Chemistry 11
Notes on Heat and Calorimetry

Some chemical reactions **release** heat to the surroundings – These are ***exothermic***
Some chemical reactions **absorb** heat from the surroundings – These are ***endothermic***

Heat is a form of energy (which cannot be created or destroyed). A chemical reaction cannot create or destroy energy, it just changes it from one form to another.

All substances have a type of chemical potential energy stored in them. This is called **enthalpy**. The actual definition of enthalpy is *the total energy contained in a system*. For Chemistry 11, it's easiest to think of enthalpy as mainly **chemical potential energy** stored in a substance.

Gasoline contains more stored chemical potential energy than water. Therefore we say that gasoline has more **enthalpy** than water.

An **exothermic** chemical reaction converts the **enthalpy** stored in a substance into **heat**, which is released to the surroundings. For example, when gasoline burns, the enthalpy in the gasoline (mainly C_8H_{18}) is converted into heat:

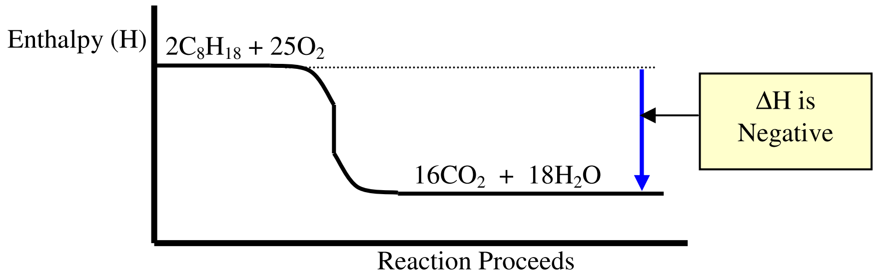
$$\underbrace{2C_8H_{18} + 25O_2}_{\text{Have } \textbf{higher} \text{ enthalpy}} \rightarrow \underbrace{16CO_2 + 18H_2O}_{\text{Have } \textbf{lower} \text{ enthalpy}} + \textbf{heat}$$

Released to the surroundings

The symbol for **enthalpy** is "**H**" (historically, enthalpy used to be called "Heat Content")
The **change in enthalpy** during a reaction is called Δ**H**. (The symbol Δ (*delta*) means "change in")

The reaction above can be shown on a graph. See the graph on the next page….

An enthalpy diagram for an **exothermic** reaction:



So, to summarize, in an **exothermic** reaction:

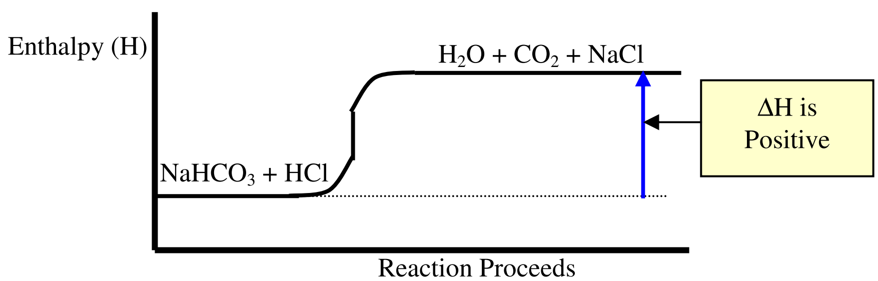
- ΔH is negative
- Products are **lower** than Reactants on the Enthalpy Diagram
- Heat is released to the surroundings

For an **endothermic** reaction:

Eg.) $NaHCO_3 + HCl + \text{heat} \rightarrow H_2O + CO_2 + NaCl$

Heat is **absorbed** from the surroundings and converted into **enthalpy**.

Here is an enthalpy graph for an **endothermic** reaction:



There are **two ways** to show that a reaction is **exothermic**:

1. A "heat term" is written on the right side of a chemical equation to show that heat is given off or produced. This can be the word "heat" or an actual amount of heat in kJ:

   eg.) $2C_8H_{18} + 25O_2 \rightarrow 16CO_2 + 18H_2O +$ **5076 kJ**

   Heat term is on the **right** side of the equation.

2. An equation is written and the **ΔH** is written **beside** it (no "+" sign between):

   eg.) $2C_8H_{18} + 25O_2 \rightarrow 16CO_2 + 18H_2O$ **$\Delta H = -5076$ kJ**

   Notice that for an **exothermic** reaction, the ΔH is NEGATIVE!

In an **endothermic** reaction, the "heat term" would be written on the LEFT side, or the **ΔH** written beside the equation would be **Positive**:

$$2C + 2H_2 + \mathbf{52.3\ kJ} \rightarrow C_2H_4$$

Heat term is on the **left** side of the equation.

$$2C + 2H_2 \rightarrow C_2H_4 \qquad \Delta H = +52.3\ kJ$$

Notice that for an **endothermic** reaction, the ΔH is POSITIVE!

Read pages 119-122 in SW and do Ex. 76-80 on p. 122.

ΔH and Moles

In a chemical equation, the coefficients can stand for "molecules" or for "moles".

If an equation contains a heat term or ΔH is shown, coefficients ALWAYS stand for moles.

For example, given the equation:

$$N_2 + \mathbf{3}H_2 \rightarrow \mathbf{2}NH_3 + \mathbf{46.2\ kJ}$$

It means that:

- If **1 mole** of N_2 is used up, 46.2 kJ of heat are released
- If **3 moles** of H_2 are used up, 46.2 kJ of heat are released
- If **2 moles** of NH_3 are produced, 46.2 kJ of heat are released

As you might guess, this can give us some useful **conversion factors** ! (goody!!!)

Eg.) $\dfrac{46.2kJ}{1molN_2}$ $\qquad \dfrac{46.2kJ}{3molH_2}$ $\qquad \dfrac{46.2kJ}{2molNH_3}$

These wonderful conversion factors can then be used to calculate the amount of heat released or absorbed in a reaction if the moles of substances are known. See the next page…

Example Question:
Find the amount of heat released if 5.0 moles of H_2 are consumed when making ammonia, given the reaction: $N_2 + \mathbf{3}H_2 \rightarrow \mathbf{2}NH_3 + \mathbf{46.2\ kJ}$

Solution:

$$5.0 \text{ mol } H_2 \times \frac{46.2kJ}{3molH_2} = \mathbf{77\ kJ}$$

Another Example:
Find the amount of heat released during the formation of 14.6 moles of NH_3, given the reaction:
$N_2 + \mathbf{3}H_2 \rightarrow \mathbf{2}NH_3 + \mathbf{46.2\ kJ}$

Solution:

$$14.6 \text{ mol } NH_3 \times \frac{46.2kJ}{2molNH_3} = \mathbf{337\ kJ}$$

Now you can do Questions 1-6 on "Hand-In Assignment #9"

Calorimetery

In order for the amount of heat to be measured in a chemical or physical change, **three** things have to be known.

1. The **temperature change** (Δ**t**) (measured with a thermometer)
2. The **mass** of the substance (**m**) (measured with a balance)
3. The "**Heat Capacity**" of the substance (**C**) (this means the heat needed to raise the temperature of 1 kg of the substance by 1 °C)
   Heat Capacity (C) will **always be given** in a problem. The Heat Capacity for water is:
   C_{H_2O} = 4180 J/kg · °C

A simple equation we use to calculate heat is:

$$\boxed{\text{Heat} = m \cdot C \cdot \Delta t}$$

But it's really important to know about the **UNITS**!

"Heat" is in **Joules (J)**, "m" is in **kilograms (kg)**, "C" is in **J/kg · °C** and Δt is in **°C**

See the next page for a couple of example problems:

Example: Given that the heat capacity of water: C_{H_2O} = 4180 J/kg · °C
Calculate the heat required to warm 400.0 g of water from 20 °C to 50 °C.

Solution:
First we have to change 400.0 g to 0.4000 kg and calculate the temperature change (Δt)
50 °C - 20 °C = 30 °C

Next, we write the equation:

Heat = m · C · Δt

Then we plug in the data:

Heat (J) = 0.4000 kg x 4180 J/kg · °C x 30 °C = **50 160 J**

Notice how the "kg" will cancel and the "°C" will cancel, leaving "J" as the unit for the answer. If you want to, you can convert the 50 160 J into 50.16 kJ.

In another type of question, you may be given the "Heat" and asked to find the final temperature. You just have to carefully use the equation and good old algebra!

Example: 75.0 kJ of heat are added to 850.0 g of water initially at 25.0 °C. Calculate the final temperature of the water. C_{H_2O} = 4180 J/kg · °C

Solution: First, the 75.0 kJ **must** be changed to 75,000 J and the 850.0 g of water must be changed to 0.8500 kg.

Now the use the equation to solve to Δt:

Heat = m · C · Δt

$$\Delta \mathbf{t} = \frac{Heat}{m \cdot C} = \frac{75{,}000 J}{0.8500 kg \cdot 4180 \frac{J}{kg \cdot {}^{o} C}} = \mathbf{21.1\ {}^{o}C}$$

Since the **initial** temperature was 25 °C, the **final** temperature will be:

$t_{final} = t_{initial} + \Delta t = 25.0\ {}^{o}C + 21.1\ {}^{o}C =$ **46.1 °C**

Now you can finish the questions on "Hand-In #9"