Announcements

More on the X-mas exam date December 17th 9 am – 12:00 pm.
Rooms:

Section 003:

Room	**Student Number**
HSB 240	60073 - 99554
AH 15	04620 - 45989

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

$q_3 = 4.184 JK^{-1}g^{-1} \times 750g \times 22K$
$= 69036 J = 69.04 kJ$

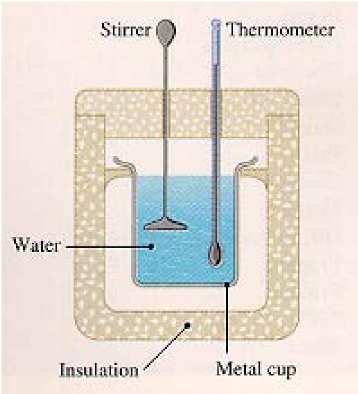
Total $q = q_1 + q_2 + q_3$
$= 22.61 + 250.42 + 69.04$

$\boxed{= 342.07 kJ}$

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: $C_C = 817\ JK^{-1}$; $C(H_2O) = 4.184\ J\ g^{-1}\ K^{-1}$

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

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

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

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

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

$$q = -3323\ kJ\ mol^{-1}$$

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)

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

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 $\Delta H \sim \Delta E$

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

For **exothermic** reactions, heat is released $\Delta H = -ve$

For **endothermic** reactions, heat is absorbed $\Delta H = +ve$

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}

$$C_7H_{16}(\ell) + 11O_2\,(g) \rightarrow 7CO_2\,(g) + 8H_2O(\ell) \qquad \Delta H^0 = -4818\ kJ$$

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:

$$C(s) + O_2(g) \xrightarrow{\Delta H_1 = -393.5\ kJmol^{-1}} CO_2(g)$$

$$C(s) + O_2(g) \xrightarrow{\Delta H_2 = -110.5\ kJmol^{-1}} CO(g) + O(g) \xrightarrow{\Delta H_3 = -283.0\ kJmol^{-1}} CO_2(g)$$

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

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

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).