Unit 3, Lesson 03: Measuring Enthalpy Changes, ΔH

Most chemical reactions are accompanied by changes in energy.
Chemical reactions can be classified as either exothermic or endothermic, depending on whether they release or absorb chemical energy.

For simplicity, in this course, we will consider the ***system*** to be the physical or chemical reaction that we are studying and the ***surroundings*** to be whatever changes temperature; usually it will be the water in the calorimeter in which the reactions take place.

Chemists always seem to want to know "how much?" By insulating a system from the rest of the universe, we can measure the temperature change of the surroundings caused by the reaction. If the reaction takes place at constant pressure (ie. no gases are used or produced by the reaction so no energy is lost or gained as work), we can calculate the change in enthalpy, ΔH. This process is called calorimetry.

If a reaction is **exothermic**:

- the total energy (including the chemical potential energy) of the products is lower than the reactants
- energy is released by the reaction
- chemical potential energy is converted to thermal kinetic energy as the particles in the system move faster
- the particles in the system are moving faster, which causes the particles in the surroundings to move faster so the temperature of the surroundings (water) goes up
- the difference in the total energy of the system (the reactants and products) is transferred to the surroundings as heat (Q)

$$\boxed{\Delta H \text{ given off by reaction} = Q \text{ gained by surroundings}}$$

If a reaction is **endothermic**:

- the total energy (including the chemical potential energy) of the products is higher than the reactants
- energy is absorbed by the reaction
- chemical potential energy is converted from thermal kinetic energy as the particles in the system move slower
- the particles in the system are moving slower, which causes the particles in the surroundings to move slower so the temperature of the surroundings (water) goes down
- the difference in the total energy of the system (the reactants and products) is absorbed from the surroundings as heat (Q)

$$\boxed{\Delta H \text{ absorbed by reaction} = Q \text{ lost from surroundings}}$$

To calculate the amount of heat lost or gained (Q) by the surroundings, use the equation:

$Q = m \cdot c \cdot \Delta T$ where m = the mass of whatever changes temperature, in grams
c = the specific heat capacity of what changes temp. (in J/g°C)
ΔT = the change in temperature

<u>Specific heat capacity</u> (c) is defined as the amount of heat that is required to raise the temperature of one gram of a substance by one degree Celsius.

- its units are J/g °C (joules per gram per °C)
- it is a physical property of every substance (see explanation on the next page)

To understand what is meant by specific heat capacity, let's go back to our model of cars on the 401:

- the average speed of the cars is represented by temperature
- if a fast-moving car hits a slow-moving car, some of the kinetic energy of the faster car will be transferred to the slower car and the slow-moving car will speed up. The transfer of energy represents heat
- if the slow moving car is a Smart car, its speed will increase a lot when it gets hit. The amount of energy it takes to increase the speed of the Smart car by 1 km/h is very low. This represents a low specific heat capacity (you don't need to put in much energy to speed it up)
- if the slow moving car is a Hummer (as if:), its speed will not increase very much when it gets hit. The amount of energy that it takes to increase the speed of a Hummer by 1 km/h is huge. This represents a high specific heat capacity (you have to put in a lot of energy to speed it up)

Every substance has its own characteristic specific heat capacity, which is a physical property of that substance.

For example:

Pure gold: c = 0.129 J/gºC, which is very low

- one gram of pure gold requires 0.129 J of energy to make 1 g of its atoms move 1ºC faster
- the specific heat capacity of gold is low for two reasons:
  - one gram of gold is not very many atoms (MM of gold is 197 g/mol) and,
  - the metallic bonding holding the atoms of gold together is quite "fluid" so it doesn't take much energy to get the gold atoms moving

Pure aluminum: c = 0.900 J/gºC, which is moderate

- one gram of pure aluminum requires 0.900 J of energy to make 1 g of its atoms move 1ºC faster
- the specific heat capacity of aluminum is higher than for gold because:
  - there are 8x more aluminum atoms in one gram of aluminum than there are in one gram of gold (MM of aluminum is 27 g/mol)
  - aluminum has 3 valence electrons to participate in metallic bonding, so the forces of attraction between atoms is quite strong, so it takes more energy to get the aluminum atoms moving

Pure water: c = 4.184 J/g ºC, which is extremely high

- one gram of liquid water requires 4.18 J of energy to make 1 g of its molecules move 1ºC faster
- the specific heat of water is high for two reasons:
  - one gram of water contains a lot of molecules because water molecules are very light (MM = 18.02 g/mol)
  - there are very strong forces of inter-molecular attraction between water molecules because of the strong polarity of the molecule and hydrogen bonding, so it takes a lot of energy to get water molecules moving
- another way of looking at this is that water has the ability to absorb a lot of energy with only a very small increase in temperature (which is what makes life on Earth possible)

In this course, the reactions are carried out in aqueous solution in a coffee cup calorimeter (2 nested coffee cups). Because the reaction takes place in water, it is water that will change in temperature. Usually the amount of reactants and products in the water is quite small, so the specific heat capacity (c) of water (4.184 J/g ºC) can be used for the surroundings.

Calculating the Amount of Heat (Q) Lost or Gained

Using specific heat capacities, we can calculate the change in thermal energy (Q, heat transferred) as objects are heated or cooled using the equation:

$$\boxed{Q = m \cdot c \cdot}$$

where:
m is the mass of what is changing temperature, in grams
c is the specific heat capacity of what is changing temperature, in J/g ºC
ΔT is the temperature change, in ºC or K

eg. How much heat is required to heat 100.0 g of water from 30.0 ºC to 50.0 ºC? The specific heat capacity of water is 4.184 J/gºC.

Given:
m = 100.0 g
c = 4.184 J/gºC
$\Delta T = T_2 - T_1$
= 50.0 ºC – 30.0 ºC
= 20.0 ºC

Q = $m \cdot c \cdot \Delta T$
= 100.0 g x 4.184 J/gºC x 20.0 ºC
= 8368 J
= 8.37 kJ (3 sig digs)

Therefore, it will take 8.37 kJ of thermal energy (heat) to heat the water from 30.0 ºC to 50.0 ºC.

eg. The temperature of an aluminum engine block goes from 22.0ºC to 118.0 ºC while a car is running. The engine block weighs 16.44 kg. The specific heat of aluminum is 0.900 J/gºC. How much heat was absorbed by the engine block?

Given:
m = 16.44 kg or 16440 g
c = 0.900 J/gºC
$\Delta T = T_2 - T_1$
= 118.0 ºC – 22.0 ºC
= 96.0 ºC

Q = $m \cdot c \cdot \Delta T$
= 16440 g x 0.900 J/gºC x 96.0 ºC
= 1420416 J
= 1.42×10^6 J or 1.42×10^3 kJ (3 sig digs)

Therefore, the engine block absorbed 1.42×10^6 J of thermal energy (3 sig digs)

Homework:
1. Read pages 234 – 235.
2. Answer questions on page 235, do them in the order: 8, 7, 5, 6.