WKS 20.1Thermochemistry Even More Review..... since you asked :)

Key concept: Calorimeters are closed systems. Energy is conserved.

Soo.... if the calorimeter cools, that is the amount of energy that the reaction or phase change "ate". If the calorimeter heats up, that energy came from the reaction or phase change.

Formulas: **ΔH = m Cp ΔT**   heating and cooling w/out phase change.

                         ΔH = m Hf energy required for fusion (freezing would be – in sign).

                         ΔH = m Hv energy required for vaporization (condensing – in sign).

Example: The experimental determination of the heat of fusion of ice.

Point: Calorimeter cools, going through a 1 step process, but the ice goes through a two step process. First it melts (ΔH = m Hf), and then the melted ice warms to meet the calorimeter at T_2

Data: Sally Super student conducts the following experiment: She adds 20 g ice to 100 g water in calorimeter and obtains a final temperature of 6.0 °C.

    mass water in calorimeter: m_{water}= 100.0 g

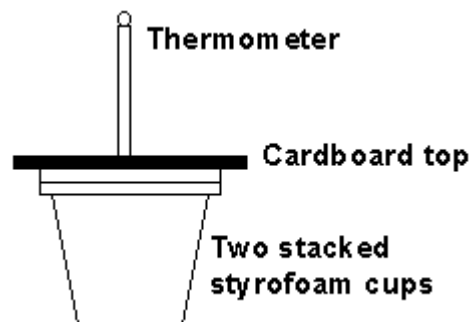   initial temp of calorimeter: T_1= 22.0 °C

    final temp of calorimeter: T_2 = 6.00 °C

    ΔT water = -16.0 °C

    mass of ice: mice = 20.0 g.

    ΔT melted ice = +6.00 °C

1.Heat loss of calorimeter: **ΔH = m Cp ΔT**

   100 g x 4.18 J/g °C x -16.0 °C = -6688 joules *The energy budget is now known, but the signs switch. Substitute these joules below.*

2.Heat gained in phase change (this is a two-step problem: melt ice, warm to T2):

$$m\ Hf\ \ +\ m\ Cp\ ΔT$$

  +6688 joules = (20.0 g x Hf) + (20.0 g x 4.18 J/g °C x 6.00 °C)

   6186.4 J = 20.0 Hf

   Hf = 309.32 J

How does this compare to the actual value of Hf? Calculate percent error:

 Measured-Actual/Actual x 100 %

 (309.2 - 333.6/333.6) x 100 = - 7.3% within our normal limits of accuracy.

Yay!!!! Sally super student did swell!!!

Now for enthalpies of reaction. Again, remember that the heat (enthalpy) lost by one component is gained by the other. We can measure one, and do minor calculations to figure out the other.

Example: Suppose you dissolve 20.0 grams of NaOH in a calorimeter with 100.0 grams of water. Given the specific heat capacity of water is 4.18 J/g C, what will the end temperature of the calorimeter be if its initial temperature is 17 C? The $\Delta H_{solution}$ of NaOH is -44.4 KJ/mole.

1. Notice we need a preliminary step. Our chemical is in grams, but our energy ($\Delta H_{solution}$) is on a per mole basis. We need to know how many moles NaOH are to be dissolved?

20.0 g NaOH	1mole NaOH
	40.0 g Na OH

 = 0.500 mole NaOH.

2. Now we can find the Heat released by the NaOH:

   0.500 mole x -44.4 kJ/mole= -22.2 kJ is released into water (These will be + joules for the calorimeter.

3. Heat (enthalpy) absorbed by water in the calorimeter.

   $\Delta H = m\ C_p\ \Delta T$

+22200 J = 100.0 g x 4.18 J/g °C x ΔT (notice also we need to covert kJ to J to use Cp)

ΔT = +53.1 °C

$T_{final} = T_{initial} + \Delta T$

T_{final} = 70.1 °C

DANGER SODIUM HYDROXIDE

If it weren't so darned caustic, NaOH would make a pretty nice heat pack!!!
Now try some on your own:

1. Compound A is burned in a bomb calorimeter that contains 2.50 liters of water. If the combustion of 0.175 moles of this compound causes the temperature of the water to rise 45.0⁰ C, what is the molar heat of combustion of compound A? The heat capacity of water is 4.184 J / g⁰C.

2. Compound B is burned in a bomb calorimeter that contains 1.50 liters of water. When I burned 50.0 grams of compound B in the calorimeter, the temperature rise of the water in the calorimeter was 35.0^0 C. If the heat of combustion of compound B is 2,150 kJ/mol, what is the molar mass of compound B?

3. The molar heat of combustion of compound C is 1,250 kJ/mol. If I were to burn 0.115 moles of this compound in a bomb calorimeter with a reservoir that holds 2.50 L of water, what would the expected temperature increase be?

Answers: 1: 2.70 x 10^3 kJ/mol, 2: 539 g/mol., 3: 13.8^0 C