Announcements

1.) The X-mas exam date has been finalized for December 17th 9 am – 12:00 pm. Room to be announced.

2.) This is the week of November 12th to 16th: Group 1 is doing the Gases tutorial while Group 2 is doing the Molar Volume lab

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Energy is required for a phase change to occur

1 g ice at 0° C (melting point) → 1 g liquid water at 0° C
Latent heat of fusion (melting) q = +333 Jg^{-1}

1 g liquid water at 100° C (boiling point) → 1 g steam at 100° C
Latent heat of vaporization (evaporation) q = +2260 Jg-1

Example:
Suppose 750 g block of ice at -15°C is allowed to melt, and the resulting water warms up to 22°C. How much heat is absorbed in this process?

Useful data:
Specific heat of ice: 2.01 J K^{-1} g^{-1}
Heat of fusion of ice: 6.01 kJ mol^{-1}
Specific heat of water: 4.184 J K^{-1} g^{-1}

Answer

Calculate in 3 steps:

1.) heat the ice

$q_1 = 2.01 JK^{-1}g^{-1} \times 750g \times 15K$
   $= 22612.5 J = 22.61 kJ$

2.) melt the ice (at fixed temperature)

$q_2 = 6.01 kJmol^{-1} \times (750 g/18gmol^{-1})$
   $= 250.42 kJ$

3.) heat the water

$q_3 = 4.184 JK^{-1}g^{-1} \times 750g \times 22K$
   $= 69036 J = 69.04 kJ$
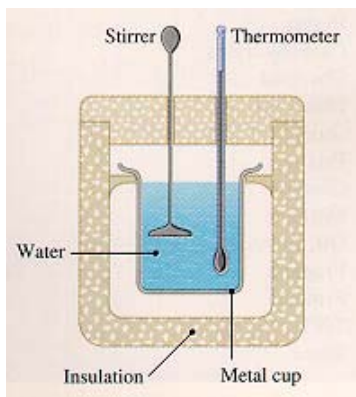
Total $q = q_1 + q_2 + q_3$
   $= 22.61 + 250.42 + 69.04$

$= 342.07 kJ$

Calorimetry (measuring heat flow)

Experimental measurement of q is termed calorimetry.

The apparatus used is called a calorimeter: usually an enclosed metal vessel surrounded by water in an insulated container

For **exothermic** reactions, heat is released   q = -ve

For **endothermic** reactions, heat is absorbed   q = +ve

For an exothermic reaction carried out inside a calorimeter the heat evolved raises the temperature of the calorimeter and the water

$$\therefore \quad -q_{reaction} = q_{calorimeter} + q_{water}$$

For water: $q_{water} = C(water)m\Delta T$

For the calorimeter: $q_{calorimter} = C_C\Delta T$

ΔT is the temperature rise and C_C is the heat capacity of the calorimeter in $J\ K^{-1}$

Example:
1.148 g of benzoic acid ($C_7H_6O_2$) is burned in a calorimeter immersed in 1215 g H_2O.
The water temperature increases from 25.12 to 30.26° C.
What is the q evolved by the combustion reaction in kJ mol^{-1} for benzoic acid?
Useful data: $C_C = 817\ JK^{-1}$; $C(H_2O) = 4.184\ J\ g^{-1}\ K^{-1}$

$$\text{Since } -q = C_C\Delta T + C(H_2O)m\Delta T$$

$\Longrightarrow$ $-q = (817\ JK^{-1})(5.14\ K) + (4.184\ Jg^{-1}K^{-1})(1215\ g)(5.14\ K)$

$$= 30330\ J = 30.33\ kJ \text{ per } 1.148\ g \text{ benzoic acid}$$

Molar mass $C_7H_6O_2 = 122.0\ gmol^{-1}$

$\therefore$ Molar $-q = 30.33\ kJ/(1.148\ g/122.0\ g\ mol^{-1}) = 3223\ kJ\ mol^{-1}$

$$q = -3223\ kJ\ mol^{-1}$$

Enthalpy

If heat is added to a system, for example, a gas, the heat absorbed (q) must equal the increase in energy of the system, ΔE, plus the work done by the system (w)

$$\therefore \qquad q= \Delta E + w$$

or

$$\Delta E = q - w$$

This is the **first law of thermodynamics**
says energy is conserved in all reactions

When chemical reactions are carried out at constant pressure
(for example, at 1 atmosphere = 1 bar = 100 kPa),
the heat term is called the enthalpy (q_P) and has the symbol ΔH.

For most chemical reactions ΔH ~ ΔE

ΔH, the enthalpy of a reaction is a measure of the energy difference between reagents and products, for reactions carried out at constant pressure

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For **exothermic** reactions, heat is released $\qquad$ ΔH = -ve

For **endothermic** reactions, heat is absorbed $\qquad$ ΔH = +ve

ΔH values are customarily measured at 25^0C (298K) and P = 1 bar.
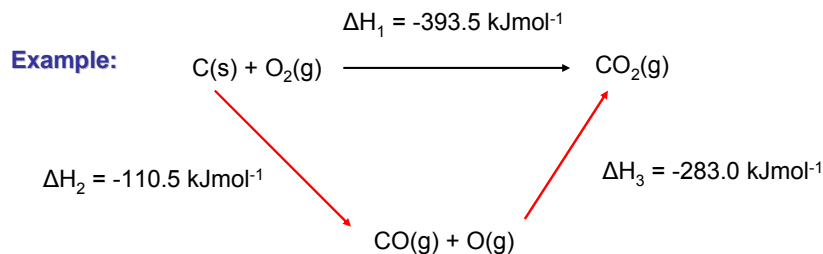They are labelled ΔH^0, the standard enthalpy of reaction

Example: For the combustion of heptane, C_7H_{16}

$$C_7H_{16}(\ell) + 11O_2 \text{ (g)} \rightarrow 7CO_2 \text{ (g)} + 8H_2O(\ell) \qquad \Delta H^0 = -4818 \text{ kJ}$$
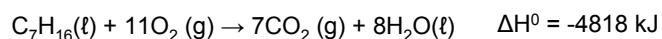
This exothermic reaction releases 4.818×10^6 J energy / mole C_7H_{16}, a molar enthalpy of reaction

Hess's Law

For a given reaction, the enthalpy change is the same for a reaction performed in one step or in many steps

Example:

$$\Delta H_1 = -393.5 \text{ kJmol}^{-1}$$

$$C(s) + O_2(g) \xrightarrow{\hspace{3cm}} CO_2(g)$$

$$\Delta H_2 = -110.5 \text{ kJmol}^{-1} \qquad \Delta H_3 = -283.0 \text{ kJmol}^{-1}$$

$$CO(g) + O(g)$$

that is, by Hess's Law: $\Delta H_1 = \Delta H_2 + \Delta H_3$

$$\longrightarrow \quad -393.5 = -110.5 + (-283.0) \text{ kJ mol}^{-1}$$

Who was Hess?

•Germain Henri Hess was a Swiss-born Russian chemist born in 1802 and died in 1850

•While he studied chemistry he was a practicing doctor in Irtusk Russia

•In addition to Hess' Law he was the first to analyze silver telluride (Ag_2Te) which is called Hessite in his honor, and to study the oxidation of sugars.

•He and his family performed at the local circus in a unicycle act which they called the Messy Hesses (really).