

CH 204
Experiment 8 Post-Lab Hints
Spring 2007

For **problems 1 – 3**, we're looking at a situation where something hot loses heat to something cold (actually, *two* somethings cold – the water and the calorimeter). In all these cases,

$$\text{heat gained by the cold stuff} = \text{heat lost by the hot stuff}$$

The amount of heat lost is always a negative number, so to correct for the fact that the hot side is always negative and the cold side is always positive, we put a negative sign on the hot side of the equation to turn that side positive as well:

$$q_{\text{gained}} = -q_{\text{lost}} \quad (1)$$

(The heat lost is always negative because $\Delta T = T_{\text{final}} - T_{\text{initial}}$, and since the hot material loses heat and cools down, its final temperature is always lower than its initial temperature.) We use the symbol q for heat, and the units are Joules (J) or kilojoules (kJ).

In all three of these problems, **the heat lost by the hot object** is equal to its mass (m_H) times its specific heat capacity (c_H) times its change in temperature (ΔT_H):

$$q_{\text{lost}} = m_H c_H \Delta T_H \quad (2)$$

The heat gained by the water in the calorimeter is equal to the mass of the water (m_C) times its specific heat capacity (c_s) times its change in temperature (ΔT_C): *heat gained by water* = $m_C c_s \Delta T_C$

The subscripts C and H in these equations refer to *Cold* and *Hot*. The subscript s on C_s refers to *solvent*, which in our case is always water. Note that the change in temperature for the hot side (ΔT_H) and for the cold side (ΔT_C) will be different because the hot and cold sides have different T_{initial} values.

When you put something hot into a calorimeter, the calorimeter itself also gains some heat and goes up in temperature. **The amount of heat absorbed by the calorimeter** is equal to its heat capacity times its temperature change, or *heat gained by calorimeter* = $C_{\text{cal}} \Delta T_C$. The ΔT_C for the calorimeter is the same as the ΔT_C for the cold water. (We don't include the mass of the calorimeter in the calculation because the mass of the calorimeter is already taken into account in the C_{cal} value.)

So the total heat gained by the cold side is equal to the heat gained by the water plus the heat gained by the calorimeter, or

$$q_{\text{gained}} = m_C c_s \Delta T_C + C_{\text{cal}} \Delta T_C \quad (3)$$

If you combine equations (2) and (3) into equation (1), you get

$$m_C c_s \Delta T_C + C_{\text{cal}} \Delta T_C = -m_H c_H \Delta T_H \quad (4)$$

Equation (4) is used to solve the first three post-lab problems. In each of these problems you will know (or be able to determine) all the variables except one, and that is what you are solving for.

In **problem 1** you are solving for the *heat absorbed by the calorimeter*:

$$m_C c_s \Delta T_C + \underline{C_{cal}} \Delta T_C = -m_H c_H \Delta T_H$$

In **problem 2**, you are solving for the *specific heat capacity of the metal*:

$$m_C c_s \Delta T_C + C_{cal} \Delta T_C = -m_H \underline{c_H} \Delta T_H$$

In **problem 3**, you are solving for the *mass of water in the calorimeter*:

$$\underline{m_C} c_s \Delta T_C + C_{cal} \Delta T_C = -m_H c_H \Delta T_H$$

Problem 4: The *molar enthalpy of formation* means you are forming *one mole* of NH_3 from its component elements, so the only thing you can have on the right side of the equation is 1 mole of NH_3 , and the only thing you can have on the left side are nitrogen and hydrogen in their elemental forms ($\text{N}_2(\text{g})$ and $\text{H}_2(\text{g})$). So you want to manipulate the chemical equations that you are given so that you have 1 NH_3 on the right, and nothing except N_2 and H_2 on the left. Everything else should cancel out of the equation. Remember that you can have only 1 mole of product, even if that means the reactants have fractional coefficients.

To determine the ΔH for the final reaction, just make sure that whatever you do to the chemical equations, you also do the same thing to the ΔH for that reaction. If you multiply an equation by 2, multiply the ΔH by 2. If you reverse the equation for the reaction, reverse the sign (+/-) on the ΔH . If you add two reactions to get a new reaction, add their ΔH values to get the ΔH for the new reaction.

Problem 5: You're doing the same thing here that you did in problem 4. Make sure you find the *molar* heats of formation for both Fe_3O_4 and for Fe_2O_3 . – one mole of product in each case, and nothing but $\text{Fe}(\text{s})$ and $\text{O}_2(\text{g})$ as reactants.