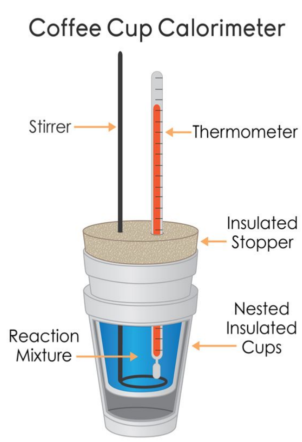
Measuring and Expressing Enthalpy Changes

Calorimetry

Calorimetry is the precise measurement of the heat flow into or out of a system for chemical and physical processes.

Keeping in mind the law of conservation of energy, whatever quantity of heat is given off by the system is equally absorbed by the surroundings. The same applies when the system absorbs heat.

A calorimeter measures the absorption of release of heat in a system.



Because most chemical and physical changes carried out here are open to the atmosphere, these changes occur at constant pressure.

The heat content of a system at constant pressure is the same as a property called the **enthalpy** (H) of the system.

When heat is released or absorbed by the system, the enthalpy changes (ΔH). Because we will be using constant pressure, the terms *heat and enthalpy are the same.*

$$q = \Delta H$$

To measure the enthalpy change for a reaction in an aqueous solution in a calorimeter, you dissolve the reacting chemicals (the system) in known volumes of water (the surroundings).

By measuring the initial and final temperatures (before and after the reaction), and knowing the heat capacity of water, we can calculate the heat absorbed by the surroundings (q_{surr})

$$q_{surr} = mC\Delta T$$

mass of water specific heat of water

Because the heat absorbed by the surroundings is equal to, but has opposite sign of, the heat lost by the system, the enthalpy change looks like:

$$-q_{sys} = -\Delta H = q_{surr} = mC\Delta T$$

The sign of ΔH is negative for an exothermic reaction and positive for an endothermic reaction.

Example

When 25.0 mL of water containing a given amount of HCl at 25.0 ^{o}C is added to 25.0 mL of water containing a given amount of NaOH at 25 ^{o}C in a calorimeter, a reaction occurs. Calculate the enthalpy change (in kJ) during this reaction if the highest temperature observed is 32.0 ^{o}C. Assume the specific heat is that of water.

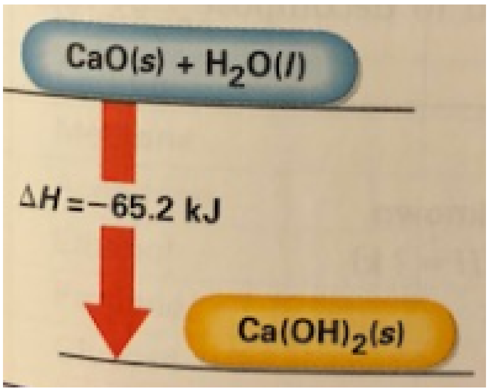
Thermochemical Equations

If you mix calcium oxide with water, the water in the mixture becomes warm.

Is this exothermic or endothermic?

When 1 mol of calcium oxide reacts with 1 mol of water, 1 mol of calcium hydroxide forms and 65.2 kJ of heat is released.

$$CaO(s) + H_2O \longrightarrow Ca(OH)_2(s) + 65.2\ kJ$$



A chemical equation that includes the enthalpy change is called a **thermochemical equation**.

The **heat of reaction** is the enthalpy change for the chemical equation.

You will usually see heats of reaction reported as ΔH.

$$CaO(s) + H_2O \longrightarrow Ca(OH)_2(s) \quad \Delta H = -65.2 \text{ kJ}$$

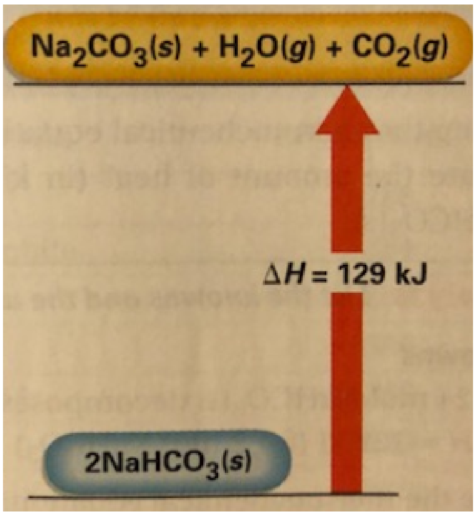
Other reactions absorb heat from their surroundings. For example, baking soda (sodium bicarbonate) decomposes when it is heated. The carbon dioxide released causes a cake to rise while baking. This process is endothermic.

$$2NaHCO_3(s) + 129\text{ kJ} \longrightarrow Na_2CO_3(s) + H_2O(g) + CO_2(g)$$

or

$$2NaHCO_3(s) \longrightarrow Na_2CO_3(s) + H_2O(g) + CO_2(g)$$

$\Delta H = 129\text{ kJ}$



The physical state of reactions also needs to be given in each case.

For example:

$$H_2O(l) \longrightarrow H_2(g) + \frac{1}{2} O_2(g) \qquad \Delta H = 285.8 \text{ kJ}$$

$$H_2O(g) \longrightarrow H_2(g) + \frac{1}{2} O_2(g) \qquad \Delta H = 241.8 \text{ kJ}$$

The vaporization of 1 mol of liquid water to water vapor at 25^{o}C requires an extra 44.0 kJ of heat.

Example

Using the thermochemical equation:

$$2NaHCO_3(s) \longrightarrow Na_2CO_3(s) + H_2O(g) + CO_2(g)$$

$\Delta H = 129$ kJ

Calculate the amount of heat (in kJ) required to decompose 2.24 mol $NaHCO_3(s)$

The **heat of combustion** is the heat of reaction for the complete burning of one mole of substance.

Table 17.2 on page 517 lists heats of combustion for some common substances.

CHEM 11 Reminder

How many moles is 3.5 grams of water?

How many moles is 13 grams of carbon dioxide?

Try questions Pg. 535 # 38-47 and
#12-20 on pages 513-517