



MICHAEL'S REVIEW OF ACTUAL STUDENT RESPONSES:

Day One Practice Test (Monday, May 4th, 2020)



(b) When a sample of C_2H_5OH was combusted, the volume of $CO_2(g)$ produced was 18.0 L when measured at $21.7^\circ C$ and 1.03 atm.

(i) Determine the number of moles of $CO_2(g)$ that was produced.

Handwritten calculation: $n = \frac{PV}{RT} = \frac{1.03(18)}{(0.08206)(21.7+273)} = 0.77 \text{ mol } CO_2(g)$

(ii) Determine the volume of $C_2H_5OH(l)$, in mL, that was combusted to produce the volume of $CO_2(g)$ collected. (The density of $C_2H_5OH(l)$ is 0.79 g/mL.)

Handwritten calculation: $0.77 \text{ mol } CO_2 \times \frac{1 \text{ mol } C_2H_5OH}{2 \text{ mol } CO_2} \times \frac{46 \text{ g/mol}}{1 \text{ mol } C_2H_5OH} = 17.71 \text{ g } C_2H_5OH$
 $\frac{17.71 \text{ g}}{0.79 \text{ g/mL}} = 22.41 \text{ mL } C_2H_5OH$

(iii) Determine the amount of heat, in kJ, that was released by the combustion reaction.

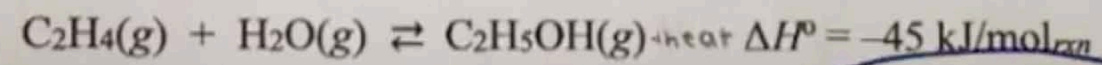
Handwritten calculation: $q = \Delta H(\text{mols}) = 1270 \text{ kJ/mol} \times (0.385 \text{ mol}) = 488.95 \text{ kJ heat released by the combustion reaction}$

The combustion reaction described in part (b) occurred in a closed room containing $5.56 \times 10^4 \text{ g}$ of air originally at $21.7^\circ C$. Assume that all of the heat produced by the reaction was absorbed by the air (specific heat = $1.005 \text{ J/(g} \cdot ^\circ C)$) in the room.

Handwritten calculation: $q = mca + 488950 \text{ J} = 5.56 \times 10^4 \text{ g} (1.005 \text{ J/g} \cdot ^\circ C)(\Delta T - 21.7^\circ C)$
 $\Delta T = \text{final temperature} = 30.45^\circ C$

(c) Determine the final temperature of the air in the room after the combustion.

One method of producing ethanol is by the reaction of ethene and water, as represented below.



(d) A chemist wants to run the reaction and maximize the amount of $C_2H_5OH(g)$ produced. Identify two ways the chemist could change the reaction conditions (other than adding or removing any chemical species) to favor the formation of more product. Justify your answer.

Handwritten justification: The chemist can decrease the temperature of the reaction and can also decrease the volume of the container. By decreasing the temperature, the reaction will shift to the right to reestablish equilibrium (increasing the amount of $C_2H_5OH(g)$). By decreasing the volume of the container, the reaction will reestablish equilibrium by shifting to the products side.

The concentration of $C_2H_5OH(aq)$ in a solution can be determined using a spectrophotometer, based on the reaction represented below.

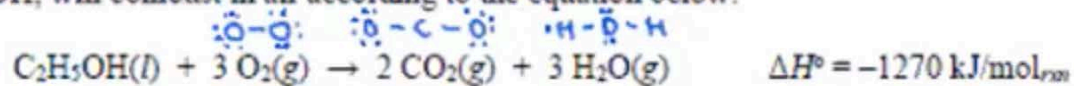




Monday Sample FRQ Q #1

Answer the following questions that are related to ethanol, C₂H₅OH.

Ethanol, C₂H₅OH, will combust in air according to the equation below.



(a) Is O₂(g) oxidized in the reaction, or is it reduced? Justify your answer in terms of oxidation numbers.

O₂ is reduced because it goes from 0 to -2 oxidation.

(b) When a sample of C₂H₅OH was combusted, the volume of CO₂(g) produced was 18.0 L when measured at 21.7°C and 1.03 atm.

+273 / 294.7 K

PV = nRT n = PV / RT

(1.03 atm)(18 L) / (0.08206)(294.7) = 18.54 / 24.183082 K = 0.767 moles CO₂

(i) Determine the number of moles of CO₂(g) that was produced.

0.767 moles CO₂

(ii) Determine the volume of C₂H₅OH(l), in mL, that was combusted to produce the volume of CO₂(g) collected. (The density of C₂H₅OH(l) is 0.79 g/mL.)

D = m / L

0.767 mol CO₂ (1 mol C₂H₅OH / 2 mol CO₂) = 0.3835 mol C₂H₅OH (46.08 g / 1 mol) = 17.67 g = 0.79 g/mL

(iii) Determine the amount of heat, in kJ, that was released by the combustion reaction.

0.3835 mol C₂H₅OH (-1270 kJ/mol) = -487.045 kJ

X = 22.37 mL

The combustion reaction described in part (b) occurred in a closed room containing 5.56 × 10⁴ g of air originally at 21.7°C. Assume that all of the heat produced by the reaction was absorbed by the air (specific heat = 1.005 J/(g · °C)) in the room.

Q = mcΔT (5560 g)(1.005 J/g·C)

(c) Determine the final temperature of the air in the room after the combustion.

One method of producing ethanol is by the reaction of ethene and water, as represented below.



(d) A chemist wants to run the reaction and maximize the amount of C₂H₅OH(g) produced. Identify two ways the chemist could change the reaction conditions (other than adding or removing any chemical species) to favor the formation of more product. Justify your answer.

1) Decrease temp because the forward is exothermic. 2) Decrease the volume of the container because forward is less moles.

The concentration of C₂H₅OH(aq) in a solution can be determined using a spectrophotometer, based on the



1.

a) O₂ was reduced in this reaction. Its oxidation number dropped from 0 to -4.

b)

i) 0.767 mols

ii) 22.4 mL

iii) 488 kJ was released

c) 30.4 degrees C

d) One thing the chemist could do is lower the temperature. Heat is a product of the reaction, and if you lower the temperature, the system will shift towards the products to fill in the missing heat. The chemist could also allow the reaction to occur in a smaller container. Less volume would mean greater pressure, and the system will shift to the side with less moles of gas (here the products) to relieve the stress.

e) 7.0×10^{-4} M

f) Yes I agree. Absorption is directly proportional to concentration. Half-lives occur for this compound as the absorbance is halved at 0.300 min and halved again at 0.600 min, meaning that the concentration also demonstrates characteristics of a half-life. Such



zoom



- a) O is reduced in this reaction; its oxidation number changes from 0 to -2
- b) i) $n = PV/RT = 1.03(18)/0.08206(294.7) = 0.767 \text{ mol CO}_2$
 ii) $0.767 \text{ mol CO}_2 \times 1 \text{ mol C}_2\text{H}_5\text{OH}/2 \text{ mol CO}_2 \times 46.08\text{g}/1 \text{ mol C}_2\text{H}_5\text{OH} \times 1 \text{ mL}/0.79 \text{ g} = 22.4 \text{ mL C}_2\text{H}_5\text{OH}$
 iii) $q = m\Delta H = 0.383\text{mol}(-1270\text{kJ/mol}) = -486 \text{ kJ}$
- c) $q = mC\Delta T$ $\Delta T = q/mC = (-486\text{kJ} \times 1000\text{J}/1\text{kJ})/5.56 \times 10^4 \text{ g}/1.005\text{J/gC} = -8.71 \text{ C}$
 $21.7 \text{ C} - (-8.71 \text{ C}) = 30.4 \text{ C}$
- d) Decrease the volume of the container; the equilibrium will adjust for this by producing more products, since there are less moles of products than reactants. Since the reaction is exothermic, decreasing the temperature will cause the equilibrium to shift towards the products. Lower temperature means less energy and correct collisions for the reverse reaction will occur, and more product will be produced.
- e) $0.552/0.780 \times 1.0 \times 10^{-3} = 7.08 \times 10^{-4} \text{ M}$



zoom



b.

i. K temperature = $21.7^{\circ}\text{C} + 273 = 294.7 \text{ K}$

$$PV = nRT, n = \frac{PV}{RT} = \frac{(1.03 \text{ atm})(18.0 \text{ L})}{\left(0.08206 \frac{\text{L}\cdot\text{atm}}{\text{K}\cdot\text{mol}}\right)(294.7 \text{ K})}$$

$$= 0.767 \text{ mol CO}_2(g) \text{ produced}$$

ii. $0.767 \text{ mol CO}_2 \times \frac{1 \text{ mol C}_2\text{H}_5\text{OH}}{2 \text{ mol CO}_2} \times 46.068 \frac{\text{g}}{\text{mol}} \times \frac{1 \text{ mL}}{0.79 \text{ g}} = 22 \text{ C}_2\text{H}_5\text{OH (l)}$

iii. $-1270 \frac{\text{kJ}}{\text{mol}_{\text{rxn}}} \times \frac{1 \text{ mol}_{\text{rxn}}}{2 \text{ mol CO}_2} \times 0.767 \text{ mol CO}_2 = -487 \text{ kJ}$

The combustion reaction released 487 kJ of heat.

c. $q = mc\Delta T, \Delta T = \frac{q}{mc} = \frac{(487 \text{ kJ})\left(\frac{1000 \text{ J}}{1 \text{ kJ}}\right)}{(5.56 \times 10^4 \text{ g})\left(1.005 \frac{\text{J}}{\text{g}\cdot^{\circ}\text{C}}\right)} = 8.71^{\circ}\text{C}$

Final temperature = $21.7^{\circ}\text{C} + 8.71^{\circ}\text{C} = 30.4^{\circ}\text{C}$

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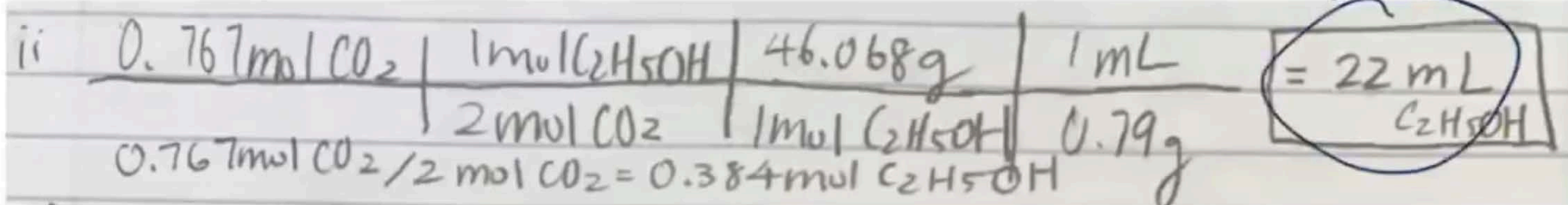


Michael Farab...

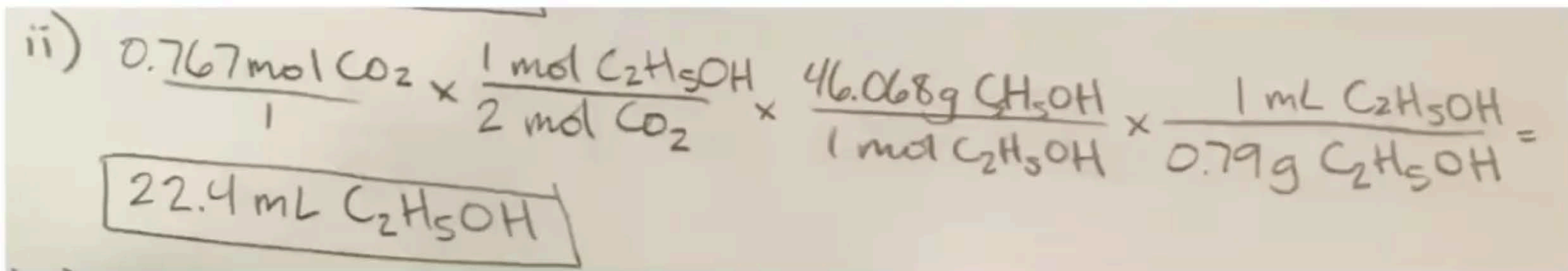
$$\text{iii)} \quad \frac{1270 \text{ kJ}}{\text{mol rxn}}, 0.767 \text{ mols} = \boxed{974.09 \text{ kJ released}}$$

$$\text{ii)} \quad 0.767 \text{ mol CO}_2 \times \frac{3 \text{ mol H}_2\text{O}}{2 \text{ mol CO}_2} = 1.15 \text{ mol H}_2\text{O}$$
$$1.916629 \text{ mol} \times \frac{-1270 \text{ kJ}}{\text{mol rxn}} = \boxed{-2434.1 \text{ kJ}}$$

zoom



ii. $0.776 \text{ mol CO}_2 \times (1 \text{ mol C}_2\text{H}_5\text{OH} / 2 \text{ mol CO}_2) (46.08 \text{ g} / 1 \text{ mol C}_2\text{H}_5\text{OH}) (1 \text{ mL} / 0.79 \text{ g CO}_2) = 22.3 \text{ mL C}_2\text{H}_5\text{OH}$



ii) $0.766 \text{ mol CO}_2(\text{g}) * 1 \text{ mol C}_2\text{H}_5\text{OH}(\text{l}) / 2 \text{ mol CO}_2(\text{g}) * 46.1 \text{ g C}_2\text{H}_5\text{OH}(\text{l}) / 1 \text{ mol} * 1 \text{ mL} / 0.79 \text{ g} = 22.3 \text{ mL C}_2\text{H}_5\text{OH}(\text{l})$

zoom

$$n_{C_2H_5OH} = \frac{0.767}{2} \approx 0.384 \text{ mol} = \frac{m_{C_2H_5OH}}{M_{C_2H_5OH}}$$

$$0.384 = \frac{m_{C_2H_5OH}}{12+12+5+16+1}$$

$$m_{C_2H_5OH} \approx 17.7 = d \times V$$

$$V = \frac{17.7}{0.79} \approx 22.4 \text{ L}$$

$$\text{ii) } .766 \text{ mol } CO_2 \left(\frac{1 \text{ mol } C_2H_5OH}{2 \text{ mol } CO_2} \right) \left(\frac{46.068 \text{ g } C_2H_5OH}{1 \text{ mol } C_2H_5OH} \right) = 17.644044 \approx 17.6 \text{ g } C_2H_5OH$$

$$d = \frac{m}{V}$$

$$.79 = \frac{17.6}{V}$$

$$\frac{17.6}{.79} = \frac{.79V}{.79}$$

$$V = 22.27848101$$

$$V = 22.3 \text{ L}$$



zoom



bi. $PV = nRT$
 $n = \frac{PV}{RT}$ $n = \frac{(1.03 \text{ atm})(18.0 \text{ L})}{(0.08206)(294.7)}$ $n = 0.767 \text{ mol}$

ii. $\frac{1.03 \text{ mol CO}_2}{2 \text{ mol CO}_2} \left| \frac{1 \text{ mol C}_2\text{H}_5\text{OH}}{1 \text{ mol C}_2\text{H}_5\text{OH}} \right| \frac{62.0802 \text{ g}}{0.79 \text{ g}} \left| \frac{1 \text{ mL}}{1 \text{ mL}} \right| = \boxed{40.5 \text{ mL}}$

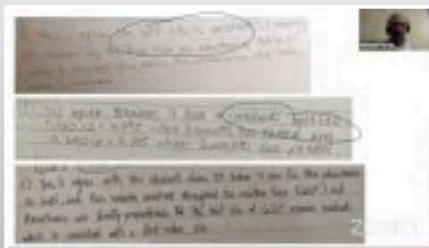
(n) $0.766 \text{ mol CO}_2 \left(\frac{1 \text{ mol C}_2\text{H}_5\text{OH}}{2 \text{ mol CO}_2} \right) \left(\frac{34 \text{ g}}{\text{mol}} \right) \left(\frac{1 \text{ mL}}{0.79 \text{ g}} \right)$
 $= 16.5 \text{ mL C}_2\text{H}_5\text{OH}$



Michael Farab...

$$\text{half life: } \frac{\ln 2}{0.23} = \underline{\underline{3 \text{ min}}}$$

I agree that the reaction is first order because it has a half life of 3 minutes (a unique feature of first order reactions). Also a graph of $\ln[\text{CrO}_7^{2-}]$ vs time yields a negatively sloped linear line.





i. $K_{subp} = (P_{SO_2})^2(P_{O_2})/(P_{SO_3})^2$

ii. $K_{subp} = (.34atm)^2(.16atm)/(.5)^2$

$K_{subP} = 0.074$

e. i. $K_p = (P_{SO_2}^2 * P_{O_2})/(P_{SO_3}^2)$

ii. $K_p = (P_{SO_2}^2 * P_{O_2})/(P_{SO_3}^2) = (0.34^2 * 0.16)/(0.50^2) = 0.074$

e)

i) $K_p = (P_{SO_2})^2(P_{O_2})/(P_{SO_3})^2$

ii) Using the values $P_{SO_2} = 0.33 \text{ atm}$, $P_{O_2} = 0.17 \text{ atm}$, $P_{SO_3} = 0.50 \text{ atm}$ at 3

• min, $K_p = 0.074$

e) i. $K_p = (P_{O_2})(P_{SO_2})^2 / (P_{SO_3})^2$

ii. $K_p = (0.165)(0.34)^2 / (0.50)^2 \rightarrow 0.0763$

zoom