

Chapter 6: Thermochemistry (Chemical Energy) (Ch6 in Chang, Ch6 in Jespersen)

Energy is defined as the **capacity to do work**, or **transfer heat**.

Work (w) - *force* (F) applied through a *distance*.

Force - any kind of *push* or *pull* on an object.

$$F = ma \quad \left(\text{kg} \cdot \frac{\text{m}}{\text{s}^2} \right)$$

$$w = F \times d \quad \left(\text{kg} \cdot \frac{\text{m}}{\text{s}^2} \times \text{m} \right)$$

Chemists define **work** as *directed* energy change resulting from a process.

The SI unit of **work** is the **Joule** $1 \text{ J} = 1 \text{ kg m}^2/\text{s}^2$

(Also the **calorie** $1 \text{ cal} = 4.184 \text{ J}$ (exactly); and the **Nutritional calorie** $1 \text{ Cal} = 1000 \text{ cal} = 1 \text{ kcal}$.)

A **calorie** is the energy needed to increase the temperature of 1 g of water by 1 °C at standard atmospheric pressure. But since this depends on the atmospheric **pressure** and the starting **temperature**, there are several different *definitions* of the “**calorie**”).

There are many *different* types of **Energy**:

Radiant Energy is energy that comes from the *sun* (heating the Earth's surface and the atmosphere).

Thermal Energy is the energy associated with the random *motion* of **atoms** and **molecules**.

Chemical Energy is *stored* within the *structural units* of *chemical substances*. It is determined by the *type* and *arrangement* of the **atoms** of the substance.

Nuclear Energy is the energy stored within the collection of *protons* and *neutrons* of an **atom**.

Kinetic Energy is the energy of *motion*. Chemists usually relate this to **molecular** and **electronic** *motion* and *movement*. It depends on the mass (m), and speed (v) of an object.

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

Potential Energy is due to the *position relative* to other *objects*.

It is “*stored energy*” that results from **attraction** or **repulsion** (e.g. gravity or electrostatic attraction/repulsion).

$$PE = mgh \quad (\text{due to } \text{gravity})$$

The **force** of electrostatic attraction (or repulsion) between *charged* particles (**Coulombs Law**):

$$F = k \frac{Q_1 Q_2}{d^2}$$

When one form of **energy** *disappears*, some other form of **energy** (of **equal magnitude**) must appear.

⇒ **Law of Conservation of energy**: the *total* quantity of energy in the universe is assumed *constant*.

Transferring Energy: Work and Heat

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

(Remember that **temperature** is related to the *average kinetic energy* of an object's particles).

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

Systems and Surroundings

To discuss the energy changes associated with chemical reactions we must first define the specific part of the universe that is of interest to us, which we call the “*system*”.

System - what we *study* (e.g. a beaker or flask).

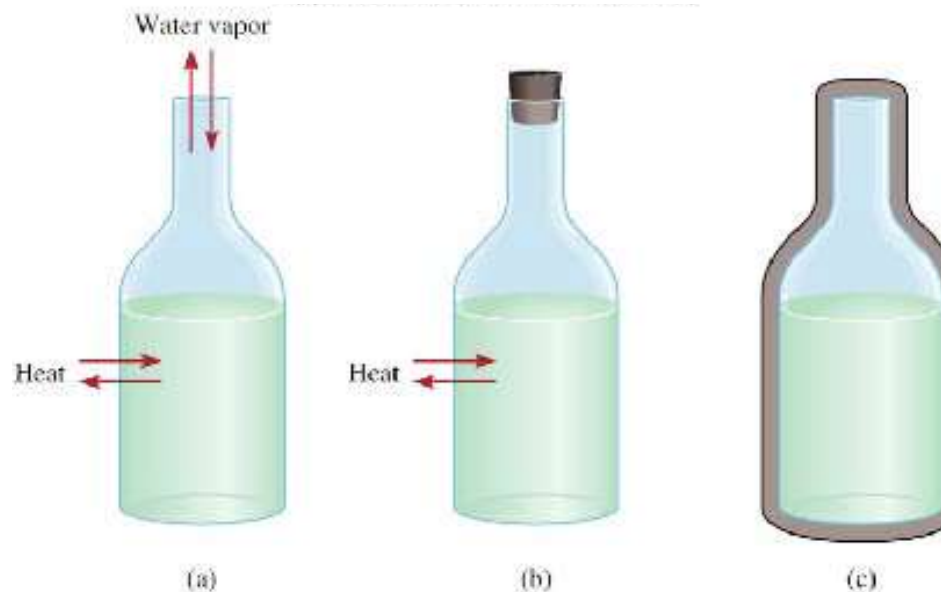
Surroundings - the *rest* of the universe *outside* of our system (e.g. the room where the experiment is performed).

We will study/use **three** types of **systems** (*open, closed and isolated*)...

Open system - can exchange **mass** (matter) and **energy** with surroundings (most common).

Closed system - can exchange **energy** but **not matter** with surroundings.

Isolated system - does **not** allow the transfer of either **mass** or **energy**.
(E.g. sealed, vacuum coated thermos flask)



Adiabatic Process is defined as a process that occurs in an **isolated system**, meaning where neither energy nor matter crosses the system/surroundings boundary.

(When permitted) **Energy** can be exchanged with the *surroundings* as **heat** and/or **work**.

Energy Changes and Energy Conservation

Thermochemistry is a subset of **Thermodynamics**, which is the scientific study of the **interconversion** of **heat** and other kinds of **energy**.

The First Law of Thermodynamics: **Energy is neither created nor destroyed.**

Energy is *conserved*. Any energy *lost* by the **system** is *gained* by the **surroundings**.

Internal Energy (E) – is the *sum* of all *kinetic* and *potential energy* of *all* components of the *system* (translational, rotational, vibrational, electronic, nuclear, etc).

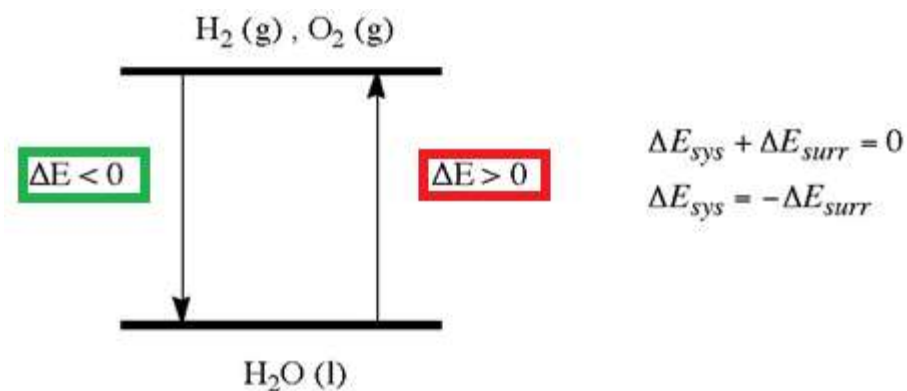
It is very **difficult** to determine or measure the *absolute total energy* of entities, but it is more *convenient* to measure the *change* in **Internal Energy (ΔE)** for a given process:

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

Thermodynamic quantities have three parts:
number, unit and sign (+ or -).

$+\Delta E$ means $E_f > E_i$
system *gained* E from surroundings
(energy **coming in**)

$-\Delta E$ means $E_f < E_i$
system *lost* E to surroundings
(energy **going out**)

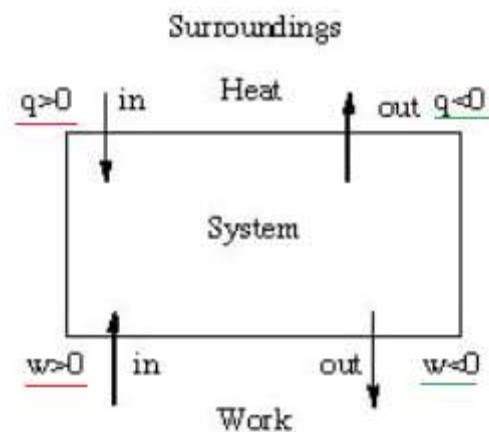


Heat (q) added to the system, and/or (w) work done on the system, **increases** the internal energy.

Notice that **+ve** means energy **in**

and

-ve means energy **out** of the system.



Endothermic and Exothermic Processes

Endothermic – the system **absorbs** heat. Heat flows **into** system from surroundings. ($+\Delta E$)

Exothermic – the system **evolves** heat. Heat flows **out of** system and into surroundings. ($-\Delta E$)

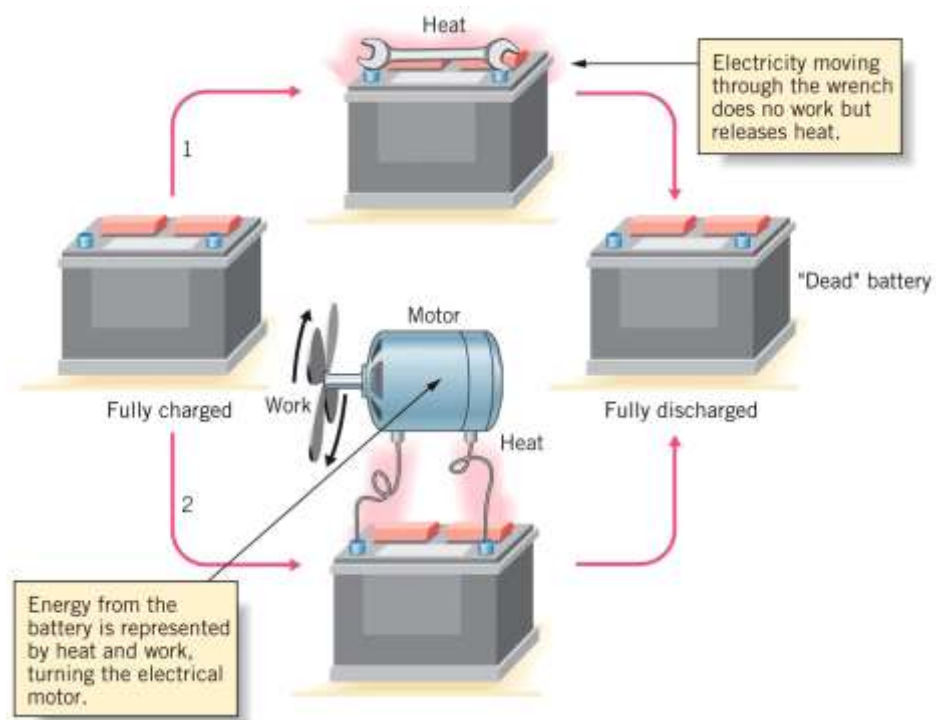
Relating ΔE to Heat (q) and Work (w)

Energy is defined as the capacity to *do work*, or *transfer heat*.

Energy can be *exchanged* with the surroundings as *heat* (q) and/ or *work* (w).

So we can write this as:

$$\Delta E = q + w$$



The battery can be *drained* either by heating the wrench, or by work and heat by powering the fan.

Internal Energy is a State Function

The **state of a system** is defined by the *values of all relevant macroscopic properties that specify an object's current condition*. E.g. composition, energy, temperature, pressure and volume, etc.

A *state function* is a property of a system that is determined by specifying its *condition* (in terms of temperature, pressure, internal energy, etc.).

The **value** of a **state function** does **not** depend on the **particular history** of the sample, only its **present condition**.

The **change** in the **state function** depends only on the *initial and final states of the system*, **not** how the change occurs.



State functions do **not** depend on the path by which the system arrived at its present state.

A state function describes the **equilibrium state** of a system.

The climbers gain the **same** gravitational potential energy whether they take the **red** or **blue** path.

(Note that **work** is **not** a state function, since the **red** path is **longer** and would take **more energy**.

Work is a *process* function).

Work by Expansion

Previously we have seen that (one form of) **work** can be defined as **force** times **distance** (*mechanical work*).

Mechanical work can also be expressed by *expansion* or *compression* of a gas.

Pressure-volume work (or **PV** work) occurs when the **volume** (**V**) of a system *changes*:

$$P \times V = \frac{F}{d^2} \times \underset{\text{volume}}{d^3} = Fd = w \qquad 1 \text{ L}\cdot\text{atm} = 101.3 \text{ J}$$

pressure

So at *constant pressure* (such as *atmospheric*):

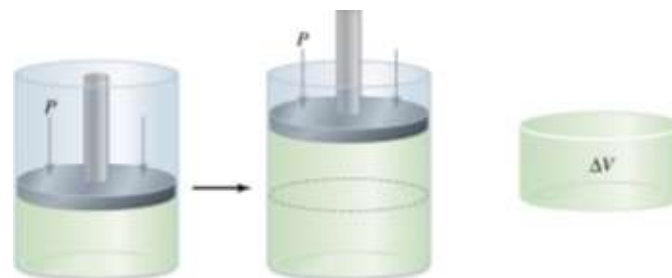
$$\text{Work, } w = -P\Delta V$$

Where P = *opposing* pressure against which the system (e.g. a *piston*) *pushes*.

ΔV = change in volume of gas during expansion = $V_{\text{final}} - V_{\text{initial}}$.

For **expansion**, since $V_{\text{final}} > V_{\text{initial}}$, then ΔV must be **positive**.

So **expansion work** is **negative** (meaning work is done **by** system. The system is *pushing the piston back*).



Problem: If a gas expands in volume from 1.0 to 4.0 L at constant temperature against a constant pressure of 2.4 atm, how much work is done by the gas?

$$\begin{aligned}\text{Work} &= -P\Delta V \\ &= -2.4 \text{ atm} \times (4.0 - 1.0) \text{ L} \\ &= -7.2 \text{ L atm}\end{aligned}$$

Enthalpy

The *change* in **Internal Energy** (ΔE) has two components, **work** (**w**) and **heat** (**q**).

If we operate under a *constant volume*, then $\Delta V=0$ (and so *no PV work is done*):

$$\begin{aligned}\Delta E &= q + w \\ &= q - P\Delta V \\ \Delta E &= q_v \quad \text{at constant } V \text{ (as indicated by the subscript).}\end{aligned}$$

Practically it is **inconvenient** to operate under constant **volume**, it is much easier to operate under constant **pressure** (e.g. at atmospheric pressure)...

Heat (q) transferred under *constant-pressure* conditions is called **enthalpy, H** .

Enthalpy is a state function.

$$E = q + w \quad \text{so} \quad q = E - W \quad \text{but} \quad W = -PV$$

$$\mathbf{H = E + PV}$$

E = internal energy; P = pressure; V = volume of the system

The *change* in **enthalpy, ΔH** , equals the **heat, q_p** gained or lost by the **system** when the process occurs under *constant pressure* (indicated by the subscript p).

$$\begin{aligned}\Delta H &= H_{\text{final}} - H_{\text{initial}} \\ &= q_p\end{aligned}$$

We know $\Delta E = q + w = q - P\Delta V$

At *constant P* $\Delta E = q_p - P\Delta V$

Rearrange to give $q_p = \Delta E + P\Delta V$

But also $q_p = \Delta H$

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

- If ΔH is **positive** the *system* has **gained heat** from the *surroundings*, and is an **endothermic** process.
- If ΔH is **negative**, the *system* has **released heat** to the *surroundings*, and is an **exothermic** process.

To recap - for a reaction:

At *constant volume*, the **heat** change, $q_v = \Delta E$.

At *constant pressure*, the **heat** change, $q_p = \Delta H$ (*this is the more common scenario*).

The Thermodynamic Standard State

We generally deal with *changes* in (**not absolute values** of) **energy**, so we need some type of *reference point*.

The **standard state** of a material (*pure substance*, *mixture* or *solution*) is the **reference point** used to *calculate* its **properties** under *different conditions*.

It is that of a **pure substance** in a *specified* state (its **most stable form** and *physical state*), at *1 atm pressure* and at *25 °C*; and **1 M concentration** for all substances in **solution**.

It is indicated by the superscript °:

E.g. ΔH° (means the **standard change** in **Enthalpy**) ΔS° ΔG° E°

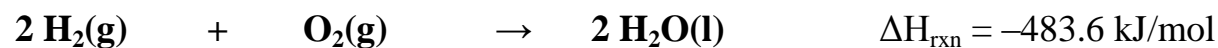
Now we are ready to talk about *reactions*...

Enthalpies of Physical and Chemical Change -Thermochemical Equations

For a chemical reaction, the **Enthalpy of Reaction** is the *change* in **enthalpy** going *from reactants to products*.

$$\Delta H = H_{\text{products}} - H_{\text{reactants}}$$

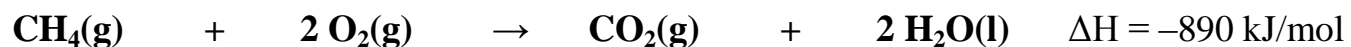
The **heat of reaction** is ΔH_{rxn} .



Enthalpy can be viewed as a measure of *how much heat is stored as potential energy in the system*, or as “*heat content*”.

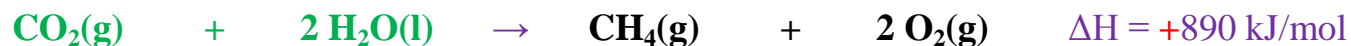
Three helpful guidelines

1. **Enthalpy** is an **extensive** property. The magnitude of ΔH is directly proportional to *amount* of reactant consumed in the process.

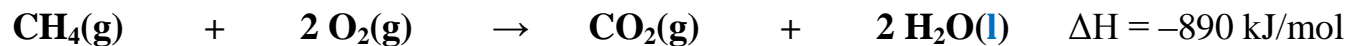


The combustion of **two** moles of CH_4 would release **double** the energy, meaning -1780 kJ/mol .

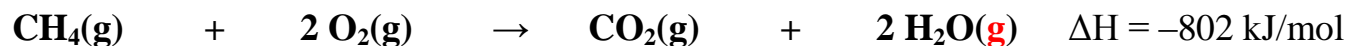
2. The **enthalpy change** for a reaction is **equal** in magnitude but **opposite** in sign to ΔH for the **reverse reaction**.



3. The **enthalpy change** for a reaction depends on the **physical state** of the *reactants* and *products*.



Sum and cancel to give...



Problem: The combustion of butane produces heat according to the equation:



What is the heat of combustion per gram of butane?

2 moles of C_4H_{10} generate -5314 kJ/mol

Molar mass $\text{C}_4\text{H}_{10} = 4(12.01) + 10(1.008) = 58.12 \text{ g/mol}$

So 2 x 58.12 g/mol generates -5314 kJ/mol

$$\begin{aligned} \text{Therefore heat of combustion per g} &= \frac{-5314 \text{ kJ/mol}}{116.24 \text{ g/mol}} \\ &= -45.72 \text{ kJ/g} \end{aligned}$$

Calorimetry and Heat Capacity

Calorimetry is the measurement of *heat changes*. (Apparatus used is a *calorimeter*).

Understanding calorimetry depends on the concepts of **specific heat**, and **heat capacity**.

Heat Capacity and Specific Heat

Heat capacity (**C**) is the amount of *heat* required to raise the *temperature* of a given *quantity* of substance by *1 K* (or *1 °C*).

Specific heat (**s**) or **specific heat capacity** is the *heat capacity* of **one gram** of substance.

Molar heat capacity is the **heat capacity** of **1 mol** of substance.

There is a *practical* issue - we **cannot** measure heat *directly*, but **heat gained** or **lost** (**q**) is directly proportional to the **temperature change** (**Δt**).

$$q = C \Delta t \quad \text{where } C \text{ is the } \textit{heat capacity}, \text{ with units of } J/^{\circ}C, \text{ or } J/K.$$

$$C = s m \quad \text{where } m = \text{mass in grams, } s \text{ has units of } J/g \text{ } ^{\circ}C.$$

So combining these: **q = s m Δt** (*Heat exchanged = Specific Heat × mass × temperature change*)

Table of [Specific Heats](#) for some Substances

Substance	Specific Heat ($\text{J g}^{-1} \text{ }^\circ\text{C}^{-1}$)
Carbon (graphite)	0.711
Copper	0.387
Ethyl alcohol	2.45
Gold	0.129
Granite	0.803
Iron	0.4498
Lead	0.128
Olive oil	2.0
Silver	0.235
Water (liquid)	4.184

Larger specific heats mean it takes more energy to heat them up; and conversely, they release a larger amount of heat as they cool.

Notice the large value for water – great for a cooling system like your body; or a heating system like radiators; also explains maritime climates.

Problem: The specific heat of ethylene glycol is 2.42 J/gK. How many Joules of heat are necessary to raise the temperature of 62.0 g of ethylene glycol from 15.2 °C to 40.8 °C?

REMEMBER Δt in °C = Δt in K. $\Delta t = t_{\text{final}} - t_{\text{initial}}$

$$\begin{aligned}q &= m s \Delta t = (62.0 \text{ g}) \left(2.42 \frac{\text{J}}{\text{g K}}\right) (40.8 - 15.2 \text{ }^\circ\text{C} \left(\frac{1 \text{ K}}{1 \text{ }^\circ\text{C}}\right)) \\&= 3841 \text{ J} \\&= 3.84 \text{ kJ}\end{aligned}$$

Problem: A piece of a newly synthesized material of mass 12.0 g at 88.0 °C is placed in a calorimeter containing 100.0 g of water at 20.0 °C. If the final temperature of the system is 24.0 °C, what is the specific heat capacity of this material?

$$q_{\text{H}_2\text{O}} = -q_{\text{material}} \quad s_{\text{H}_2\text{O}} = 4.184 \frac{\text{J}}{\text{g} \cdot ^\circ\text{C}} \quad \Delta t = t_{\text{final}} - t_{\text{initial}}$$

$$m_{\text{H}_2\text{O}} s_{\text{H}_2\text{O}} \Delta t_{\text{H}_2\text{O}} = - (m_{\text{mat}} s_{\text{mat}} \Delta t_{\text{mat}})$$

$$\begin{aligned} s_{\text{mat}} &= \frac{m_{\text{H}_2\text{O}} s_{\text{H}_2\text{O}} \Delta t_{\text{H}_2\text{O}}}{- m_{\text{mat}} \Delta t_{\text{mat}}} \\ &= \frac{(100.0 \text{ g}) (4.184 \frac{\text{J}}{\text{g} \cdot ^\circ\text{C}}) (24.0 - 20.0 \text{ } ^\circ\text{C})}{- (12.0 \text{ g}) (24.0 - 88.0 \text{ } ^\circ\text{C})} \\ &= 2.18 \frac{\text{J}}{\text{g} \cdot ^\circ\text{C}} \end{aligned}$$

There are two common forms of **Calorimeter**: constant *pressure* and constant *volume*.

Constant Pressure Calorimetry is the simplest.

E.g. a Styrofoam cup calorimeter.

Since it is open to the atmosphere, it is *constant Pressure*.

But we **approximate** this as an **isolated** system.

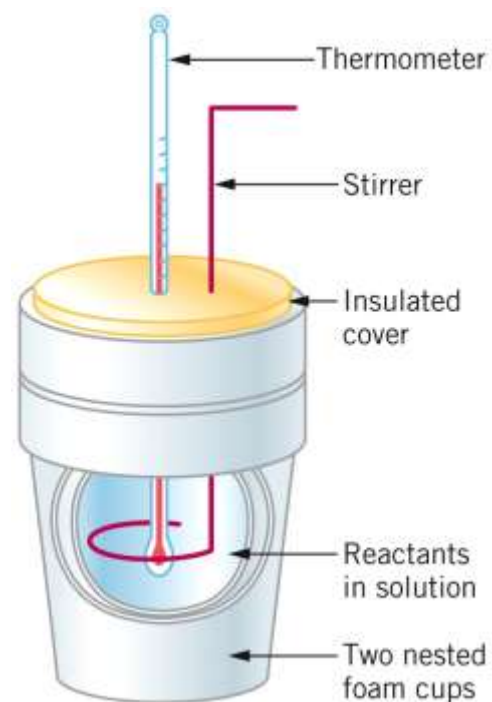
Heat is exchanged between the *reaction* and the *solution* (water).

The thermometer measures the change in temperature.

The heat of the reaction, $-q_p = q_{\text{solution}}$

$$= C \Delta t$$

$$= s m \Delta t$$



Problem: When a 9.55 g sample of solid sodium hydroxide dissolves in 100.0 g of water in a Styrofoam cup calorimeter, the temperature rises from 23.6 °C to 47.4 °C.

Calculate ΔH (in kJ/mol NaOH) for the solution process:



(Assume the specific heat of the solution is the same as that of pure water).

$$q_{\text{soln}} = m s \Delta t = -q_{\text{rxn}} \quad s_{\text{H}_2\text{O}} = 4.184 \frac{\text{J}}{\text{g}^\circ\text{C}} \quad \Delta t = t_{\text{final}} - t_{\text{initial}}$$

$$m_{\text{soln}} = 9.55 + 100.0 \text{ g} = 109.55 \text{ g}$$

$$q_{\text{soln}} = (109.55 \text{ g}) \left(4.184 \frac{\text{J}}{\text{g}^\circ\text{C}} \right) (47.4 - 23.6^\circ\text{C}) = 10908.9 \text{ J}$$

$$9.55 \text{ g NaOH} \times \frac{1 \text{ mol NaOH}}{40.00 \text{ g NaOH}} = 0.23875 \text{ mol}$$

$$\Delta H = \frac{-10908.9 \text{ J} \cdot \frac{\text{kJ}}{1000 \text{ J}}}{0.23875 \text{ mol}} = -45.7 \frac{\text{kJ}}{\text{mol}}$$

Constant Volume Calorimetry

A *bomb calorimeter* is used for *constant volume* reactions.

A *known amount* of compound is placed the bomb.

The **sturdy** bomb does **not** change its volume.

Gas is added to achieve the desired pressure.

The bomb is immersed in a *known amount* of water.

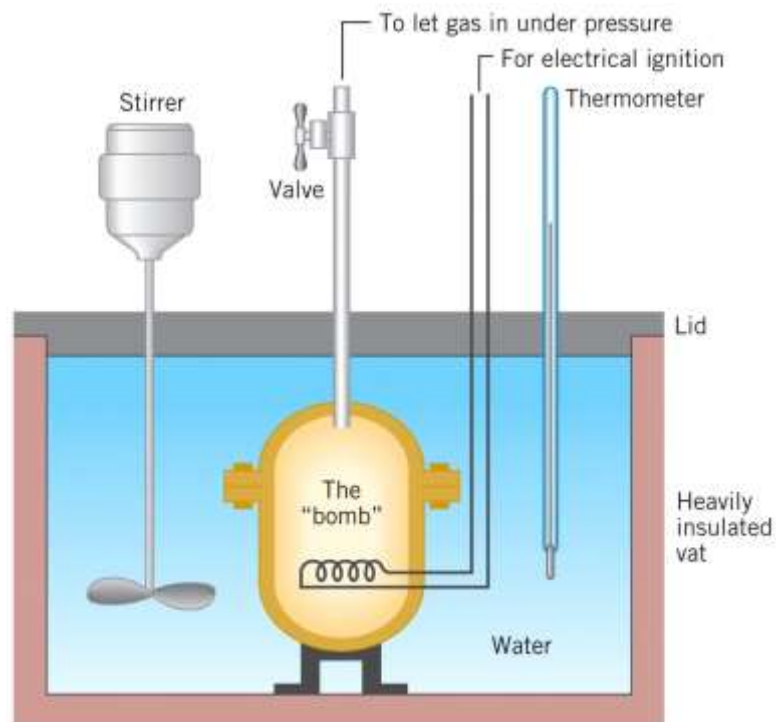
Ignition causes the reaction, and the *heat released raises* the *temperature* of the *water*.

The bomb and the water create an isolated system.

$\Delta E = q - P\Delta V = q_v$ for constant volume.

$$-q_{\text{rxn}} = q_{\text{calorimeter}} = C_{\text{cal}}\Delta t$$

where C_{cal} = heat capacity of the calorimeter (*determined* by known samples, calibration).



Standard Heats of Formation

There are some **standard enthalpy changes** associated with certain processes:

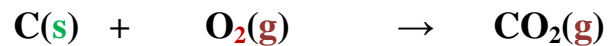
- Enthalpies of **vaporization**, ΔH_{vap} (*liquid to gas*)
- Enthalpies of **fusion**, ΔH_{fus} (*solid to liquid*)
- Enthalpies of **combustion**
- Enthalpy of **formation**, ΔH_{f} , (or *heat of formation*), the formation of a compound from its *constituent elements*.

One of the most useful pieces of information for **chemists** is...

The **standard enthalpy of formation** of a compound, $\Delta H_{\text{f}}^{\circ}$ is the **change in enthalpy** for the **reaction** that forms **1 mol** of the **compound** from its **elements**, with all *substances* in their *standard states*.

(Recall that **standard state** means the *most stable form* and *physical state* of the element at *1 atm* and *25 °C*).

E.g. the formation of CO_2 uses **solid** C, **gaseous** O_2 , and makes **gaseous** CO_2 ; and for oxygen, we use O_2 not O_3 .



The **standard enthalpy of formation** of the *most stable form* of any **element** is **zero**.

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

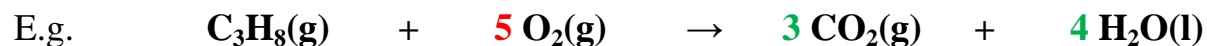
Substance	ΔH_f° (kJ/mol)	Substance	ΔH_f° (kJ/mol)
Ag(s)	0	H ₂ O ₂ (l)	-187.6
AgCl(s)	-127.0	Hg(l)	0
Al(s)	0	I ₂ (s)	0
Al ₂ O ₃ (s)	-1669.8	HI(g)	25.9
Br ₂ (l)	0	Mg(s)	0
HBr(g)	-36.2	MgO(s)	-601.8
C(graphite)	0	MgCO ₃ (s)	-1112.9
C(diamond)	1.90	N ₂ (g)	0
CO(g)	-110.5	NH ₃ (g)	-46.3
CO ₂ (g)	-393.5	NO(g)	90.4
Ca(s)	0	NO ₂ (g)	33.85
CaO(s)	-635.6	N ₂ O(g)	81.56
CaCO ₃ (s)	-1206.9	N ₂ O ₄ (g)	9.66
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 ₂ S(g)	-20.15
H ₂ O(g)	-241.8	Zn(s)	0
H ₂ O(l)	-285.8	ZnO(s)	-348.0

Using Enthalpies of Formation to Calculate Enthalpies of Reaction (**Direct Method**)

We can **calculate** the **standard enthalpy change** for *any* reaction for which we *know* the ΔH_f° values for *all reactants and products*.

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

(**m** and **n** are the equation *coefficients*)



<u>Substance</u>	<u>ΔH_f° (kJ/mol)</u>
$\text{C}_3\text{H}_8(\text{g})$	-103.85
$\text{O}_2(\text{g})$	0
$\text{CO}_2(\text{g})$	-393.50
$\text{H}_2\text{O}(\text{l})$	-285.83

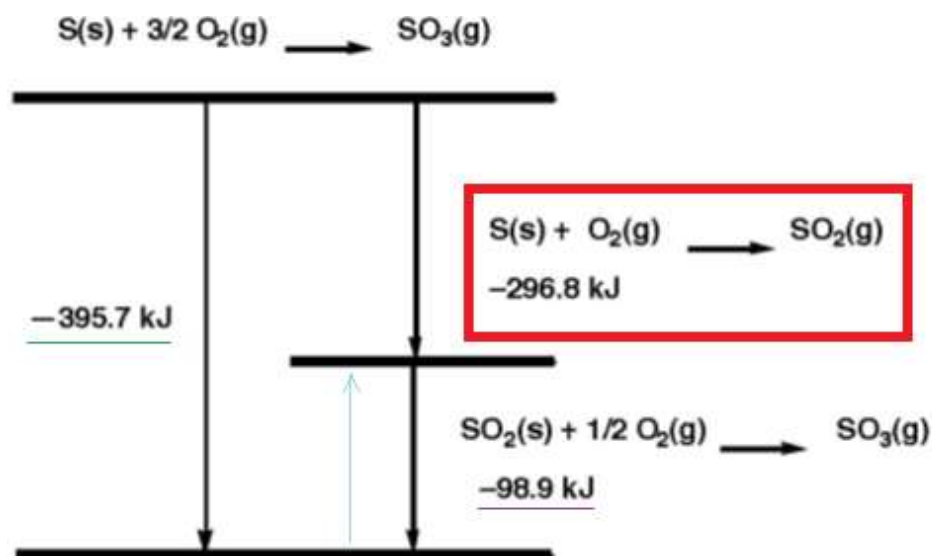
$$\begin{aligned}\Delta H_{\text{rxn}}^\circ &= [3 \Delta H_f^\circ(\text{CO}_2) + 4 \Delta H_f^\circ(\text{H}_2\text{O})] - [\Delta H_f^\circ(\text{C}_3\text{H}_8) + 5 \Delta H_f^\circ(\text{O}_2)] \\ &= [3(-393.50 + 4(-285.83))] - [-103.85 - 5(0)] \\ &= -2220.05 \text{ kJ/mol} \quad (\text{addition, so answer to 2 dec. pl.})\end{aligned}$$

Hess's Law (Indirect Method)

Hess's law states that if a reaction is carried out in a *series of steps*, ΔH for the **reaction** will be *equal* to the *sum* of the **enthalpy changes** for the *individual steps*.

Hess pointed out that the heat *absorbed* (or *evolved*) in a given chemical reaction is the **same** whether the process takes *one step* or *several steps*.

This is *useful* when a reaction **cannot** be measured **directly**.



Trying to calculate: $\text{S(s)} + \text{O}_2(\text{g}) \rightarrow \text{SO}_2(\text{g})$

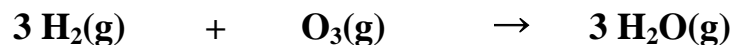
- (a) $\text{S} + \frac{3}{2} \text{O}_2 \rightarrow \text{SO}_3$ $\Delta H = -395.7 \text{ kJ/mol}$
(b) $\text{SO}_2 + \frac{1}{2} \text{O}_2 \rightarrow \text{SO}_3$ $\Delta H = -98.9 \text{ kJ/mol}$
(c) $\text{SO}_3 \rightarrow \text{SO}_2 + \frac{1}{2} \text{O}_2$ $\Delta H = +98.9 \text{ kJ/mol}$

Adding (a) and (c) gives:

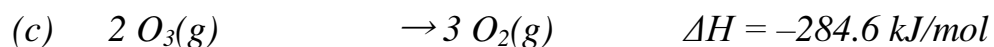
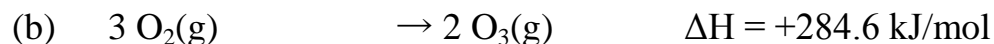
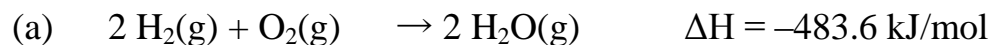


$$\text{So } \Delta H = -395.7 + 98.9 = -296.8 \text{ kJ/mol}$$

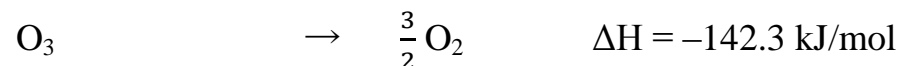
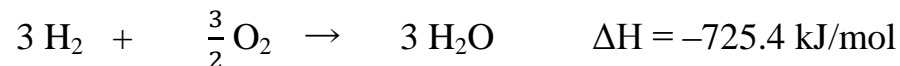
Problem: Calculate the enthalpy change for the reaction:



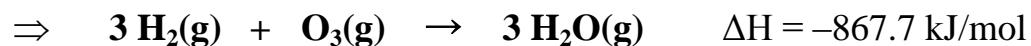
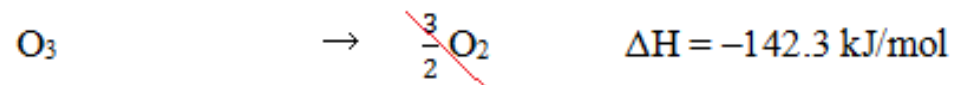
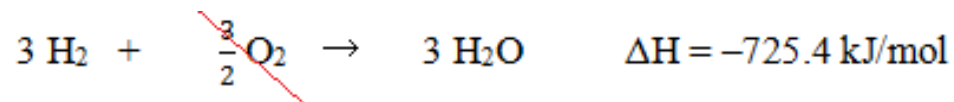
Given the following enthalpies of reaction:



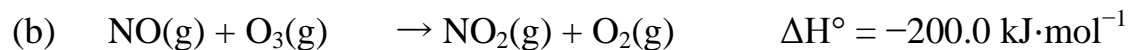
To generate our desired equation, we must multiply (a) by $\frac{3}{2}$, and (c) by $\frac{1}{2}$.



Adding these gives:



Problem: Given the following standard reaction enthalpies:

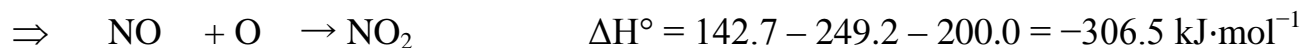


If the standard enthalpy of formation of ozone = $+142.7 \text{ kJ}\cdot\text{mol}^{-1}$, calculate the standard reaction enthalpy for the reaction:



So, (c) The standard enthalpy of formation of ozone: $\frac{3}{2} \text{O}_2 \rightarrow \text{O}_3 \quad \Delta H = 142.7 \text{ kJ}\cdot\text{mol}^{-1}$

So to get our desired equation, we can combine: (c) - $\frac{1}{2}$ (a) + (b).



Many **physical** processes also involve the *absorption* or *release* of **heat**.

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

The heat of **hydration** is the enthalpy change associated with the **hydration** process.

The heat of **dilution** is the heat change associated with the **dilution** process.

The heat of **solution** represents the difference between the enthalpy of the **final solution** and the enthalpies of its **original components** (i.e. the *solute* and the *solvent*).

The heat of **solution** can be regarded as the *sum* of the enthalpy changes of three *intermediate steps*:

1. The *breaking of bonds* within the **solute**, such as the *electrostatic attraction* between two **ions** (**endothermic**)
2. The *breaking of intermolecular attractive* forces within the solvent, such as **hydrogen bonds** (**endothermic**)
3. The formation of **new attractive solute-solvent** bonds in solution (**exothermic**).

The value of the *overall heat of solution* is the **sum** of these steps.

$$\Delta H_{\text{soln}} = U + \Delta H_{\text{hydr}}$$

Where **U** = **Lattice energy** (the *energy* required to completely *separate one mole* of a *solid ionic compound* into *gaseous ions*).

The *gaseous* ions then must become **hydrated**, which is the **heat of hydration**.

Heats of Solution of Some Ionic Compounds

Compound	(kJ/mol)
LiCl	-37.1
CaCl ₂	-82.8
NaCl	4.0
KCl	17.2
NH ₄ Cl	15.2
NH ₄ NO ₃	26.2

Notice this can be overall **endo-** or **exothermic** depending on the species.

(If you dissolve NaCl in water, it actually takes energy, and cools down the solution).