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VCE Chemistry $\frac{3}{4}$
Gas Calculations & Stoichiometry [1.3]
Test Solutions

20 Marks. 2 Minutes Reading. 18 Minutes Writing.

Results:

Test	_____ / 15
Extension	_____ / 5



Section A: Test (15 Marks)

INSTRUCTION: 15 Marks. 1 Minute Reading. 12 Minutes Writing.



Question 1 (4 marks)

Tick whether the following statements are **true** or **false**.

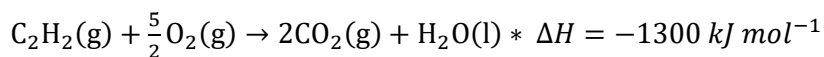
Statement	True	False
a. In complete combustion, the oxygen is assumed to be the limiting reagent.		<input checked="" type="checkbox"/>
b. If the efficiency of a spirit burner experiment is 30%, then 70% of the energy released by the fuel was lost to the environment.	<input checked="" type="checkbox"/>	
c. Experimental ΔH values are typically greater than theoretical ones.		<input checked="" type="checkbox"/>
d. The energy released by a fuel in a spirit burner is never less than the energy absorbed by the water being heated.	<input checked="" type="checkbox"/>	
e. Volume-volume stoichiometry may only be used if the temperature and pressure are constant at SLC.		<input checked="" type="checkbox"/>
f. At Standard Laboratory Conditions (SLC), one mole of every ideal gas occupies 24.8 L of volume.	<input checked="" type="checkbox"/>	
g. In a chemical reaction, the limiting reagent determines the maximum amount of product that can be formed.	<input checked="" type="checkbox"/>	
h. If 5 moles of hydrogen gas reacts with 2 moles of oxygen gas, the reaction will produce 4 moles of water.	<input checked="" type="checkbox"/>	

Space for Personal Notes

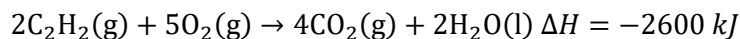
Question 2 (4 marks)

Ethyne, C_2H_2 , is a chemical compound that can be used as a fuel. Upon combustion of one mole of ethyne at SLC, 1300 kJ of energy is released.

- a. Write the balanced thermochemical equation for the complete combustion of C_2H_2 . (2 marks)



or



- b. Calculate the volume of CO_2 produced when 200.0 L of C_2H_2 is completely combusted at standard laboratory conditions (SLC). Justify your working with appropriate steps or reasoning. (2 marks)

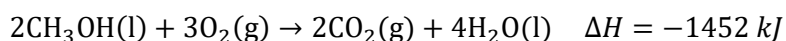
Using volume-volume stoich (as T and P are constant) (1)

$$V(C_2H_2) = 2 \times n(C_2H_2) = 400.0 \text{ L (4 s.f.) (2)}$$

If molar volume approach was used, $V(\text{ethyne}) = 400 \text{ L (3 s.f.)}$

Question 3 (1 mark)

How many kilojoules of energy would be released from the production of 49.6 L of carbon dioxide at 25°C and 100 kPa based on the following thermochemical equation?



$$n(CO_2) = V/V_m = 49.6/24.8 = 2.0 \text{ mol (1)}$$

1452 kJ of energy would be released when 2 mol of CO_2 is evolved, directly observed from the equation provided.

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Question 4 (6 marks)

A sample of butan-1-ol (C_4H_9OH) on a spirit burner which initially weighs 78.70 g undergoes complete combustion. After the combustion is complete, it is found that the spirit burner weighs 74.92 g . The heat energy released is used to heat 350 mL of water at SLC. The temperature of the water rises to 53.20°C .

- a. Calculate the molar heat of combustion of butan-1-ol. (4 marks)

$$\begin{aligned}\Delta m &= 78.70 - 74.92 = 3.78\text{ g} \\ n(C_4H_9OH) &= \frac{m}{M} = \frac{3.78}{4 \times 12 + 10 + 16} = 5.11 \times 10^{-2}\text{ mol} \\ q_{\text{absorbed}} &= mc\Delta T = 350 \times 4.18 \times (53.20 - 25.00) \\ &= 4.13 \times 10^4\text{ J} \\ &= 41.3\text{ kJ} \\ \Delta H &= \frac{q}{n} = \frac{41.3\text{ kJ}}{5.11 \times 10^{-2}\text{ mol}} = -809\text{ kJ/mol (3 s.f.)}\end{aligned}$$

- b. Given that the process carried out is known to be 65.0% efficient, find the new calculated heat of combustion. (1 mark)

$$\Delta H_{\text{new}} = \frac{809}{0.650} = -1.24 \times 10^3\text{ kJ/mol (3 s.f.)}$$

- c. Propose a major reason as to why the experiment was not 100% efficient. (1 mark)

Heat loss to the surroundings \rightarrow not all energy released by butan-1-ol was absorbed by the water.

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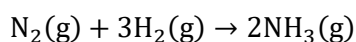
Section B: Extension (5 Marks)

INSTRUCTION: 5 Marks. 1 Minute Reading. 6 Minutes Writing.



Question 5 (5 marks)

A reaction between nitrogen gas and hydrogen gas produces ammonia (NH_3) according to the balanced equation:



If 5.00 moles of nitrogen and 12.0 moles of hydrogen are reacted:

- a. Calculate the volume of NH_3 produced. (3 marks)

H_2 is limiting as only 4 mol of N_2 would be needed to react with it. (1)

$$n(\text{NH}_3) = \frac{2}{3} \times n(\text{H}_2) = \frac{2}{3} \times 12 = 8 \text{ mol (2)}$$

$$V(\text{NH}_3) = n \times V_m = 8 \times 24.8 = 198.4 \text{ L} = 198 \text{ L (3 s.f.) (3)}$$

- b. Determine the amount, in g, of the excess reagent left over. (2 marks)

$$n(\text{N}_2)_{\text{used}} = 4 \text{ mol} \rightarrow n(\text{N}_2)_{\text{excess}} = 5 - 4 = 1 \text{ mol (1)}$$

$$m(\text{N}_2)_{\text{excess}} = n \times M = 1 \times (14 \times 2) = 28.0 \text{ g (2)}$$

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