

## Chapter 6

### Thermochemistry

**Understanding Heats of Reaction - *thermodynamics*** is the science of the relationships between heat and other forms of energy. *Thermochemistry* refers to the heat absorbed or evolved during a chemical reaction.

### Law of Conservation of Energy

Energy can be converted from one form to another but can neither be created nor destroyed.

( $E_{\text{universe}}$  is constant)

### Energy

- **Energy** is defined as *the capacity to move matter*.

**Energy can be in many forms:**

- > **Radiant Energy** - Electromagnetic radiation.
- > **Thermal Energy** - Associated with random motion of a molecule or atom.
- > **Chemical Energy** - Energy stored within the structural limits of a molecule or atom.

### CONCEPT CHECK 6.1

A solar-powered water pump has photovoltaic cells that protrude from top panels. These cells collect energy from sunlight, storing it momentarily in a battery, which later runs an electric motor that pumps water up to a storage tank on a hill. What energy conversions are involved in using sunlight to pump water into the storage tank?

**Energy and Its Units** - energy is the potential or capacity to move matter

- 1 **Kinetic energy.**  $E_k = \frac{1}{2} mv^2$   
SI unit of energy is the joule (J) which has units  $\text{kg} \cdot \text{m}^2/\text{s}^2$ . 1 calorie (cal) is equal to the amt of heat necessary to raise 1 g of  $\text{H}_2\text{O}$  by  $1^\circ\text{C}$ .  $1 \text{ cal} = 4.184 \text{ J}$

Potential energy and kinetic energy. Photo Courtesy of Tennessee Valley Authority.



### A Problem to Consider

- Consider the kinetic energy of a person whose mass is 130 lb (59.0 kg) traveling in a car at 60 mph (26.8 m/s).

$$E_k = \frac{1}{2}(59.0 \text{ kg}) \times (26.8 \text{ m/s})^2$$

$$E_k = 2.12 \times 10^4 \text{ J}$$

- The SI unit of energy,  $\text{kg} \cdot \text{m}^2/\text{s}^2$ , is given the name **Joule**.

## Energy

- **Potential Energy:** This energy depends on the “position” (such as height) in a “field of force” (such as gravity).
- For example, water of a given mass  $m$  at the top of a dam is at a relatively high “position”  $h$  in the “gravitational field”  $g$  of the earth.

$$E_p = mgh$$

## A Problem to Consider

- Consider the potential energy of 1000 lb of water (453.6 kg) at the top of a 300 foot dam (91.44 m).

$$E_p = (453.6 \text{ kg}) \times (9.80 \text{ m/s}^2) \times (91.44 \text{ m})$$

$$E_p = 4.06 \times 10^5 \text{ kg} \cdot \text{m}^2 / \text{s}^2$$

$$E_p = 4.06 \times 10^5 \text{ J}$$

## Energy

- **Internal Energy** is the energy of the *particles making up a substance*.
- The *total energy* of a system is the sum of its kinetic energy, potential energy, and internal energy,  $U$ .

$$E_{\text{tot}} = E_k + E_p + U$$

## First Law

### First Law of Thermodynamics:

The energy of the universe is constant.

## First Law

$$\Delta E = q + w$$

$\Delta E$  = change in system's internal energy

$q$  = heat

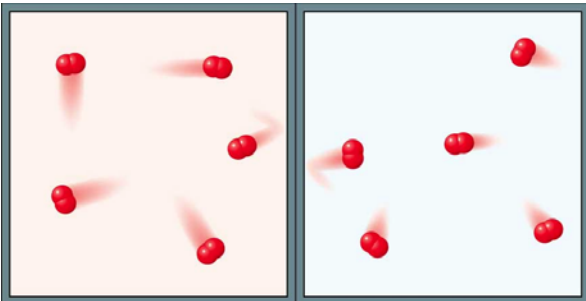
$w$  = work

## Temperature v. Heat

Temperature reflects random motions of particles, therefore related to kinetic energy of the system.

Heat involves a transfer of energy between 2 objects due to a temperature difference

Figure 6.5: A kinetic-theory



## Work

work = force  $\times$  distance

since pressure = force / area,

work = pressure  $\times$  volume

$$w_{\text{system}} = -P\Delta V$$

Where  $\Delta V = V_{\text{final}} - V_{\text{initial}}$

## First Law of Thermodynamics

Both work & heat are ways in which energy can be transferred.

The way that energy transfer is divided between work and heat depends on the conditions of transfer - and is called the "pathway". However, the total amount of energy transferred ( $w + q$ ) will always remain constant.

## State Function

Depends only on the present state of the system - not how it arrived there.

It is independent of pathway.

**Internal Energy is a State Function**

## Campsite to illustrate altitude.



Internal Energy is a State Function

But work and heat are not!

## Heat of Reaction

- **Heat** is defined as the energy that flows into or out of a system because of a difference in temperature between the system and its surroundings.
  - Heat flows from a region of higher temperature to one of lower temperature; once the temperatures become equal, heat flow stops.

## System and Surroundings

System: That on which we focus attention

Surroundings: Everything else in the universe

$$\text{Universe} = \text{System} + \text{Surroundings}$$

## Heat of Reaction

- In chemical reactions, heat is often transferred from the “system” to its “surroundings,” or vice versa.
- The substance or mixture of substances under study in which a change occurs is called the *thermodynamic system* (or simply *system*.)
- The *surroundings* are everything in the vicinity of the thermodynamic system.

Illustration of a thermodynamic system.

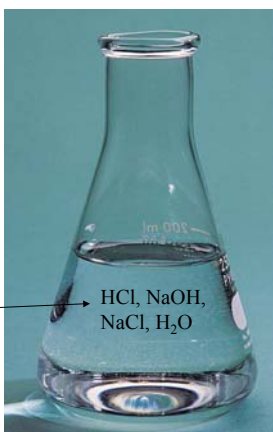
Photo courtesy of American Color.

System = reaction between HCl and NaOH

**System** →

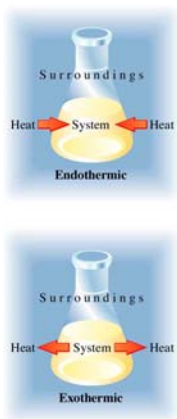
HCl, NaOH,  
NaCl, H<sub>2</sub>O

The surroundings = flask, air, solvent (if not involved in the reaction) - “everything else”!



1. An **exothermic** process is a chemical change in which heat is evolved ( $q$  is negative)
2. An **endothermic** process is a chemical reaction or physical change in which heat is absorbed ( $q$  is positive)

## Endothermic and exothermic.



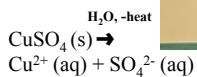
A test tube that contains anhydrous copper (II) sulfate and a thermometer that registers 26.1° C. Photo courtesy of American Color.



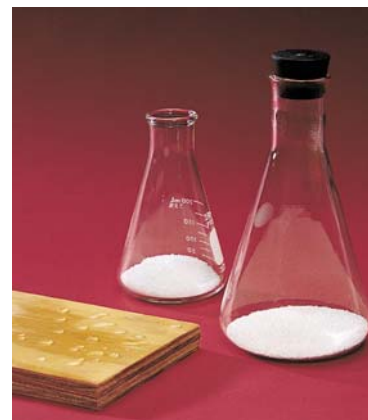
Water has been added. The thermometer now registers 90.2° C.

Photo courtesy of American Color.

This dissolving process is exothermic!

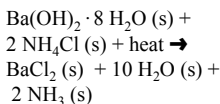


Barium hydroxide octahydrate and ammonium salt are mixed. Photo courtesy of American Color.



The flask and board are frozen solidly together.  
Photo courtesy of American

Color.



An endothermic reaction!



## Exo and Endothermic

Heat exchange accompanies chemical reactions.

Exothermic: Heat flows out of the system (to the surroundings).

Endothermic: Heat flows into the system (from the surroundings).

## Enthalpy and Enthalpy Change

- The heat absorbed or evolved by a reaction depends on the conditions under which it occurs - the "pathway".
  - ♦ Usually, a reaction takes place in an open vessel, and therefore at the constant pressure of the atmosphere.
  - ♦ The heat of this type of reaction is denoted  $q_p$ , the heat at constant pressure.

## Enthalpy and Enthalpy Change

- **Enthalpy and Internal Energy**
  - The internal energy of a system,  $U$ , is precisely defined as the heat at constant pressure plus the work done by the system:

$$U = q_p + w$$

- In chemical systems, work is defined as a change in volume at a given pressure, that is:

$$w = -P\Delta V$$

(if volume of system expands – it does work on surroundings)



**Internal energy (U).** The sum of kinetic & potential energies of the atoms making up a substance as well as the subatomic particles of each atom.

$$E_{\text{Total}} = E_k + E_p + U$$

## C Enthalpy and Enthalpy Change

**Enthalpy is a state function.** A state function is a property of a system that depends only on its present state, which is determined by variables such as temperature and pressure, and is independent of any previous history of the system. **Thus, changes in enthalpy,  $\Delta H$ , depend only on the initial and final states of the system not on the path (or way) in which the final state is reached.**

## Enthalpy

Enthalpy =  $H = E + PV$  (all state functions)

$$\Delta E = \Delta H - \Delta(PV)$$

$$\Delta H = \Delta E + \Delta(PV)$$

At constant pressure,

$$q_p = \Delta E + P\Delta V, \text{ or}$$

$$\Delta E = q_p - P\Delta V$$

where  $q_p = \Delta H$  at constant pressure,  $\Delta E = U_f - U_i$

$\Delta H$  = energy flow as heat (at constant pressure)

## Enthalpy and Enthalpy Change

- **Enthalpy**, denoted **H**, is an extensive property of a substance that can be used to obtain the heat absorbed or evolved in a chemical reaction.
  - An **extensive property** is one that depends on the quantity of substance. (**other examples - volume, mass, energy**). An **intensive property** does not depend on amount of substance - e.g. temperature, vapor pressure.
  - Enthalpy is also a **state function**, a property of a system that depends only on its present state and is independent of any previous history of the system.

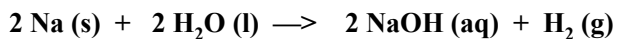
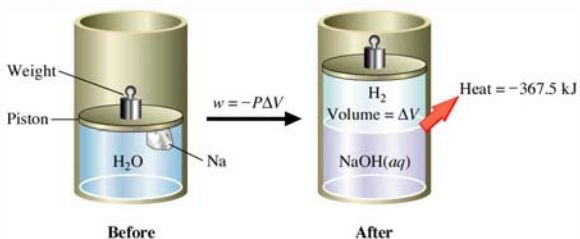
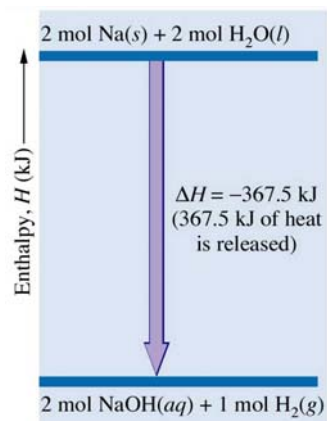
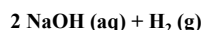
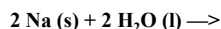
**Enthalpy of Reaction** ( $\Delta H$ ) is the change in enthalpy for a reaction at a given temperature and pressure.

$$\Delta H = H(\text{products}) - H(\text{reactants})$$

$$= q_p \text{ (at constant pressure)}$$

So  $\Delta H$  is essentially the heat evolved or absorbed by the system (reaction) in an open vessel where the work portion of  $\Delta U$  is unmeasured.

An enthalpy diagram.



$$\Delta E = (-367.5 \text{ kJ} - P \Delta V)$$

## Measuring Heats of Reaction

- To see how **heats of reactions** are measured, we must look at the heat required to raise the temperature of a substance, because a thermochemical measurement is based on the relationship between heat and temperature change.
- The heat required to raise the temperature of a substance is its **heat capacity**.

## Measuring Heats of Reaction

- **Heat Capacity and Specific Heat**

- The **heat capacity**,  $C$ , of a sample of substance is the quantity of heat required to raise the temperature of one mole of the sample of substance one degree Celsius.
- Changing the temperature of the sample requires heat equal to:

$$q = nC\Delta T$$

## Heat Capacity

$$C = \frac{\text{heat absorbed}}{\text{mole substance increase in temperature}}$$
$$= \frac{\text{J}}{\text{mole}^\circ\text{C}} \text{ or } \frac{\text{J}}{\text{mole K}}$$

## Some Heat Exchange Terms

specific heat capacity,  $s$

heat capacity per gram =  $\text{J}^\circ\text{C g}$  or  $\text{J/K g}$

$q = s m \Delta T$ , where  $m$  is in g,  $\Delta T = T_f - T_i$

molar heat capacity,  $C$

heat capacity per mole =  $\text{J}^\circ\text{C mol}$  or  $\text{J/K mol}$

$q = nC \Delta T$ , where  $n = \text{no. moles}$ ,  $\Delta T = T_f - T_i$

**Measurement of Heat of Reaction** may be done using a calorimeter. A calorimeter is a device used to measure the heat absorbed or evolved during a physical or chemical change

## Coffee-cup calorimeter.

Reaction is done at constant pressure, so we are measuring  $q_p$  or  $\Delta H$



## Heats of Reaction: Calorimetry

- A **calorimeter** is a device used to measure the heat absorbed or evolved during a physical or chemical change. (see [Figure 6.11](#))
- The heat absorbed by the calorimeter and its contents is the negative of the heat of reaction.

$$q_{\text{calorimeter}} = -q_{\text{rxn}}$$

## A Problem to Consider

- Suppose one mole of iron requires 6.70 J of heat to raise its temperature by one degree Celsius. The quantity of heat required to raise the temperature of the piece of iron from 25.0 °C to 35.0 °C is: since  $C \times n = (6.70 \text{ J/mole } ^\circ\text{C}) \times 1 \text{ mole}$

$$q = nC\Delta T = (6.70 \text{ J/}^\circ\text{C}) \times (35.0^\circ\text{C} - 25.0^\circ\text{C})$$

$$q = 67.0 \text{ J}$$

## A Problem to Consider

- Calculate the heat absorbed by a calorimeter when the temperature of 15.0 grams of water is raised from 20.0 °C to 50.0 °C. (The specific heat of water is 4.184 J/g·°C.)

$$q = s \times m \times \Delta T$$

$$q = (4.184 \frac{\text{J}}{\text{g} \cdot ^\circ\text{C}}) \times (15.0\text{g}) \times (50.0 - 20.0^\circ\text{C})$$

$$q = 1.88 \times 10^3 \text{ J}$$

## A Problem to Consider

- When 23.6 grams of calcium chloride,  $\text{CaCl}_2$ , was dissolved in water in a calorimeter, the temperature rose from  $25.0^\circ\text{C}$  to  $38.7^\circ\text{C}$ .

If the heat capacity of the solution and the calorimeter is  $1258 \text{ J}/^\circ\text{C}$ , **what is the enthalpy change per mole** of calcium chloride?

## Heats of Reaction: Calorimetry

- First, let us calculate the heat absorbed by the calorimeter.

$$q_{\text{cal}} = (1258 \text{ J}/^\circ\text{C})(38.7^\circ\text{C} - 25.0^\circ\text{C})$$

$$q_{\text{cal}} = 1.72 \times 10^4 \text{ J}$$

- Now we must calculate the heat per mole of calcium chloride.

## Heats of Reaction: Calorimetry

- Calcium chloride has a molecular mass of  $111.1 \text{ g}$ , so

$$(23.6 \text{ g CaCl}_2) \times \frac{(1 \text{ mol CaCl}_2)}{111.1 \text{ g}} = 0.212 \text{ mol CaCl}_2$$

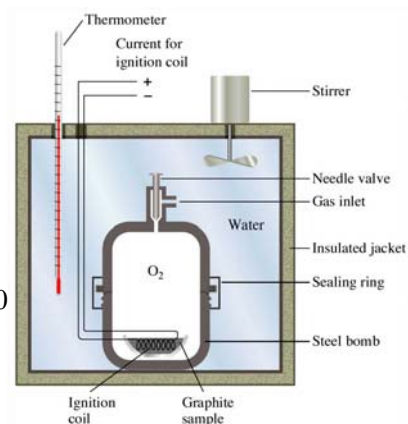
- Now we can calculate the heat per mole of calcium chloride, using  $q_{\text{cal}} = -q_{\text{rxn}}$ .

$$\Delta H = \frac{q_{\text{rxn}}}{\text{mol CaCl}_2} = \frac{-17.2 \text{ kJ}}{0.212 \text{ mol}} = -81.1 \text{ kJ/mol}$$

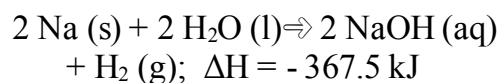
A bomb calorimeter

$$q_v = \Delta E$$

since  $w = 0$

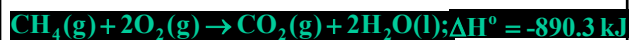


**Thermochemical Equations** - the chemical equation for a reaction (including phase labels) in which the equation is given a molar interpretation, and the enthalpy of reaction for these molar amounts is written directly after the equation



## Applying Stoichiometry and Heats of Reactions

- Consider the reaction of methane,  $\text{CH}_4$ , burning in the presence of oxygen at constant pressure. Given the following equation, how much heat could be obtained by the combustion of 10.0 grams  $\text{CH}_4$ ?



$$10.0 \text{ g CH}_4 \times \frac{1 \text{ mol CH}_4}{16.0 \text{ g}} \times \frac{-890.3 \text{ kJ}}{1 \text{ mol CH}_4} = -556 \text{ kJ}$$

## Hess's Law

- Hess's law of heat summation** states that for a chemical equation that can be written as the sum of two or more steps, the enthalpy change for the overall equation is the sum of the enthalpy changes for the individual steps.

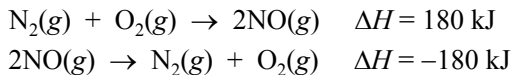
## Hess's Law

Reactants  $\rightarrow$  Products

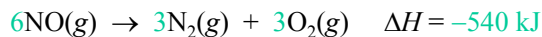
The change in enthalpy is the same whether the reaction takes place in one step or a series of steps.

## Calculations via Hess's Law

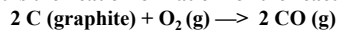
1. If a reaction is reversed,  $\Delta H$  is also reversed.



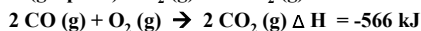
2. If the coefficients of a reaction are multiplied by an integer,  $\Delta H$  is multiplied by that same integer.



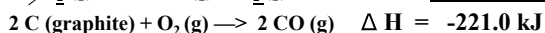
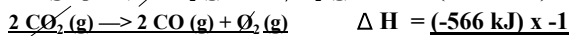
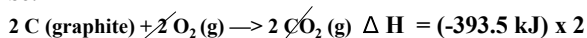
What is the heat of formation for the reaction?



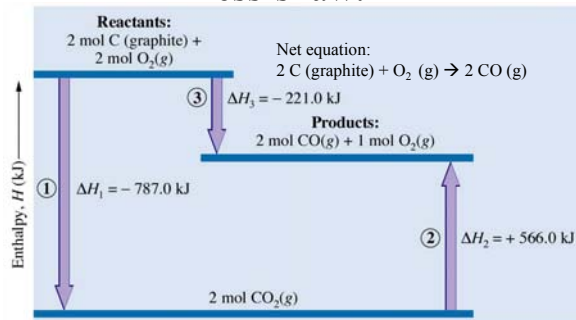
Given:



So:

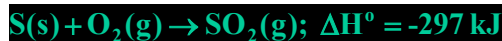


## Enthalpy diagram illustrating Hess's law.



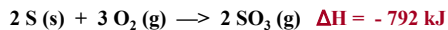
## Hess's Law

- For example, suppose you are given the following data:



- Could you use these data to obtain the enthalpy change for the following reaction?





We have multiplied the equation for the formation of sulfur dioxide by 2 and we have multiplied the heat of formation for sulfur dioxide by 2.

We have reversed the reaction for the formation of sulfur trioxide from sulfur dioxide and oxygen (and changed the sign for the corresponding heat of reaction)

We have added the reactions and the heats of reaction to obtain the values for the heat of formation of sulfur trioxide.

## Standard Enthalpies of Formation

- The standard enthalpy of formation of a substance, denoted  $\Delta H_f^\circ$ , is the enthalpy change for the formation of one mole of a substance in its standard state from its component elements in their standard state.
- Note **that** the standard enthalpy of formation for a pure element in its standard state is zero.

## Standard States

### Compound

- For a **gas**, pressure is exactly **1 atmosphere**.
- For a **solution**, concentration is exactly **1 molar**.
- Pure substance (liquid or solid), it is the pure liquid or solid.

### Element

- The form  $[\text{N}_2(\text{g}), \text{K}(\text{s})]$  in which it exists at **1 atm and 25°C**.

Table 6.2  
Standard Enthalpies of Formation (at 25°C)\*

Formula	$\Delta H_f^\circ$ (kJ/mol)	Formula	$\Delta H_f^\circ$ (kJ/mol)
e <sup>-</sup> (g)	0	C <sub>2</sub> H <sub>4</sub> (g)	52.5
<b>Hydrogen</b>		C <sub>2</sub> H <sub>6</sub> (g)	-84.7
H <sup>+</sup> (aq)	0	C <sub>2</sub> H <sub>2</sub> (g)	-89.0
H <sub>2</sub> (g)	218.0	HCHO(g)	-116
H <sub>2</sub> (g)	0	CH <sub>3</sub> OH(l)	-238.6
<b>Sodium</b>		CS <sub>2</sub> (l)	117
Na <sup>+</sup> (g)	609.8	CS <sub>2</sub> (l)	87.9
Na <sup>+</sup> (aq)	-239.7	HCN(g)	135
Na(g)	107.8	HCN(l)	105
Na(s)	0	CCl <sub>4</sub> (g)	-96.0
NaCl(s)	-411.1	CCl <sub>4</sub> (l)	-139
NaHCO <sub>3</sub> (s)	-947.7	CH <sub>3</sub> CHO(g)	-166
Na <sub>2</sub> CO <sub>3</sub> (s)	-1130.8	C <sub>2</sub> H <sub>5</sub> OH(l)	-277.6
<b>Calcium</b>		<b>Silicon</b>	
Ca <sup>2+</sup> (aq)	-543.0	Si(s)	0
Ca(s)	0	SiO <sub>2</sub> (s)	-910.9
CaO(s)	-635.1	SiF <sub>4</sub> (g)	-1548
CaCO <sub>3</sub> (s) (calcite)	-1206.9	<b>Lead</b>	
<b>Carbon</b>		Pb(s)	0
C(g)	715.0	PbO(s)	-219
C(graphite)	0	PbS(s)	-98.3
C(diamond)	1.9	<b>Nitrogen</b>	
CO(g)	-110.5	N <sub>2</sub> (g)	473
CO <sub>2</sub> (g)	-393.5	N <sub>2</sub> (l)	0
HCO <sub>3</sub> <sup>-</sup> (aq)	-691.1	NH <sub>3</sub> (g)	-45.9
CH <sub>4</sub> (g)	-74.9		

(continued)



What is  $\Delta H_f^\circ$  for  
HCl (aq)?

Remember that  
HCl is a strong acid  
and  
 $\text{HCl (aq)} \rightarrow \text{H}^+(\text{aq}) + \text{Cl}^-(\text{aq})$

So to use the  $\Delta H_f^\circ$   
for  $\text{Cl}^-(\text{aq})$  and  
 $\text{H}^+(\text{aq})$  (-167.5 kJ +  
0 kJ = -167.5 kJ)

Table 6.2 (Continued)			
Formula	$\Delta H_f^\circ$ (kJ/mol)	Formula	$\Delta H_f^\circ$ (kJ/mol)
$\text{NH}_4^+(\text{aq})$	-132.8	$\text{HF(g)}$	-273
$\text{NO(g)}$	90.3	<b>Chlorine</b>	
$\text{NO}_2(\text{g})$	33.2	$\text{Cl}^-(\text{aq})$	-167.5
$\text{HNO}_3(\text{aq})$	-206.6	$\text{Cl}_2(\text{g})$	121.0
<b>Oxygen</b>		$\text{Cl}_2(\text{l})$	0
$\text{O(g)}$	249.2	$\text{HCl(g)}$	-92.3
$\text{O}_2(\text{g})$	0	<b>Bromine</b>	
$\text{O}_3(\text{g})$	143	$\text{Br}^-(\text{g})$	-218.9
$\text{OH}^-(\text{aq})$	-229.9	$\text{Br}^-(\text{aq})$	-120.9
$\text{H}_2\text{O(g)}$	-241.8	$\text{Br}_2(\text{l})$	0
$\text{H}_2\text{O(l)}$	-285.8	<b>Iodine</b>	
<b>Sulfur</b>		$\text{I}^-(\text{g})$	-194.7
$\text{S(g)}$	279	$\text{I}^-(\text{aq})$	-55.9
$\text{S}_2(\text{g})$	129	$\text{I}_2(\text{s})$	0
$\text{S}_8(\text{rhombic})$	0	<b>Silver</b>	
$\text{S}_8(\text{monoclinic})$	2	$\text{Ag}^+(\text{g})$	1026.4
$\text{SO}_2(\text{g})$	-296.8	$\text{Ag}^+(\text{aq})$	105.9
$\text{H}_2\text{S(g)}$	-20	$\text{Ag(s)}$	0
<b>Fluorine</b>		$\text{AgF(s)}$	-203
$\text{F}^-(\text{g})$	-255.6	$\text{AgCl(s)}$	-127.0
$\text{F}^-(\text{aq})$	-329.1	$\text{AgBr(s)}$	-99.5
$\text{F}_2(\text{g})$	0	$\text{AgI(s)}$	-62.4

\*See Appendix C for additional values.

Allotropes of sulfur.

Photo courtesy of James Scherer.

An **allotrope** is one of two or more distinct forms of an element in the same physical state. Shown here two allotropes of  $\text{S}_8$ , rhombic (left) and monoclinic (right). The rhombic form is more stable at room temperature and is the **reference form**.

## Standard Enthalpies of Formation

- The **law of summation of heats of formation** states that the enthalpy of a reaction is equal to the total formation energy of the products minus that of the reactants.

$$\Delta H^\circ = \sum n\Delta H_f^\circ(\text{products}) - \sum m\Delta H_f^\circ(\text{reactants})$$

- $\Sigma$  is the mathematical symbol meaning “the sum of”, and  $m$  and  $n$  are the coefficients of the substances in the chemical equation.

**TABLE 6.2**  
Standard Enthalpies of Formation for Several Compounds at 25°C

Compound	$\Delta H_f^\circ$ (kJ/mol)
$\text{NH}_3(\text{g})$	-46
$\text{NO}_2(\text{g})$	34
$\text{H}_2\text{O(l)}$	-286
$\text{Al}_2\text{O}_3(\text{s})$	-1676
$\text{Fe}_2\text{O}_3(\text{s})$	-826
$\text{CO}_2(\text{g})$	-394
$\text{CH}_3\text{OH(l)}$	-239
$\text{C}_8\text{H}_{18}(\text{l})$	-269

## A Problem to Consider

- Large quantities of ammonia are used to prepare nitric acid according to the following equation:



- What is the standard enthalpy change for this reaction? Use Table 6.2 for data.

## A Problem to Consider

- You record the values of  $\Delta H_f^\circ$  under the formulas in the equation, multiplying them by the coefficients in the equation.



- You can calculate  $\Delta H^\circ$  by subtracting the values for the reactants from the values for the products.

## A Problem to Consider

- Using the summation law:

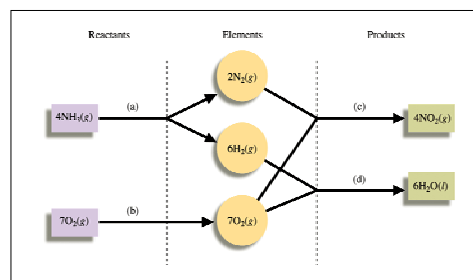
$$\Delta H^\circ = \sum n\Delta H_f^\circ(\text{products}) - \sum m\Delta H_f^\circ(\text{reactants})$$

$$\Delta H^\circ = [4(90.3) + 6(-241.8)] \text{ kJ} - [4(-45.9) + 5(0)] \text{ kJ}$$

$$\Delta H^\circ = -906 \text{ kJ}$$

- Be careful of arithmetic signs as they are a likely source of mistakes.

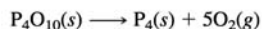
Figure 6.10: A pathway for the combustion of ammonia.



**6.45** When white phosphorus burns in air, it produces phosphorus(V) oxide.

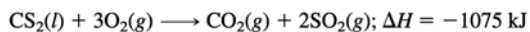


What is  $\Delta H$  for the following equation?

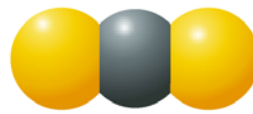
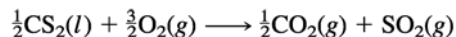


$\text{P}_4$

**6.46** Carbon disulfide burns in air, producing carbon dioxide and sulfur dioxide.



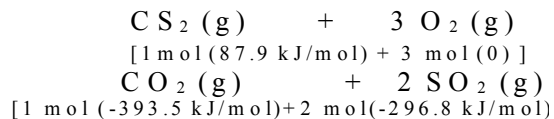
What is  $\Delta H$  for the following equation?



$\text{CS}_2$

$\text{CS}_2$  is used in the manufacture of plastics and cellophane.

Calculate  $\Delta H^\circ$  for the reaction below using the  $\Delta H_f^\circ$  for the reactants and products shown below:



$$\begin{aligned} \Delta H^\circ &= [\sum m \Delta H_f^\circ(\text{products})] - [n \Delta H_f^\circ(\text{reactants})] \\ &= [(-393.5 \text{ kJ}) + (-593.6 \text{ kJ})] - [87.9 \text{ kJ}] \\ &= -1075.0 \text{ kJ} \end{aligned}$$

Keep in mind the following concepts.

**Keep in Mind the Following Key Concepts When Doing Enthalpy Calculations:**

- When a reaction is reversed, the magnitude of  $\Delta H$  remains the same, but its sign changes.
- When the balanced equation for a reaction is multiplied by an integer, the value of  $\Delta H$  for that reaction must be multiplied by the same integer.
- The change in enthalpy for a given reaction can be calculated from the enthalpies of formation of the reactants and products:
 
$$\Delta H^\circ_{\text{reaction}} = \sum n_p \Delta H_f^\circ(\text{products}) - \sum n_r \Delta H_f^\circ(\text{reactants})$$
- Elements in their standard states are not included in the  $\Delta H_{\text{reaction}}$  calculations. That is,  $\Delta H_f^\circ$  for an element in its standard state is zero.