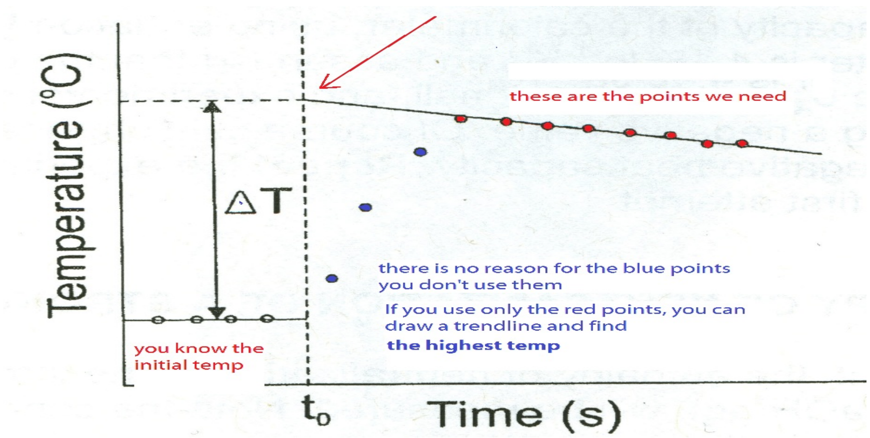
Chem 116: Experiment 10 Thermochemistry

You've done a thermochemistry lab and this one is similar. The main differences are:

1. You will take into account that the calorimeter itself (the holder, thermometer, chemicals) absorbs or gives out heat.

2. You will also take into account the fact that the calorimeter is not perfectly insulated. Heat is lost to or gained from the surroundings.

Let's deal with '2' first as it comes up in all the other parts of the lab. This paragraph concerns exothermic reactions (endothermic reactions are the 'opposite'). We mix the chemicals, the temp goes up and we measure the maximum temperature. But all the time the temperature of the chemicals is above room temp, the calorimeter is losing heat to the outside. What we need to find is the temperature reached if no heat is lost. Here's how we do it: We mix the chemicals (quickly) and measure the temperature while it rises and then falls. We make temp/time plot and extrapolate back to what the temp would be a time zero. Neat eh?



Part A: the Heat Capacity of the Calorimeter

To find this you put some cold water in the calorimeter, and stir in some hot water. You make a temp/time plot and find the temperature at the time the two liquids were mixed

In the manual it says

heat lost by the hot water = heat gained by the cold water + heat gained by the calorimeter

this is wrong: heat gain can never equal heat lost. The lab manual uses $|\Delta T_{hot}|$ which is the absolute change in temp. We are going to do it properly:

$\Delta T = (T_{final} - T_{initial})$ or to make it easier to write: $\Delta T = (T_f - T_i)$

heat lost by the hot water = -(heat gained by the cold water + heat gained by the calorimeter)

$q_{(hot)}$ = -($q_{(cold)}$ + q_p) $q_p = C_p\Delta T$

$mass_{(hot)}$ x sh of water x ΔT = -($mass_{(cold)}$ x sh of water x ΔT + **C_p**ΔT) C_p is the heat capacity
4.184 J/g K 4.184 J/gK at const pressure (J/K)
or (J/K)
C_v is heat capacity
at constant volume

The only unknown is the heat capacity of the calorimeter, **Cp**.

Part B: Enthalpy of the Neutralization of a Strong Acid

HCl + NaOH ---> NaCl + H_2O

You mix 50 mL of the two and make a temp/time plot. You must stir very quickly as the reaction is very fast. You need to figure out how many moles you are reacting and calculate the enthalpy change per mole.

Here's the calculation

$q_{(reaction)}$ = -($q_{(contents)}$ + $q_{(calorimeter)}$)
= -($mass_{(solution)}$ x sh H_2O x ΔT + $C_p\Delta T$)

note: we are assuming the specifc heat of the solution = sh H_2O

You may write an equation for the conclusion: HCl + NaOH ---> NaCl + H_2O ΔH = -xxxx kJ/mol
(if its exothermic)

Part C: Enthalpy of Reaction Between Zinc and Copper(II)sulphate

$Zn_{(s)} + CuSO_{4(aq)}$ ---> $ZnSO_{4(aq)}$ + Cu

We start with 95 mL of water (which weighs ??) and add 5g (2 decimals) $CuSO_4$. Stir until dissolved. Then you add 6.5 g Zn(you might want to check the AW of Zn). Stir and take time/temp measurements until there is no blue color in the solution. There will be excess Zn left.

Make a temp/time plot

So far we've been dealing with solutions which we assume have the same sh as water, but here we have a solution of $ZnSO_4$ (we are told the *specific* heat capacity is 3.97 J/*g*K) as well as solid(Zn and Cu) and we are told the *molar* heat capacity of each is 25 J/*mol*K All the units are in harmony

Note: For every mol of Zn that reacts, a mole of Cu forms. The total number of moles of metal stays the same.

Here's our calculation:

$q_{(reaction)}$ = -($q_{(contents)}$ + $q_{(calorimeter)}$)

= -($mass_{(solution)}$ x specific heat capacity$_{(solution)}$ x ΔT + $mol_{(metal)}$ x molar heat capacity$_{(metal)}$ x ΔT + $C_p\Delta T$)

You may write you conclusion as an equation

I've put up a spreadsheet