Name _____ KEY _____ Lab TA/time _____

1.   (4 pts) The work done when a gas is compressed in a cylinder is 580.7 J. A heat transfer of 85.68 kJ occurs from the gas to the surroundings. Calculate ΔE of the gas in kJ.

w > 0 (positive):   work done **ON** **system** by the surroundings
w < 0 (negative):   work done **BY** **system** on the surroundings
q > 0 (positive):   system gains heat (heat goes into the system from the surroundings)
                                  - **endo**thermic
q < 0 (negative):   system loses heat (heat leaves the system to the surroundings)
                                  - **exo**thermic

$$\Delta E \;=\; q + w \;=\; q - P\Delta V \qquad \text{(for pressure-volume work at constant volume)}$$

If the piston is pushed down on the system (compressed) the **surroundings** is doing work **on** the **system**.

This means **w > 0, positive** (energy is entering the system).

The process is **exo**thermic, energy is **leaving** the **system** and therefore q < 0.

$$\Delta E \;=\; q \;+\; w \;=\; -85.6\underline{\mathbf{8}}\ \text{kJ} + 0.580\underline{7}\ \text{kJ} \;=\; -85.0\underline{9}93\ \text{kJ} \;=\; -85.10\ \text{kJ}$$

2.   (4 pts) A piston is pushed down on a reaction vessel containing a gas phase reaction.

w > 0 (positive):   work done **ON** **system** by the surroundings
w < 0 (negative):   work done **BY** **system** on the surroundings
q > 0 (positive):   system gains heat (heat goes into the system from the surroundings)
                                  - **endo**thermic
q < 0 (negative):   system loses heat (heat leaves the system to the surroundings)
                                  - **exo**thermic

$$\Delta E \;=\; q + w \;=\; q - P\Delta V \qquad \text{(for pressure-volume work at constant volume)}$$

a) Does the system do work on the surroundings or is work done on the system by the surroundings?

If the piston is pushed down on the system the **surroundings** is **doing work on** the **system**.

2) (cont.)

b) What would be the sign of the work, w?

If a **surroundings does work ON** the **system** energy is entering the system. Anytime energy enters the system it's positive.

w > 0, positive (energy is entering the system)

c) If the process is also endothermic is heat being added to or removed from the system?

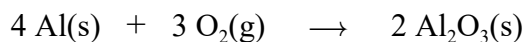Since the process is **endo**thermic, energy is **entering** the **system** and therefore q > 0.

d) What would be the sign of ΔE or can that not be determined based on the information given? Explain!

What is the sign of ΔE if w > 0 and q > 0?

$$\Delta E = q + w$$
$$(+) \quad (+) + (+)$$

ΔE > 0 (+) if both w > 0 (+) and q > 0 (+).

3.   (3 pts) The reaction below was carried out in a bomb calorimeter.

$$4 \, Al(s) \; + \; 3 \, O_2(g) \; \longrightarrow \; 2 \, Al_2O_3(s)$$

A thermometer in the calorimeter registered an increase in temperature. Choose words correctly to complete the following statements for this system.

The reaction is _____. (exothermic, endothermic)
The products have _____(lower, higher) heat content than the reactants.
The quantity of heat determined is a measure of _____. (ΔH, ΔE)

a) exothermic	lower	ΔH
b) exothermic	higher	ΔE
c)* exothermic	**lower**	**ΔE**
d) endothermic	lower	ΔH
e) endothermic	higher	ΔE

$\Delta H_{rxn} > 0$ **endo**thermic, energy is **absorbed** from the surroundings (required)
        Think of heat as a reactant (required). It might feel cold to the touch.

$\Delta H_{rxn} < 0$ **exo**thermic, energy is **released** to the surroundings (given off)
        Think of heat as a product (produced or given off). It might feel hot to the touch.

A thermometer in the **calorimeter** registered an **iucrease** in **temperature**. That means the **surroundings gained** heat. That means the **system lost** heat. Thus, the process (rxn) is **exothermic**.

For the above reaction, $\Delta H_{rxn} < 0$ since the reaction is **exo**thermic and **heat** is **released**.

$$\Delta H_{rxn} = H_{final} - H_{initial}$$

Since $\Delta H_{rxn} < 0$, $H_{products} < H_{reactants}$

        Thus, **products** have **lower enthalpy** (**heat content**) **than** the **reactants**

The rxn. is carried out in a bomb calorimeter at **constant volume** so the **heat** is **equal** to **internal energy**. The above explanation about ΔH applies to ΔE as well.

        $\Delta H = q_p$ enthalpy is heat at constant <u>pressure</u>

        $\Delta E = q_v$ **internal energy is heat at constant <u>volume</u>**

4. (5 pts) Given the following equation, how many **grams** of **acetylene** gas are required to produce 1208 kcal of heat energy by combustion with oxygen? (Atomic weights: C = 12.01, H = 1.008, O = 16.00; Mol. Wts.: C_2H_2 = 26.04, O_2 = 32.00, CO_2 = 44.01, H_2O = 18.02)

$$2\ C_2H_2\ (g)\ +\ 5\ O_2\ (g)\ \longrightarrow\ 4\ CO_2\ (g)\ +\ 2\ H_2O\ (g)\qquad \Delta H\ =\ -302\ kcal$$
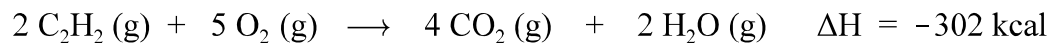
This is an exothermic rx ($\Delta H < 0$) and 302 kcal of heat are released (as product) per mole of rxn (i.e. for eqn. with quantities as written)

$$
\begin{aligned}
2\ \text{mol}\ C_2H_2\ &=\ 302\ \text{kcal released}\\
5\ \text{mol}\ O_2\ &=\ 302\ \text{kcal released}\\
4\ \text{mol}\ CO_2\ &=\ 302\ \text{kcal released}\\
2\ \text{mol}\ H_2O\ &=\ 302\ \text{kcal released}
\end{aligned}
$$

$$?\ g\ C_2H_2\ =\ 1208\ kcal\ \times\ \frac{2\ \text{mol}\ C_2H_2}{302\ \text{kcal}}\ \times\ \frac{26.04\ g\ C_2H_2}{1\ \text{mol}\ C_2H_2}$$

$$=\ 208\ g\ C_2H_2$$

208 g C_2H_2 will produce 1208 kcal

5.   (6 pts) A 28.2 g sample of a metal was heated to 99.81 °C. It was placed in 150.0 g of water at 23.50 °C. After the metal cools, the final temperature of the metal and water is 25.011°C. Calculate the **specific heat** of the **metal** (in J/g•°C), assuming no heat was lost to the calorimeter. Specific heat of water = 4.184 J/g•°C

For calorimetry problems you should know what heat capacity, C, is and how molar heat capacity, C_m, and specific heat, C_s, are related to it and how they all related to heat transfer and temperature change.

Heat Capacity: The amount of heat required to raise the temperature of a substance by 1 °C. Often used when you don't have pure or uniform substances (such as a styrofoam cup in calorimetry lab). This is an extensive property (depends on how much is present, how big the cup is).

$$C = q/\Delta T$$

Molar Heat Capacity: Heat Capacity of one mole of a substance (heat cap. per mol). The amount of heat required to raise the temp. of 1 g of a substance by 1 °C. This is most often used for a pure substance and is an intensive property (while the heat added is extensive).

$$C_m = C/mol$$

Specific Heat: Heat Capacity of one gram of a substance (heat cap. per gram). The amount of heat required to raise the temp. of 1 g of a substance by 1 °C. This is most often used for a pure substance and is an intensive property (while the heat added is extensive).

$$C_s = C/gram$$

This problem is very similar to what you do in the calorimetry lab except the problem assumes no heat is lost to the calorimeter (it has a heat capacity of zero). Heat was gained by the H_2O since it's temp. increased (from 23.5 °C to 25.0 °C) and heat was lost by the Ni since it's temp. decreases (from 99.8 °C to 25.0 °C).

$$q = m \times C_s \times \Delta T \quad \text{Heat} = (\text{mass}) \cdot (\text{specific heat}) \cdot (\text{temperature change})$$

$$q_{\text{gained by H2O}} = {}^- q_{\text{lost by metal}}$$

$C_s = 4.184$ J/g·°C (the specific heat of H_2O to 4 s.f. - from the first sentence)

$$q_{H2O} = m \cdot C_s \cdot \Delta T = 150.0 \text{ g} \times (4.184 \text{ J/g·°C}) \times (25.01 \text{ °C} - 23.51 \text{ °C})$$

$q_{H2O} = 94\underline{1}.4$ J   (2 s.f. since 25.01 − 23.51 = 1.54, 3 s.f.)

$q_{metal} = {}^- 94\underline{1}.4$ J & $q_{metal} = m \cdot C_s \cdot \Delta T$

$- 94\underline{1}.4$ J $= (28.2 \text{ g}) \bullet C_{s,metal} \bullet (25.01 \text{ °C} - 99.81 \text{ °C}) = (-21\underline{0}9.36 \text{ g} \bullet \text{°C}) \bullet C_{s,metal}$

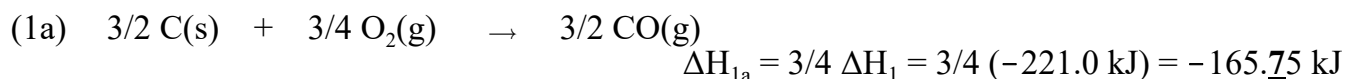$C_{s,metal} = {}^- 94\underline{1}.4$ J $/ (-21\underline{0}9.36 \text{ g} \bullet \text{°C}) = 0.446\underline{2}$ J/g·°C (technically only 3 s.f.)

$C_{s,metal} = 0.446$ J/g•°C

6. (6 pts) Using reactions (1) and (2) calculate the ΔH_{rxn} for reaction (3). **ALSO** is reaction (3) endothermic or exothermic? **Show your work**.

(1) $2\ C(s)\ +\ O_2(g)\ \rightarrow\ 2\ CO(g)$ $\Delta H_1 = -221.0\ kJ$

(2) $2\ C(s)\ +\ O_2(g)\ +\ 4\ H_2(g)\ \rightarrow\ 2\ CH_3OH(g)$ $\Delta H_2 = -402.4\ kJ$

(3) $3/2\ CH_3OH(g)\ \rightarrow\ 3/2\ CO(g)\ +\ 3\ H_2(g)$ $\Delta H = ?$
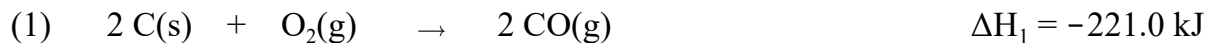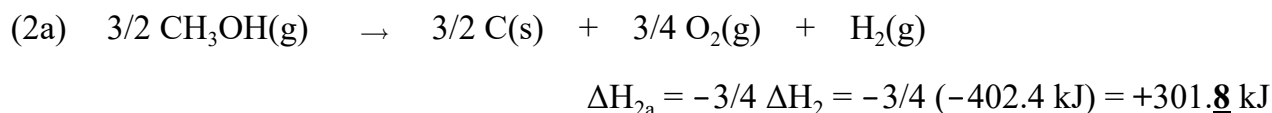
When you reverse an eqn. the ΔH changes sign. Reversing an eqn. is the same as multiplying the original eqn by -1. When you multiply by some factor the ΔH gets multiplied by that factor.
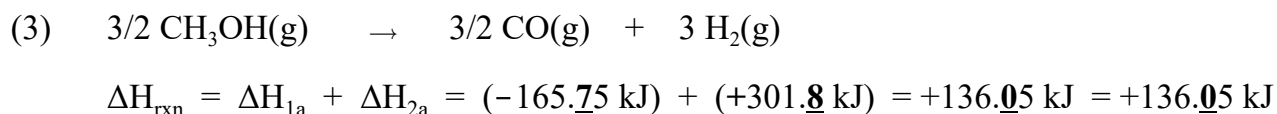
Need 3/2 CO on the right and multiply eqn (1) by 3/4 (multiply by 3/4).

(1a) $3/2\ C(s)\ +\ 3/4\ O_2(g)\ \rightarrow\ 3/2\ CO(g)$
$$\Delta H_{1a} = 3/4\ \Delta H_1 = 3/4\ (-221.0\ kJ) = -165.\underline{7}5\ kJ$$

Need 3/2 C(s) on the right so multiply eqn (2) by $-3/4$. This also then gives the proper numbers of O_2 on the right to cancel those in (1a) and the proper number of H_2 and CH_3OH molecules in the correct locations (right and left, respectively).

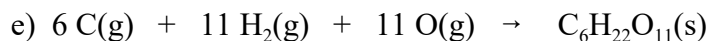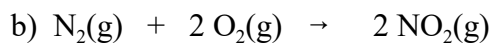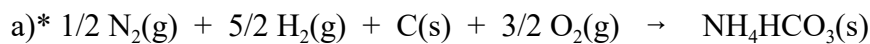(2a) $3/2\ CH_3OH(g)\ \rightarrow\ 3/2\ C(s)\ +\ 3/4\ O_2(g)\ +\ H_2(g)$

$$\Delta H_{2a} = -3/4\ \Delta H_2 = -3/4\ (-402.4\ kJ) = +301.\underline{8}\ kJ$$

Now add eqns 1a and 2a to get eqn 3. The ΔH_{rxn} values are also added.

(3) $3/2\ CH_3OH(g)\ \rightarrow\ 3/2\ CO(g)\ +\ 3\ H_2(g)$

$\Delta H_{rxn}\ =\ \Delta H_{1a}\ +\ \Delta H_{2a}\ =\ (-165.\underline{7}5\ kJ)\ +\ (+301.\underline{8}\ kJ)\ =+136.\underline{0}5\ kJ\ =+136.\underline{0}5\ kJ$

Endothermic (a <u>positive</u> ΔH; energy is required, as a reactant)

7.   (2 pts) Which of the following reactions corresponds to a heat of formation, $\Delta H°_f$?

a)* $1/2\ N_2(g)\ +\ 5/2\ H_2(g)\ +\ C(s)\ +\ 3/2\ O_2(g)\ \rightarrow\ NH_4HCO_3(s)$

b) $N_2(g)\ +\ 2\ O_2(g)\ \rightarrow\ 2\ NO_2(g)$

c) $6\ C(s)\ +\ 12\ H(g)\ \rightarrow\ C_6H_{12}(\ell)$

d) $N(g)\ +\ 4\ H(g)\ +\ Br(g)\ \rightarrow\ NH_4Br(\ell)$

e) $6\ C(g)\ +\ 11\ H_2(g)\ +\ 11\ O(g)\ \rightarrow\ C_6H_{22}O_{11}(s)$

Standard Enthalpy (heat) of Formation: Change in enthalpy for a reaction which forms ONE mole of a compound from its elements with all substances in their standard states.

$$\Delta H_f°$$

Standard state: pure form at atmospheric pressure (and some temperature of interest). A specific temp. is NOT part of the definition. Thus for H_2O at 1 atm and 25 °C it's standard state would be a liquid. For H_2O at 1 atm and 110 °C it's standard state would be a gas.

The only reaction given which produces 1 mole of a substance in its standard state from it's elements is

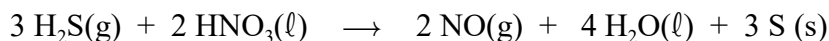a) $1/2\ N_2(g)\ +\ 5/2\ H_2(g)\ +\ C(s)\ +\ 3/2\ O_2(g)\ \rightarrow\ NH_4HCO_3(s)$
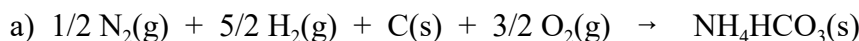
Note: You may see fractions in these balanced eqns. for $\Delta H_f°$ since this is for molar quantitites.

8.   (4 pts) Determine $\Delta H°$ (kJ) for the following reaction using the listed heats of formation.

Use $\Delta H_f°$ to determine the ΔH_{rxn}. Based on Hess's Law.

$$\Delta H° \ =\ \sum_{products} n\,H^0_{f,\ products}\ -\ \sum_{reactants} m\,H^0_{f,\ reactants}$$

$3\ H_2S(g)\ +\ 2\ HNO_3(\ell)\ \longrightarrow\ 2\ NO(g)\ +\ 4\ H_2O(\ell)\ +\ 3\ S\ (s)$

$H_2S\ (g)$	$\Delta H_f° =$	-20.6 kJ/mol
$HNO_3(\ell)$	$\Delta H_f° =$	-174.1 kJ/mol
$NO(g)$	$\Delta H_f° =$	90.25 kJ/mol
$H_2O\ (\ell)$	$\Delta H_f° =$	-285.6 kJ/mol

$\Delta H° =$ [(2 mol)(90.25 kJ/mol) + (4 mol)(-285.6 kJ/mol) + (3 mol)(0 kJ/mol)] −
[(3 mol)(-20.6 kJ/mol) + (2 mol)(-174.1 kJ/mol)]

= [− 961.9 kJ] − [− 410 kJ]

= − 551.9 kJ

9. (3 pts) What is the wavelength of a photon with an energy of 2.69×10^{-19} J ?

The energy of a photon is related to it's frequency (which is related to wavelength):

$$E_{photon} = h\nu = \frac{h\,c}{\lambda} \qquad\qquad h = \text{Planck's constant} = 6.626 \times 10^{-34} \text{ J}\bullet\text{s}$$

What is the energy of a photon with a wavelength of 740 nm?

$$\lambda = \frac{h\,c}{E_{photon}} = \frac{(6.626 \times 10^{-34} \text{ J}\bullet\text{s})\,(3.00 \times 10^{8} \text{ m/s})}{2.69 \times 10^{-19} \text{ J}} = 739 \times 10^{-9} \text{ m}$$

$$\lambda = \textbf{739 } \text{nm}$$

10. (4 pts) What is the energy (kJ/mol) of one mole of photons which have a wavelength of 250 nm?

The energy of a photon is related to it's frequency (which is related to wavelength):

$$E_{photon} = h\nu = \frac{h\,c}{\lambda} \qquad\qquad h = \text{Planck's constant} = 6.626 \times 10^{-34} \text{ J}\bullet\text{s}$$

What is the energy of a single photon with a wavelength of 250 nm?

$$E_{photon} = \frac{h\,c}{\lambda} = \frac{(6.626 \times 10^{-34} \text{ J}\bullet\text{s})\,(3.00 \times 10^{8} \text{ m/s})}{(250 \times 10^{-9} \text{ m})} = 7.95 \times 10^{-19} \text{ J}$$

Want **energy** of a **_mole_** of **photons**:

$$? \text{ kJ} = \frac{7.95 \times 10^{-19} \text{ J}}{1 \text{ photon}} \times \frac{6.02 \times 10^{23} \text{ photons}}{1 \text{ mol}} \times \frac{1 \text{ kJ}}{10^{3} \text{ J}} = 47\underline{8}.66 \text{ kJ} = 479 \text{ kJ}$$