

## Pretest for January 2016 Chem Lab Exam

### Format:

#### **Part 1:**

- You and your partner will be given a procedure to do a Hess-Law type lab involving different compounds than the ones we used in a previous lab.
- The equipment (thermometer, Styrofoam cup, lid, balance, graduated cylinder) will be similar.

**Part 2:** Individually you will be given a problem that can be solved using the data from your lab along with other data that will be given to you.

You also need to report measurements and report answers to calculations while respecting sig figs.

You are also responsible for knowing how to do error analysis and how to report percent error.

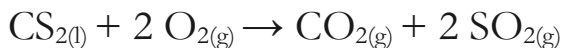
### Sample Questions:

1. a) What could easily go wrong while you are stirring an exothermic reaction?  
b) How do you minimize such an error?
2. In calorimetry experiments, what assumption do we usually (and will again) make regarding the density and specific heat of aqueous solutions.
3. Why does the accurate measurement of volumes and mass matter in a Hess Law experiment?
4. Find the percent error associated with each measurement:  
a)  $75.00 \pm 0.05$  ml  
b)  $3.00 \pm 0.05$  g
5. In #3a, what was the smallest division on the graduated cylinder?

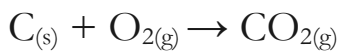
6. What's wrong with the way this student began his conclusion?

*The purpose of this experiment was to discover an efficient means of preventing our shoes from slipping off our feet.*

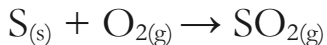
7. a) What's wrong with listing calculation errors as error sources?  
b) Saying that procedure was not followed properly—is that a good error source?  
c) What are good systemic error sources (part of the design of experiment and beyond your control) in calorimetry experiments?
8. What is the value for  $\Delta H$  for the following reaction?



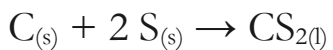
Given:



$$\Delta H_f = -393.5 \text{ kJ/mol}$$



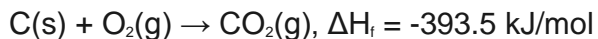
$$\Delta H_f = -296.8 \text{ kJ/mol}$$



$$\Delta H_f = 87.9 \text{ kJ/mol}$$

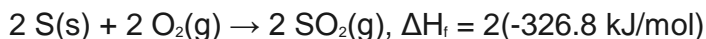
8. Solution from abooutchemistry.com:

We need one CO<sub>2</sub> and the first reaction has one CO<sub>2</sub> on the product side.

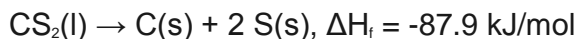


This gives us the CO<sub>2</sub> we need on the product side and one of the O<sub>2</sub> moles we need on the reactant side.

To get two more O<sub>2</sub> moles, use the second equation and multiply it by two. Remember to multiply the ΔH<sub>f</sub> by two as well.



Now we have two extra S and one extra C molecule on the reactant side we don't need. The [third reaction](#) also has two S and one C [on the reactant side](#). Reverse this reaction to bring the molecules to the product side. Remember to change the sign on ΔH.



When all three reactions are added, the extra two sulfur and one [extra carbon atoms](#) are cancelled out, leaving the target reaction. All that remains is adding up the values of ΔH.

$$\Delta H = -393.5 \text{ kJ/mol} + 2(-296.8 \text{ kJ/mol}) + (-87.9 \text{ kJ/mol})$$

$$\Delta H = -393.5 \text{ kJ/mol} - 593.6 \text{ kJ/mol} - 87.9 \text{ kJ/mol}$$

$$\Delta H = -1075.0 \text{ kJ/mol}$$