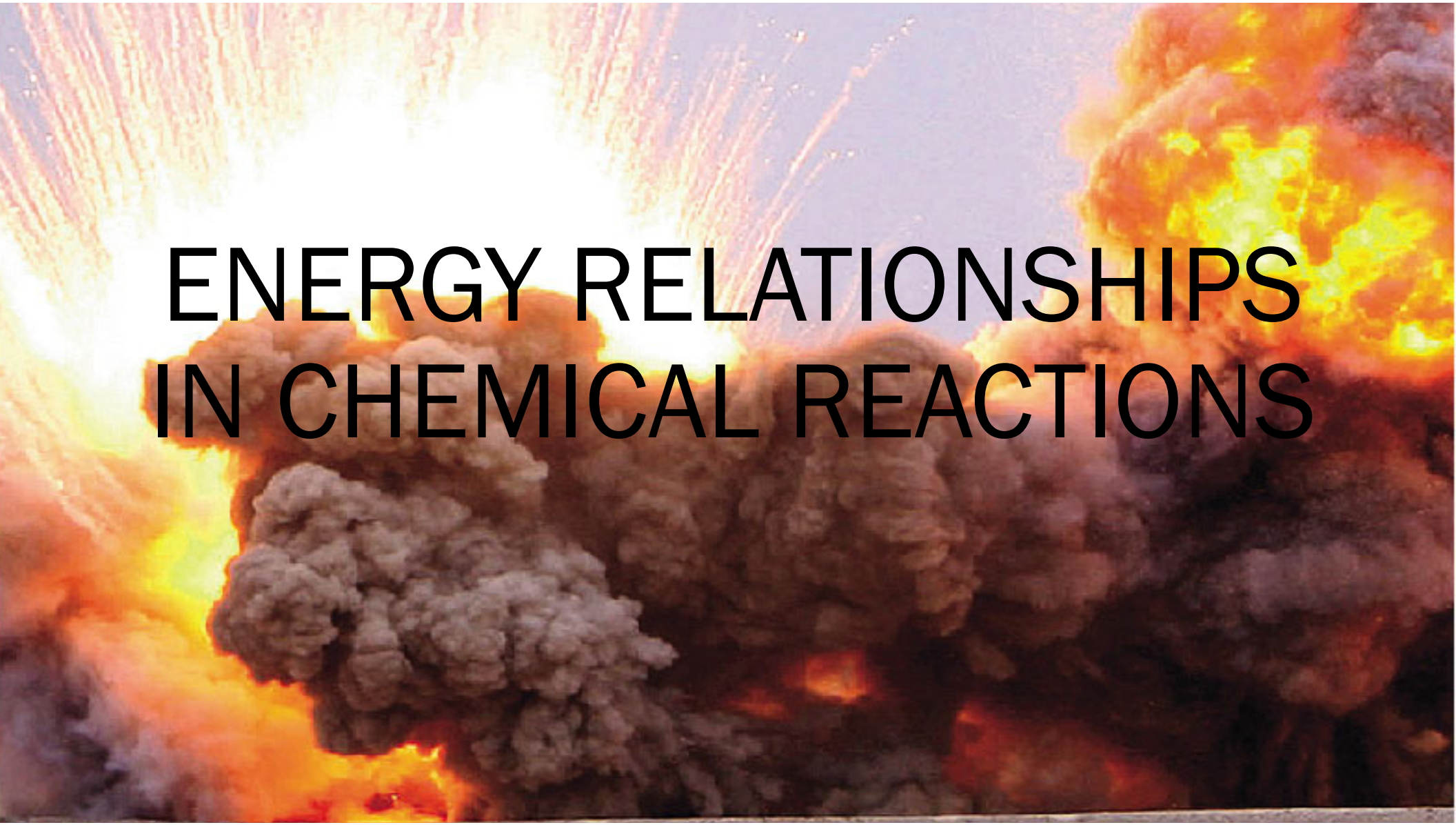


ENERGY RELATIONSHIPS IN CHEMICAL REACTIONS

Chapter 6 – Energy Relationships in Chemical Reactions

- The nature of energy and types of energy (6.1)
- Energy changes in chemical reactions (6.2)
- Introduction to thermodynamics (6.3)
- Calorimetry (6.5)
- Enthalpy of chemical reactions (6.4)
- Standard enthalpy of formation and reaction (6.6)

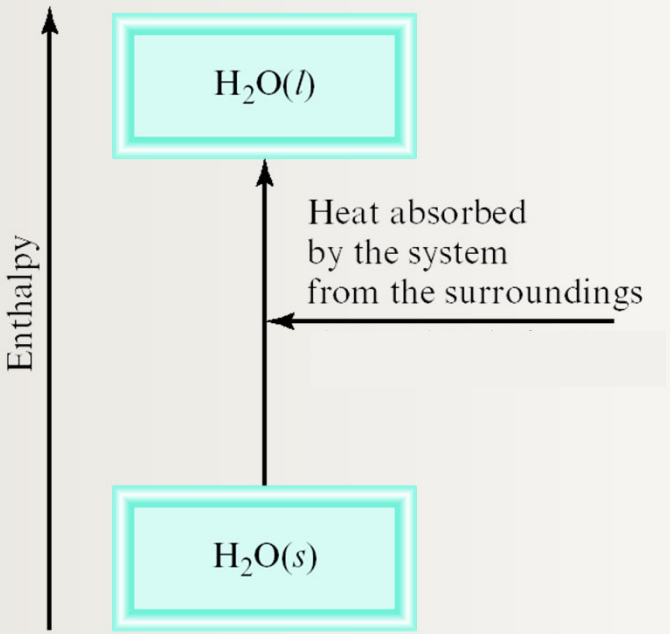
Chapter 6 – Energy Relationships in Chemical Reactions

- ~~The nature of energy and types of energy (6.1)~~
- ~~Energy changes in chemical reactions (6.2)~~
- ~~Introduction to thermodynamics (6.3)~~
- ~~Calorimetry (6.5)~~
- Enthalpy of chemical reactions (6.4)
- Standard enthalpy of formation and reaction (6.6)

6.4 and 6.6 Enthalpy

- What is heat?
- What is a state function?
- What is enthalpy?
  - *Heat content of a reaction or a quantity to describe the heat flow into or out of a system in a process that occurs at constant pressure*
- Is enthalpy a state function?
  - *What does this mean in terms of a reaction?*
- How can we visualize this?

Thermochemical Equations

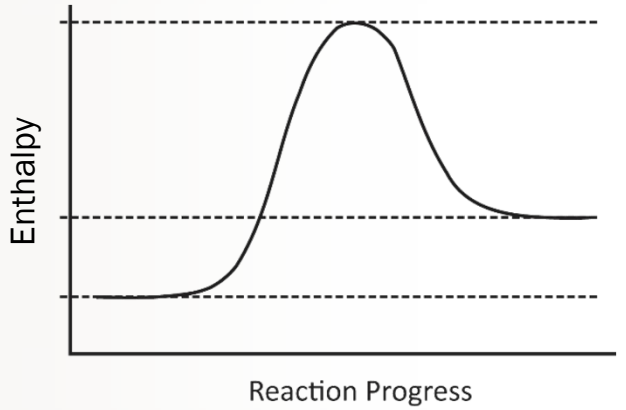


Is ΔH negative or positive?

System absorbs heat

Endothermic

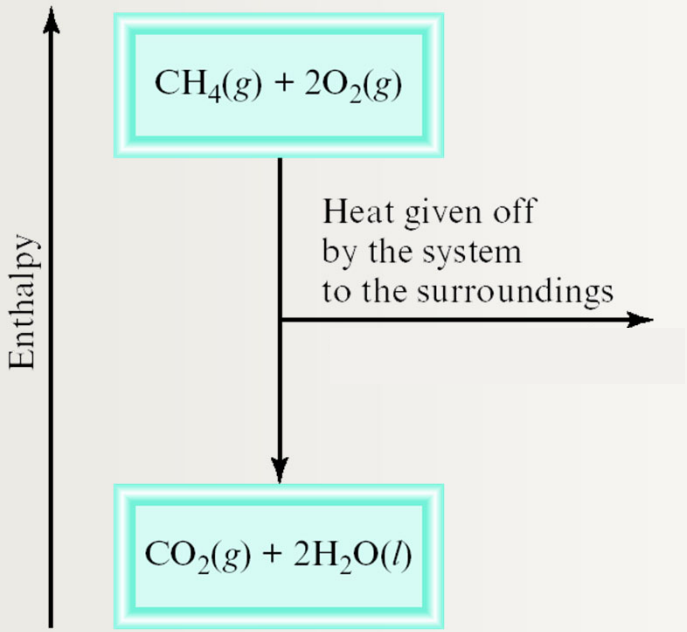
$\Delta H > 0$ $H_{reactants} < H_{products}$



6.01 kJ are absorbed for every 1 mole of ice that melts at 0^0C and 1 atm.

$H_2O(s) \longrightarrow H_2O(l)$ $\Delta H = 6.01$ kJ/mol

Thermochemical Equations

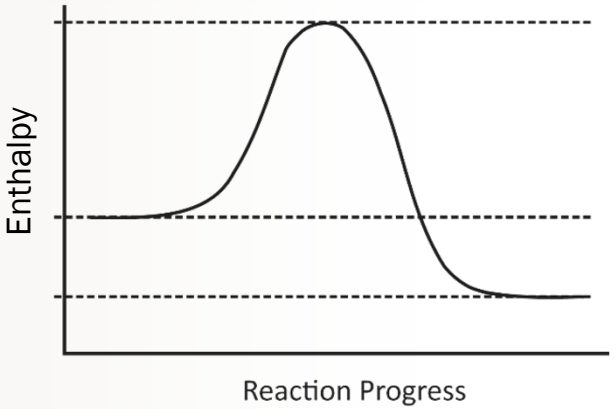


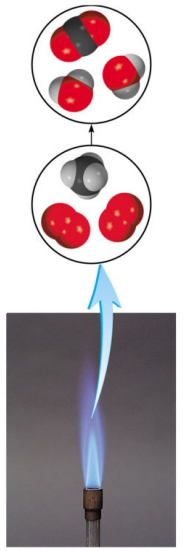
Is ΔH negative or positive?

System gives off heat

Exothermic

$\Delta H < 0$ $H_{reactants} < H_{products}$





890.4 kJ are released for every 1 mole of methane that is combusted at 25^0C and 1 atm.

$$CH_4(g) + 2O_2(g) \longrightarrow CO_2(g) + 2H_2O(l) \qquad \Delta H = -890.4 \text{ kJ/mol}$$

6.4 and 6.6 Enthalpy

- How is energy released in a reaction?
- How are enthalpy and heat related?
- How are enthalpy and energy related?
- Are all reactions that are exothermic spontaneous?
- Are all reactions that are endothermic spontaneous?

Consider the role of the surroundings

6.4 and 6.6 Enthalpy

- What are the rules for enthalpy?
  1. *Writing a thermochemical equation*
  2. *Direction of the equation (reactants vs. products)*
  3. *Extensive property – increasing or decreasing the amounts of products or reactants*
  4. *State property – one reaction or multiple reactions to an overall reaction*

6.4 and 6.6 Enthalpy

- Key Definitions:
  - *Standard conditions*
  - *Enthalpy of formation*
  - *Enthalpy of combustion*
  - *Enthalpy of solution*
  - *Enthalpy of vaporization (heat of vaporization)*
  - *Enthalpy of fusion (heat of fusion)*
- Where will you find these data?

Table 6.3 Heats of Some Typical Reactions Measured at Constant Pressure

Type of Reaction	Example	ΔH (kJ)
Heat of neutralization	$HCl(aq) + NaOH(aq) \longrightarrow NaCl(aq) + H_2O(l)$	−56.2
Heat of ionization	$H_2O(l) \longrightarrow H^+(aq) + OH^-(aq)$	56.2
Heat of fusion	$H_2O(s) \longrightarrow H_2O(l)$	6.01
Heat of vaporization	$H_2O(l) \longrightarrow H_2O(g)$	44.0*
Heat of reaction	$MgCl_2(s) + 2Na(l) \longrightarrow 2NaCl(s) + Mg(s)$	−180.2

*Measured at 25°C. At 100°C, the value is 40.79 kJ.

p. 206 of LA book

Table 6.4 Standard Enthalpies of Formation of Some Inorganic Substances at 25°C

Substance	ΔH_f° (kJ/mol)	Substance	ΔH_f° (kJ/mol)
$Ag(s)$	0	$H_2O_2(l)$	−187.6
$AgCl(s)$	−127.04	$Hg(l)$	0
$Al(s)$	0	$I_2(s)$	0
$Al_2O_3(s)$	−1669.8	$HI(g)$	25.94
$Br_2(l)$	0	$Mg(s)$	0
$HBr(g)$	−36.2	$MgO(s)$	−601.8
C(graphite)	0	$MgCO_3(s)$	−1112.9
C(diamond)	1.90	$N_2(g)$	0
$CO(g)$	−110.5	$NH_3(g)$	−46.3
$CO_2(g)$	−393.5	$NO(g)$	90.4
$Ca(s)$	0	$NO_2(g)$	33.85
$CaO(s)$	−635.6	$N_2O_4(g)$	9.66
$CaCO_3(s)$	−1206.9	$N_2O(g)$	81.56
$Cl_2(g)$	0	$O(g)$	249.4
$HCl(g)$	−92.3	$O_2(g)$	0
$Cu(s)$	0	$O_3(g)$	142.2
$CuO(s)$	−155.2	S(rhombic)	0
$F_2(g)$	0	S(monoclinic)	0.30
$HF(g)$	−268.61	$SO_2(g)$	−296.1
$H(g)$	218.2	$SO_3(g)$	−395.2
$H_2(g)$	0	$H_2S(g)$	−20.15
$H_2O(g)$	−241.8	$ZnO(s)$	−347.98
$H_2O(l)$	−285.8	ZnS(s)	−202.9

More complete listing in Appendix 2 of your textbook.

Appendix 2 (or the information from here) will be provided on all tests or quizzes.

Table 6.4, p. 199

6.4 and 6.6 Enthalpy

- What is an energy diagram?

Practice:

- What is the thermochemical equation and energy diagram for the enthalpy of formation of ammonium nitrate?
- What is the thermochemical equation and energy diagram for the enthalpy of solution of ammonium nitrate?

6.4 and 6.6 Enthalpy

- How do we measure change in enthalpy for a reaction?
  - *Direct method*
  - *Indirect methods*
    - Using a series of reactions (Hess's Law)
    - Using tables of data (enthalpy of formation)

p. 209 of LA book

Table 6.4 Standard Enthalpies of Formation of Some Inorganic Substances at 25°C

Substance	ΔH_f° (kJ/mol)	Substance	ΔH_f° (kJ/mol)
$Ag(s)$	0	$H_2O_2(l)$	−187.6
$AgCl(s)$	−127.04	$Hg(l)$	0
$Al(s)$	0	$I_2(s)$	0
$Al_2O_3(s)$	−1669.8	$HI(g)$	25.94
$Br_2(l)$	0	$Mg(s)$	0
$HBr(g)$	−36.2	$MgO(s)$	−601.8
C(graphite)	0	$MgCO_3(s)$	−1112.9
C(diamond)	1.90	$N_2(g)$	0
$CO(g)$	−110.5	$NH_3(g)$	−46.3
$CO_2(g)$	−393.5	$NO(g)$	90.4
$Ca(s)$	0	$NO_2(g)$	33.85
$CaO(s)$	−635.6	$N_2O_4(g)$	9.66
$CaCO_3(s)$	−1206.9	$N_2O(g)$	81.56
$Cl_2(g)$	0	$O(g)$	249.4
$HCl(g)$	−92.3	$O_2(g)$	0
$Cu(s)$	0	$O_3(g)$	142.2
$CuO(s)$	−155.2	S(rhombic)	0
$F_2(g)$	0	S(monoclinic)	0.30
$HF(g)$	−268.61	$SO_2(g)$	−296.1
$H(g)$	218.2	$SO_3(g)$	−395.2
$H_2(g)$	0	$H_2S(g)$	−20.15
$H_2O(g)$	−241.8	$ZnO(s)$	−347.98
$H_2O(l)$	−285.8	ZnS(s)	−202.9

More complete listing in Appendix 2 of your textbook.

Appendix 2 (or the information from here) will be provided on all tests or quizzes.

Table 6.4, p. 199

Chapter 6 – Practice

6.4 and 6.6 Enthalpy

Practice

- What is the enthalpy of combustion for methane?
- What volume of methane at 1 atm and 25°C is needed to heat 500 mL of water from 5°C to 95°C?
- If the current price of natural gas is $0.10 per cubic foot, (and we assume natural gas is all methane) how much does it cost to heat the water?

p. 211 of LA book

6.4 and 6.6 Enthalpy

Practice

- What is the thermochemical equation for the molar enthalpy of solution of calcium chloride?

- What is the enthalpy of formation for dinitrogen pentoxide given:

$$2H_2(g) + O_2(g) \rightarrow 2H_2O(l) \qquad \Delta H = -571.5 \text{ kJ}$$

$$N_2O_5(g) + H_2O(l) \rightarrow 2HNO_3(l) \qquad \Delta H = -76.6 \text{ kJ}$$

$$\frac{1}{2}N_2(g) + \frac{3}{2}O_2(g) + \frac{1}{2}H_2(g) \rightarrow HNO_3(l) \qquad \Delta H = -174 \text{ kJ}$$

Chapter 6 – Practice

6.4 and 6.6 Enthalpy

Practice

- Which substance(s) will have an enthalpy of formation equal to zero?
  a. O(g)
  b. $O_2(g)$
  c. $O_2(l)$
  d. $H_2O(l)$
  e. $OH^-(aq)$

Chapter 6 – Practice

6.4 and 6.6 Enthalpy

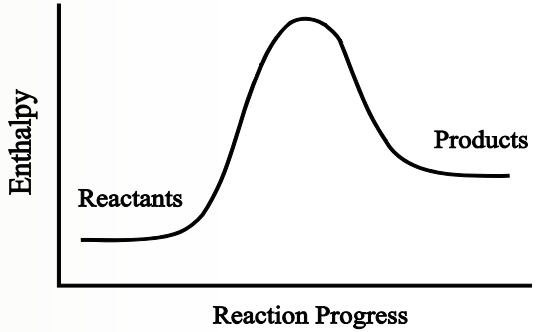
For the reaction energy diagram shown to the right, the reaction is

A. exothermic with $\Delta H_{reactants} > \Delta H_{products}$.

B. exothermic with $\Delta H_{reactants} < \Delta H_{products}$.

C. endothermic with $\Delta H_{reactants} > \Delta H_{products}$.

D. endothermic with $\Delta H_{reactants} < \Delta H_{products}$.



Which reaction corresponds to the reaction for the enthalpy of formation for methanol, $CH_3OH(l)$?

A. $2CH_3OH(l) + 3O_2(g) \rightarrow 2CO_2(g) + 4H_2O(l)$

B. $CH_3OH(l) + {}^3/_2O_2(g) \rightarrow CO_2(g) + 2H_2O(l)$

C. $2C(s, graphite) + O_2(g) + 4H_2(g) \rightarrow 2CH_3OH(l)$

D. $C(s, graphite) + ½\ O_2(g) + 2H_2(g) \rightarrow CH_3OH(l)$

Chapter 6 – Practice

6.4 and 6.6 Enthalpy

Hess's Law states when reactants are converted to products, the change in enthalpy is the same whether the reaction takes place in one step or in a series of steps. This is valid because change in enthalpy is

A. extensive.

B. a state function.

C. a type of energy.

D. equal to energy when no work is done.