

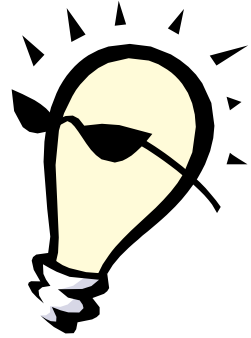
# THERMOCHEMISTRY

## Chapter

## 7

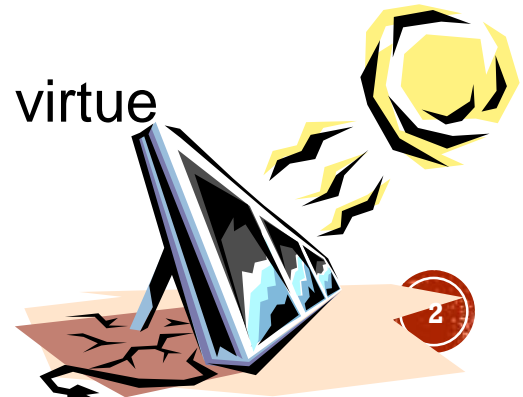
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# The Nature & Types of Energy



**Energy** is the capacity to do work

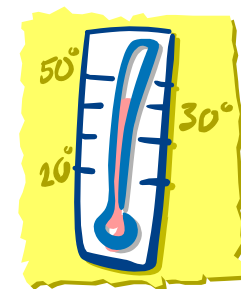
- **Thermal energy** is the energy associated with the random motion of atoms and molecules
- **Chemical energy** is the energy stored within the bonds of chemical substances
- **Nuclear energy** is the energy stored within the collection of neutrons and protons in the atom
- **Electrical energy** is the energy associated with the flow of electrons
- **Potential energy** is the energy available by virtue of an object's position



# Energy Changes in Chemical Reactions

**Heat** is the transfer of **thermal energy** between two bodies that are at different temperatures.

**Temperature** is a measure of the **thermal energy**.



Temperature  $\neq$  Thermal Energy



90°C



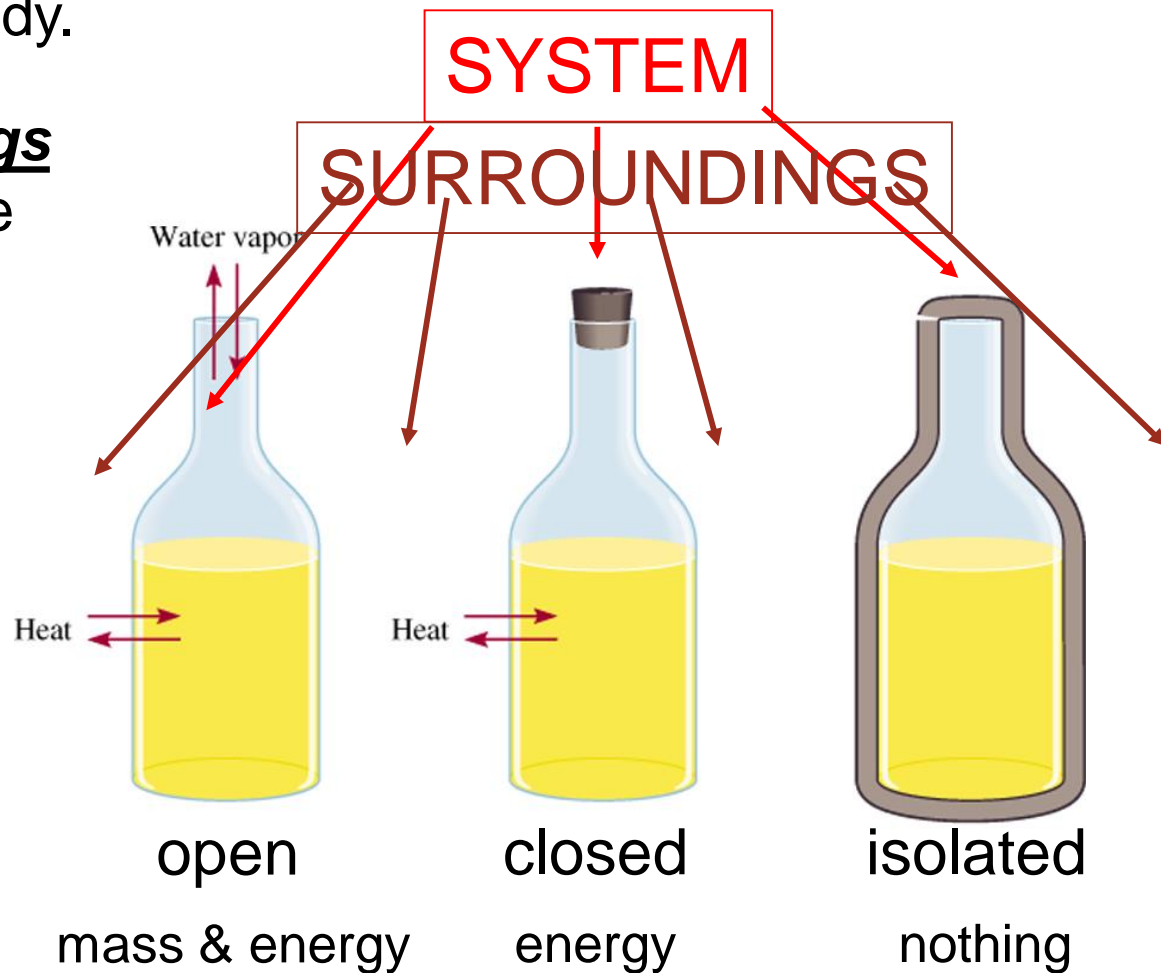
40°C

greater thermal energy

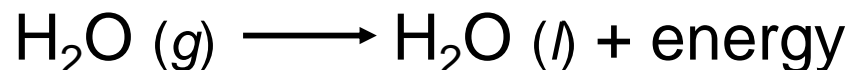
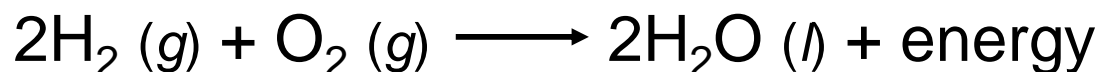
**Thermochemistry** is the study of heat change in chemical reactions.

The **system** is the specific part of the universe that is of interest in the study.

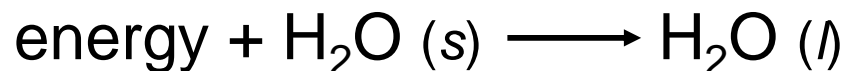
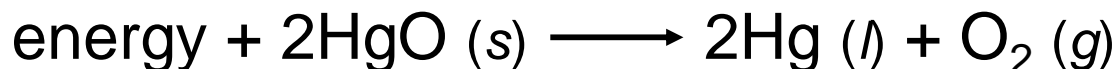
The **surroundings** are the rest of the universe outside the system.



**Exothermic process** is any process that gives off heat – transfers thermal energy from the system to the surroundings.



**Endothermic process** is any process in which heat has to be supplied to the system from the surroundings.

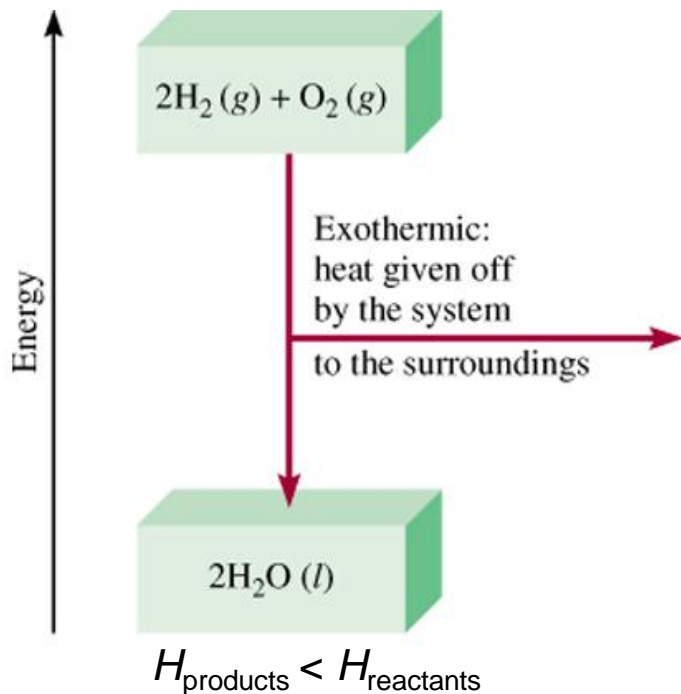


# Enthalpy

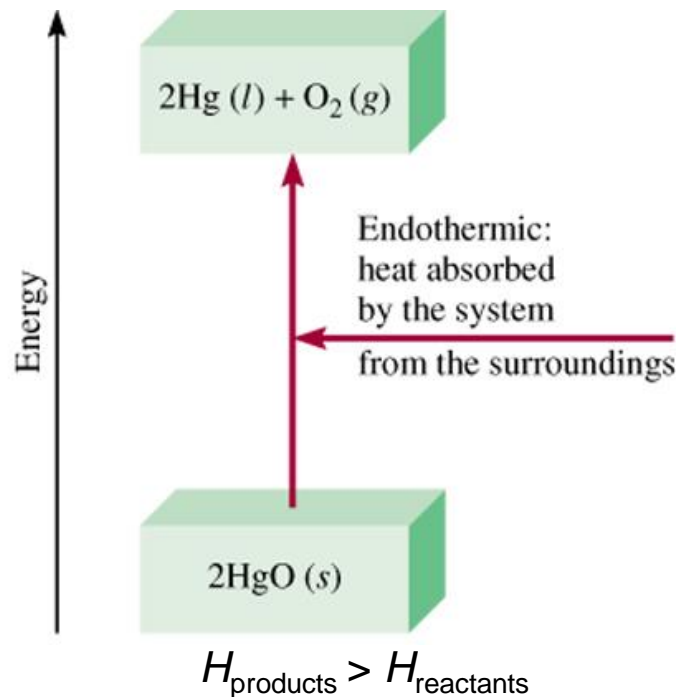
**Enthalpy ( $H$ )** is used to quantify the heat flow into or out of a system in a process that occurs at constant pressure.

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

$\Delta H$  = heat given off or absorbed during a reaction **at constant pressure**

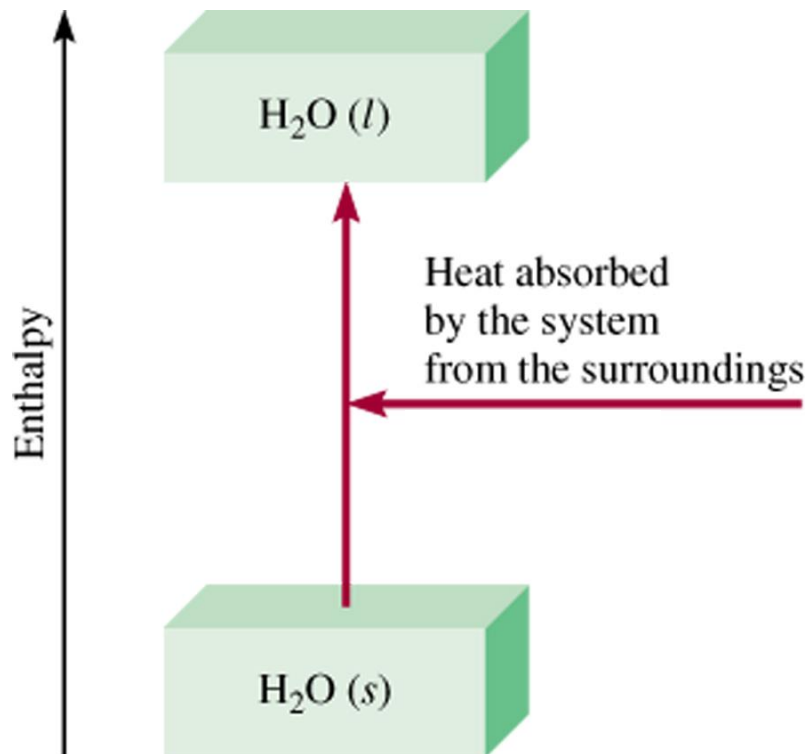


$$\Delta H < 0$$



$$\Delta H > 0$$

# Thermochemical Equations



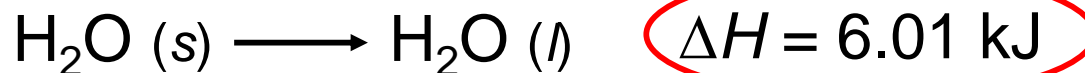
Is  $\Delta H$  negative or positive?

System absorbs heat

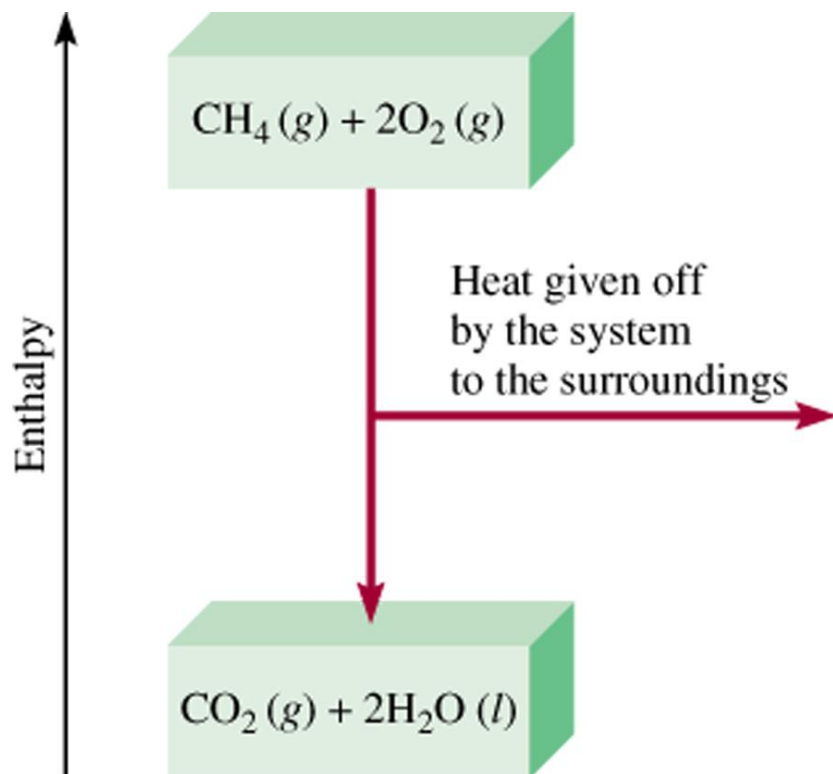
Endothermic

$$\Delta H > 0$$

6.01 kJ are absorbed for every 1 mole of ice that melts at  $0^\circ\text{C}$  and 1 atm.



# Thermochemical Equations



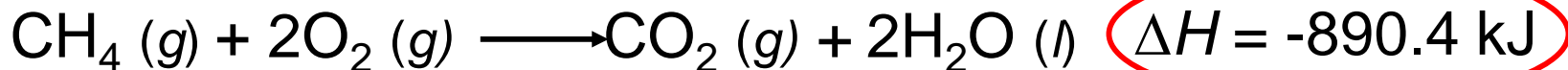
Is  $\Delta H$  negative or positive?

System gives off heat

Exothermic

$$\Delta H < 0$$

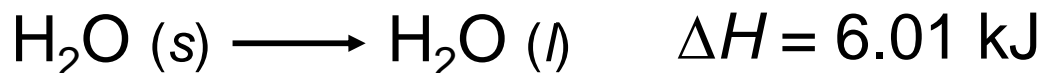
890.4 kJ are released for every 1 mole of methane that is combusted at 25°C and 1 atm.





# Thermochemical Equations

- The stoichiometric coefficients always refer to the number of moles of a substance



- If you reverse a reaction, the sign of  $\Delta H$  changes

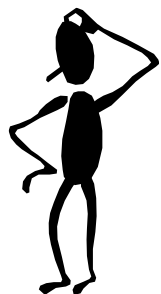
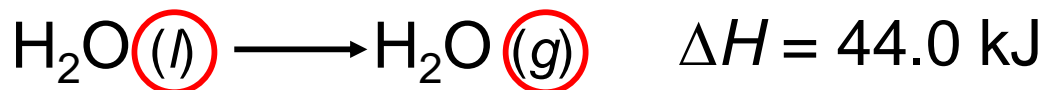


- If you multiply both sides of the equation by a factor  $n$ , then  $\Delta H$  must change by the same factor  $n$ .



# Thermochemical Equations

- The physical states of all reactants and products must be specified in thermochemical equations.



How much heat is evolved when 266 g of white phosphorus ( $\text{P}_4$ ) burn in air?

# Calorimetry

The **specific heat** ( $s$ ) of a substance is the amount of heat ( $q$ ) required to raise the temperature of **one gram** of the substance by **one degree** Celsius.

The **heat capacity** ( $C$ ) of a substance is the amount of heat ( $q$ ) required to raise the temperature of **a given quantity** ( $m$ ) of the substance by **one degree** Celsius.

$$C = ms$$

Heat ( $q$ ) absorbed or released:

$$q = ms\Delta t$$

$$q = C\Delta t$$

$$\Delta t = t_{\text{final}} - t_{\text{initial}}$$

**Table 6.1** The Specific Heats of Some Common Substances

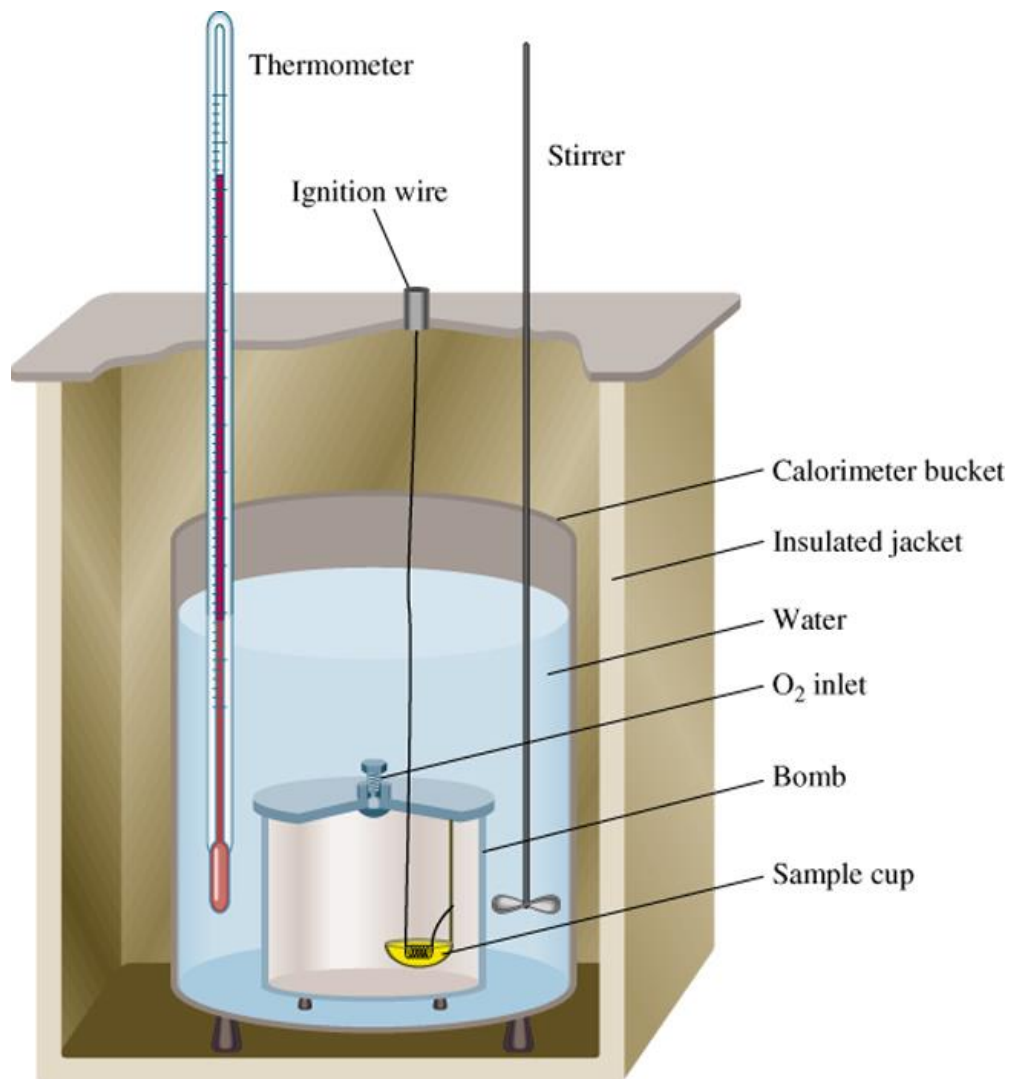
Substance	Specific heat (J/g · °C)
Al	0.900
Au	0.129
C (graphite)	0.720
C (diamond)	0.502
Cu	0.385
Fe	0.444
Hg	0.139
H <sub>2</sub> O	4.184
C <sub>2</sub> H <sub>5</sub> OH (ethanol)	2.46



How much heat is given off when an 869 g iron bar cools from  $94^{\circ}\text{C}$  to  $5^{\circ}\text{C}$ ?



# Constant-Volume Calorimetry



$$q_{\text{sys}} = q_{\text{water}} + q_{\text{bomb}} + q_{\text{rxn}}$$

$$q_{\text{sys}} = 0$$

$$q_{\text{rxn}} = - (q_{\text{water}} + q_{\text{bomb}})$$

$$q_{\text{water}} = ms\Delta t$$

$$q_{\text{bomb}} = C_{\text{bomb}}\Delta t$$

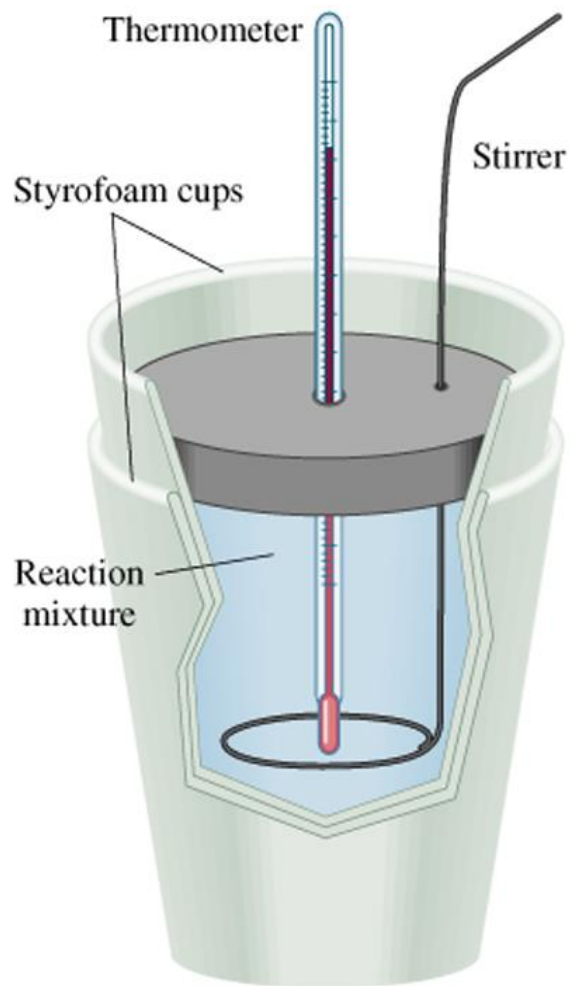
Reaction at Constant  $V$

$$\Delta H \neq q_{\text{rxn}}$$

$$\Delta H \sim q_{\text{rxn}}$$

No heat enters or leaves!

# Constant-Pressure Calorimetry



$$q_{\text{sys}} = q_{\text{water}} + q_{\text{cal}} + q_{\text{rxn}}$$

$$q_{\text{sys}} = 0$$

$$q_{\text{rxn}} = - (q_{\text{water}} + q_{\text{cal}})$$

$$q_{\text{water}} = ms\Delta t$$

$$q_{\text{cal}} = C_{\text{cal}}\Delta t$$

Reaction at Constant  $P$

$$\Delta H = q_{\text{rxn}}$$

No heat enters or leaves!

# Standard Enthalpy of Formation & Reaction

Because there is no way to measure the absolute value of the enthalpy of a substance, must I measure the enthalpy change for every reaction of interest?

Establish an arbitrary scale with the **standard enthalpy of formation** ( $\Delta H_f^0$ ) as a reference point for all enthalpy expressions.

**Standard enthalpy of formation** ( $\Delta H_f^0$ ) is the heat change that results when **one mole** of a compound is formed from its **elements** at a pressure of 1 atm.

The standard enthalpy of formation of any element in its most stable form is zero.

$$\Delta H_f^0 (\text{O}_2) = 0$$

$$\Delta H_f^0 (\text{C, graphite}) = 0$$

$$\Delta H_f^0 (\text{O}_3) = 142 \text{ kJ/mol}$$

$$\Delta H_f^0 (\text{C, diamond}) = 1.90 \text{ kJ/mol}$$

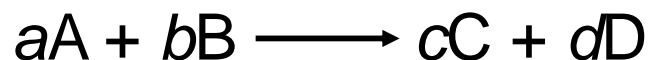
**Table 6.3** Standard Enthalpies of Formation of Some Inorganic Substances at 25°C

Substance	$\Delta H_f^\circ$ (kJ/mol)	Substance	$\Delta H_f^\circ$ (kJ/mol)
Ag(s)	0	H <sub>2</sub> O <sub>2</sub> (l)	-187.6
AgCl(s)	-127.04	Hg(l)	0
Al(s)	0	I <sub>2</sub> (s)	0
Al <sub>2</sub> O <sub>3</sub> (s)	-1669.8	HI(g)	25.94
Br <sub>2</sub> (l)	0	Mg(s)	0
HBr(g)	-36.2	MgO(s)	-601.8
C(graphite)	0	MgCO <sub>3</sub> (s)	-1112.9
C(diamond)	1.90	N <sub>2</sub> (g)	0
CO(g)	-110.5	NH <sub>3</sub> (g)	-46.3
CO <sub>2</sub> (g)	-393.5	NO(g)	90.4
Ca(s)	0	NO <sub>2</sub> (g)	33.85
CaO(s)	-635.6	N <sub>2</sub> O <sub>4</sub> (g)	9.66
CaCO <sub>3</sub> (s)	-1206.9	N <sub>2</sub> O(g)	81.56
Cl <sub>2</sub> (g)	0	O(g)	249.4
HCl(g)	-92.3	O <sub>2</sub> (g)	0
Cu(s)	0	O <sub>3</sub> (g)	142.2
CuO(s)	-155.2	S(rhombic)	0
F <sub>2</sub> (g)	0	S(monoclinic)	0.30
HF(g)	-268.61	SO <sub>2</sub> (g)	-296.1
H(g)	218.2	SO <sub>3</sub> (g)	-395.2
H <sub>2</sub> (g)	0	H <sub>2</sub> S(g)	-20.15
H <sub>2</sub> O(g)	-241.8	ZnO(s)	-347.98
H <sub>2</sub> O(l)	-285.8		

[Back to Q & A](#)

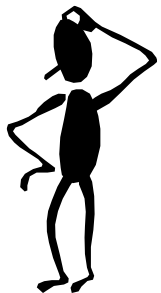


The **standard enthalpy of reaction** ( $\Delta H_{\text{rxn}}^0$ ) is the enthalpy of a reaction carried out at 1 atm.



$$\Delta H_{\text{rxn}}^0 = [c\Delta H_{\text{f}}^0(\text{C}) + d\Delta H_{\text{f}}^0(\text{D})] - [a\Delta H_{\text{f}}^0(\text{A}) + b\Delta H_{\text{f}}^0(\text{B})]$$

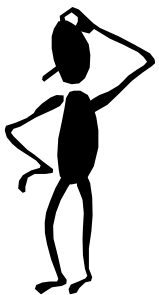
$$\Delta H_{\text{rxn}}^0 = \Sigma n\Delta H_{\text{f}}^0(\text{products}) - \Sigma m\Delta H_{\text{f}}^0(\text{reactants})$$



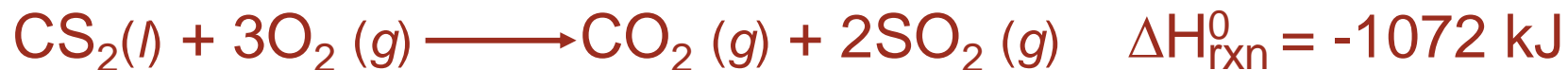
Benzene ( $\text{C}_6\text{H}_6$ ) burns in air to produce carbon dioxide and liquid water. How much heat is released **per mole** of benzene combusted? The standard enthalpy of formation of benzene is 49.04 kJ/mol.

**Hess's Law:** 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.

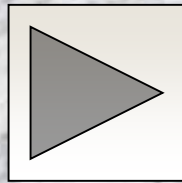
(Enthalpy is a state function. It doesn't matter how you get there, only where you start and end.)



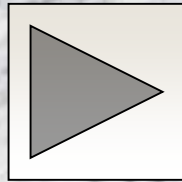
Calculate the standard enthalpy of formation of  $\text{CS}_2(l)$  given that:



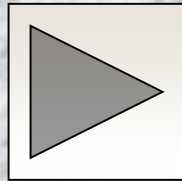
# Q & A session



Question 1

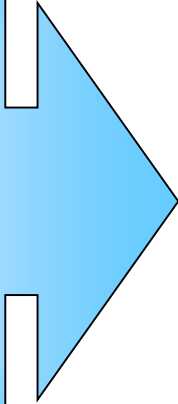


Question 2

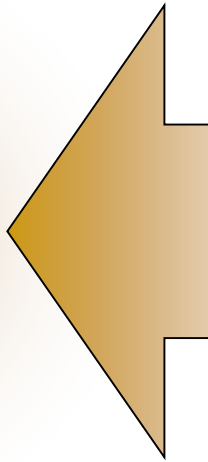


Question 3

**Calculate the quantity of heat required to raise the temperature of 2.10 gram of water from 2.0 °C to 6.0 °C.**



**Calculate the specific heat if a 36.9 gram sample of a metal at 57.2 °C is immersed in 50.1 gram of water at 16.0 °C, warming the water to 20.5 °C.**



**Calculate the enthalpy change for the reaction of one mole of CH<sub>4</sub> with oxygen gas to yield carbon dioxide and water.**

$$\Delta H_f^\circ \text{CH}_4(\text{g}) = -74.5 \text{ kJ}$$

Table 6.3

**Calculate the enthalpy of the following reaction based on the enthalpies of combustion in equation given:**

