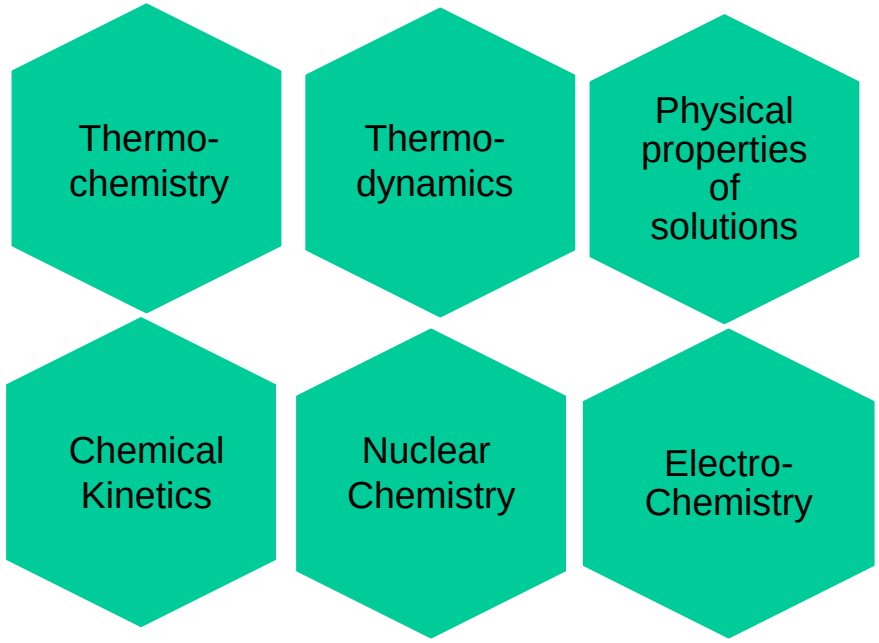
Chemistry 202

Course title: GCII [Physical chemistry I (Fundamentals)]



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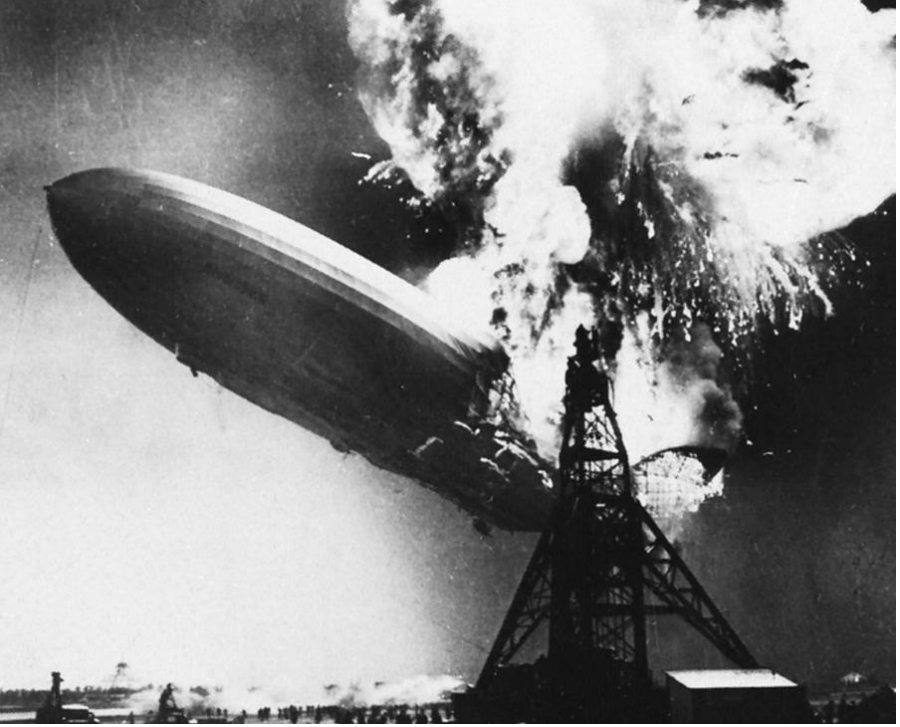
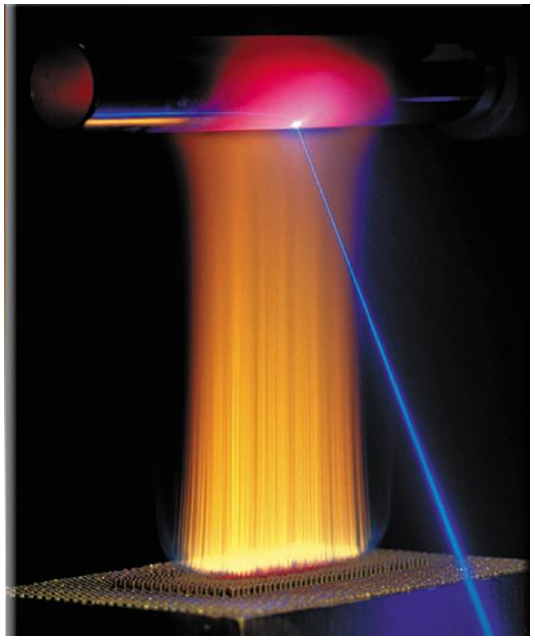


Thermochemistry

Chapter 6



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Thermochemistry

Chapter One Contains:

1.1 The Nature of Energy and Types of Energy

1.2 Energy Changes in Chemical Reactions

1.3 Introduction to Thermodynamics

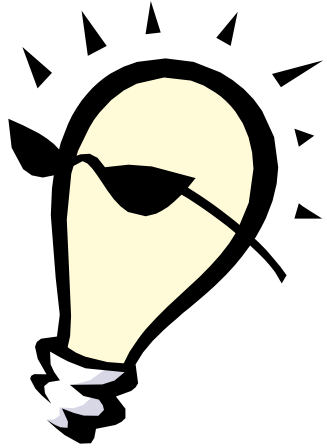
1.4 Enthalpy of Chemical Reactions

1.5 Calorimetry

1.6 Standard Enthalpy of Formation and Reaction

1.7 Heat of Solution and Dilution

Energy is the capacity to do work

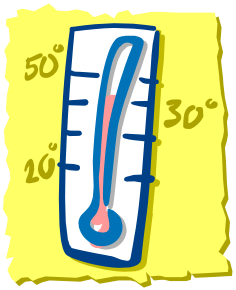


- ***Radiant energy*** comes from the sun and is earth's primary energy source
- ***Thermal energy*** is the energy associated with the random motion of atoms and molecules
- ***Chemical energy*** is the energy stored within the bonds of chemical substances
- ***Nuclear energy*** is the energy stored within the collection of neutrons and protons in the atom
- ***Potential energy*** is the energy available by virtue of an object's position

1.2 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 ≠ Thermal Energy

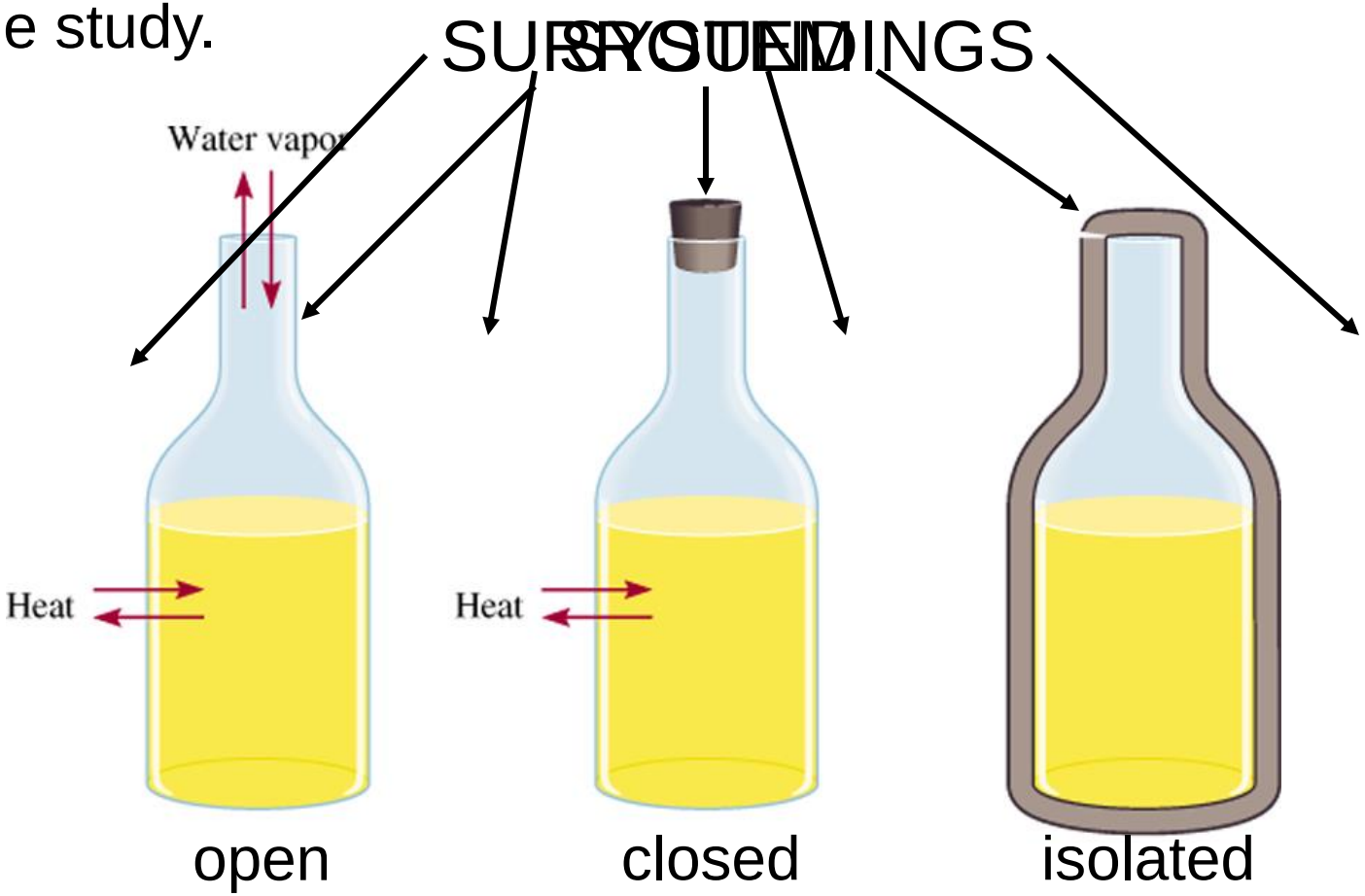
90^0C

40^0C

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.



	open	closed	isolated
Exchange:	mass & energy	energy	nothing

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

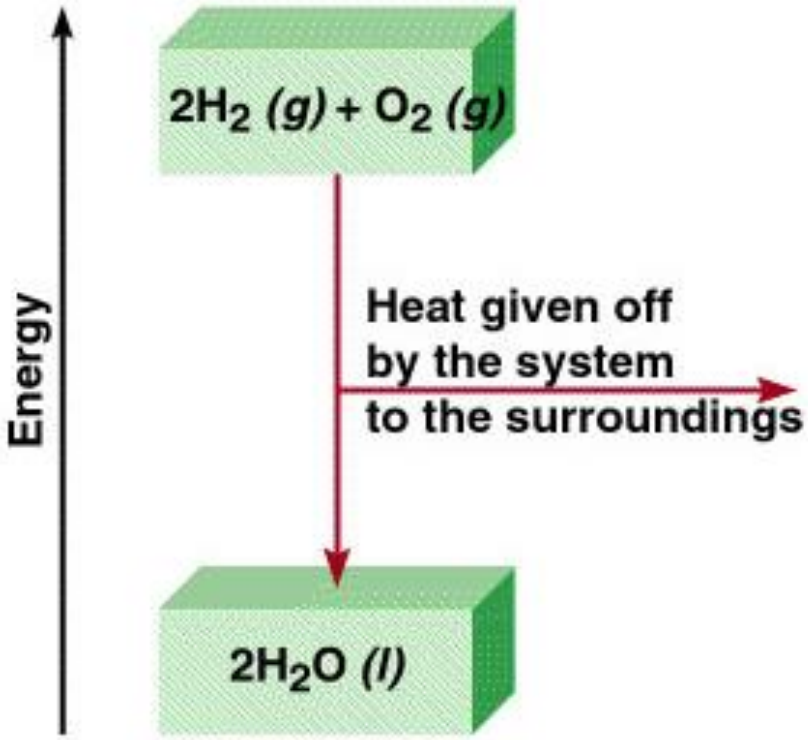
$$2H_2\,(g) + O_2\,(g) \longrightarrow 2H_2O\,(l) + \text{energy}$$

$$H_2O\,(g) \longrightarrow H_2O\,(l) + \text{energy}$$

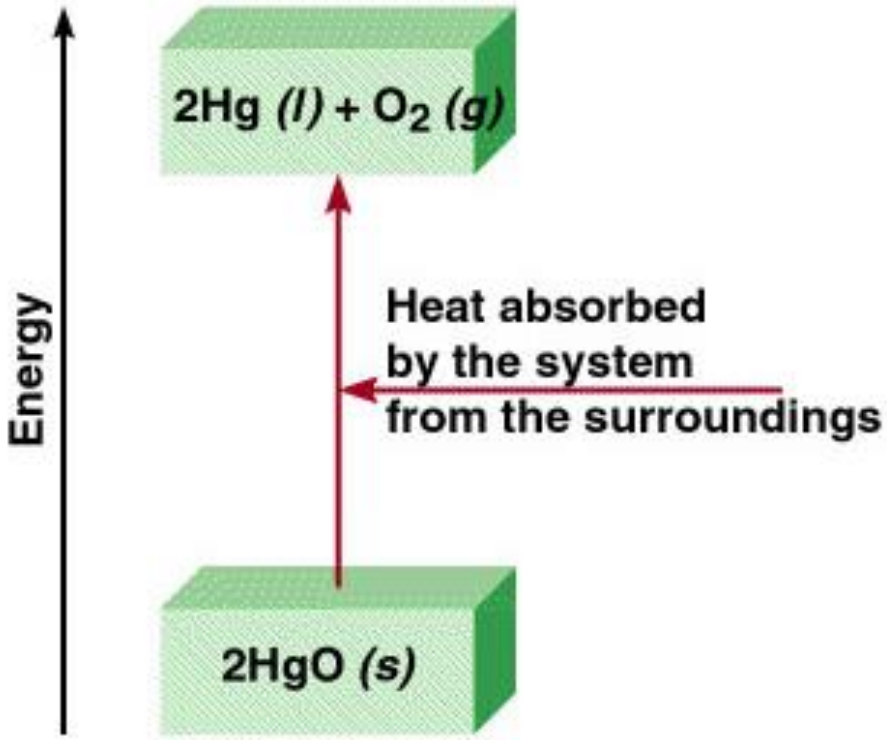
Endothermic process is any process in which heat has to be supplied to the system from the surroundings.

$$\text{energy} + 2HgO\,(s) \longrightarrow 2Hg\,(l) + O_2\,(g)$$

$$\text{energy} + H_2O\,(s) \longrightarrow H_2O\,(l)$$



Exothermic

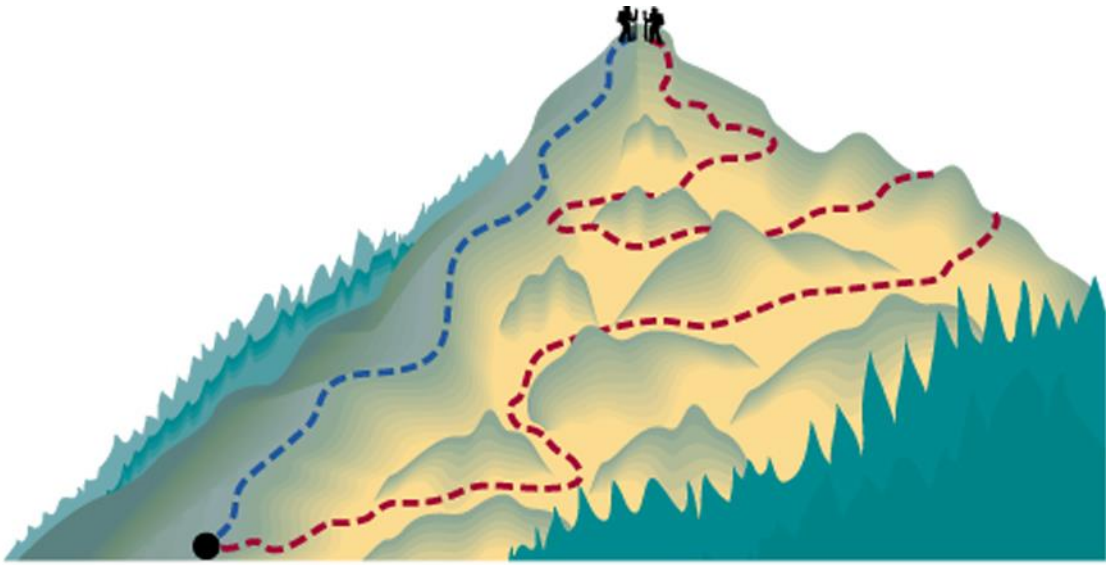


Endothermic

1.3 Introduction to the Thermodynamics

State functions are properties that are determined by the state of the system, regardless of how that condition was achieved.

energy, pressure, volume, temperature



$$\Delta E = E_{final} - E_{initial}$$

$$\Delta P = P_{final} - P_{initial}$$

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

$$\Delta T = T_{final} - T_{initial}$$

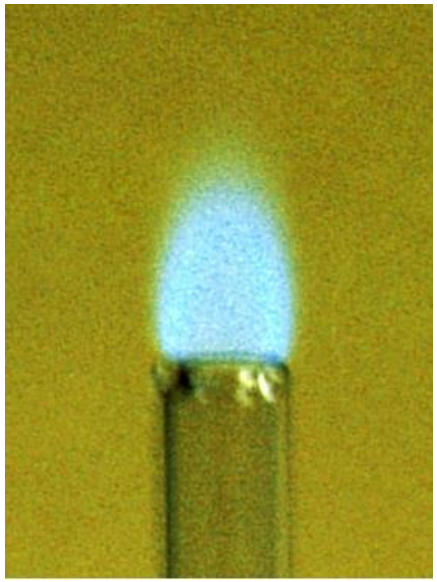
Potential energy of hiker 1 and hiker 2 is the same even though they took different paths.

First law of thermodynamics – energy can be converted from one form to another, but cannot be created or destroyed.

$$\Delta E_{system} + \Delta E_{surroundings} = 0$$

or

$$\Delta E_{system} = -\Delta E_{surroundings}$$



$$C_3H_8 + 5O_2 \longrightarrow 3CO_2 + 4H_2O$$

Exothermic chemical reaction!

Chemical energy **lost** by combustion = Energy **gained** by the surroundings

system surroundings

- What is the internal energy of the system (E)?
- The internal energy of a system has two components: kinetic energy of all motions and potential energy of electrons and nuclei.
- It is impossible to measure all these contributions accurately, so we cannot calculate the total energy of a system with any certainty.
- The change in E only can be measured, for a system when moves from initial state to the final state.

$$\Delta E = E_f - E_i$$

Another form of the ***first law*** for ΔE_{system}

$\Delta E = q + w$

ΔE is the change in internal energy of a system

q is the heat exchange between the system and the surroundings

w is the work done on (or by) the system

$w = -P\Delta V$ when a gas expands against a constant external pressure

TABLE 6.1 Sign Conventions for Work and Heat

Process	Sign
Work done by the system on the surroundings	−
Work done on the system by the surroundings	+
Heat absorbed by the system from the surroundings (endothermic process)	+
Heat absorbed by the surroundings from the system (exothermic process)	−

Work Done On the System

$w = Fd$

$w = -P\,\Delta V$

$$P \times V = \frac{F}{d^2} \times d^3 = Fd = w$$

$\Delta V > 0$

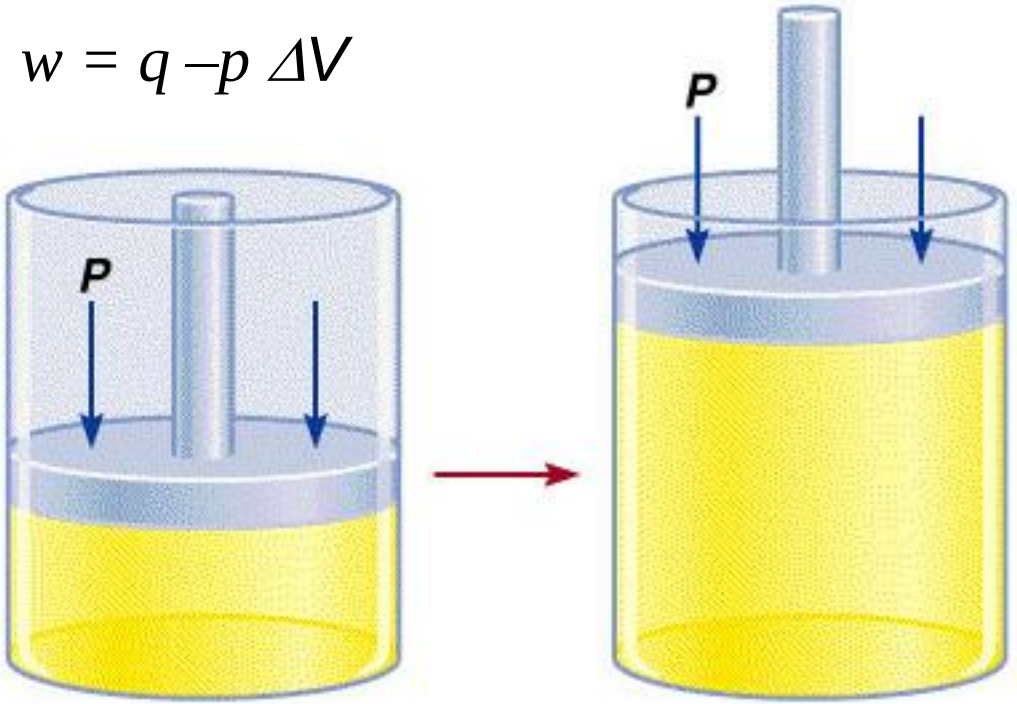
$-P\Delta V < 0$

$w_{sys} < 0$

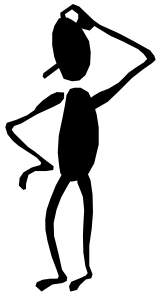
$$\Delta E = q + w = q - p\,\Delta V$$

Work is not a state function!

$\Delta w \neq w_{final} - w_{initial}$



initial final



Example 1.1 A sample of nitrogen gas expands in volume from 1.6 L to 5.4 L at constant temperature. What is the work done in joules if the gas expands (a) against a vacuum and (b) against a constant pressure of 3.7 atm?

$$w = -P\,\Delta V$$

(a) $\Delta V = 5.4\text{ L} - 1.6\text{ L} = 3.8\text{ L}$ $P = 0$ atm

$$W = -0\text{ atm} \times 3.8\text{ L} = 0\text{ L}\bullet\text{atm} = 0\text{ joules}$$

(b) $\Delta V = 5.4\text{ L} - 1.6\text{ L} = 3.8\text{ L}$ $P = 3.7$ atm

$$w = -3.7\text{ atm} \times 3.8\text{ L} = -14.1\text{ L}\bullet\text{atm}$$

$$w = -14.1\text{ L}\bullet\text{atm} \times \frac{101.3\text{ J}}{1\text{L}\bullet\text{atm}} = -1430\text{ J}$$

Enthalpy and the First Law of Thermodynamics

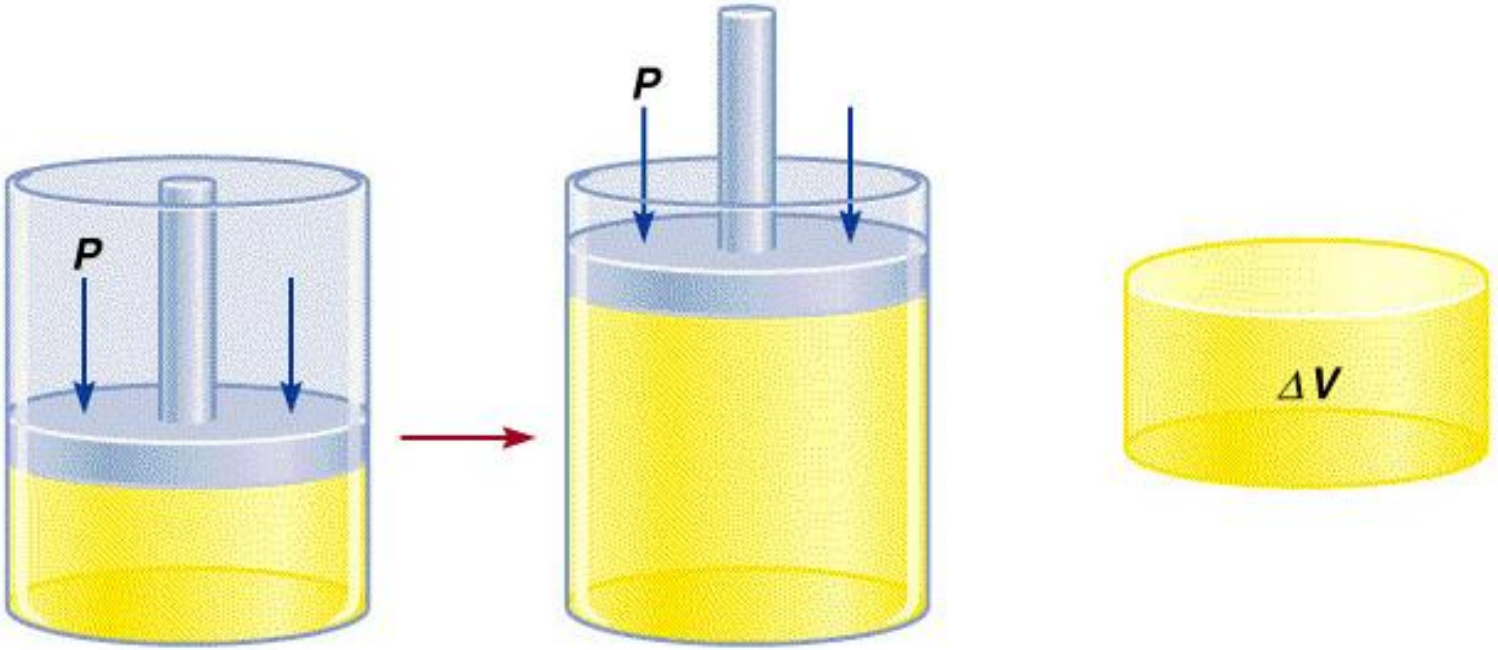
$\Delta E = q + w$ At Constant Temperature

At constant pressure:

$q = \Delta H$ and $w = -P\Delta V$

$\Delta E = \Delta H - P\Delta V$ *At Constant Pressure*

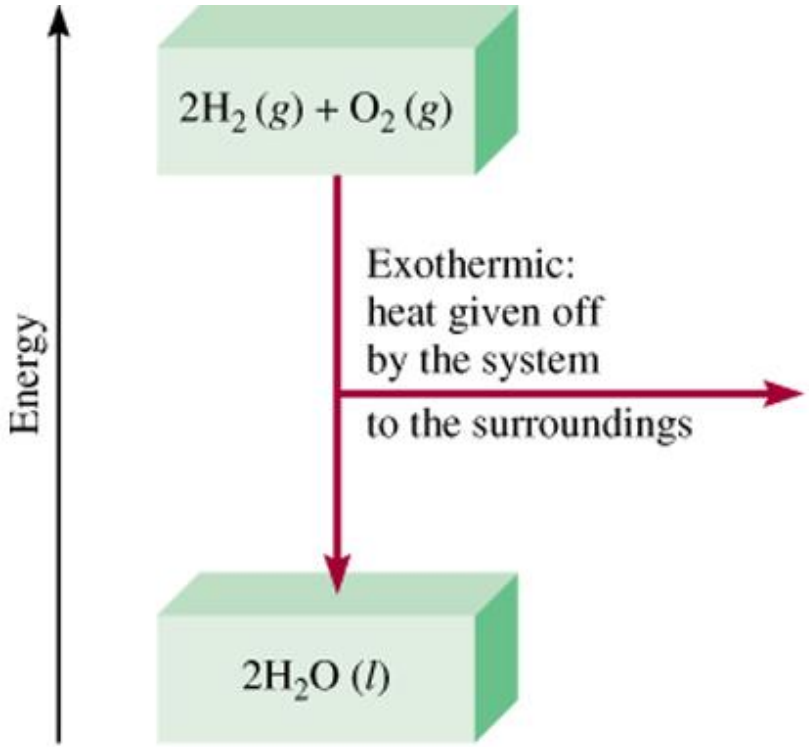
$\Delta H = \Delta E + P\Delta V$



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

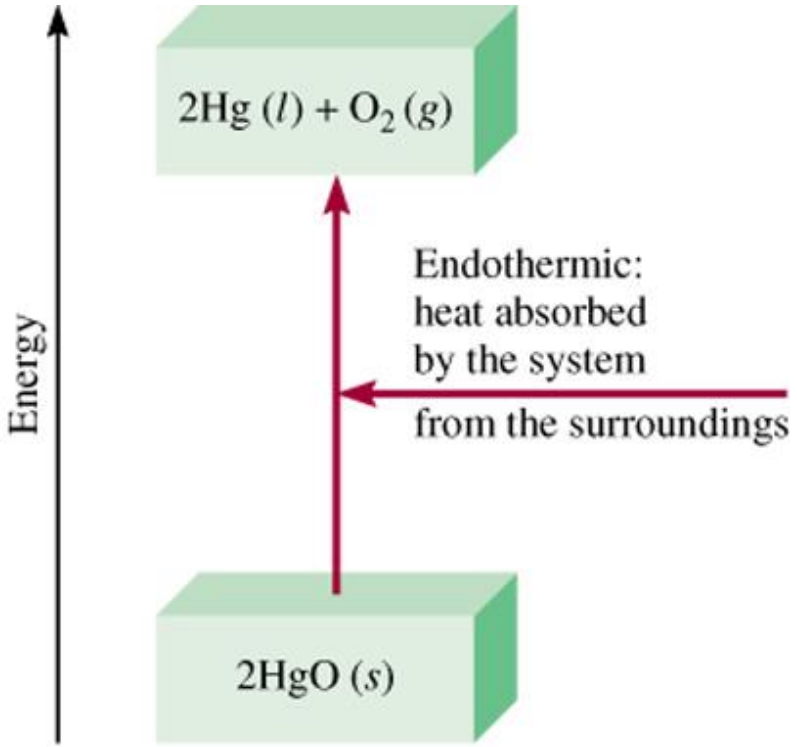
$$\Delta H = H \text{ (products)} - H \text{ (reactants)}$$

ΔH = heat given off or absorbed during a reaction **at constant pressure**



$H_{products} < H_{reactants}$

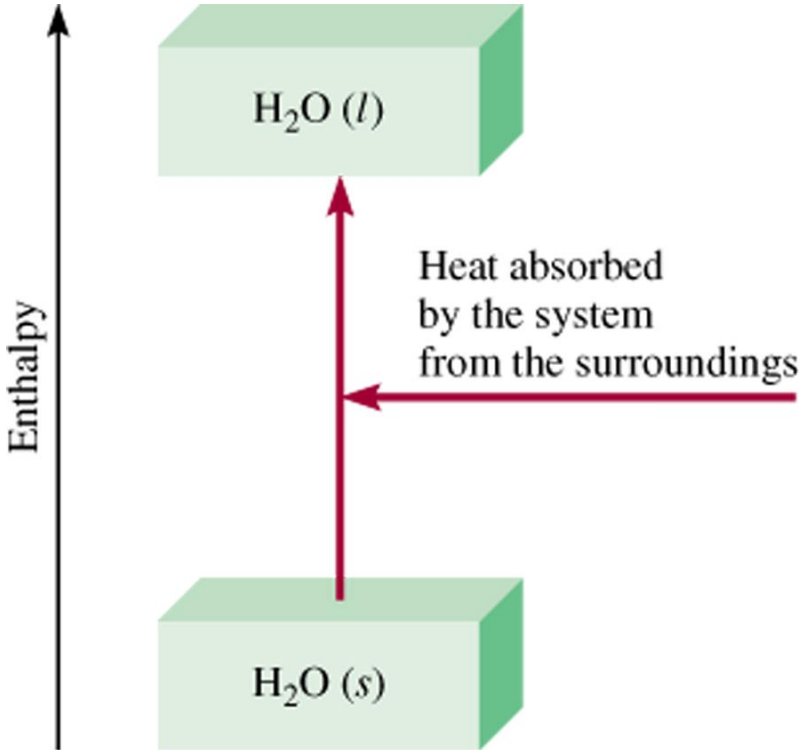
$\Delta H < 0$



$H_{products} > H_{reactants}$

$\Delta H > 0$

Thermochemical Equations



Is ΔH negative or positive?

System absorbs heat

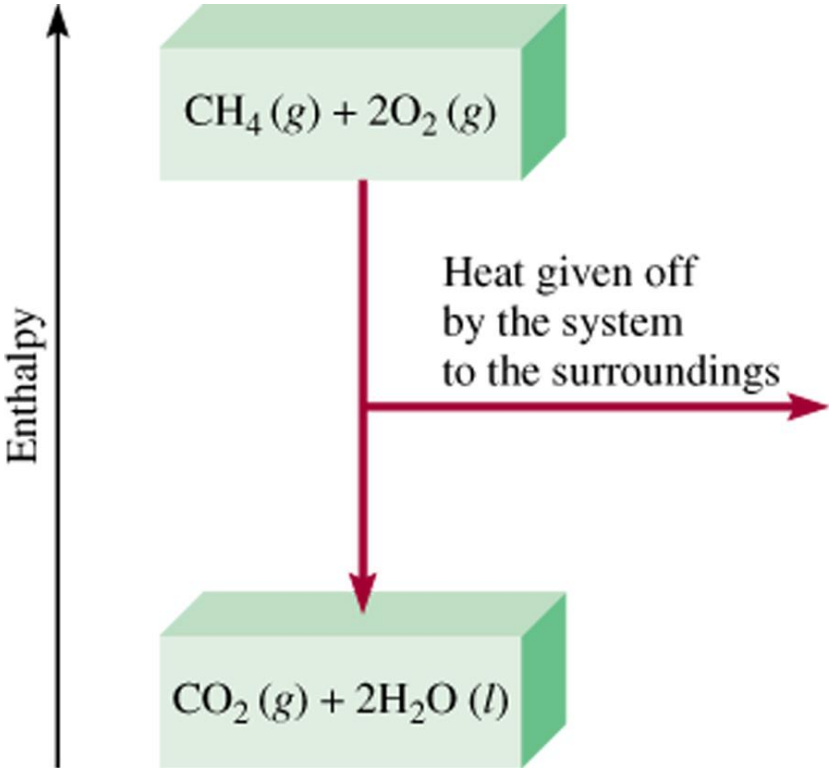
Endothermic

$\Delta H > 0$

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

$$H_2O\ (s) \longrightarrow H_2O\ (l) \qquad \Delta H = 6.01\ \text{kJ}$$

Thermochemical Equations



Is ΔH 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 25^0C and 1 atm.

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

Thermochemical Equations

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

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

- If you reverse a reaction, the sign of ΔH changes

$$H_2O\ (l) \longrightarrow H_2O\ (s) \qquad \Delta H = -6.01\ kJ$$

- If you multiply both sides of the equation by a factor *n*, then ΔH must change by the same factor *n*.

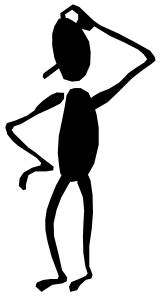
$$2H_2O\ (s) \longrightarrow 2H_2O\ (l) \qquad \Delta H = 2 \times 6.01 = 12.0\ kJ$$

Thermochemical Equations

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

$$H_2O(s) \longrightarrow H_2O(l) \qquad \Delta H = 6.01 \text{ kJ}$$

$$H_2O(l) \longrightarrow H_2O(g) \qquad \Delta H = 44.0 \text{ kJ}$$



Example 1.2 How much heat is evolved when 266 g of white phosphorus (P_4) burn in air?

$$P_4\,(s) + 5O_2\,(g) \longrightarrow P_4O_{10}\,(s) \qquad \Delta H = -3013 \text{ kJ}$$

n = m (g) /molar mass (g/mol) = 266 g /(4x31) (g/mol)

1 mol $\longrightarrow$ 3013

2.15 mol $\longrightarrow$ x

2.15 mol x 3013 kJ / 1 mol = 6470 kJ

For reaction involves ideal gases

$PV = nRT$

$P\Delta V = \Delta nRT$

where Δn is the number of moles of gases in products minus the number of moles of gases in reactants .

$$\Delta H = \Delta E + \Delta nRT$$

$R = 8.314\ JK^{-1}\ mol^{-1}$

PLEASE REMEMBER:

ΔH and ΔE given in kJ

The term ΔnRT in J

Example1.3: The heat of combustion at constant volume of CH_4 (g) is measured in bomb calorimeter at 25°C and is found to be –88.5 KJ / mol, what is ΔH?

$$CH_4\,(g) + 2O_2\,(g) \rightarrow CO_2\,(g) + 2HO_2\,(l),\ \Delta E = -88.5\ KJ / mol$$

$$\Delta n = 1-(2+1) = -2$$

$$\Delta H = \Delta E + \Delta n\, RT$$

$$\Delta H = -88.5 + \frac{(-2)(8.314X298)}{1000}$$

$$= -93.5\ KJ$$

Example 1.4 Find ΔE *for the reaction:*

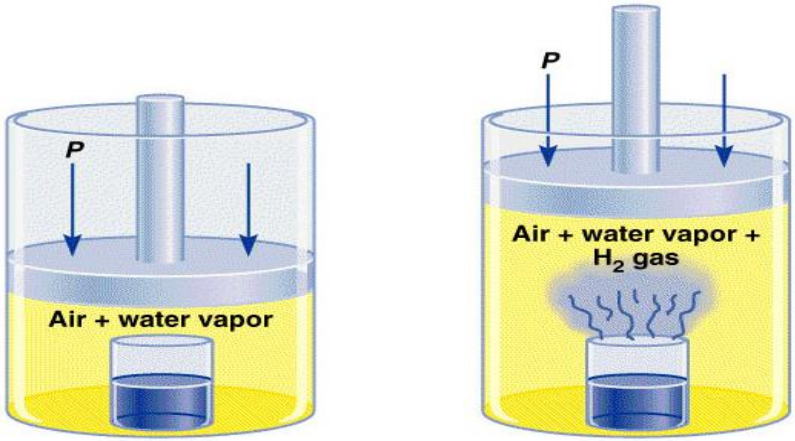
$2Na\ (s) + 2H_2O\ (l) \longrightarrow 2NaOH\ (aq) + H_2\ (g);\ \Delta H = -367.5\ kJ/mol$

$\Delta E = \Delta H - P\Delta V$ at 25 ^{0}C, 1 mole H_2 = 24.5 L at 1 atm

$P\Delta V$ = 1 atm x 24.5 L = 24.5 atm. L

= 1 atm x 24.5 Lx 101.3 / 1000 = 2.5 kJ

ΔE = -367.5 kJ/mol – 2.5 kJ/mol = -370.0 kJ/mol



1.4 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.

TABLE 6.2 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_2O	4.184
C_2H_5OH (ethanol)	2.46

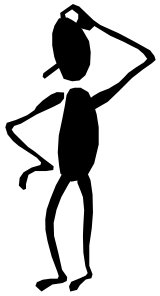
$$C = ms$$

Heat (*q*) absorbed or released:

$$q = ms\Delta t$$

$$q = C\Delta t$$

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



Example 1.5 How much heat is given off when an 869 g iron bar cools from 94^0C to 5^0C?

s of Fe = 0.444 J/g • ^{0}C

$\Delta t = t_{final} - t_{initial} = 5^0\text{C} - 94^0\text{C} = -89^0\text{C}$

$q = ms\Delta t$ = 869 ~~g~~ x 0.444 J/~~g~~ • ^{0}~~C~~ x -89^0~~C~~ = -34,000 J

Find the final temperature when 34000 j of heat evolved (given off) on cooling 869 g of Iron bar at 94 ^{o}C?

s of Fe = 0.444 J/g • ^{0}C

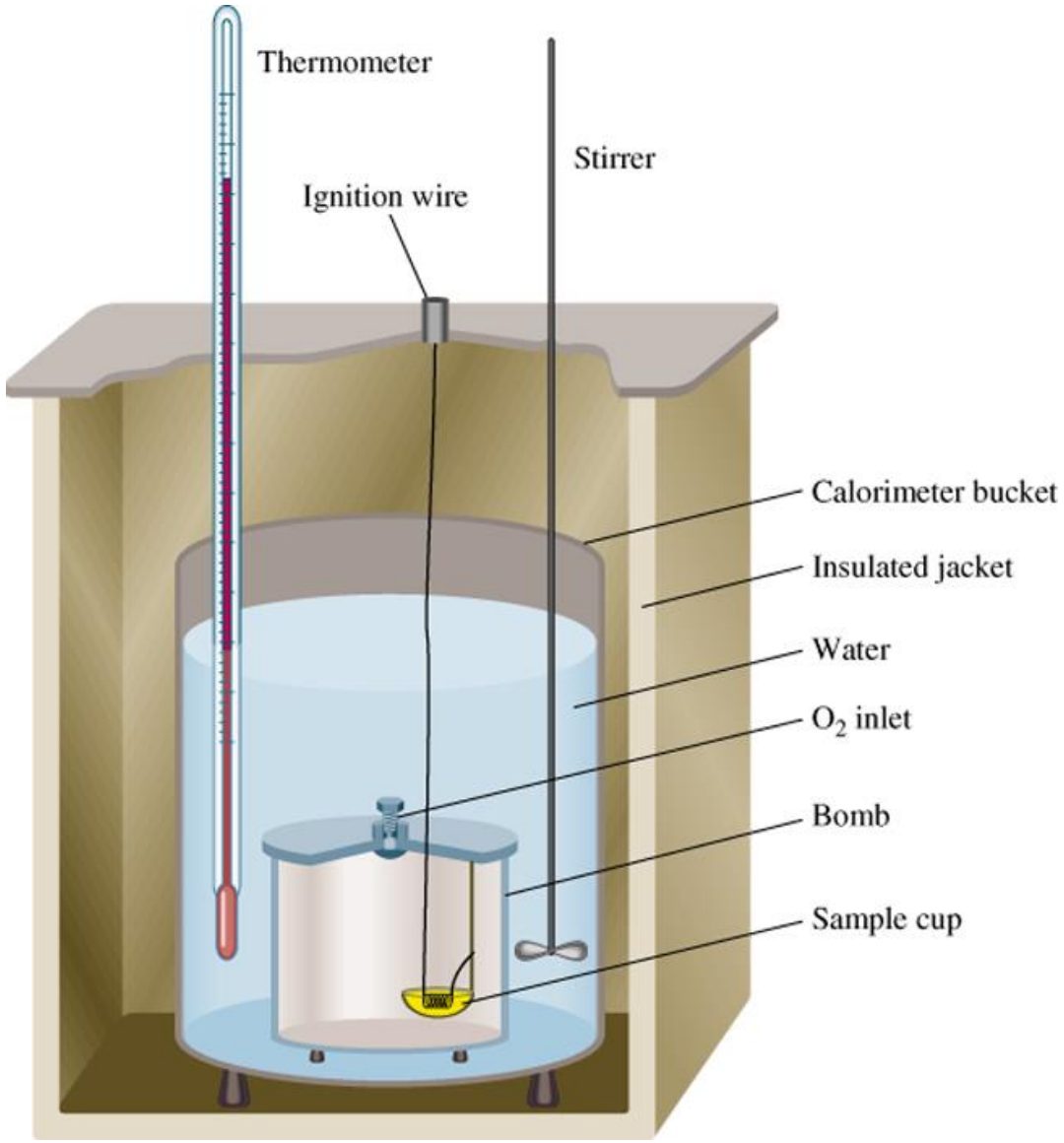
$q = mxsx\Delta T$

$-34000 = 869\ g\ x\ 0.444$ J/g • ^{0}C x ΔT

$\Delta T = -34000/\ 869\ x\ 0.444 = -89\ ^oC$

$T_f = T_i + \Delta T = 94 - 89 = 5\ ^oC$

Constant-Volume Calorimetry



No heat enters or leaves!

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

$q_{sys} = 0$

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

$q_{water} = ms\Delta t$

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

Reaction at Constant V

$\Delta H \neq q_{rxn}$

$\Delta H \sim q_{rxn}$

Example1.6: A bomb type calorimeter used to measure the heat evolved by the combustion of glucose $C_6H_{12}O_6$

$$C_6H_{12}O_6\,(s) + 6O_2(g) \rightarrow 6CO_2(g) + 6\,H_2O\,(l)$$

A 3.0 g sample of glucose is placed in bombs which is then filled with oxygen gas under pressure . The bomb is placed in wall insulated colorimeter vessel that is filled with 1.20 kg of water the initial temperature of the assembly is 19.00°C . The reaction mixture is ignited by the electrical heating of a wire within the bomb . The reaction causes the temperature of the calorimeter and its contents to increase to 25.50°C . The heat capacity of the bomb is 2.21 kJ / ˚C and the specific heat of water is 4.184 J/g.°C. How much heat is evolved by the combustion of 1 mol of glucose ?

Mass of $C_6H_{12}O_6$ =3 g, mass of water = 1.2kg,
T_1=19 °C, T_2= 25.5 °C,
C_{bom}= 2.21 kJ
s_{H2O} = 4.184 J/g °C, q_{rxn} at 1mol of $C_6H_{12}O_6$ =?

Mass of $C_6H_{12}O_6$ =3 g, mass of water = 1.2kg,

T_1=19 °C, T_2= 25.5 °C, C_{bom}= 2.21 kJ/°C

s_{H2O} = 4.184 J/g °C, qrxn at 1mol of $C_6H_{12}O_6$ =?

$q_{water} = m \times s \times \Delta t$

$= (1.2 \times 1000) \times 4.184 \times (25.5-19) = 32635.2$ J = 32.6 kJ

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

$= 2.21 \times (25.5-19) = 14.4$ kJ

$q_{rxn} = -(q_{water} + q_{bomb}) = -(32.6 + 14.4) = -47$ kJ

This q is for 3 g of $C_6H_{12}O_6$ what is q for 1 mole

n = g/mwt = 3 / 180 = 0.017 mol

0.017 mol --------□ -47 kJ

1 mol ----------□ ? kJ

1 x 47 /0.017 = -2764.7 kJ

Constant-Volume Calorimetry

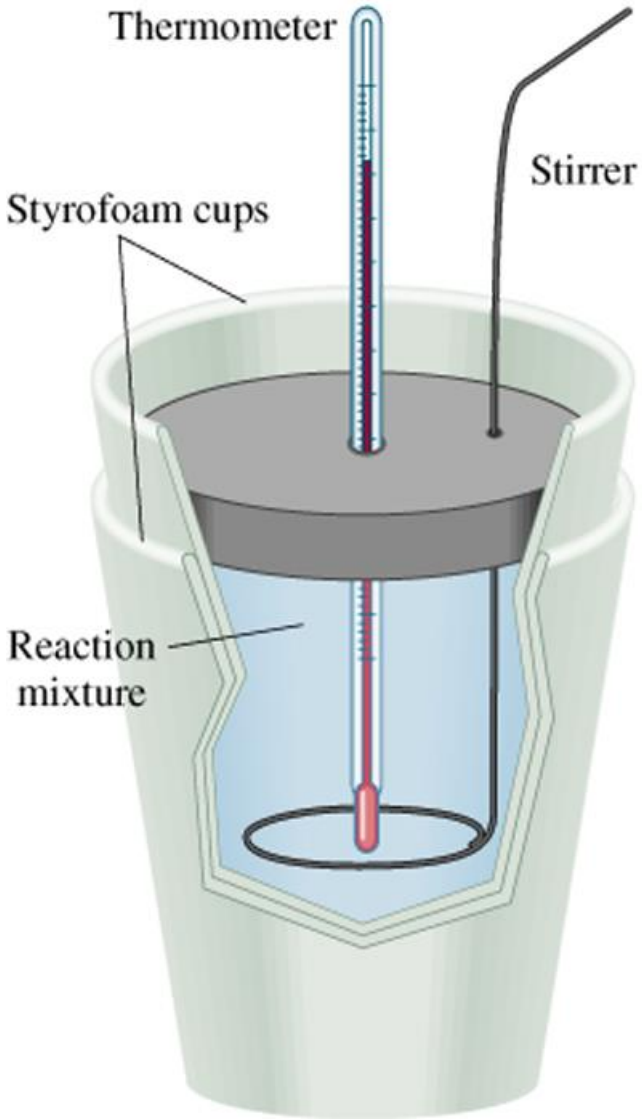
Example1.7: A quantity of 1.435 g of naphthalene ($C_{10}H_8$), a pungent-smelling substance used in moth repellents, was burned in a constant-volume bomb calorimeter. Consequently, the temperature of the water rose from 20.28°C to 25.95°C. If the heat capacity of the bomb plus water was 10.17 kJ/°C, calculate the heat of combustion of naphthalene?

$$q_{cal} = C_{cal} x\, \Delta t$$

$$= (10.17 \text{ kJ/°C})\ (25.95°C - 20.28°C)$$
$$= 57.66 \text{ kJ}$$

$$q_{rxn} = -\ (q_{water} + q_{bomb})$$
$$= -\ q_{cal}$$
$$= -57.66 \text{ kJ}$$

Constant-Pressure Calorimetry



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

$q_{sys} = 0$

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

$q_{water} = ms\Delta t$

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

Reaction at Constant P

$\Delta H = q_{rxn}$

No heat enters or leaves!

Constant-Pressure Calorimetry

Example1.8: A lead (Pb) pellet having a mass of 26.47 g at 89.98°C was placed in a constant-pressure calorimeter of negligible heat capacity containing 100.0 g of water. The water temperature rose from 22.50°C to 23.17°C. What is the specific heat of the lead pellet, if the specific heat of water is 4.184 J/g.°C ?

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

$q_{sys} = 0$

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

$q_{rxn} = -q_{water}$

$q_{Pb} = -q_{water}$

$q_{water} = m \times s \times \Delta t$

$= (100.0\ g)(4.184\ J/g.°C)(23.17°C - 22.50°C)$

$= 280.3\ J$

$q_{Pb} = -280.3\ J$

$q_{Pb} = m \times s \times \Delta t$

$$s = \frac{q_{Pb}}{m \times \Delta t}$$

$$s = \frac{-280.3\ J}{26.47g \times (23.17\ °C - 89.98\ °C)}$$

$s = 0.158\ J/g.°C$

Example1.9: The thermic reaction below is highly exothermic

$$2Al\ (s) + Fe_2O_3\ (s) \longrightarrow 2Fe\ (s) + Al_2O_3\ (s)\ ,\ \Delta H = -848\ KJ$$

How much heat is liberated when 36.0 g of Al, reacts with excess Fe_2O_3 ?

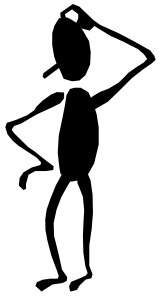
n = g/mwt =36/ 27 = 1.33 mol

From equation
2 mole Al -----→ -848 kJ
1.33 mol -----→ ? kJ
(1.33 x (-848)) / 2 = 563.9 kJ

TABLE 6.3 Heats of Some Typical Reactions Measured at Constant Pressure

Type of Reaction	Example	ΔH (kJ/mol)
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.



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** (ΔH_f^0) as a reference point for all enthalpy expressions.

Standard enthalpy of formation (ΔH_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.

$\Delta H_f^0 (O_2) = 0$ $\qquad$ ΔH_f^0 (C, graphite) = 0

$\Delta H_f^0 (O_3) = 142$ kJ/mol $\qquad$ ΔH_f^0 (C, diamond) = 1.90 kJ/mol

Example 1.10 Which of the following reactions represent the heat of formation of CO_2 !?

$$CaCO_3 = CaO + CO_2 , \ \Delta H^o_f \ ???$$

$$C_{graphite} + O_2 (g) = CO_2 (g) , \Delta H^o_f$$

Example 1.11 Which of the following reactions represent heat of formation of nitric acid HNO_3 ?

$H (g) + N (g) + O_3 (g) = HNO_3 (l)$
$H_2 (g) + N_2 (g) + O_2 (g) = HNO_3 (l)$
$H_2 (g) + N_2 (g) + 3O_2 (g) = 2HNO_3 (l)$

Appendix 1.1

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
$AgCl(s)$	−127.0	$Hg(l)$	0
$Al(s)$	0	$I_2(s)$	0
$Al_2O_3(s)$	−1669.8	$HI(g)$	25.9
$Br_2(l)$	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
$HF(g)$	−271.6	$SO_2(g)$	−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	$ZnO(s)$	−348.0
$H_2O(l)$	−285.8		

The standard enthalpy of reaction (ΔH^0_{rxn}) is the enthalpy of a reaction carried out at 1 atm.

$$aA + bB \longrightarrow cC + dD$$

$$\Delta H^0_{rxn} = [\, c\Delta H^0_f (C) + d\Delta H^0_f (D)\,] - [\, a\Delta H^0_f (A) + b\Delta H^0_f (B)\,]$$

$$\Delta H^0_{rxn} = S\, n\Delta H^0_f (\text{products}) - S\, m\Delta H^0_f (\text{reactants})$$

Example 1.12: Use enthalpies of formation to calculate ΔH^o for the reaction .

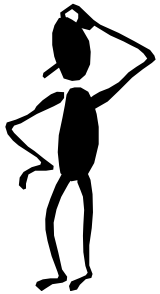
$Fe_2O_3(s) + 3CO\,(g) \rightarrow 2Fe(s) + 3CO_2\,(g)$

IF $\Delta H^\circ_f\, F_2O_3\,(s) = -822.2$ KJ / mol, ΔH°_f CO(g) = -110.5 KJ/ mol, $\Delta H^\circ_f\, CO_2\,(g) = -393.5$ KJ / mol

$\Delta H^\circ = 3\Delta H^\circ_f\, CO_2\,(g) - [\Delta H^\circ_f\, Fe_2\, O_3(s) + 3\, \Delta H^\circ_f\, CO(g)\,]$

$= 3\,(\,-393.5\text{ KJ}\,) - [\,-(822.2\text{ KJ}\,) + 3\,(\,-110.5\text{ KJ})]$

$= -\,26.8$ KJ



Example 1.13 Benzene (C_6H_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.

$$2C_6H_6\,(l) + 15O_2\,(g) \longrightarrow 12CO_2\,(g) + 6H_2O\,(l)$$

$$\Delta H^0_{rxn} = \Sigma n \Delta H^0_f \text{ (products)} - \Sigma m \Delta H^0_f \text{ (reactants)}$$

$$\Delta H^0_{rxn} = [12\Delta H^0_f\,(CO_2) + 6\Delta H^0_f\,(H_2O)] - [\,2\Delta H^0_f\,(C_6H_6)]$$

$$\Delta H^0_{rxn} = [\,12x{-}393.5 + 6x{-}187.6\,] - [\,2x49.04\,] = -5946 \text{ kJ}$$

$$\frac{-5946 \text{ kJ}}{2 \text{ mol}} = -2973 \text{ kJ/mol } C_6H_6$$

Example 1.14 : Calculate ΔH and ΔE for the reaction

$$OF_2(g) + H_2O(g) \rightarrow O_2(g) + 2HF(g)$$

IF $\Delta H^\circ_f\ OF_2(g) = 23$ KJ / mol, $\Delta H^\circ_f\ H_2O(g) = -241.8$ KJ/ mol, $\Delta H^\circ_f\ HF(g) = -268.6$ KJ / mol

$$\Delta H^\circ = \Sigma \Delta H^\circ_f\ P - \Sigma \Delta H^\circ_f\ R$$

$$= 2\ \Delta H^\circ_f\ HF(g) - [\ \Delta H^\circ_f\ OF2(g) + \Delta H^\circ_f\ H_2O\ (g)]$$

$$= 2(\ -268.6) - [\ 23 + (-241.8)]$$

$$= -318.4 KJ$$

$$\Delta H = \Delta E + \Delta nRT$$

$$\Delta n = 3 - 2 = 1$$

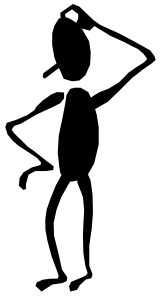
$$-318.4 = \Delta E + (\ 1 \times 8.314 \times 298) / 1000$$

$$= \Delta E + 2.48$$

$$\Delta E = -318.4 - 2.48 = -320.9 KJ$$

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.)



Example 1.15 Calculate the standard enthalpy of formation of CS_2 (*l*) given that:

$C(graphite) + O_2\ (g) \longrightarrow CO_2\ (g)$ $\quad \Delta H^0_{rxn} = -393.5$ kJ

$S(rhombic) + O_2\ (g) \longrightarrow SO_2\ (g)$ $\quad \Delta H^0_{rxn} = -296.1$ kJ

$CS_2(l) + 3O_2\ (g) \longrightarrow CO_2\ (g) + 2SO_2\ (g)$ $\quad \Delta H^0_{rxn} = -1072$ kJ

1. Write the enthalpy of formation reaction for CS_2

$C(graphite) + 2S(rhombic) \longrightarrow CS_2\ (l)$

2. Add the given rxns so that the result is the desired rxn.

$C(graphite) + O_2\ (g) \longrightarrow CO_2\ (g)$ $\quad \Delta H^0_{rxn} = -393.5$ kJ

$2S(rhombic) + 2O_2\ (g) \longrightarrow 2SO_2\ (g)$ $\quad \Delta H^0_{rxn} = -296.1 \times 2$ kJ

$+ \quad CO_2(g) + 2SO_2\ (g) \longrightarrow CS_2\ (l) + 3O_2\ (g)$ $\quad \Delta H^0_{rxn} = +1072$ kJ

$C(graphite) + 2S(rhombic) \longrightarrow CS_2\ (l)$

$\Delta H^0_{rxn} = -393.5 + (2x-296.1) + 1072 = 86.3$ kJ

The ***enthalpy of solution*** (ΔH_{soln}) is the heat generated or absorbed when a certain amount of solute dissolves in a certain amount of solvent.

$$\Delta H_{soln} = H_{soln} - H_{components}$$

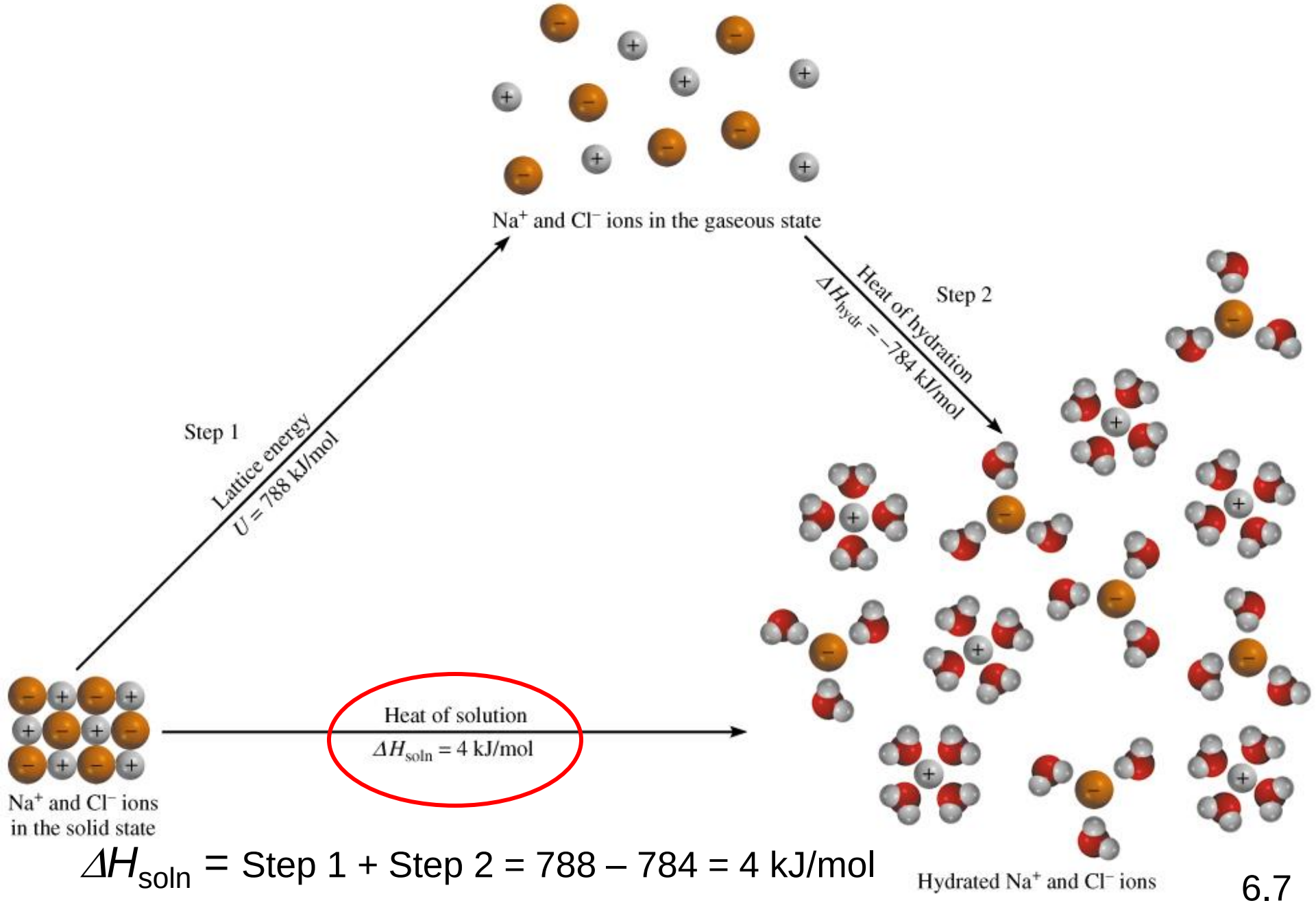
TABLE 6.5 Heats of Solution of Some Ionic Compounds

Compound	ΔH_{soln} (kJ/mol)
LiCl	−37.1
$CaCl_2$	−82.8
NaCl	4.0
KCl	17.2
NH_4Cl	15.2
NH_4NO_3	26.2

Which substance(s) could be used for melting ice?

Which substance(s) could be used for a cold pack?

The Solution Process for NaCl



ΔH_{soln} = Step 1 + Step 2 = 788 – 784 = 4 kJ/mol