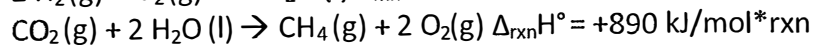
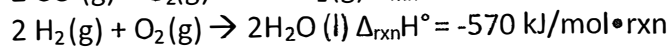
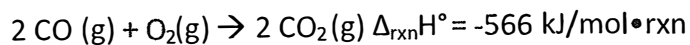


Using the following information, determine the $\Delta_{\text{rxn}}H^\circ$ for $\text{CH}_4(\text{g}) + \text{H}_2\text{O}(\text{l}) \rightarrow \text{CO}(\text{g}) + 3\text{H}_2(\text{g})$



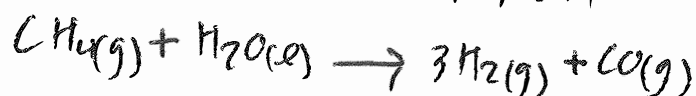
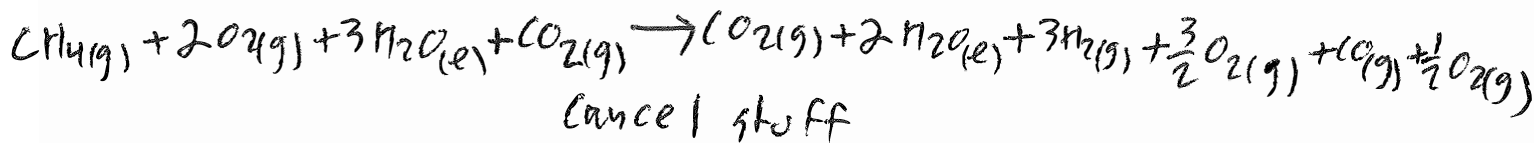
$$\Delta_r H^\circ = \frac{-890 \text{ kJ}}{\text{mol rxn}}$$



$$\Delta_r H^\circ = \frac{+570 \text{ kJ}}{\text{mol rxn}} \times \frac{3}{2}$$



$$\Delta_r H^\circ = \frac{+566 \text{ kJ}}{\text{mol rxn}} \times \frac{1}{2}$$

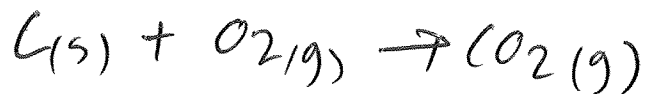


$$\Delta_r H^\circ = -890 + \left(\frac{3}{2} \times 570\right) + \left(\frac{1}{2} \times 566\right) = +248 \text{ kJ/mol rxn}$$

How much energy is required to make 10 kg of $\text{H}_2(\text{g})$

$$10000 \text{ g H}_2 \times \frac{1 \text{ mol}}{2 \text{ g}} = 5000 \text{ mol H}_2 \times \frac{1 \text{ mol rxn}}{3 \text{ mol H}_2} \times \frac{248 \text{ kJ}}{\text{mol rxn}} = 4.1 \times 10^5 \text{ kJ}$$

Assume that coal is pure carbon. Write the balanced chemical equation for the burning of coal.



$$\Delta_r H^\circ = -393 \text{ kJ/mol rxn}$$

How many joules of heat can be generated when 100 pounds of coal is burned? The $\Delta_f H^\circ$ of $\text{H}_2\text{O}(l)$ = -285 kJ/mol and the $\Delta_f H^\circ$ of $\text{CO}_2(g)$ = -393 kJ/mol

$$100 \text{ lb C} \times \frac{1 \text{ kg}}{2.2 \text{ lb}} \times \frac{1000 \text{ g}}{1 \text{ kg}} \times \frac{1 \text{ mol}}{12 \text{ g}} \times \frac{-393 \text{ kJ}}{1 \text{ mol rxn}} = -1.49 \times 10^6 \text{ kJ}$$

How many kilograms of CO_2 will be produced in this process?

$$100 \text{ lb C} \times \frac{1 \text{ kg}}{2.2 \text{ lb}} \times \frac{1000 \text{ g}}{1 \text{ kg}} \times \frac{1 \text{ mol C}}{12 \text{ g}} \times \frac{1 \text{ mol CO}_2}{1 \text{ mol C}} \times \frac{44 \text{ g CO}_2}{1 \text{ mol CO}_2} \times \frac{1 \text{ kg}}{1000 \text{ g}} = 166.7 \text{ kg CO}_2$$

Write the balanced equation for the combustion of natural gas - CH₄ (g)



$$\Delta_r H^\circ [(2 \times -285 \text{ kJ/mol}) + (-393 \text{ kJ/mol})] - (-74 \text{ kJ/mol}) = -889 \frac{\text{kJ}}{\text{mol rxn}}$$

How many joules of heat can be generated when 100 pounds of natural gas is burned? The $\Delta_f H^\circ$ of H₂O (l) = -285 kJ/mol, the $\Delta_f H^\circ$ of CO₂ (g) = -393 kJ/mol and the $\Delta_f H^\circ$ of CH₄ (g) = -74 kJ/mol.

$$100 \text{ lb} \times \frac{\text{CH}_4}{2.21 \text{ lb}} \times \frac{1 \text{ kg}}{1 \text{ kg}} \times \frac{1000 \text{ g}}{1 \text{ kg}} \times \frac{1 \text{ mol}}{16 \text{ g}} \times \frac{-889 \text{ kJ}}{\text{mol rxn}} = -2.53 \times 10^6 \text{ kJ}$$

How many kilograms of CO₂ will be produced in this process?

$$100 \text{ lb} \times \frac{\text{CH}_4}{2.21 \text{ lb}} \times \frac{1 \text{ kg}}{1 \text{ kg}} \times \frac{1000 \text{ g}}{1 \text{ kg}} \times \frac{1 \text{ mol CH}_4}{16 \text{ g CH}_4} \times \frac{1 \text{ mol CO}_2}{1 \text{ mol CH}_4} \times \frac{44 \text{ g CO}_2}{1 \text{ mol CO}_2} \times \frac{1 \text{ kg}}{1000 \text{ g}} = 125 \text{ kg CO}_2$$

Compare coal to natural gas

$$\begin{array}{l} \text{coal} \\ \hline \frac{-1.49 \times 10^6 \text{ kJ}}{166.7 \text{ kg CO}_2} = \frac{-8938 \text{ kJ}}{\text{kg CO}_2} \end{array}$$

$$\begin{array}{l} \text{CH}_4 \\ \hline \frac{-2.53 \times 10^6 \text{ kJ}}{125 \text{ kg CO}_2} = \frac{-20240 \text{ kJ}}{\text{kg CO}_2} \end{array}$$

Consider the combustion of gasoline C_8H_{18} .

Write the balanced chemical equation for this process



You combust 2.8 grams of gasoline in a calorimeter that contains 2.00 L of water and the temperature of the water increases by $15.97^\circ C$. The $\Delta_f H^\circ$ of $H_2O(l) = -285 \text{ kJ/mol}$ and the $\Delta_f H^\circ$ of $CO_2(g) = -393 \text{ kJ/mol}$. What is the $\Delta_f H^\circ$ of C_8H_{18} ?

$$q = (2000g)(4.184 \frac{J}{g^\circ C})(15.97^\circ C) = 133637 \text{ J given off}$$

$$2.8g C_8H_{18} \times \frac{1 \text{ mol}}{114g} = 0.0246 \text{ mol } C_8H_{18}$$

$$\Delta_r H^\circ = \frac{-133637 \text{ J}}{0.0246 \text{ mol}} = -5441 \frac{\text{kJ}}{\text{mol}}$$

$$\left[(8 \times -393 \frac{\text{kJ}}{\text{mol}}) + (9 \times -285 \frac{\text{kJ}}{\text{mol}}) \right] - \Delta_f H^\circ_{C_8H_{18}} = -5441 \frac{\text{kJ}}{\text{mol}}$$

$$-5709 \frac{\text{kJ}}{\text{mol}} - \Delta_f H^\circ_{C_8H_{18}} = -5441 \frac{\text{kJ}}{\text{mol}}$$

$$\Delta_f H^\circ_{C_8H_{18}} = -268 \frac{\text{kJ}}{\text{mol}}$$