Test 4 Study Guide - Answers

1. a) CH_4

i- molar mass (CH_4) = 1 × molar mass (C) + 4 × molar mass (H)

$$= 1\,(12.011) + 4\,(1.0079)$$

$$= 16.0426 \text{ g/mol}$$

iii- $$CH_4 + \frac{3}{2} O_2 \rightarrow CO_2 + 2H_2O$$

ii- $$\%(C) = \frac{(\#\text{ atoms C/molecule } CH_4)(\text{molar mass C})}{\text{molar mass } CH_4} = \frac{1 \times 12.011}{16.0426} = 0.7487 = 74.87\%$$

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

-or-

$$\%(H) = 1 - (\%\,C) = 1 - (0.7487) = 0.2513 = 25.13\%$$

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

$$20\text{g } CH_4 \times \frac{1 \text{ mol } CH_4}{16.0426 \text{ g } CH_4} \times \frac{1 \text{ mol } CO_2}{1 \text{ mol } CH_4} \times \frac{44.011 \text{ g } CO_2}{1 \text{ mol } CO_2} = 54.87 \text{ g } CO_2$$

$$30\text{g } O_2 \times \frac{1 \text{ mol } O_2}{32 \text{ g } O_2} \times \frac{1 \text{ mol } CO_2}{\frac{3}{2} \text{ mol } O_2} \times \frac{44.011 \text{ g } CO_2}{1 \text{ mol } CO_2} = \boxed{27.51 \text{ g } CO_2 \text{ produced}}$$

O_2 is the limiting reagent.

$$30\text{g } O_2 \times \frac{1 \text{ mol } O_2}{32 \text{ g } O_2} \times \frac{2 \text{ mol } H_2O}{\frac{3}{2} \text{ mol } O_2} \times \frac{18.0158 \text{ g } H_2O}{1 \text{ mol } H_2O} = \boxed{22.52 \text{ g } H_2O \text{ produced}}$$

$$30\text{g } O_2 \times \frac{1 \text{ mol } O_2}{32 \text{ g } O_2} \times \frac{1 \text{ mol } CH_4}{\frac{3}{2} \text{ mol } O_2} \times \frac{16.0426 \text{ g } CH_4}{1 \text{ mol } CH_4} = 10.0266 \text{ g } CH_4 \text{ consumed}$$

$$20\text{g} - 10.0266\text{ g} = \boxed{9.97 \text{ g } CH_4 \text{ remaining}}$$

v– $30\,g\ CH_4 \times \frac{1\ mol\ CH_4}{16.0426\ g\ CH_4} \times \frac{1\ mol\ CO_2}{1\ mol\ CH_4} \times \frac{44.011\ g\ CO_2}{1\ mol\ CO_2} = 82.30\ g\ CO_2$ produced

$\%\ yield = \frac{actual\ yield}{theoretical\ yield} \times 100\% = \frac{2.5\ g\ CO_2}{82.30\ g\ CO_2} \times 100\% = \boxed{3.04\%\ yield}$

* 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) $C_4H_7O_5$

i – 135.10 g/mol

ii – $\%(C) = 35.56\%$

$\%(H) = 5.22\%$

$\%(O) = 59.22\%$

iii – $C_4H_7O_5 + \frac{13}{4} O_2 \rightarrow 4\,CO_2 + \frac{7}{2} H_2O$

–or–

$4\,C_4H_7O_5 + 13\,O_2 \rightarrow 16\,CO_2 + 14\,H_2O$

iv – $C_4H_7O_5$ 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

c) $CH_3(CH_2)_5COOH \Rightarrow C_7H_{14}O_2$

i – 130.19 g/mol

ii – $\%(C) = 64.58\%$

$\%(H) = 10.84\%$

$\%(O) = 24.58\%$

iii – $C_7H_{14}O_2 + \frac{19}{2} O_2 \rightarrow 7 CO_2 + 7 H_2O$

–or–

$2\, CH_3(CH_2)_5COOH + 19\, O_2 \rightarrow 14\, CO_2 + 14\, H_2O$

iv – O_2 is the limiting reagent.

30.40 g CO_2 produced

12.45 g H_2O produced

7.15 g $C_7H_{14}O_2$ remaining

v – 3.52 % yield

2. $45\,g\ C_xH_yO_z + O_2 \rightarrow 20\,g\ CO_2 + 10\,g\ H_2O$

$$20\,g\ CO_2 \times \frac{1\ mol\ CO_2}{44.011\ g\ CO_2} \times \frac{1\ mol\ C}{1\ mol\ CO_2} \times \frac{12.011\ g\ C}{1\ mol\ C} = \boxed{5.46\ g\ C\ in\ C_xH_yO_z}$$

$$\boxed{0.45\ mol\ C\ in\ C_xH_yO_z}$$

$$10\,g\ H_2O \times \frac{1\ mol\ H_2O}{18.0158\ g\ H_2O} \times \frac{2\ mol\ H}{1\ mol\ H_2O} \times \frac{1.0079\ g\ H}{1\ mol\ H} = \boxed{1.12\ g\ H\ in\ C_xH_yO_z}$$

$$\boxed{1.11\ mol\ H\ in\ C_xH_yO_z}$$

$$45\,g - (5.46\,g + 1.12\,g) = \boxed{38.42\,g\ O\ in\ C_xH_yO_z} \times \frac{1\ mol\ O}{16\ g\ O} = \boxed{2.40\ mol\ O\ in\ C_xH_yO_z}$$

$$C_{0.45}H_{1.11}O_{2.40} \rightarrow CH_{2.5}O_{5.3} \rightarrow \boxed{\text{empirical formula}\ C_6H_{15}O_{32}}$$

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

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

$$20\,g\,CO_2 \times \frac{1\,mol\,CO_2}{44.011\,g\,CO_2} \times \frac{1\,mol\,C}{1\,mol\,CO_2} \times \frac{12.011\,g\,C}{1\,mol\,C} = 5.45818091\,g\,C$$

$\parallel$

0.454431847 mol C

$$10\,g\,H_2O \times \frac{1\,mol\,H_2O}{18.0158\,g\,H_2O} \times \frac{2\,mol\,H}{1\,mol\,H_2O} \times \frac{1.0079\,g\,H}{1\,mol\,H} = 1.11890674\,g\,H$$

$\parallel$

1.11013666 mol H

$$45\,g - (5.45818091\,g + 1.11890674)\,g = 38.4229123\,g\,O = 2.40143202\,mol\,O$$

empirical formula

$$C_{0.454431847}H_{1.11013666}O_{2.40143202} \rightarrow CH_{2.44291109}O_{5.28447123} \rightarrow C_4H_{10}O_{21}$$

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

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

$$C_{12}H_{30}O_{63}$$

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

$$\underset{(0.66\text{ moles})}{200\,g\,C_xH_yO_z} + O_2 \rightarrow \underset{(80\%\text{ yield})}{37.6\,g\,CO_2} + \underset{(80\%\text{ yield})}{11.2\,g\,H_2O}$$

theoretical yield CO_2 = 37.6/0.8 = 47 g CO_2

theoretical yield H_2O = 11.2/0.8 = 14 g H_2O

* see units in problem (2) *

$$47\,g\,CO_2 \times \left(\frac{1}{44.011}\right)\times\left(\frac{1}{1}\right) = \boxed{1.06791484\,mol\,C} = \boxed{12.8267251\,g\,C}$$

$$14\,g\,H_2O \left(\frac{1}{18.0158}\right)\left(\frac{2}{1}\right) = \boxed{1.55419132\,mol\,H} = \boxed{1.56646943\,g\,H}$$

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

$C_{1.067914484}H_{1.55419132}O_{11.6064253} \rightarrow CH_{1.45535136}O_{10.8626876}$

$\rightarrow C_2H_3O_{22}$

empirical molar mass = 379.0457 g/mol

molecular molar mass = 303.0303 g/mol = 200g/0.66 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.