

## Chapter 5. Thermochemistry

### 5.1 The Nature of Energy

- **Thermodynamics** = study of energy and its transformations.
- **Thermochemistry** = the study of the relationships between chemical reactions and energy changes involving heat.

#### Kinetic Energy and Potential Energy

- **Kinetic energy** = energy of motion:

$$E_k = \frac{1}{2}mv^2$$

- **Potential energy** = the energy an object possesses by virtue of its position or composition.
  - Electrostatic energy ( $E_d$ ) is an example
    - It arises from interactions between charged particles.  $K = 8.99 \times 10^9 \text{ J-m/C}^2$

$$e_l = \frac{kQ_1 Q_2}{2r}$$

- Potential energy can be converted into kinetic energy.
  - Example: A ball of clay dropped off a building.

#### Units of Energy

- SI unit is the **joule**, J.
- From  $E_k = \frac{1}{2}mv^2$ ,  $1\text{J} = 1\text{kg} \times \frac{\text{m}^2}{\text{s}^2}$
- Traditionally, we use the **calorie** as a unit of energy.
  - $1 \text{ cal} = 4.184 \text{ J}$  (exactly)
- The nutritional Calorie, Cal = 1,000 cal.

#### System and Surroundings

- A **system** is the part of the universe we are interested in studying.
- **Surroundings** are the rest of the universe (i.e., the surroundings are the portions of the universe not involved in the system).
- Example: If we are interested in the interaction between hydrogen and oxygen in a cylinder, then the  $\text{H}_2$  and  $\text{O}_2$  in the cylinder form a system.

#### Transferring Energy: Work and Heat

- From physics:
  - **Force** = a push or pull on an object.
  - **Work** = the energy used to move an object against a force.
$$w = F \times d$$
  - **Heat** = the energy transferred from a hotter object to a colder one.
  - **Energy** = the capacity to do work or to transfer heat.
  - **Internal Energy** = total energy of a system
  - **Absolute** internal energy **cannot be measured**.
  - **Change in internal energy**:  $\Delta E = E_{\text{final}} - E_{\text{initial}}$

## 5.2 The First Law of Thermodynamics

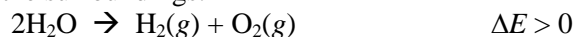
- **The first law of thermodynamics:** Energy cannot be created or destroyed.
- The first law of thermodynamics = the law of conservation of energy.
  - That is, the energy of (system + surroundings) is constant.
  - Thus any energy transferred from a system must be transferred to the surroundings (and vice versa).

### Internal Energy

- The total energy of a system = **internal energy**.
  - = sum of all the kinetic and potential energies of all components of the system.
- Absolute internal energy **cannot** be measured, only changes in internal energy.
- **Change in internal energy,  $\Delta E = E_{\text{final}} - E_{\text{initial}}$ .**
- Example: A mixture of  $\text{H}_2(\text{g})$  and  $\text{O}_2(\text{g})$  has a higher internal energy than  $\text{H}_2\text{O}(\text{g})$ .
- Going from  $\text{H}_2(\text{g})$  and  $\text{O}_2(\text{g})$  to  $\text{H}_2\text{O}(\text{g})$  results in a negative change in internal energy, indicating that the system has lost energy to the surroundings:



- Going from  $\text{H}_2\text{O}(\text{g})$  to  $\text{H}_2(\text{g})$  and  $\text{O}_2(\text{g})$  results in a positive change in internal energy, indicating that the system has gained energy from the surroundings:

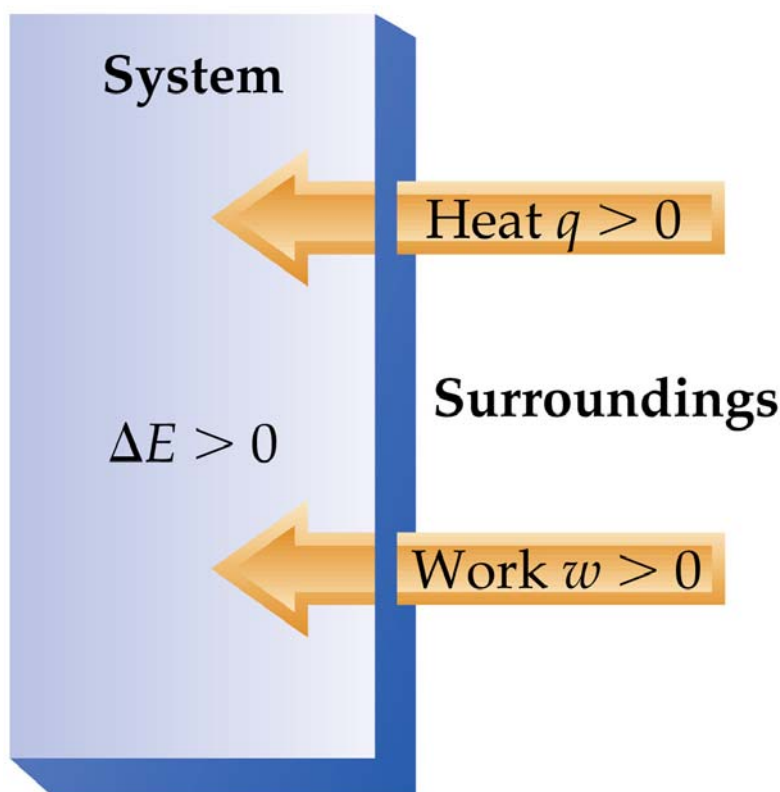


### Relating $\Delta E$ to Heat and Work

- From the first law of thermodynamics:
  - When a system undergoes a physical or chemical change, the change in internal energy is given by the heat added to or liberated from the system plus the work done on or by the system:

$$\Delta E = q + w$$

- Heat flowing from the surroundings to the system is positive,  $q > 0$ .
- Work done by the surroundings on the system is positive,  $w > 0$ .



**TABLE 5.1 Sign Conventions Used and the Relationship Among  $q$ ,  $w$ , and  $\Delta E$**

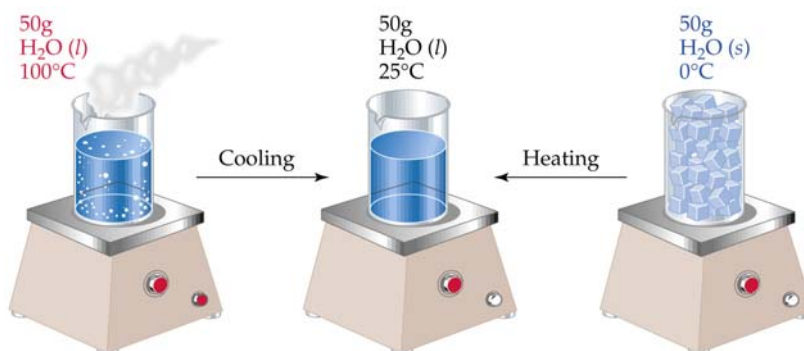
Sign Convention for $q$ :	Sign of $\Delta E = q + w$
$q > 0$ : Heat is transferred from the surroundings to the system	$q > 0$ and $w > 0$ : $\Delta E > 0$
$q < 0$ : Heat is transferred from the system to the surroundings	$q > 0$ and $w < 0$ : The sign of $\Delta E$ depends on the magnitudes of $q$ and $w$
<b>Sign Convention for <math>w</math>:</b>	$q < 0$ and $w > 0$ : The sign of $\Delta E$ depends on the magnitudes of $q$ and $w$
$w > 0$ : Work is done by the surroundings on the system	$q < 0$ and $w < 0$ : $\Delta E < 0$
$w < 0$ : Work is done by the system on the surroundings	

### Endothermic and Exothermic Processes

- An **endothermic** process is one that *absorbs* heat from the surroundings.
  - An endothermic reaction feels cold.
- An **exothermic** process is one that *transfers* heat to the surroundings.
  - An exothermic reaction feels hot.

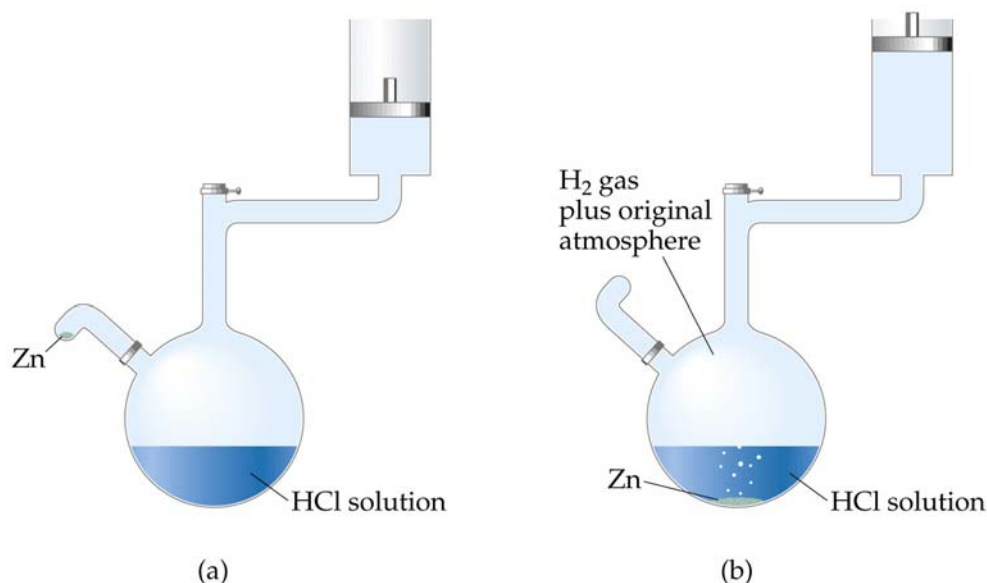
### State Functions

- A **state function** depends **only** on the initial and final states of a system.
  - e.g. The altitude difference between Denver and Chicago does not depend on whether you fly or drive, only on the elevation of the two cities above sea level.
  - Similarly, the internal energy of 50 g of  $\text{H}_2\text{O}(l)$  at  $25^\circ\text{C}$  does not depend on whether we cool 50 g of  $\text{H}_2\text{O}(l)$  from  $100^\circ\text{C}$  to  $25^\circ\text{C}$  or heat 50 g of  $\text{H}_2\text{O}(l)$  at  $0^\circ\text{C}$  to  $25^\circ\text{C}$



## 5.3 Enthalpy

- Chemical changes may involve release or absorption of heat.
- Many also involve work done on or by the system.
  - Work is often either electrical or mechanical work.
  - Mechanical work done by a system involving expanding gases is called **pressure-volume work** or  $P\text{-}V$  work.



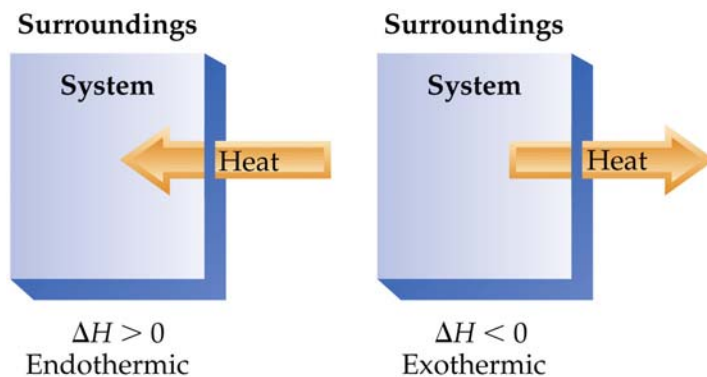
- The heat transferred between the system and surroundings during a chemical reaction carried out under constant pressure is called **enthalpy**,  $H$ .
- Again, we can only measure the change in enthalpy,  $\Delta H$ .
- Mathematically,

$$\Delta H = H_{\text{final}} - H_{\text{initial}} = \Delta E + P\Delta V$$

$$w = -P\Delta V; \Delta E = q + w$$

$$\Delta H = \Delta E + P\Delta V = q_p + w - w = q_p$$

- For most reactions  $P\Delta V$  is small thus  $\Delta H = \Delta E$
- Heat transferred from surroundings to the system has a positive enthalpy (i.e.,  $\Delta H > 0$  for an endothermic reaction).
- Heat transferred from the system to the surroundings has a negative enthalpy (i.e.,  $\Delta H < 0$  for an exothermic reaction).
- Enthalpy is a state function.



## 5.4 Enthalpies of Reaction

- For a reaction,  $\Delta H_{\text{rxn}} = H(\text{products}) - H(\text{reactants})$ .
  - The enthalpy change that accompanies a reaction is called the **enthalpy of reaction** or *heat of reaction* ( $\Delta H_{\text{rxn}}$ ).
  - Consider the thermochemical equation for the production of water:
$$2\text{H}_2(g) + \text{O}_2(g) \rightarrow 2\text{H}_2\text{O}(g) \quad \Delta H = -483.6 \text{ kJ}$$
    - The equation tells us that 483.6 kJ of energy are released to the surroundings when water is formed.
    - $\Delta H$  noted at the end of the balanced equation depends on the number of moles of reactants and products associated with the  $\Delta H$  value.
    - These equations are called *thermochemical equations*.
  - Enthalpy diagrams are used to represent enthalpy changes associated with a reaction.
  - In the enthalpy diagram for the combustion of  $\text{H}_2(g)$ , the reactants,  $2\text{H}_2(g) + \text{O}_2(g)$ , have a higher enthalpy than the products  $2\text{H}_2\text{O}(g)$ ; this reaction is exothermic.
- Enthalpy is an extensive property.
    - Therefore, the *magnitude* of enthalpy is directly proportional to the amount of reactant consumed.
    - Example: If one mol of  $\text{CH}_4$  is burned in oxygen to produce  $\text{CO}_2$  and water, 890 kJ of heat is released to the surroundings. If two mol of  $\text{CH}_4$  is burned, then 1780 kJ of heat is released.

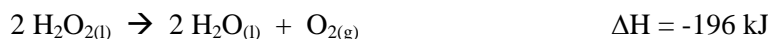
**Sample Exercise 5.5 (p. 167)**

How much heat is released when 4.50 g of methane gas is burned in a constant-pressure system?  
(Use the information given in Equation 5.18.)

(-250 kJ)

**Practice Exercise 5.5**

Hydrogen peroxide can decompose to water and oxygen by the following reaction:



Calculate the value of  $q$  when 5.00 g of  $\text{H}_2\text{O}_{2(l)}$  decomposes at constant pressure.

(-14.4 kJ)

- The sign of  $\Delta H$  depends on the direction of the reaction.
  - The enthalpy change for a reaction is equal in magnitude but opposite in sign to  $\Delta H$  for the reverse reaction.
  - Example:  $\text{CH}_4(g) + 2\text{O}_2(g) \rightarrow \text{CO}_2(g) + 2\text{H}_2\text{O}(l) \quad \Delta H = -890 \text{ kJ}$ ,
  - But  $\text{CO}_2(g) + 2\text{H}_2\text{O}(l) \rightarrow \text{CH}_4(g) + 2\text{O}_2(g) \quad \Delta H = +890 \text{ kJ}$ .
- Enthalpy change depends on state.
  - $2\text{H}_2\text{O}(g) \rightarrow 2\text{H}_2\text{O}(l) \quad \Delta H = -88 \text{ kJ}$

**5.5 Calorimetry**

- Calorimetry** is a measurement of heat flow.
- Calorimeter** is an apparatus that measures heat flow.

**Heat Capacity and Specific Heat**

- Heat capacity** is the amount of energy required to raise the temperature of an object by  $1^\circ\text{C}$ .
  - Molar heat capacity** is the heat capacity of 1 mol of a substance.
  - Specific heat**, or specific heat capacity is the heat capacity of 1 g of a substance.
- Heat,  $q = (\text{specific heat}) \times (\text{grams of substance}) \times \Delta T$ .
- Be careful of the sign of  $q$ .

**Table 5.2 Specific Heats of Some Substances at 298 K**

Elements		Compounds	
Substance	Specific Heat (J/g-K)	Substance	Specific Heat (J/g-K)
$\text{N}_{2(g)}$	1.04	$\text{H}_2\text{O}_{(l)}$	4.18
$\text{Al}_{(s)}$	0.90	$\text{CH}_{4(g)}$	2.20
$\text{Fe}_{(s)}$	0.45	$\text{CO}_{2(g)}$	0.84
$\text{Hg}_{(l)}$	0.14	$\text{CaCO}_{3(s)}$	0.82

**Sample Problem 5.6 (p. 170)**

- a) How much heat is needed to warm 250 g of water from 22°C to near its boiling point, 98°C? The specific heat of water is 4.18 J/g-K.

( $7.9 \times 10^4$  J)

- b) What is the molar heat capacity of water?

(75.2 J/mol-K)

**Practice Problem 5.6**

- a) Large beds of rocks are used in some solar-heated homes to store heat. Assume that the specific heat of the rocks is 0.82 J/g-K. Calculate the quantity of heat absorbed by 50.0 kg of rocks if their temperature increases by 12.0°C.

( $4.9 \times 10^5$  J)

- b) What temperature change would these rocks undergo if they emitted 450 kJ of heat?

(11k = 11°C decrease)

**Constant-Pressure Calorimetry**

- Most common technique: use atmospheric pressure as the constant pressure.
- Recall  $\Delta H = q_p$ .
- Easiest method: use a coffee cup calorimeter.  
 $q_{\text{soln}} = (\text{specific heat of solution}) \times (\text{grams of solution}) \times \Delta T = -q_{\text{rxn}}$
- For dilute aqueous solutions, the specific heat of the solution will be close to that of pure water.

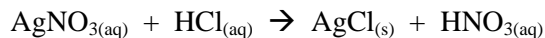
**Sample Exercise 5.7 (p.171)**

When a student mixes 50 mL of 1.0 M HCl and 50 mL of 1.0 M NaOH in a coffee-cup calorimeter, the temperature of the resultant solution increases from 21.0°C to 27.5°C. Calculate the enthalpy change for the reaction, assuming that the calorimeter loses only a negligible quantity of heat, that the total volume of the solution is 100 mL, that its density is 1.0 g/mL, and that its specific heat is 4.18 J/g·K.

(-54 kJ/mol)

**Practice Exercise 5.7**

When 50.0 mL of 0.100 M AgNO<sub>3</sub> and 50.0 mL of 0.100 M HCl are mixed in a constant-pressure calorimeter, the temperature of the mixture increases from 22.30°C to 23.11°C. The temperature increase is caused by the following reaction:



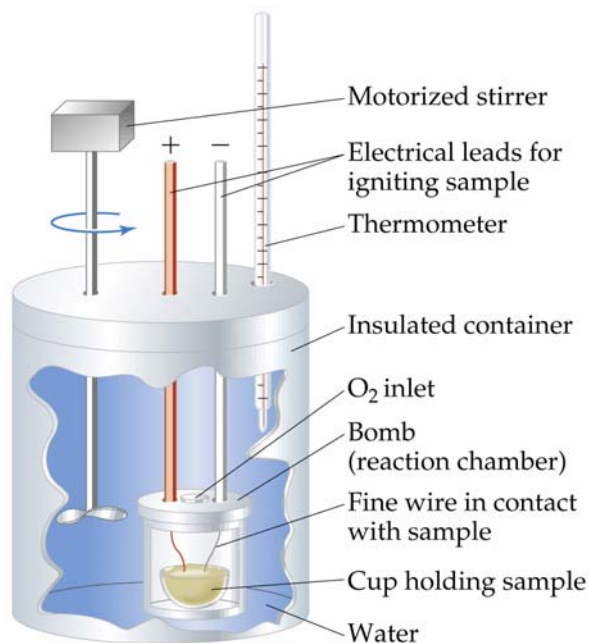
Calculate  $\Delta H$  for this reaction, assuming that the combined solution has a mass of 100.0 g and a specific heat of 4.18 J/g·°C.

(-68 kJ/mol)



**Bomb Calorimetry (Constant-Volume Calorimetry)**

- Reactions can be carried out under conditions of constant volume instead of constant pressure.
- Constant volume calorimetry is carried out in a **bomb calorimeter**.
- The most common type of reaction studied under these conditions is combustion.
- If we know the heat capacity of the calorimeter,  $C_{\text{calorimeter}}$ , then the heat of reaction,  
$$q_{\text{rxn}} = -C_{\text{calorimeter}} \times \Delta T.$$
- Since the reaction is carried out under constant volume,  $q$  relates to  $\Delta E$ .

**Sample Exercise 5.8 (p. 172)**

Methylhydrazine ( $CH_6N_2$ ) is commonly used as a liquid rocket fuel. The combustion of methylhydrazine with oxygen produces  $N_{2(g)}$ ,  $CO_{2(g)}$ , and  $H_2O_{(l)}$ :



When 4.00 g of methylhydrazine is combusted in a bomb calorimeter, the temperature of the calorimeter increases from  $25.00^\circ\text{C}$  to  $39.50^\circ\text{C}$ . In a separate experiment the heat capacity of the bomb calorimeter is measured to be  $7.794 \text{ kJ}/^\circ\text{C}$ . What is the heat of reaction for the combustion of a mole of  $CH_6N_2$  in this calorimeter?

$(-1.30 \times 10^3 \text{ kJ/mol } CH_6N_2)$

**Practice Exercise 5.8**

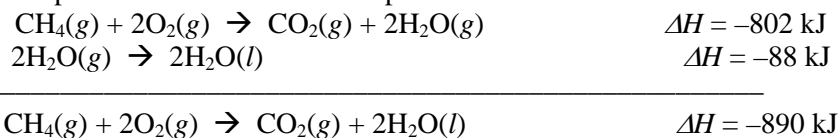
A 0.5865-g sample of lactic acid ( $\text{HC}_3\text{H}_5\text{O}_3$ ) is burned in a calorimeter whose heat capacity is  $4.812 \text{ kJ/}^\circ\text{C}$ . The temperature increases from  $23.10^\circ\text{C}$  to  $24.95^\circ\text{C}$ . Calculate the heat of combustion of

a) lactic acid per gram ( $-15.2 \text{ kJ/g}$ ) and

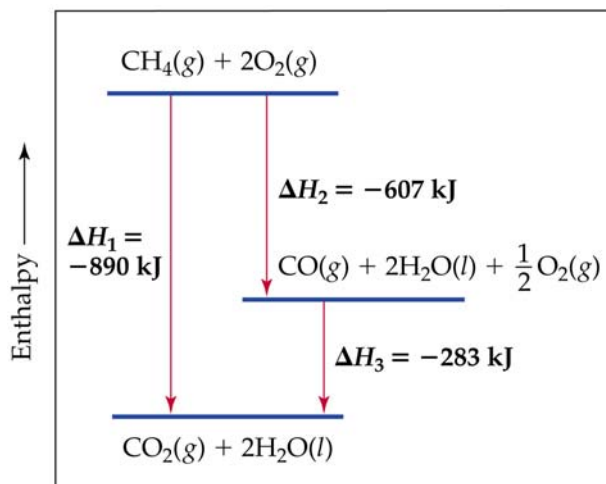
b) per mole ( $-1370 \text{ kJ/mol}$ )

**5.6 Hess's Law**

- Hess's Law:** If a reaction is carried out in a series of steps,  $\Delta H$  for the reaction is the sum of  $\Delta H$  for each of the steps.
- The total change in enthalpy is independent of the number of steps.
- Total  $\Delta H$  is also independent of the nature of the path.



- Therefore, for the reaction  $\text{CH}_4(\text{g}) + 2\text{O}_2(\text{g}) \rightarrow \text{CO}_2(\text{g}) + 2\text{H}_2\text{O}(\text{l})$ ,  $\Delta H = -890 \text{ kJ}$ .
- Note that  $\Delta H$  is sensitive to the states of the reactants and products.
- Hess's law allows us to calculate enthalpy data for reactions which are difficult to carry out directly:  $\text{C}(\text{s}) + \text{O}_2(\text{g})$  produces a *mixture* of  $\text{CO}(\text{g})$  and  $\text{CO}_2(\text{g})$ .

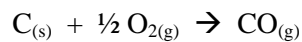


**Sample Exercise 5.9 (p. 175)**

The enthalpy of combustion of C to CO<sub>2</sub> is -393.5 kJ/mol C, and the enthalpy of combustion of CO to CO<sub>2</sub> is -283.0 kJ/mol CO:



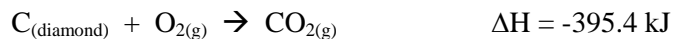
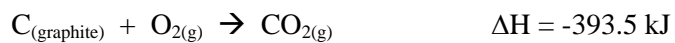
Using these data, calculate the enthalpy of combustion of C to CO:



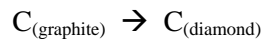
(-110.5 kJ)

**Practice Exercise 5.9**

Carbon occurs in two forms, graphite and diamond. The enthalpy of combustion of graphite is -393.4 kJ/mol and that of diamond is -395.4 kJ:



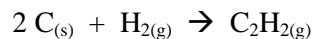
Calculate  $\Delta H$  for the conversion of graphite to diamond:



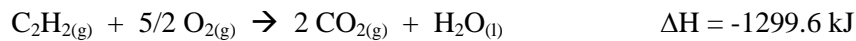
(+1.9 kJ)

**Sample Exercise 5.10 (p. 175)**

Calculate  $\Delta H$  for the reaction



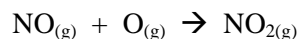
Given the following reactions and their respective enthalpy changes:



(226.8 kJ)

**Practice Exercise 5.10**

Calculate  $\Delta H$  for the reaction



given the following reactions and their respective enthalpy changes:

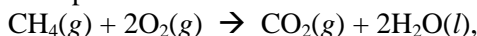


(-304.1 kJ)

## 5.7 Enthalpies of Formation

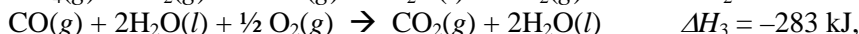
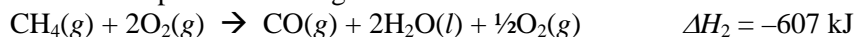
- Hess's law states that if a reaction is carried out in a number of steps,  $\Delta H$  for the overall reaction is the sum of the  $\Delta H$ s for each of the individual steps.
- Consider the formation of  $\text{CO}_2(\text{g})$  and  $2\text{H}_2\text{O}(\text{l})$  from  $\text{CH}_4(\text{g})$  and  $2\text{O}_2(\text{g})$ .

- If the reaction proceeds in one step:



then  $\Delta H_1 = -890 \text{ kJ}$ .

- However, if the reaction proceeds through a CO intermediate:



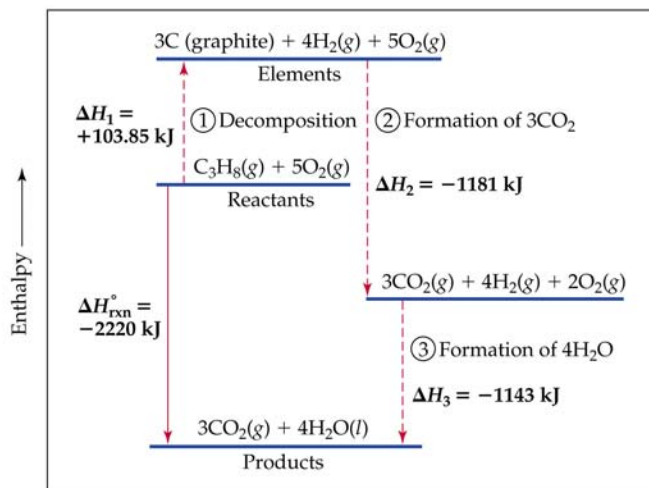
Then  $\Delta H$  for the overall reaction is:

$$\Delta H_2 + \Delta H_3 = -607 \text{ kJ} - 283 \text{ kJ} = -890 \text{ kJ} = \Delta H_1$$

- If a compound is formed from its constituent elements, then the enthalpy change for the reaction is called the **enthalpy of formation**,  $\Delta H_f$ .
- Standard state* (standard conditions) refer to the substance at:
  - 1 atm and  $25^\circ\text{C}$  (298 K).
- Standard enthalpy**,  $\Delta H^\circ$ , is the enthalpy measured when everything is in its standard state.
- Standard enthalpy of formation** of a compound,  $\Delta H_f^\circ$  is the enthalpy change for the formation of 1 mol of compound with all substances in their standard states.
- If there is more than one state for a substance under standard conditions, the more stable one is used.  
Example: When dealing with carbon we use graphite because graphite is more stable than diamond or  $\text{C}_{60}$ .
- The standard enthalpy of formation of the most stable form of an element is zero.

TABLE 5.3 Standard Enthalpies of Formation,  $\Delta H_f^\circ$ , at 298 K

Substance	Formula	$\Delta H_f^\circ$ (kJ/mol)	Substance	Formula	$\Delta H_f^\circ$ (kJ/mol)
Acetylene	$\text{C}_2\text{H}_2(\text{g})$	226.7	Hydrogen chloride	$\text{HCl}(\text{g})$	-92.30
Ammonia	$\text{NH}_3(\text{g})$	-46.19	Hydrogen fluoride	$\text{HF}(\text{g})$	-268.6
Benzene	$\text{C}_6\text{H}_6(\text{l})$	49.0	Hydrogen iodide	$\text{HI}(\text{g})$	25.9
Calcium carbonate	$\text{CaCO}_3(\text{s})$	-1207.1	Methane	$\text{CH}_4(\text{g})$	-74.8
Calcium oxide	$\text{CaO}(\text{s})$	-635.5	Methanol	$\text{CH}_3\text{OH}(\text{l})$	-238.6
Carbon dioxide	$\text{CO}_2(\text{g})$	-393.5	Propane	$\text{C}_3\text{H}_8(\text{g})$	-103.85
Carbon monoxide	$\text{CO}(\text{g})$	-110.5	Silver chloride	$\text{AgCl}(\text{s})$	-127.0
Diamond	$\text{C}(\text{s})$	1.88	Sodium bicarbonate	$\text{NaHCO}_3(\text{s})$	-947.7
Ethane	$\text{C}_2\text{H}_6(\text{g})$	-84.68	Sodium carbonate	$\text{Na}_2\text{CO}_3(\text{s})$	-1130.9
Ethanol	$\text{C}_2\text{H}_5\text{OH}(\text{l})$	-277.7	Sodium chloride	$\text{NaCl}(\text{s})$	-410.9
Ethylene	$\text{C}_2\text{H}_4(\text{g})$	52.30	Sucrose	$\text{C}_{12}\text{H}_{22}\text{O}_{11}(\text{s})$	-2221
Glucose	$\text{C}_6\text{H}_{12}\text{O}_6(\text{s})$	-1273	Water	$\text{H}_2\text{O}(\text{l})$	-285.8
Hydrogen bromide	$\text{HBr}(\text{g})$	-36.23	Water vapor	$\text{H}_2\text{O}(\text{g})$	-241.8



**Sample Exercise 5.11 (p. 177)**

For which of the following reactions at 25°C would the enthalpy change represent a standard enthalpy of formation? For those where it does not, what changes would need to be made in the reaction conditions?

- a)  $2 \text{Na}_{(s)} + \frac{1}{2} \text{O}_{2(g)} \rightarrow \text{Na}_2\text{O}_{(s)}$   
 b)  $2 \text{K}_{(l)} + \text{Cl}_{2(g)} \rightarrow 2 \text{KCl}_{(s)}$   
 c)  $\text{C}_6\text{H}_{12}\text{O}_{6(s)} \rightarrow 6 \text{C}_{(\text{diamond})} + 6 \text{H}_{2(g)} + 3 \text{O}_{2(g)}$

**Practice Exercise 5.11**

Write the equation corresponding to the standard enthalpy of formation of liquid carbon tetrachloride ( $\text{CCl}_4$ ).

**Using Enthalpies of Formation to Calculate Enthalpies of Reaction**

- Use Hess's law!
- Example: Calculate  $\Delta H$  for
 
$$\text{C}_3\text{H}_8(g) + 5\text{O}_2(g) \rightarrow 3\text{CO}_2(g) + 4\text{H}_2\text{O}(l)$$
- We start with the reactants, decompose them into elements, then rearrange the elements to form products. The overall enthalpy change is the sum of the enthalpy changes for each step.
  - Decomposing into elements (note  $\text{O}_2$  is already elemental, so we concern ourselves with  $\text{C}_3\text{H}_8$ ):
 
$$\text{C}_3\text{H}_8(g) \rightarrow 3\text{C}(s) + 4\text{H}_2(g) \quad \Delta H_1 = -\Delta H_f^\circ[\text{C}_3\text{H}_8(g)]$$
  - Next we form  $\text{CO}_2$  and  $\text{H}_2\text{O}$  from their elements:
 
$$3\text{C}(s) + 3\text{O}_2(g) \rightarrow 3\text{CO}_2(g) \quad \Delta H = 3 \Delta H_f^\circ[\text{CO}_2(g)]$$

$$4\text{H}_2(g) + 2\text{O}_2(g) \rightarrow 4\text{H}_2\text{O}(l) \quad \Delta H_3 = 4 \Delta H_f^\circ[\text{H}_2\text{O}(l)]$$
  - We look up the values and add:
 
$$\Delta H_{\text{rxn}}^\circ = -1(-103.85 \text{ kJ}) + 3(-393.5 \text{ kJ}) + 4(-285.8 \text{ kJ}) = -2220 \text{ kJ}$$
- In general:
 
$$\Delta H_{\text{rxn}}^\circ = \sum n \Delta H_f^\circ(\text{products}) - \sum m \Delta H_f^\circ(\text{reactants})$$
  - Where  $n$  and  $m$  are the stoichiometric coefficients.

**Sample Exercise 5.12 (p. 179)**

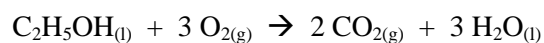
- a) Calculate the standard enthalpy change for the combustion of 1 mol of benzene,  $\text{C}_6\text{H}_{6(l)}$ , to  $\text{CO}_{2(g)}$  and  $\text{H}_2\text{O}_{(l)}$ .  
 (-3267 kJ)

- b) Compare the quantity of heat produced by combustion of 1.00 g propane ( $\text{C}_3\text{H}_8$ ) to that produced by 1.00 g benzene.

( $\text{C}_3\text{H}_{8(g)}$ : -50.3 kJ/g;  $\text{C}_6\text{H}_{6(l)}$ : -41.8 kJ/g)

### Practice Exercise 5.12

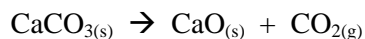
Using the standard enthalpies of formation listed in Table 5.3, calculate the enthalpy change for the combustion of 1 mol of ethanol:



(-1367kJ)

### Sample Exercise 5.13 (p. 180)

The standard enthalpy change for the reaction



is 178.1 kJ. From the values for the standard enthalpies of formation of  $\text{CaO}_{(s)}$  and  $\text{CO}_{2(g)}$  given in Table 5.3, calculate the standard enthalpy of formation of  $\text{CaCO}_{3(s)}$ .

(-1207.1 kJ/mol)