

## Notes Chapter 11 – Thermochemistry – Heat and Chemical Change

**Thermochemistry** – the study of heat changes in chemical reactions.

Energy is the capacity for doing work and is detected only because of its effects.

**Chemical potential energy** – energy stored within the bonds of chemical substances.

Energy changes in chemical reactions due to the *making and breaking* of chemical bonds.

**Heat content ( $\Delta H$ )** – known as **enthalpy**

**Heat ( $q$  or  $\Delta H$ )** – energy that transfers from one object to another because of a temperature difference. It is *not measured directly*, only changes caused by heat are detected.

**Endothermic** process – heat flows *into* the system from the surroundings. **Absorbs** heat so heat ( $q$  or  $\Delta H$ ) has a **positive value**. (Feels cold)

**Exothermic** process – heat flows *from* the system to the surroundings. **Releases** heat so heat ( $q$  or  $\Delta H$ ) has a **negative value**. (Feels hot)

**calorie** – quantity of heat needed to raise the temperature of 1g of pure water 1°C.

**1000cal. = 1 kcal. = 1 Cal.**

Food labels use the Calorie so 10g of sugar = 41Cal = 41000 cal.

**1 cal = 4.184J** (J or joule is the SI unit of heat and energy)

**Heat capacity** – amount of heat needed to increase the temperature of an object exactly 1°C. *Depends on its mass and chemical composition.*

**Specific heat (c)** – amount of heat needed to raise the temperature of *1g of the substance* 1°C. (chart on p.296)

So heat capacity = its specific heat times its mass in grams.

$$\Delta H = q = (m)(c)(\Delta T)$$

where H is cal or J, m is grams, c is cal/g°C or J/g°C, and T is °C

Therefore, units of specific heat are J/g°C or cal/g°C

**Examples:**

**1) The temperature of a piece of Cu with a mass of 95.4g increases from 25.0°C to 48.0°C when the metal absorbs 849J of heat. What is the specific heat of copper?**

$$849\text{J} = (95.4\text{g})(c)(23^\circ\text{C}) \quad ; \text{ therefore specific heat} = .387\text{J/g}^\circ\text{C}$$

**2.) How much heat is required to raise the temperature of 125.0g of mercury 25°C? (Specific heat chart on p.296)**

$$\Delta H = (125.0\text{g})(.14\text{J/g}^\circ\text{C})(25^\circ\text{C}) = 437.5\text{J}$$

**Other conversion factors that are useful to remember:**

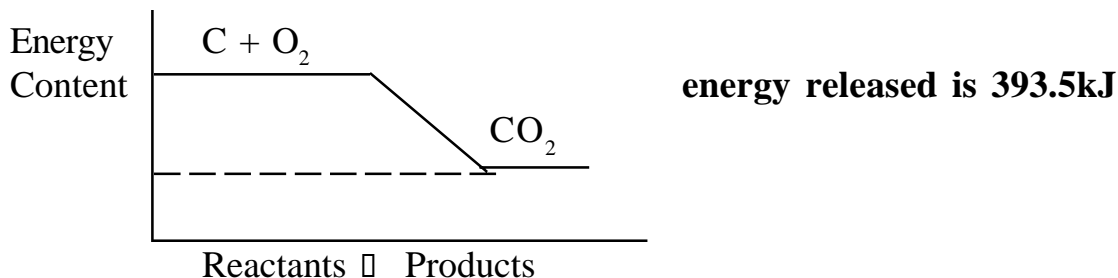
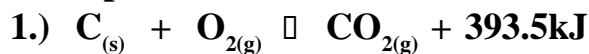
1kJ = 1000J; 1kcal(Cal) = 1000cal; 1mL H<sub>2</sub>O = 1g H<sub>2</sub>O; 1L = 1000mL

**Calorimetry** – accurate and precise measurement of heat change. The processes must be carried out in an insulated device called a **calorimeter**.

**Thermochemical equations** – an equation that includes the heat change.

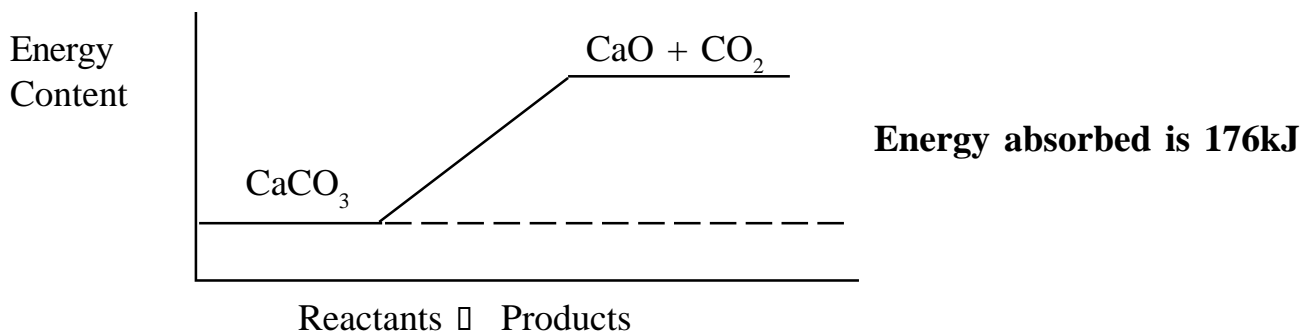
A heat of reaction is the heat change for the equation, reported as  $\Delta H$ .

**Examples:**



**Exothermic**  $\Delta H = -393.5kJ$

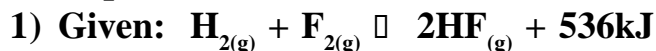
$\Delta H$  is *negative* because less heat energy is stored in the  $CO_2$  bonds than was stored in the C and  $O_2$  bonds.



**Endothermic**  $\Delta H = 176kJ$

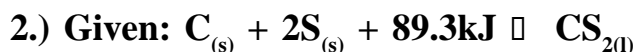
$\Delta H$  is *positive* because there is more heat energy stored in the CaO and  $CO_2$  bonds than was stored in the  $CaCO_3$  bonds.

**Examples:**



**Calculate the heat change (in kJ) for the conversion of 11.00g of hydrogen gas to hydrogen fluoride gas at constant pressure.**

$$\frac{11.00g H_2}{2.0g H_2} \times \frac{1 mol H_2}{1 mol H_2} \times \frac{-536kJ}{1 mol H_2} = -2948kJ$$



Calculate the amount of heat (in kJ) absorbed when 4.25g of carbon disulfide is formed.

$$\frac{4.25\text{g CS}_2}{76.2\text{g CS}_2} \times \frac{1\text{ mol CS}_2}{1\text{ mol CS}_2} \times 89.3\text{kJ} = 4.98\text{kJ}$$

3.) Given:  $2\text{NaHCO}_3(\text{s}) + 129\text{kJ} \rightarrow \text{Na}_2\text{CO}_3(\text{s}) + \text{H}_2\text{O}(\text{g}) + \text{CO}_2(\text{g})$   
 Calculate the kilojoules of heat required to decompose 1.45mol  $\text{NaHCO}_3(\text{s})$ .

$$\frac{1.45\text{ mol NaHCO}_3}{2\text{ mol NaHCO}_3} \times 129\text{kJ} = 93.5\text{kJ}$$

**Heat of fusion** – heat *absorbed* by one mole of a substance in melting from a solid to a liquid at a constant temperature ( $\Delta H_f$ )

**Heat of solidification** – heat *released* when one mole of a liquid solidifies at a constant temperature ( $\Delta H_s$ )

$$\text{So } \Delta H_f = -\Delta H_s$$

To **melt** one mole of ice at  $0^\circ\text{C}$  it requires the *absorption* of 6.01KJ of heat. Likewise, the conversion of one mole of water at  $0^\circ\text{C}$  to ice at  $0^\circ\text{C}$  *releases* 6.01KJ of heat

**Heat of vaporization** – amount of heat necessary to vaporize one mole of a given liquid ( $\Delta H_v$ ).

**Heat of condensation** – amount of heat released when one mole of vapor condenses. ( $\Delta H_c$ )

$$\text{So } \Delta H_v = -\Delta H_c$$

