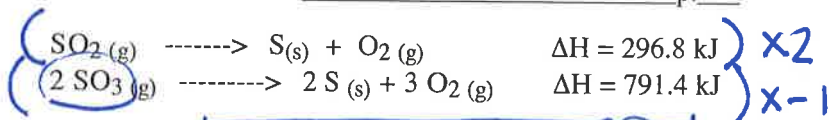


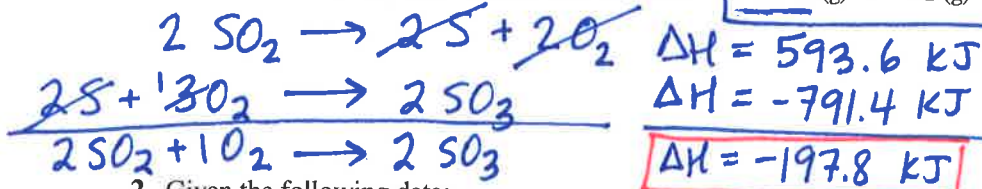
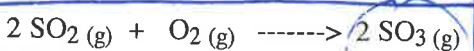
WS 11.3 Hess's Law

Name: _____ p. _____

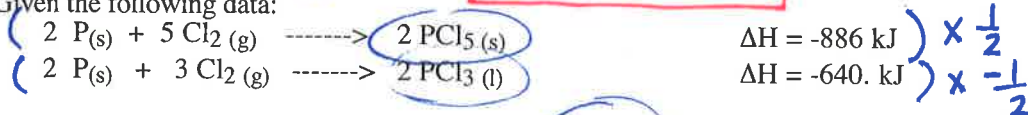
1. Given the following data:



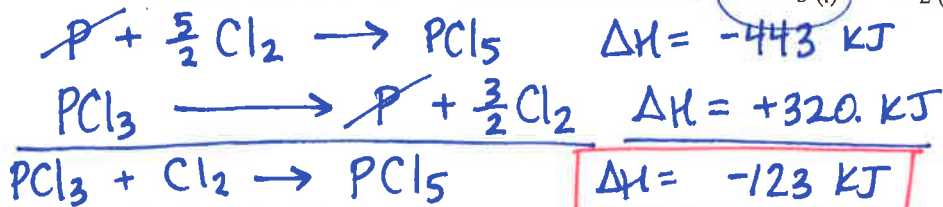
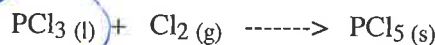
Use Hess's Law to calculate ΔH for this reaction:



2. Given the following data:



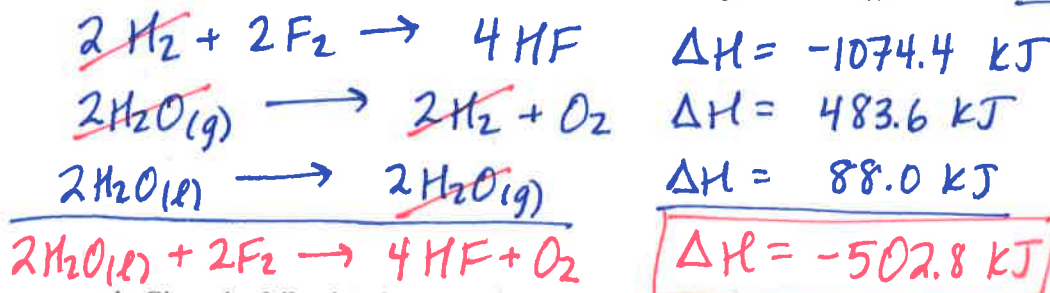
Use Hess's Law to calculate ΔH for this reaction:



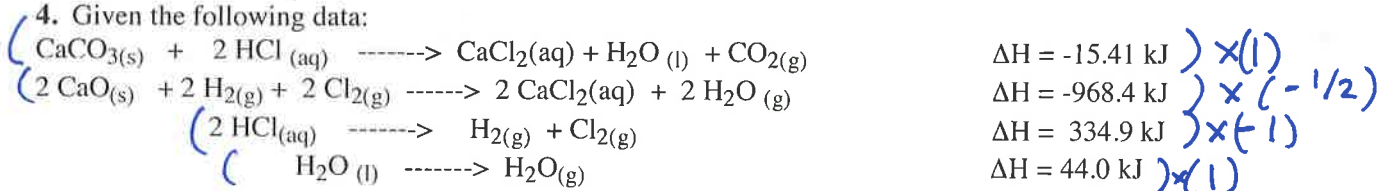
3. Given the following data:



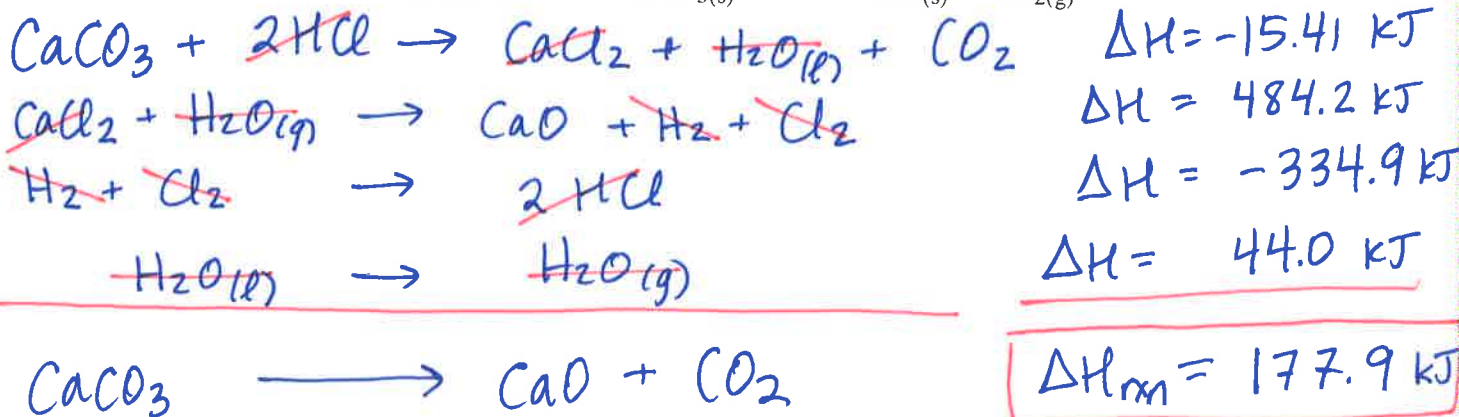
Use Hess's Law to calculate ΔH for this rxn: $2 \text{F}_2(\text{g}) + 2 \text{H}_2\text{O}(\text{l}) \text{ -----} \rightarrow 4 \text{HF}(\text{g}) + \text{O}_2(\text{g})$



4. Given the following data:



Use Hess's Law to calculate ΔH for this rxn: $\text{CaCO}_3(\text{s}) \text{ -----} \rightarrow \text{CaO}(\text{s}) + \text{CO}_2(\text{g})$



5. Olive oil contains many types of fat molecules. A typical molecular formula for olive oil is $C_{57}H_{104}O_6$, which has a molecular weight of 885.4 amu. When you metabolize ("burn") olive oil in your body, the following reaction occurs:



a. Calculate the energy (in kilojoules) produced when 1.00 gram of $C_{57}H_{104}O_6$ is metabolized.

$$(1.00 \text{ g } C_{57}H_{104}O_6) \left(\frac{1 \text{ mole fat}}{885.4 \text{ g}} \right) \left(\frac{33429 \text{ kJ}}{1 \text{ mole fat}} \right) = 37.7558 \rightarrow \boxed{37.8 \text{ kJ}}$$

b. Convert the answer in (a) to kilocalories (1 kcal = 4.184 kJ)

$$(37.7558 \text{ kJ}) \left(\frac{1 \text{ kcal}}{4.184 \text{ kJ}} \right) = \boxed{9.02 \text{ kcal}} \text{ (or 9.02 "food calories")}$$

c. 1 gram of fat typically contains about 9 dietary calories. How does olive oil compare?

It has the same "energy content" as a typical fat.

d. If 1.00 grams of $C_{57}H_{104}O_6$ react, how many grams of carbon dioxide gas would be produced (exhaled)?

$$(1.00 \text{ g } C_{57}H_{104}O_6) \left(\frac{1 \text{ mole}}{885.4 \text{ g}} \right) \left(\frac{57 \text{ mole } CO_2}{1 \text{ mole } C_{57}H_{104}O_6} \right) \left(\frac{44.0098 \text{ g}}{1 \text{ mole}} \right) = \boxed{2.83 \text{ g } CO_2}$$

6. Ethanol (ethyl alcohol) has the formula $C_2H_5OH(l)$.

a. Write a chemical equation for ethanol burning. Balance and include phase subscripts.



b. Without looking at any numbers, predict whether the above reaction will be exo- or endo-thermic, and explain your answer.

It should be exothermic since it is a combustion rxn.

When 10.00 grams of ethanol are combusted under constant pressure conditions, the heat released is enough to increase the temperature of 2000. g of water from 10.0°C to 42.0°C .

c. Determine the heat energy absorbed/released (which is it?) by the water, in calories.

$$q = mc\Delta T = (2000. \text{ g}) (1.000 \frac{\text{cal}}{\text{g}^\circ\text{C}}) (32.0^\circ\text{C}) = 64000 \rightarrow \boxed{6.40 \times 10^4 \text{ cal}}$$

d. Convert your answer in (c) to kilojoules.

$$(64000 \text{ cal}) \left(\frac{4.184 \text{ J}}{1 \text{ cal}} \right) \left(\frac{1 \text{ kJ}}{1000 \text{ J}} \right) = 267.776 \text{ kJ} \rightarrow \boxed{268 \text{ kJ}}$$

e. Determine how many moles of ethanol burned in the experiment.

$$(10.00 \text{ g}) \left(\frac{1 \text{ mole}}{46.0688 \text{ g}} \right) = 0.2170666 \rightarrow \boxed{0.2171 \text{ moles ethanol}}$$

f. Determine ΔH_{rxn} for the combustion of ethanol in kJ/mole.

$$267.776 \text{ kJ} / 0.2170666 \text{ mole} = 1233.61 \frac{\text{kJ}}{\text{mole}} \rightarrow \boxed{-1230 \frac{\text{kJ}}{\text{mole}}}$$

g. Write the heat term into the equation in (a).

must be negative since the rxn caused H₂O to gain heat so rxn is exo and it is combustion!