text{ moles Fe}$$

$$(5.01 \text{ g O}) \left( \frac{1 \text{ mole}}{15.9994 \text{ g}} \right) = 0.31314 \text{ moles O}$$

$$\frac{0.31314}{0.20878} = 1.49984 \approx 1.5 \quad (\text{Fe}_1\text{O}_{1.5}) \times 2 = \boxed{\text{Fe}_2\text{O}_3} \text{ (b)}$$

(c) 40.0% C, 6.7% H, 53.3% O by mass.

Assume we have 100 grams total ..

$$(40.0 \text{ g C}) \left( \frac{1 \text{ mole}}{12.011 \text{ g}} \right) = 3.3303 \text{ moles C}$$

$$\frac{3.3314}{3.3303} = 1.0003 \approx 1$$

$$(6.7 \text{ g H}) \left( \frac{1 \text{ mole}}{1.0079 \text{ g}} \right) = 6.6475 \text{ moles H}$$

$$\frac{6.6475}{3.3303} = 1.996 \approx 2$$

$$(53.3 \text{ g O}) \left( \frac{1 \text{ mole}}{15.9994 \text{ g}} \right) = 3.3314 \text{ moles O}$$

$$\boxed{\text{CH}_2\text{O}} \text{ (c)}$$

- (47) Empirical Formula  $\text{XF}_3$ , and is 65% Fluorine by mass.  
 (a) find atomic mass of X.

$$\frac{3(18.9984)}{X + 3(18.9984)} = 0.65$$

$$X + 3(18.9984)$$

↑  
 F atomic mass is 18.9984

$$\frac{56.9952}{X + 56.9952} = 0.65$$

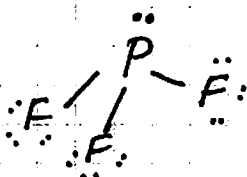
$$56.9952 = 0.65X + 37.04688$$

$$19.94832 = 0.65X$$

$$X = 30.6897$$

- (b) acc. to atomic mass  
 the element is phosphorus.

$$X = 31 \text{ g/mole or amu} \quad (a)$$



← could work, if the  
 empirical formula is same as molecular formula.

- (49) (a) empirical formula (given) =  $\text{CH}_2$   
 molar mass (given) = 84 g/mole

$$\text{CH}_2 \approx 14 \text{ g/mole} \quad 14x = 84$$

$$x = 6$$

$$(\text{CH}_2)_6 = \boxed{\text{C}_6\text{H}_{12}} \text{ molecular formula}$$

- (b) emp:  $\text{NH}_2\text{Cl}$   
 m. mass: 51.5 g/mole

$\text{NH}_2\text{Cl}$  empirical mass is 51.5 g/mole so empirical and  
 molecular mass are  
 the same.  
 ( $\approx 14 + 2(1) + 35.5 = 51.5$ )

$$\boxed{\text{NH}_2\text{Cl}} \text{ molecular formula}$$

- (90) (b) Hemoglobin (this protein transports  $\text{O}_2$  in blood.  
 is found in red blood cells)

has 4 iron atoms per molecule, and is 0.340% Fe by mass.  
 Find Hb molar mass.

$$\frac{4(55.847) \frac{\text{g}}{\text{mole}}}{M} = 0.00340$$

M

$$.00340 M = 4(55.847) \frac{\text{g}}{\text{mole}}$$

$$M = 65702 \text{ g/mole}$$

$$\boxed{M = 65700 \text{ g/mole}} \quad (!)$$

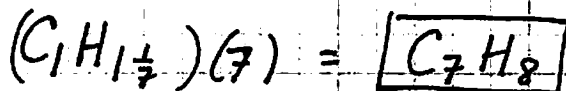
53 (a) Toluene (a hydrocarbon) combusts to form 5.86 mg CO<sub>2</sub> and 1.37 mg H<sub>2</sub>O. Find emp. formula

Since Toluene only contains C and H, we won't need to figure out the mass of oxygen in the compound. So we can convert CO<sub>2</sub>, H<sub>2</sub>O into moles C and H right away

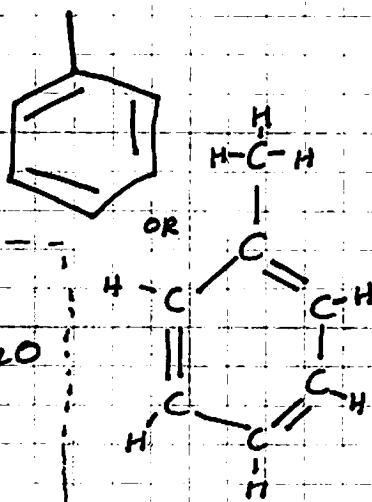
$$(.00586 \text{ g CO}_2) \left( \frac{1 \text{ mole CO}_2}{44.0098 \text{ g}} \right) \left( \frac{1 \text{ mole Carbon}}{1 \text{ mole CO}_2} \right) = .00013315 \text{ moles C}$$

$$(.00137 \text{ g H}_2\text{O}) \left( \frac{1 \text{ mole H}_2\text{O}}{18.0152 \text{ g}} \right) \left( \frac{2 \text{ mole H}}{1 \text{ mole H}_2\text{O}} \right) = .00015209 \text{ moles H}$$

$$\frac{.00015209 \text{ mole H}}{.00013315 \text{ mole C}} = 1.142 \approx \frac{1}{7}$$



(Toluene is aka "methyl benzene")



(b) Menthol: C<sub>x</sub>H<sub>y</sub>O<sub>z</sub>

.1005 g menthol burns → .2829 g CO<sub>2</sub>, .1159 g H<sub>2</sub>O

$$(0.2829 \text{ g CO}_2) \left( \frac{12.011 \text{ g C}}{44.0098 \text{ g CO}_2} \right) = 0.077208 \text{ g C}$$

$$(.1159 \text{ g H}_2\text{O}) \left( \frac{2(1.0079) \text{ g H}}{18.0152 \text{ g H}_2\text{O}} \right) = .012969 \text{ g H}$$

$$.1005 \text{ g} - .077208 \text{ g C} - .012969 \text{ g H} = 0.01032 \text{ grams oxygen}$$

$$(.077208 \text{ g C}) \left( \frac{1 \text{ mole}}{12.011 \text{ g}} \right) = 0.006428 \text{ moles C}$$

$$(.012969 \text{ g H}) \left( \frac{1 \text{ mole}}{1.0079 \text{ g}} \right) = 0.012867 \text{ moles H}$$

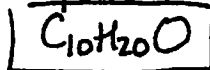
$$(.01032 \text{ g O}) \left( \frac{1 \text{ mole}}{15.9994 \text{ g}} \right) = 0.0006452 \text{ moles O}$$

$$\frac{.006428}{.0006452} = 9.963 \text{ moles C per mole O } (\approx 10)$$

$$\frac{.012867}{.0006452} = 19.943 \text{ moles H per mole O } (\approx 20)$$



and, since the molar mass is given ( $\approx 156 \text{ g/mole}$ ) and it matches the empirical mass of  $156 \text{ g/mole}$ , The molecular formula must match the empirical formula:



54. (a) pineapple odor: ethyl butyrate,  $C_xH_yO_z$

2.78 mg ethyl butyrate  $\rightarrow$  6.32 mg  $CO_2$  + 2.58 mg  $H_2O$   
+  $O_2$  (combustion)

$$(6.32 \text{ mg } CO_2) \left( \frac{12.011 \text{ mg C}}{44.0098 \text{ mg } CO_2} \right) = 1.7248 \text{ mg C}$$

$$(2.58 \text{ mg } H_2O) \left( \frac{2 \times 1.0079 \text{ mg H}}{18.0152 \text{ mg } H_2O} \right) = 0.28869 \text{ mg H}$$

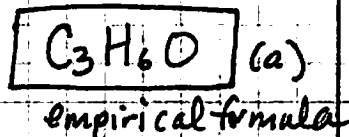
$$2.78 \text{ mg compound} - 1.7248 \text{ mg C} - 0.28869 \text{ mg H} = 0.7665 \text{ mg Oxygen}$$

$$(1.7248 \text{ mg C}) \left( \frac{1 \text{ mole}}{12.011 \text{ g}} \right) = 0.1436 \text{ mmole C} \leftarrow (\text{millimole} = 10^{-3} \text{ mole})$$

$$(0.28869 \text{ mg H}) \left( \frac{1 \text{ mole}}{1.0079 \text{ g}} \right) = 0.2864 \text{ mmole H}$$

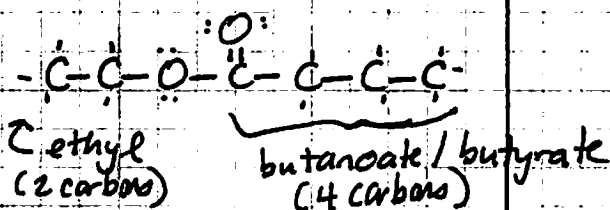
$$(0.7665 \text{ mg O}) \left( \frac{1 \text{ mole}}{15.9994 \text{ g}} \right) = 0.047908 \text{ mmole O}$$

$$\frac{.2864}{.047908} = 5.98 \approx 6 \text{ moles H per mole O}$$



$$\frac{.1436}{.047908} = 2.9974 \approx 3 \text{ moles C per mole O}$$

(If you remember organic naming,  
and the functional group of  
ester:



(b) Nicotine  $C_xH_yN_z$   
5.250 mg nicotine

$\rightarrow$  14.242 mg  $CO_2$  + 4.083 mg  $H_2O$

$$(14.242 \text{ mg } CO_2) \left( \frac{12.011 \text{ mg C}}{44.0098 \text{ mg } CO_2} \right) = 3.8869 \text{ mg C}$$

$$(4.083 \text{ mg } H_2O) \left( \frac{2 \times 1.0079 \text{ mg H}}{18.0152 \text{ mg } H_2O} \right) = 0.45686 \text{ mg H}$$

$$5.250 \text{ mg} - 3.8869 \text{ mg} - 0.45686 \text{ mg} = 0.9062 \text{ mg N}$$

$$\left( \frac{3.8869 \text{ mg C}}{12.011 \text{ g/mole}} \right) = 0.3236 \text{ mmoles C}$$

$$\left( \frac{0.45686 \text{ mg H}}{1.0079 \text{ g/mole}} \right) = 0.4533 \text{ mmoles H}$$

$$\left( \frac{.9062 \text{ mg N}}{14.0067 \text{ g/mole}} \right) = 0.06470 \text{ mmoles N}$$

$$\frac{0.32361}{.06470} = 5.002 \approx 5$$

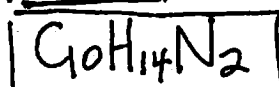
$$\frac{.4533}{.06470} = 7.006 \approx 7$$

$C_5H_7N$  = empirical formula

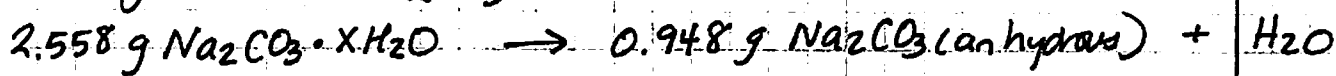
empirical mass = 81 amu.

given: molar mass  $\approx$  160 amu

So must double it to get  
molecular formula of



57. Washing Soda:  $\text{Na}_2\text{CO}_3 \cdot x\text{H}_2\text{O}$  Find  $x$ , if:



$$2.558 \text{ g hydrate} - 0.948 \text{ g anhydrous salt} = 1.610 \text{ g } \text{H}_2\text{O} \text{ "lost"}$$

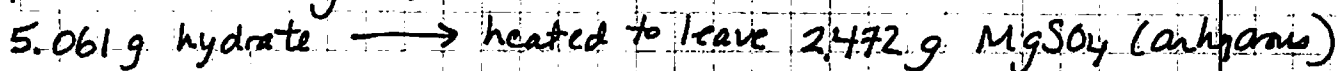
$$(1.610 \text{ g } \text{H}_2\text{O}) \left( \frac{1 \text{ mole}}{18.0152 \text{ g}} \right) = 0.08937 \text{ moles } \text{H}_2\text{O}$$

$$(0.948 \text{ g } \text{Na}_2\text{CO}_3) \left( \frac{1 \text{ mole}}{105.9888 \text{ g}} \right) = 0.008944 \text{ moles } \text{Na}_2\text{CO}_3$$

$$\frac{0.08937 \text{ moles } \text{H}_2\text{O}}{0.008944 \text{ moles } \text{Na}_2\text{CO}_3} = 9.992 \approx 10 \text{ moles } \text{H}_2\text{O} \text{ per mole } \text{Na}_2\text{CO}_3.$$

$$\boxed{x = 10} \quad (\text{hydrate formula is } \text{Na}_2\text{CO}_3 \cdot 10 \text{ H}_2\text{O})$$

58. epsom salts:  $\text{MgSO}_4 \cdot x\text{H}_2\text{O}$



$$\begin{array}{r} 5.061 \text{ g } \text{MgSO}_4 \cdot x\text{H}_2\text{O} \\ - 2.472 \text{ g } \text{MgSO}_4 \\ \hline \end{array}$$

$$2.589 \text{ g } \text{H}_2\text{O} \text{ lost}$$

$$(2.589 \text{ g } \text{H}_2\text{O}) \left( \frac{1 \text{ mole}}{18.0152 \text{ g}} \right) = 0.14371 \text{ moles } \text{H}_2\text{O}$$

$$(2.472 \text{ g } \text{MgSO}_4) \left( \frac{1 \text{ mole}}{120.3666 \text{ g}} \right) = 0.02054 \text{ moles } \text{MgSO}_4$$

$$\frac{0.14371 \text{ mole } \text{H}_2\text{O}}{0.02054 \text{ mole } \text{MgSO}_4} = 6.9975 \approx 7 \text{ mole } \text{H}_2\text{O} / \text{mole } \text{MgSO}_4$$

$$\boxed{x = 7} \quad (\text{hydrate formula is } \text{MgSO}_4 \cdot 7 \text{ H}_2\text{O})$$





a)  $(0.400 \text{ mole KO}_2) \left( \frac{3 \text{ mole O}_2}{4 \text{ mole KO}_2} \right) = \boxed{0.300 \text{ mole O}_2 \text{ produced}} \text{ (a)}$

b)  $(7.50 \text{ g O}_2) \left( \frac{1 \text{ mole}}{31.9988 \text{ g}} \right) \left( \frac{4 \text{ mole KO}_2}{3 \text{ mole O}_2} \right) \left( \frac{71.0971 \text{ g}}{1 \text{ mole}} \right) = \boxed{22.2 \text{ g KO}_2 \text{ must react}} \text{ (b)}$

c)  $(7.50 \text{ g O}_2) \left( \frac{1 \text{ mole}}{31.9988 \text{ g}} \right) \left( \frac{2 \text{ mole CO}_2}{3 \text{ mole O}_2} \right) \left( \frac{44.0098 \text{ g}}{1 \text{ mole}} \right) = \boxed{6.88 \text{ g CO}_2 \text{ must react}} \text{ (c)}$

d)  $(7.50 \text{ g O}_2) \left( \frac{1 \text{ mole}}{31.9988 \text{ g}} \right) \left( \frac{2 \text{ mole CO}_2}{3 \text{ mole O}_2} \right) \left( \frac{6.02 \times 10^{23} \text{ molecules}}{1 \text{ mole}} \right) = \boxed{9.41 \times 10^{22} \text{ molecules CO}_2 \text{ must react}} \text{ (d)}$

OR, start w/  $6.8768 \text{ g CO}_2$  (answer to part c)

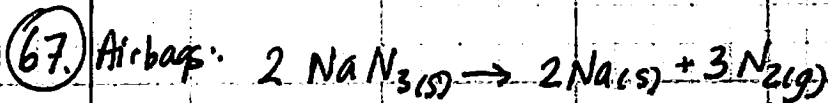
$(6.8768 \text{ g CO}_2) \left( \frac{1 \text{ mole}}{44.0098 \text{ g}} \right) \left( \frac{6.02 \times 10^{23} \text{ molecules}}{1 \text{ mole}} \right) = 9.41 \times 10^{22} \text{ molecules CO}_2$

e)  $(5.0 \text{ moles O}_2) \left( \frac{6.02 \times 10^{23} \text{ molecules}}{1 \text{ mole}} \right) = 3.01 \times 10^{24} \text{ molecules}$

$\rightarrow \boxed{3.0 \times 10^{24} \text{ molecules O}_2}$

f)  $(5.0 \text{ moles O}_2) \left( \frac{31.9988 \text{ g}}{\text{mole}} \right) = 159.994 \rightarrow \boxed{160 \text{ g O}_2}$

g)  $(5.0 \text{ moles O}_2) \left( \frac{4 \text{ mole KO}_2}{3 \text{ mole O}_2} \right) \left( \frac{71.0971 \text{ g}}{1 \text{ mole}} \right) = 473.98 \rightarrow \boxed{470 \text{ g KO}_2}$



a)  $(1.50 \text{ mole NaN}_3) \left( \frac{3 \text{ mole N}_2}{2 \text{ mole NaN}_3} \right) = \boxed{2.25 \text{ moles N}_2}$

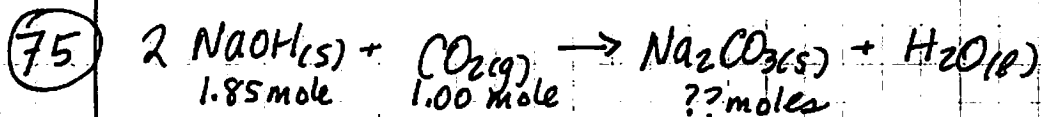
b)  $(10.0 \text{ g N}_2) \left( \frac{1 \text{ mole}}{28.0134 \text{ g}} \right) \left( \frac{2 \text{ mole NaN}_3}{3 \text{ mole N}_2} \right) \left( \frac{65.0099 \text{ g}}{1 \text{ mole}} \right) = \boxed{15.5 \text{ g NaN}_3}$

c)  $(10.0 \text{ ft}^3) \left( \frac{12 \text{ in}}{\text{ft}} \right)^3 \left( \frac{2.54 \text{ cm}}{\text{in}} \right)^3 \left( \frac{1 \text{ Liter}}{1000 \text{ cm}^3} \right) \left( \frac{1.25 \text{ g}}{\text{L}} \right) \left( \frac{1 \text{ mole}}{28.0134 \text{ g}} \right) \left( \frac{2 \text{ mole NaN}_3}{3 \text{ mole N}_2} \right) \left( \frac{65.0099 \text{ g}}{1 \text{ mole}} \right)$

this part gets you to the mass of  $\text{N}_2$ , in grams

$= 547.62 \rightarrow \boxed{548 \text{ g NaN}_3}$

needed to produce about an hour's worth of oxygen gas.



$$(1.85 \text{ mole NaOH}) \left( \frac{1 \text{ mole CO}_2}{2 \text{ mole NaOH}} \right) = 0.925 \text{ mole CO}_2 \text{ needed.}$$

$$1.00 \text{ mole CO}_2 > 0.925 \text{ mole CO}_2$$

(have)                      (needed)

So  $\text{CO}_2$  is the excess reactant, and  $\text{NaOH}$  is the limiting reactant.

$$\textcircled{\text{OR}}: (1.00 \text{ mole CO}_2) \left( \frac{2 \text{ mole NaOH}}{1 \text{ mole CO}_2} \right) = 2.00 \text{ mole NaOH needed.}$$

$$1.85 \text{ mole} < 2.00 \text{ mole}$$

(have)                      (needed)

So  $\text{NaOH}$  is the limiting reactant.

Since  $\text{NaOH}$  is limiting reactant, all the  $\text{NaOH}$  will react. The 1.85 moles  $\text{NaOH}$  will react with 0.925 moles  $\text{CO}_2$ , as shown above. So the amount of  $\text{CO}_2$  that remains will be:

$$1.00 \text{ mole CO}_2 - 0.925 \text{ mole CO}_2 = 0.075 \text{ moles CO}_2$$

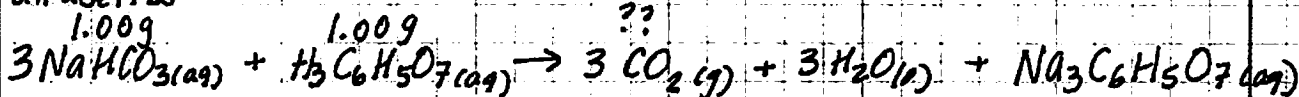
will remain.  $\rightarrow$  0.08 moles  $\text{CO}_2$  remain

When the 1.85 mole  $\text{NaOH}$  react,

$$(1.85 \text{ mole NaOH}) \left( \frac{1 \text{ mole Na}_2\text{CO}_3}{2 \text{ mole NaOH}} \right) = 0.925 \text{ moles Na}_2\text{CO}_3$$

will be produced.

$\textcircled{77}$  alkaseltzer:



$\rightarrow$  this first calc was a waste of time, Dish!

$$(1.00 \text{ g H}_3\text{C}_6\text{H}_5\text{O}_7) \left( \frac{1 \text{ mole}}{192.125 \text{ g}} \right) \left( \frac{3 \text{ mole NaHCO}_3}{1 \text{ mole H}_3\text{C}_6\text{H}_5\text{O}_7} \right) \left( \frac{84.0069 \text{ g}}{1 \text{ mole}} \right) = 1.3118 \text{ g NaHCO}_3$$

needed to react

$$\text{OR} (1.00 \text{ g NaHCO}_3) \left( \frac{1 \text{ mole}}{84.0069 \text{ g}} \right) \left( \frac{1 \text{ mole H}_3\text{C}_6\text{H}_5\text{O}_7}{3 \text{ mole NaHCO}_3} \right) \left( \frac{192.125 \text{ g}}{1 \text{ mole}} \right) = 0.7623 \text{ g H}_3\text{C}_6\text{H}_5\text{O}_7 \text{ needed to react.}$$

so  $\text{NaHCO}_3$  is limiting.

$$1.00 \text{ g} > 0.7623 \text{ g}$$

(have)                      (needed)

So  $\text{H}_3\text{C}_6\text{H}_5\text{O}_7$  is the excess reactant;  $\text{NaHCO}_3$  is limiting. (a)

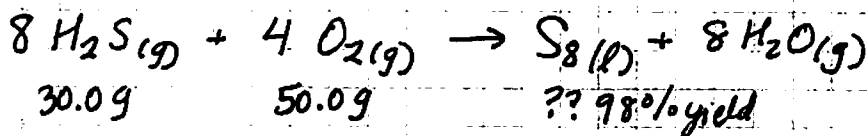
$$1.00 \text{ g H}_3\text{C}_6\text{H}_5\text{O}_7 \text{ (have)} - 0.7623 \text{ g (needed)} = 0.24 \text{ g H}_3\text{C}_6\text{H}_5\text{O}_7 \text{ left over. (c)}$$

$$(1.00 \text{ g NaHCO}_3) \left( \frac{1 \text{ mole}}{84.0069 \text{ g}} \right) \left( \frac{3 \text{ mole CO}_2}{3 \text{ mole NaHCO}_3} \right) \left( \frac{44.0098 \text{ g}}{1 \text{ mole}} \right) = 0.524 \text{ g CO}_2 \text{ can form. (b)}$$

$$\textcircled{d} \quad \% \text{ yield} = \frac{\text{actual}}{\text{expected}} \times 100 = \frac{0.517 \text{ g}}{0.52388 \text{ g}} \times 100 = 98.7 \% \text{ yield (d)}$$

83

Removing hydrogen sulfide (H<sub>2</sub>S) from natural gas:



$$(30.0 \text{g H}_2\text{S}) \left( \frac{1 \text{ mole}}{34.0798 \text{g}} \right) \left( \frac{1 \text{ mole S}_8}{8 \text{ mole H}_2\text{S}} \right) \left( \frac{256.512 \text{g}}{1 \text{ mole}} \right) = 28.226 \text{g S}_8$$

could form from H<sub>2</sub>S

$$(50.0 \text{g O}_2) \left( \frac{1 \text{ mole}}{31.9988 \text{g}} \right) \left( \frac{1 \text{ mole S}_8}{4 \text{ mole O}_2} \right) \left( \frac{256.512 \text{g}}{1 \text{ mole}} \right) = 100.2 \text{g S}_8$$

could form from O<sub>2</sub>

(or show which reactant is limiting in some other way, for example, 30.0g H<sub>2</sub>S can react with 14.1g O<sub>2</sub>, and you have way more O<sub>2</sub> than 14.1 grams...)

H<sub>2</sub>S can produce less product so H<sub>2</sub>S is limiting.

$$\% \text{ yield} = \frac{\text{actual}}{\text{theoretical}} \times 100 \quad 0.98 = \frac{x}{28.226 \text{g}} \quad x = 27.661 \text{g}$$

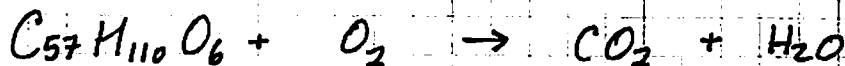
28 grams S<sub>8</sub>

(If the H<sub>2</sub>S is not removed, it will react to form SO<sub>2</sub> and SO<sub>3</sub>. These will react to form sulfuric acid rain (H<sub>2</sub>SO<sub>4</sub>))

(#90 is a few pages back, right after #49)

99

Camel Hump Fat! Burning Fat releases H<sub>2</sub>O for camel to use.



Calculate mass H<sub>2</sub>O produced per 1.0 kg fat burned

Could use stoichiometry:



$$(1.0 \times 10^3 \text{g fat}) \left( \frac{1 \text{ mole}}{891.4924 \text{g}} \right) \left( \frac{110 \text{ mole H}_2\text{O}}{2 \text{ mole fat}} \right) \left( \frac{18.0152 \text{g}}{1 \text{ mole}} \right) = 1111 \text{g H}_2\text{O}$$

OR, Find the mass of hydrogen in the fat, and figure out how much water would correspond to this amount of hydrogen:

$$\frac{1.1 \times 10^3 \text{g H}_2\text{O}}{1.1 \text{kg H}_2\text{O}}$$

$$(1.0 \times 10^3 \text{g C}_{57}\text{H}_{110}\text{O}_6) \left( \frac{110 (1.0079 \text{g H})}{891.4924 \text{g C}_{57}\text{H}_{110}\text{O}_6} \right) \left( \frac{18.0152 \text{g H}_2\text{O}}{2 (1.0079 \text{g H})} \right) = 1.1 \times 10^3 \text{g}$$

or 1.1 kg

(if you stop here, you get that 1.0 kg fat has 124 g Hydrogen)