

Test 4 Study Guide - Answers

1. a) CH₄

$$\begin{aligned} \text{i- molar mass (CH}_4\text{)} &= 1 \times \text{molar mass (C)} + 4 \times \text{molar mass (H)} \\ &= 1 (12.011) + 4 (1.0079) \\ &= 16.0426 \text{ g/mol} \end{aligned}$$



$$\text{ii- \% (C)} = \frac{(\# \text{ atoms C/molecule CH}_4)(\text{molar mass C})}{\text{molar mass CH}_4} = \frac{1 \times 12.011}{16.0426} = 0.7487 = 74.87\%$$

$$\% (\text{H}) = \frac{4 \times 1.0079}{16.0426} = 0.2513 = 25.13\%$$

- or -

$$\% (\text{H}) = 1 - (\% \text{ C}) = 1 - (0.7487) = 0.2513 = 25.13\%$$

iv- ~~Since~~ Since we don't have unlimited reactants, this is a limiting reagent problem.

$$20 \text{ g CH}_4 \times \frac{1 \text{ mol CH}_4}{16.0426 \text{ g CH}_4} \times \frac{1 \text{ mol CO}_2}{1 \text{ mol CH}_4} \times \frac{44.011 \text{ g CO}_2}{1 \text{ mol CO}_2} = 54.87 \text{ g CO}_2$$

$$30 \text{ g O}_2 \times \frac{1 \text{ mol O}_2}{32 \text{ g O}_2} \times \frac{1 \text{ mol CO}_2}{\frac{3}{2} \text{ mol O}_2} \times \frac{44.011 \text{ g CO}_2}{1 \text{ mol CO}_2} = \boxed{27.51 \text{ g CO}_2 \text{ produced}}$$

O₂ is the limiting reagent.

$$30 \text{ g O}_2 \times \frac{1 \text{ mol O}_2}{32 \text{ g O}_2} \times \frac{2 \text{ mol H}_2\text{O}}{\frac{3}{2} \text{ mol O}_2} \times \frac{18.0158 \text{ g H}_2\text{O}}{1 \text{ mol H}_2\text{O}} = \boxed{22.52 \text{ g H}_2\text{O produced}}$$

$$30 \text{ g O}_2 \times \frac{1 \text{ mol O}_2}{32 \text{ g O}_2} \times \frac{1 \text{ mol CH}_4}{\frac{3}{2} \text{ mol O}_2} \times \frac{16.0426 \text{ g CH}_4}{1 \text{ mol CH}_4} = 10.0266 \text{ g CH}_4 \text{ consumed}$$

$$20 \text{ g} - 10.0266 \text{ g} = \boxed{9.97 \text{ g CH}_4 \text{ remaining}}$$

$$V = 30 \text{ g CH}_4 \times \frac{1 \text{ mol CH}_4}{16.0426 \text{ g CH}_4} \times \frac{1 \text{ mol CO}_2}{1 \text{ mol CH}_4} \times \frac{44.011 \text{ g CO}_2}{1 \text{ mol CO}_2} = 82.30 \text{ g CO}_2 \text{ produced}$$

$$\% \text{ yield} = \frac{\text{actual yield}}{\text{theoretical yield}} \times 100\% = \frac{2.5 \text{ g CO}_2}{82.30 \text{ g CO}_2} \times 100\% = \boxed{3.04\% \text{ yield}}$$

* For (b) and (c), I will just give you the answers, Please come see me if you have questions. The work looks nearly identical to that shown in (a). *

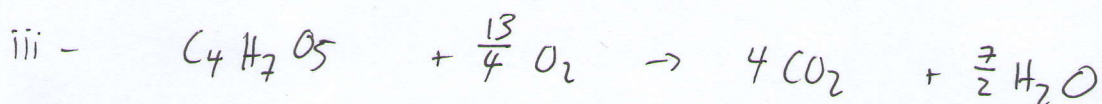
b) $\text{C}_4\text{H}_7\text{O}_5$

i - 135.10 g/mol

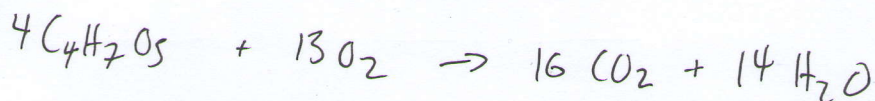
ii - % (C) = 35.56%

% (H) = 5.22%

% (O) = 59.22%



-or-



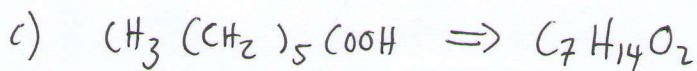
iv - $\text{C}_4\text{H}_7\text{O}_5$ is the limiting reagent.

26.06 g CO_2 produced

9.33 g H_2O produced

14.60 g O_2 remaining

v - 6.40% yield

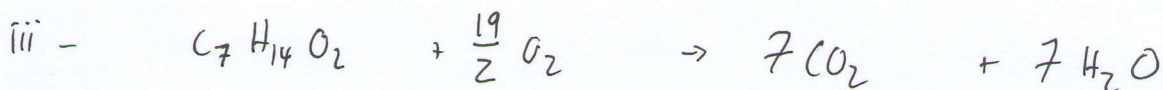


i - 130.19 g/mol

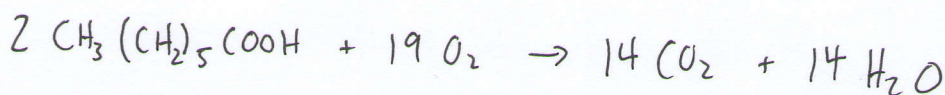
ii - %C = 64.58%

%H = 10.84%

%O = 24.58%



-or-



iv - O_2 is the limiting reagent.

30.40 g CO_2 produced

12.45 g H_2O produced

7.15 g $\text{C}_7\text{H}_{14}\text{O}_2$ remaining

v - 3.52% yield



$20\text{g CO}_2 \times \frac{1\text{mol CO}_2}{44.011\text{g CO}_2} \times \frac{1\text{mol C}}{1\text{mol CO}_2} \times \frac{12.011\text{g C}}{1\text{mol C}} = 5.46\text{g C in C}_x\text{H}_y\text{O}_z$

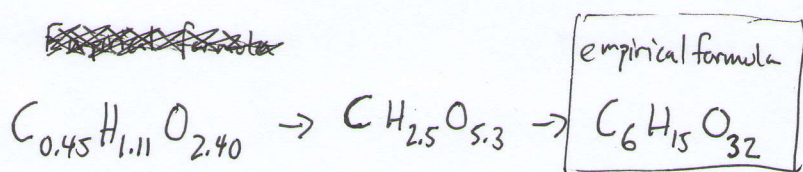
$0.45\text{mol C in C}_x\text{H}_y\text{O}_z$

$10\text{g H}_2\text{O} \times \frac{1\text{mol H}_2\text{O}}{18.015\text{g H}_2\text{O}} \times \frac{2\text{mol H}}{1\text{mol H}_2\text{O}} \times \frac{1.0079\text{g H}}{1\text{mol H}} = 1.12\text{g H in C}_x\text{H}_y\text{O}_z$

$1.11\text{mol H in C}_x\text{H}_y\text{O}_z$

$45\text{g} - (5.46\text{g} + 1.12\text{g}) = 38.42\text{g O}_2 \text{ in C}_x\text{H}_y\text{O}_z \times \frac{1\text{mol O}_2}{32\text{g O}_2} = 2.40\text{mol O}_2 \text{ in C}_x\text{H}_y\text{O}_z$

~~Empirical formula~~



... this has a mass greater than the molecular molar mass...

... let's try being more careful with rounding...

$$20 \text{ g CO}_2 \times \frac{1 \text{ mol CO}_2}{44.011 \text{ g CO}_2} \times \frac{1 \text{ mol C}}{1 \text{ mol CO}_2} \times \frac{12.011 \text{ g C}}{1 \text{ mol C}} = 5.45818091 \text{ g C}$$

$$\parallel$$

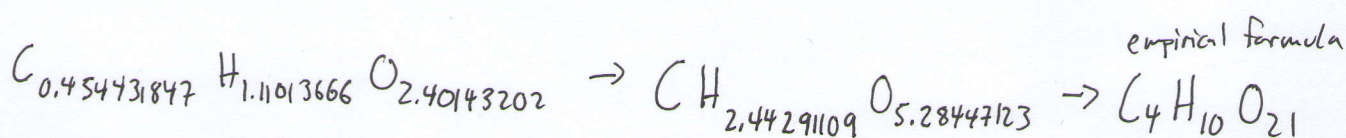
$$0.454431847 \text{ mol C}$$

$$10 \text{ g H}_2\text{O} \times \frac{1 \text{ mol H}_2\text{O}}{18.0158 \text{ g H}_2\text{O}} \times \frac{2 \text{ mol H}}{1 \text{ mol H}_2\text{O}} \times \frac{1.0079 \text{ g H}}{1 \text{ mol H}} = 1.11890674 \text{ g H}$$

$$\parallel$$

$$1.11013666 \text{ mol H}$$

$$45 \text{ g} - (5.45818091 \text{ g} + 1.11890674 \text{ g}) = 38.4229123 \text{ g O} = 2.40143202 \text{ mol O}$$



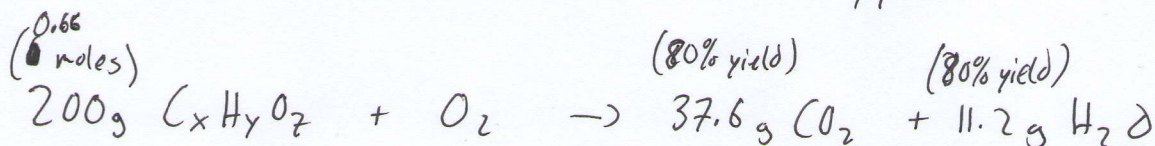
the molecular molar mass is still too big: 394.123 g/mol. Bad problem - sorry!

Assume the molecular molar mass is 1.182 kg. Then

$$\frac{\text{molecular molar mass}}{\text{empirical molar mass}} = \frac{1,182 \text{ g/mol}}{394.123 \text{ g/mol}} \approx 3; \text{ the molecular formula is then}$$

$$\text{C}_{12} \text{H}_{30} \text{O}_6$$

3. No promises this one works better... Let's see what happens.



$$\text{theoretical yield CO}_2 = 37.6 / 0.8 = 47 \text{ g CO}_2$$

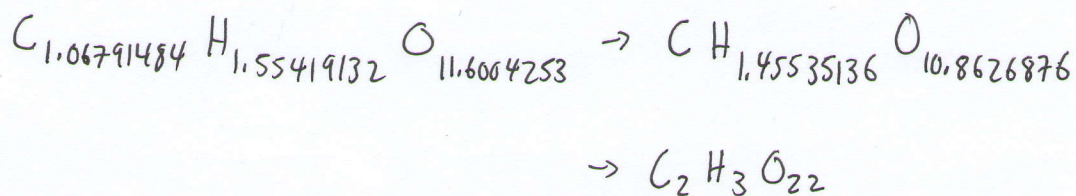
$$\text{theoretical yield H}_2\text{O} = 11.2 / 0.8 = 14 \text{ g H}_2\text{O}$$

$$47 \text{ g CO}_2 \times \left(\frac{1}{44.011}\right) \times \left(\frac{1}{1}\right) = \boxed{1.06791484 \text{ mol C}} = \boxed{12.8267251 \text{ g C}}$$

$$14 \text{ g H}_2\text{O} \times \left(\frac{1}{18.0158}\right) \times \left(\frac{2}{1}\right) = \boxed{1.55419132 \text{ mol H}} = \boxed{1.56646943 \text{ g H}}$$

$$200 \text{ g} - (12.8267251 \text{ g} + 1.56646943 \text{ g}) = \boxed{185.606805 \text{ g O}_2} = \boxed{11.6004253 \text{ mol O}}$$

* see units in problem (2) *



$$\text{empirical molar mass} = 379.0457 \text{ g/mol}$$

$$\text{molecular molar mass} = 303.0303 \text{ g/mol} = \frac{200\%}{0.66} \text{ mol}$$

... That makes no sense. Another bad problem.

I will give you another - easy and logical - review problem before the test.

The answers to the Ch. 8 and Ch. 9 review problems are in the back of your textbook. Please come see me if have any trouble getting the correct answer.