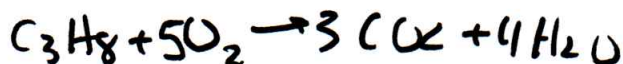


## Enthalpy Calculation and Stoichiometry

Gaseous Propane or (sold as LP or liquid propane) is the common gas used in gas grills. Over the course of grilling you produce 400g of Carbon Dioxide (product) and 12,700kJ of energy. Some of the energy is used for cooking your food and some just warms the air. The oxygen consumed from the atmosphere is considered to be in excess.

1. Write out the reaction that is occurring and balance it.



2. How many moles of  $\text{CO}_2$  were used?

$$44 \text{ g/mol} \quad 400 \text{ g} \cdot \frac{1 \text{ mol}}{44 \text{ g}} = \boxed{9.09 \text{ mol}} \left( \frac{1}{3} \right)$$

3. Determine the  $\Delta H$  for the reaction.

→ Need kJ for 1 mol  $\text{C}_3\text{H}_8$

$$\frac{12,700 \text{ kJ}}{3.03 \text{ mol}} = \boxed{-4191 \text{ kJ/mol}}$$

3.03 mol  $\text{C}_3\text{H}_8$

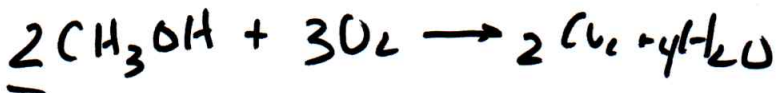
4. What is the original mass of propane used?

$$3.03 \text{ mol} \cdot \frac{44}{1 \text{ mol}} = \boxed{133 \text{ g}}$$

exothermic

In a separate experiment 500g of methanol ( $\text{CH}_3\text{OH}$ ) is used to heat up 1 gallon of water (8.3lbs = 3760g) raising the temperature of the water about 70 degrees.

5. Write out the reaction that is occurring and balance it.



6. How much energy did the water gain?

$$q = m \cdot \Delta T \cdot C$$

$$3760 \cdot 70 \cdot 4.18 = 110,000 \text{ J} \quad (1100 \text{ kJ})$$

7. Determine the  $\Delta H^\circ$  for the reaction.

$$\frac{1100 \text{ kJ}}{15.6 \text{ mol}} = 70.5 \text{ kJ/mol}$$

$$500 \text{ g} \cdot \frac{1 \text{ mol}}{32 \text{ g}} = 15.6$$

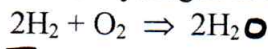
8. What is the kJ/g of methanol?

$$1100 / 500 = 2.2 \text{ kJ/g}$$

Hydrogen burning releases 141 kJ/g.

$$\rightarrow \times 2 \quad \boxed{-141 \text{ kJ } \Delta H} \quad \text{rxn uses 2 moles}$$

9. What is the enthalpy of combustion of hydrogen in the following reaction?



$$\frac{141 \text{ kJ}}{1 \text{ g}} \cdot \frac{2 \text{ g}}{1 \text{ mol}} = -\frac{282 \text{ kJ}}{\text{mol}} \times 2 = -482 \text{ kJ}$$

$\Delta H$