

Ch. 14 Notes – ENERGY AND CHEMICAL CHANGE

NOTE: Vocabulary terms are in **boldface and underlined**. Supporting details are in *italics*.

- I. **Energy**—*the capacity to do work or produce heat*
- A. two basic types of energy
- 1) *potential*—energy at rest; energy of position
 - 2) *kinetic*—energy of motion ($1/2 mv^2$)
- B. other subcategories of energy
- | | |
|-----------------------|------------|
| <i>radiant</i> —light | electrical |
| <i>thermal</i> —heat | sound |
| mechanical | atomic |
| gravitational | nuclear |
- chemical energy (chemical potential energy)**—*stored in chemical bonds*
- C. **Law of Conservation of Energy**—*in a physical or chemical change, energy cannot be created nor destroyed, it merely changes form*
- D. **heat (q)**
- 1) energy flowing from warmer to cooler objects or areas
 - 2) units: *Joule (J), calorie (cal); diet Calorie (Cal)*

- II. Heat Capacity
- A. **heat capacity**—*amount of heat required to change a substance's temperature by exactly 1 °C; common unit: J / °C*
- B. **calorie** (*calorie with a lower-case c*)—*amount of heat required to raise the temp of 1 g of pure water by 1 °C*
- C. **Calorie** (*Calorie, with an upper-case c*)—*“diet Calorie” = 1000 calories*
- D. **Joule**—*SI unit of heat and energy; amount of heat required to raise the temp of 1 g of pure water by 0.239 °C.*

UNIT CONVERSIONS:

$$1 \text{ Cal (diet Calorie)} = 1000 \text{ cal} = 1 \text{ kcal}$$

$$1 \text{ J} = 0.239 \text{ cal}$$

$$4.184 \text{ J} = 1 \text{ cal}$$

- E. heat changes can be measured with a **calorimeter**, which is a covered container

- III. Specific Heat Capacity
- A. **Specific heat** (*c or c_p*)—*amount of heat required to raise the temperature of 1 g of a substance by exactly 1 °C.*
- B. *c* deals with *heat requirements* and *heat retention* :
- 1) *low c = low requirements and retention = heats up quickly and cools down quickly*
 - 2) *high c = high requirements and retention = heats up slowly and cools down slowly*
- C. common unit: **J / g °C**
- D. heat changes can be measured with a **calorimeter**
- E. metals have low specific heats
- F. water has the highest specific heat of common substances = 4.184 J/g °C

G. equations $\Delta = \text{delta} = \text{“change in”}$ $\Delta T = (T_{\text{final}} - T_{\text{initial}})$

$$q = m c \Delta T$$

$$c = \frac{q}{m \Delta T}$$

H. examples

EXAMPLE 1) The temperature of an 89.1 g piece of metal rises from 22.0 °C to 51.1 °C when the metal absorbs 794 J of energy. What is the specific heat of the metal?

$$m = 89.1 \text{ g} \quad q = 794 \text{ J} \quad T = 51.1 \text{ }^\circ\text{C} - 22.0 \text{ }^\circ\text{C} = 29.1 \text{ }^\circ\text{C}$$

$$c = \frac{q}{m \Delta T} = \frac{794 \text{ J}}{(89.1 \text{ g})(29.1 \text{ }^\circ\text{C})} = \boxed{0.306 \frac{\text{J}}{\text{g}^\circ\text{C}}}$$

EXAMPLE 2) How much heat energy is needed to increase the temperature of 44.7 g of water from 20.0 to 36.3 °C?

$$m = 44.7 \text{ g} \quad c = 4.184 \text{ J/g }^\circ\text{C} \quad \Delta T = 36.3 \text{ }^\circ\text{C} - 20.0 \text{ }^\circ\text{C} = 16.3 \text{ }^\circ\text{C}$$

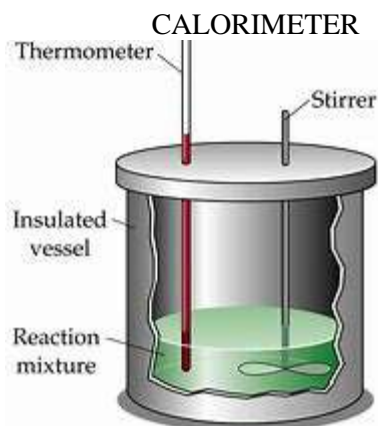
$$q = m c \Delta T \quad q = 44.7 \text{ g} (4.184 \frac{\text{J}}{\text{g}^\circ\text{C}})(16.3 \text{ }^\circ\text{C}) = \boxed{3050 \text{ J or } 3.05 \text{ kJ}}$$

IV. **calorimetry**—the accurate and precise measurement of the heat change for chemical reactions and physical changes

A. equation

$$q = \Delta H = m c \Delta T \quad (\text{rearrangement of the specific heat equation})$$

$$q = \Delta H = (\text{mass of water})(\text{specific heat of water})(\text{change in temp})$$



Source: prenhall

B. typical lab steps for determination of specific heat of an unknown metal

- 1) water is measured and poured into the calorimeter
- 2) initial temp of water is recorded
- 3) metal is heated and temperature of metal is recorded
- 4) metal is added to the calorimeter and lid is closed immediately
- 5) mixture is stirred
- 6) final temp is recorded

heat lost by metal = heat gained by water $-\Delta H_m = +\Delta H_w$ q lost by metal = q gained by H ₂ O $q_m = q_w$ $-[(m_m)(c_m)(\Delta T_m)] = (m_w)(c_w)(\Delta T_w)$ $-[(m_m)(c_m)(T_{final, m} - T_{initial, m})] = (m_w)(c_w)(T_{final, w} - T_{initial, w})$

EXAMPLE 3) A 20.00 g piece of Bauckium metal is heated thoroughly in a boiling water bath. The calorimeter is filled with 275.00 mL of water. The room temperature is 26.5 °C. The final calorimeter temperature at the end of the experiment is 28.4 °C. What is the specific heat of Bauckium?

$$-[(m_m)(c_m)(T_{final, m} - T_{initial, m})] = (m_w)(c_w)(T_{final, w} - T_{initial, w})$$

$$(c_m) = \frac{(m_w)(c_w)(T_{final, w} - T_{initial, w})}{(m_m)(T_{final, m} - T_{initial, m})}$$

$$-[(20.00 \text{ g})(c_m)(28.4 - 100.0 \text{ }^\circ\text{C})] = (275.00 \text{ g})(4.184 \text{ J/g }^\circ\text{C})(28.4 - 26.5 \text{ }^\circ\text{C})$$

$$1432 c_m = 2186.14 \quad \boxed{c_m = 1.53 \text{ J/g }^\circ\text{C}}$$

EXAMPLE 4) A 20.00 g piece of Bauckium metal is heated thoroughly in a boiling water bath. The calorimeter is filled with 275.00 mL of water. The room temperature is 26.5 °C. The specific heat of Bauckium metal is 1.53 J/g °C. What is the final temperature of the system at the end of the experiment?

$$-[(m_m)(c_m)(T_{final, m} - T_{initial, m})] = (m_w)(c_w)(T_{final, w} - T_{initial, w})$$

$$-[(20.00 \text{ g})(1.53 \text{ J/g }^\circ\text{C})(x - 100.0 \text{ }^\circ\text{C})] = (275.00 \text{ g})(4.184 \text{ J/g }^\circ\text{C})(x - 26.5 \text{ }^\circ\text{C})$$

$$-[(30.6)(x - 100)] = 1150.6(x - 26.5)$$

$$-30.6x + 3060 = 1150.6x - 30490.9$$

$$33550.9 = 1181.2x$$

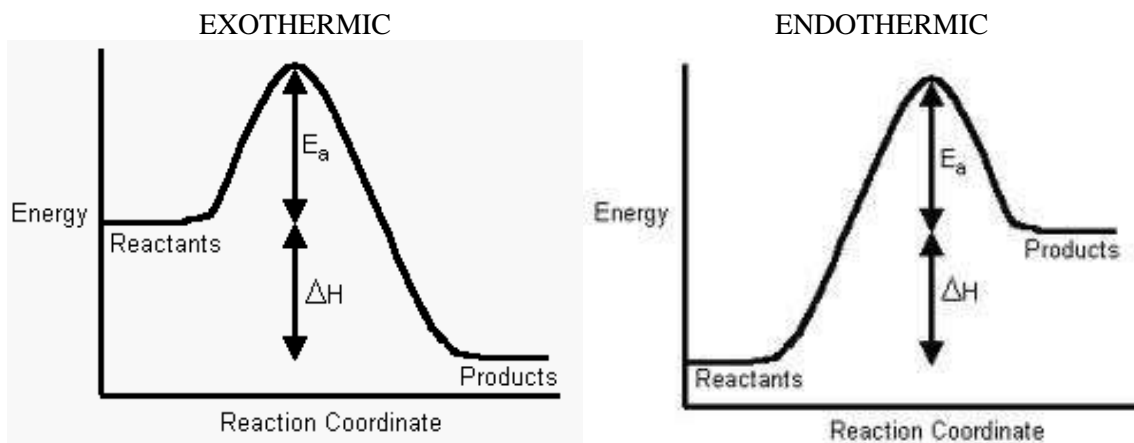
$$\boxed{28.4} = \text{ }^\circ\text{C}$$

- V. Chemical Energy and the Universe
- A. **thermochemistry**
- 1) *the study of heat changes in chemical reactions and physical changes*
 - 2) *the study of heat flow between a system and its surroundings*
 - a. **system**—specific part being analyzed
 - b. **surroundings**—everything outside the system (usually the immediate area)
 - c. **universe** = system + surroundings
 - 3) **thermochemical equations**—*equations that show heat changes*

B. **enthalpy**

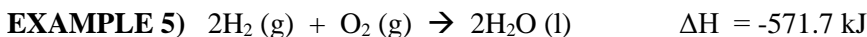
- 1) (H)—heat content of a system at constant pressure
- 2) ΔH is used interchangeably with q
- 3) change in enthalpy = ΔH ; heat change for a process at constant pressure; usually measured in kJ (kilojoules)
- 4) enthalpy (heat) of reaction— heat absorbed or released in a chemical rxn. $\Delta H_{\text{rxn}} = H_{\text{products}} - H_{\text{reactants}}$

ENDOTHERMIC:	$+\Delta H$	heat absorbed into system; surroundings cool down	$(A + B + \text{ENERGY} \rightarrow C + D)$	$+\Delta H$
EXOTHERMIC:	$-\Delta H$	heat released from system; surroundings heat up	$(A + B \rightarrow C + D + \text{ENERGY})$	$-\Delta H$



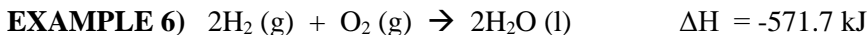
VI. Thermochemical Equations

- A. **thermochemical equation**—a balanced chemical equation including physical states of all reactants and products and the energy change (usually ΔH)
- B. ΔH relates to the coefficients of the balanced equation
- C. examples



How much energy is produced if 100. grams of H_2 are converted to water?

$$100. \text{ g H}_2 \times \frac{1 \text{ mol H}_2}{2.02 \text{ g H}_2} \times \frac{-571.7 \text{ kJ}}{2 \text{ mol H}_2} = \boxed{-14,200 \text{ or } 1.42 \times 10^4 \text{ kJ}}$$



How many grams of H_2 would be required to produce $5.00 \times 10^3 \text{ kJ}$?

$$5.00 \times 10^3 \text{ kJ} \times \frac{2 \text{ mol H}_2}{-571.7 \text{ kJ}} \times \frac{2.02 \text{ g H}_2}{1 \text{ mol H}_2} = \boxed{35.3 \text{ g H}_2} \text{ (absolute value for grams)}$$

Chemical change heat transfers:

ΔH_r or ΔH_{rxn} = generic ENTHALPY (HEAT) OF REACTION = heat absorbed or released in a chemical rxn.

ΔH_{comb} = MOLAR ENTHALPY (HEAT) OF COMBUSTION = heat released in combustion of 1 mol of substance

ΔH_f = ENTHALPY (HEAT) OF FORMATION = heat absorbed or released to make 1 mol of a compd from its elements

ΔH°_f = STANDARD ENTHALPY (HEAT) OF FORMATION = heat absorbed or released to make 1 mol of a compd from its elements in their standard states at 298 K (25 °C)

VII. Heat and Changes of State

A. math problems use ΔH related to coefficients

B. problems discuss melting, freezing, boiling, condensing, dissolving

Physical change heat transfers:

ΔH_{vap} = MOLAR ENTHALPY (HEAT) OF VAPORIZATION = heat required to vaporize 1 mol of a liquid

$$q = (\text{mass} / \text{molar mass}) (\Delta H_{vap})$$

ΔH_{fus} = MOLAR ENTHALPY (HEAT) OF FUSION = heat required to melt 1 mol of a solid

$$q = (\text{mass} / \text{molar mass}) (\Delta H_{fus})$$

ΔH_{solid} = MOLAR ENTHALPY (HEAT) OF SOLIDIFICATION = heat energy released freezing 1 mol of substance

$$\Delta H_{fus} = - \Delta H_{solid}$$

$$\Delta H_{fus} \text{ H}_2\text{O} = 6.01 \text{ kJ/mol}$$

ΔH_{cond} = MOLAR ENTHALPY (HEAT) OF CONDENSATION = heat energy released condensing 1 mol of substance

$$\Delta H_{vap} = - \Delta H_{cond}$$

$$\Delta H_{vap} \text{ H}_2\text{O} = 40.7 \text{ kJ/mol}$$

ΔH_{soln} = MOLAR ENTHALPY (HEAT) OF SOLUTION = heat energy change due to the dissolving of 1 mol of substance

C. examples

EXAMPLE 7) 53.1 g of H_2O exists as a liquid at 0 °C, but it is about to freeze. What is the heat change, in kJ, to convert the water to a solid at 0 °C? $\Delta H_{fus} \text{ H}_2\text{O} = 6.01 \text{ kJ/mol}$

This problem deals with freezing. ΔH_{solid} $\Delta H_{fus} = - \Delta H_{solid}$

$$q = 53.1 \text{ g-H}_2\text{O} \times \frac{1 \text{ mol-H}_2\text{O}}{18.02 \text{ g-H}_2\text{O}} \times \frac{-6.01 \text{ kJ}}{\text{mol}} = \boxed{-17.7 \text{ kJ}}$$

EXAMPLE 8) 49.5 g of H₂O is being boiled at its boiling point of 100 °C. How many kJ are required? $\Delta H_{\text{vap}} \text{H}_2\text{O} = 40.7 \text{ kJ/mol}$

$$q = 49.5 \text{ g H}_2\text{O} \times \frac{1 \text{ mol H}_2\text{O}}{18.02 \text{ g H}_2\text{O}} \times \frac{40.7 \text{ kJ}}{\text{mol}} = \boxed{+112 \text{ kJ}}$$

VIII. Hess' Law (Germain Hess 1802-1850)

A. **Hess' Law** of Heat Summation —if two or more thermochemical equations can be added to produce a final equation, the overall ΔH for a rxn. equals the sum of all ΔH 's for the individual steps involved

$$\Delta H^\circ_{\text{rxn}} = \Sigma [(\text{cf})(\Delta H^\circ_{\text{f}} \text{ products})] - \Sigma [(\text{cf})(\Delta H^\circ_{\text{f}} \text{ reactants})]$$

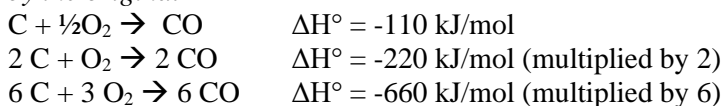
where cf = coefficient from the balanced equation

B. If a chemical equation is written in reverse order, the sign of ΔH must be changed (+ to - OR - to +)

example:



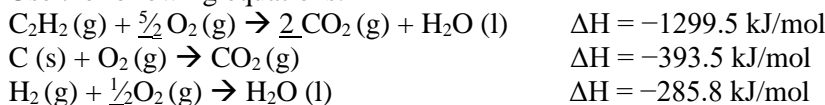
C. If coefficients are multiplied by 2, 3, 4, etc. the same number must be multiplied by the original ΔH



D. Examples

EXAMPLE 9) Calculate ΔH for the reaction: $2\text{C(s)} + \text{H}_2(\text{g}) \rightarrow \text{C}_2\text{H}_2(\text{g})$

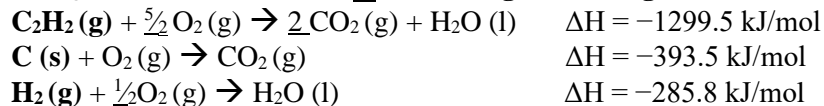
Use the following equations:



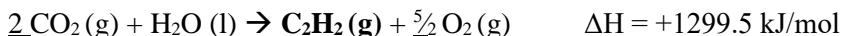
Steps...

1. Highlight, circle, or boldface the formulas in the three equations that are needed in the original equation.

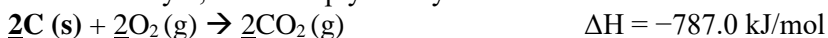
FINAL EQUATION NEEDED: $2\text{C(s)} + \text{H}_2(\text{g}) \rightarrow \text{C}_2\text{H}_2(\text{g})$



2. The first equation has C₂H₂ on the wrong side, so reverse the order and change the sign of ΔH .



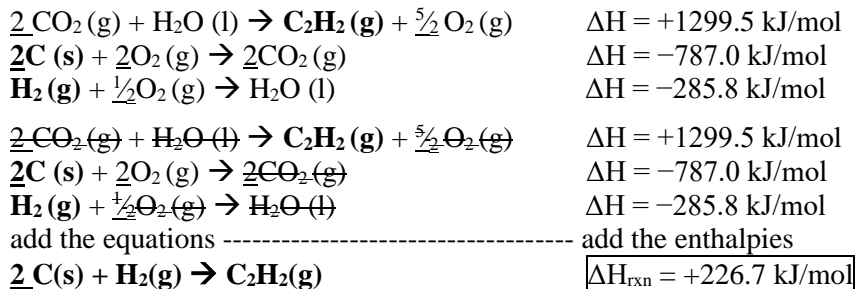
3. The second equation is in the correct order, but 2 moles of C are needed. Multiply all the coefficients by 2, and multiply ΔH by 2.



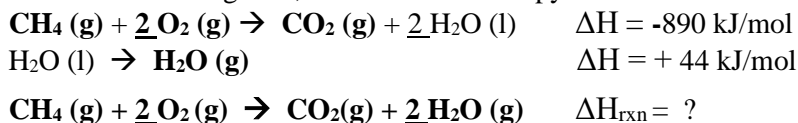
4. The third equation does not need to be changed. H₂ is on the correct side in the correct amount. Keep ΔH as is.



5. List the adjusted equations and cancel out duplicate formulas on opposite sides. Physical states and coefficients must match to be cancelled. The final result should match the equation that is needed.

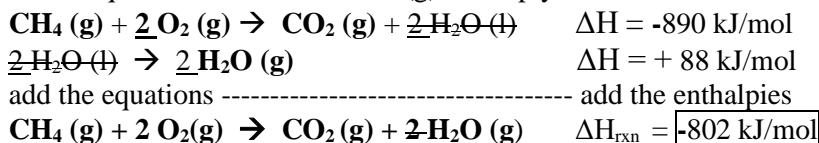


EXAMPLE 10) From the following data, calculate the enthalpy of the third reaction.



Steps...

- The first equation and its ΔH do not need to be changed.
- The second equation needs 2 mol $\text{H}_2\text{O}(\text{g})$. Multiply the coefficients and ΔH by 2.



IX. Standard Heats of Formation

A. standard state of an element

- pure, at 1 atm pressure, in the normal phase at 298 K (25 °C)
- enthalpy of an element in its standard state at 298K (25 °C) = 0

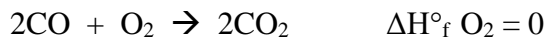
examples: Ar (g), O_2 (g), C (solid graphite), Br_2 (l); all diatomics = 0

B. values can be looked up in reference tables

C. example

EXAMPLE 11) Calculate $\Delta\text{H}_{\text{comb}}$, in kJ, from the burning of CO to form CO_2 .

from tables: $\Delta\text{H}^\circ_{\text{f}} \text{CO}_2 = -393.5 \text{ kJ/mol}$ & $\Delta\text{H}^\circ_{\text{f}} \text{CO} = -110.5 \text{ kJ/mol}$



$$\Delta\text{H}^\circ_{\text{rxn}} = \sum [(cf)(\Delta\text{H}^\circ_{\text{f}} \text{ products})] - \sum [(cf)(\Delta\text{H}^\circ_{\text{f}} \text{ reactants})]$$

$$\Delta\text{H}^\circ_{\text{rxn}} = [(2 \text{ mol CO}_2)(-393.5 \frac{\text{kJ}}{\text{mol}})] - [(2 \text{ mol CO})(-110.5 \frac{\text{kJ}}{\text{mol}})] + (0 \text{ kJ/mol}) =$$

$$-787 - -221 = \boxed{-566 \text{ kJ}}$$

X. Reaction Spontaneity

A. **spontaneous process**—a physical or chemical change that, once begun, occurs without outside manipulation (may need help getting started)

B. **entropy (S)**

- 1) disorder in a system
- 2) typical unit: J/K (Joules per Kelvin)
- 3) **Second Law of Thermodynamics**—spontaneous process show an increase in entropy
- 4) $\Delta S_{\text{rxn}} = S_{\text{products}} - S_{\text{reactants}}$
- 5) $\Delta S = +$ with increased entropy (favorable)
- 6) $\Delta S = -$ with decreased entropy (unfavorable)

C. ΔH (enthalpy) conditions

- 1) $-\Delta H$ is favorable (exothermic)
- 2) $+\Delta H$ is unfavorable (endothermic)

D. **Gibbs free energy (G)**—energy available to do work; useful energy

Gibbs free energy equation (Gibbs function, free enthalpy equation):

$$\Delta G = \Delta H - T\Delta S \quad \text{or} \quad \Delta G_{\text{system}} = \Delta H_{\text{system}} - T\Delta S_{\text{system}}$$

E. ΔG (Gibbs free energy) conditions

- 1) $-\Delta G$ is favorable (exergonic)
- 2) $+\Delta G$ is unfavorable (endergonic)

F. Examples...

EXAMPLE 12) Determine ΔG and the overall spontaneity of a system under the following conditions: $\Delta H_{\text{system}} = +452 \text{ kJ}$ $\Delta S_{\text{system}} = +55.7 \text{ J/K}$ $T = 165 \text{ K}$

$$\Delta S_{\text{system}} = 55.7 \frac{\text{J}}{\text{K}} \times \frac{1 \text{ kJ}}{1000 \text{ J}} = 0.0557 \frac{\text{kJ}}{\text{K}}$$

$$\Delta G_{\text{system}} = \Delta H_{\text{system}} - T\Delta S_{\text{system}}$$

$$\Delta G_{\text{system}} = 452 \text{ kJ} - (165 \text{ K})(0.0557 \frac{\text{kJ}}{\text{K}}) = 452 - 9.1905 =$$

$$\boxed{+443 \text{ kJ; nonspontaneous (endergonic)}}$$

.....

EXAMPLE 13) Determine ΔG and the overall spontaneity of a system under the following conditions: $\Delta H_{\text{system}} = -75.9 \text{ kJ}$ $\Delta S_{\text{system}} = +138 \text{ J/K}$ $T = 273 \text{ K}$

$$\Delta S_{\text{system}} = 138 \frac{\text{J}}{\text{K}} \times \frac{1 \text{ kJ}}{1000 \text{ J}} = 0.138 \frac{\text{kJ}}{\text{K}}$$

$$\Delta G_{\text{system}} = \Delta H_{\text{system}} - T\Delta S_{\text{system}}$$

$$\Delta G_{\text{system}} = -75.9 \text{ kJ} - (273 \text{ K})(0.138 \frac{\text{kJ}}{\text{K}}) = -75.9 - 37.674 = \boxed{-113.6 \text{ kJ; spontaneous}}$$

Summary: Spontaneous and Non-Spontaneous Reactions

	$\Delta H > 0$	$\Delta H < 0$
$\Delta S > 0$	Spontaneity depends on T (spontaneous at higher temperatures)	Spontaneous at all temperatures
$\Delta S < 0$	Nonspontaneous (proceeds only with a continuous input of energy)	Spontaneity depends on T (spontaneous at lower temperatures)

- ΔH = exothermic (favorable)

+ ΔH = endothermic (unfavorable)

+ ΔS = increased entropy (favorable)

- ΔS = decreased entropy (unfavorable)

- ΔG = exergonic: spontaneous (favorable)

+ ΔG = endergonic: nonspontaneous (unfavorable)