<table>
<thead>
<tr>
<th>Week</th>
<th>Chapters</th>
<th>Topics</th>
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<td><em>Intermolecular Forces, Liquids and Solids</em></td>
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<td><em>Intermolecular Forces, Liquids and Solids</em></td>
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</table>
Energy

Wind

Solar

Hydroelectric

Geothermal

Nuclear
Radiant energy: energy in the form of radiation. Sunlight (photons) is an example of radiant energy.

Thermal energy: energy associated with the random (kinetic) motion of atoms and molecules.

Chemical energy: energy in stored within structure of molecules. This form of energy is released in chemical reactions.

Potential energy: energy associated with an object’s position.

Nuclear energy: energy stored within the collection of neutrons and protons in the atom.
The law of **energy conservation** is the most fundamental law of nature. Energy can be converted from one form to another, but not created or destroyed.

Ball A has a higher potential energy than ball B.

Potential energy of A has been converted to frictional heating and to the increase in the potential energy of B.
6.2 – Chemical Energy

CH\(_4\) (g) + 2O\(_2\) (g) → CO\(_2\) (g) + H\(_2\)O(g)

In an **exothermic reaction**, some of the energy stored in the chemical bonds is released as heat.
In an **endothermic reaction**, the energy of the products is higher than that of the reactants. Energy must come from an external source to make the reaction go.
Definition of a "System"

The **system** is the specific part of the universe that is of interest in a specific application. The system is surround by its **surroundings** – the reminder of the universe.

---

**open**

Exchange: mass & energy

**closed**

Exchange: energy

**isolated**

Exchange: nothing
6.3 – The 1st Law of Thermodynamics

The energy of the universe is constant

Energy lost by the system = energy gained by the surroundings

\[ \Delta E_{\text{system}} + \Delta E_{\text{surroundings}} = 0 \]

\[ \Delta E_{\text{system}} = -\Delta E_{\text{surroundings}} \]
The First Law of Thermodynamics

The internal energy \((E)\) of a system is defined as the sum of the kinetic and potential energies of all parts of the system. Note that chemical energy is a form of potential energy.

The internal energy \((E)\) can be changed by a transfer of heat \((q)\) and/or work \((w)\):

\[
\Delta E = q + w
\]

\(\Delta E\) 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
Heat ≠ Temperature!

Temperature is a measure of the random molecular motion in an object or substance.

\[(KE)_{avg} = \frac{3}{2} \, RT\]

Heat is a transfer of energy between two objects often driven by a temperature difference.

- \(q\) for meat = +\(E\)
- \(q\) for grill = -\(E\)
Work performed by or on the system is energy transfer that results in macroscopic changes in the system.

Mechanical work in the form of volume compression

Mechanical work in the form of macromolecular motion (protein kinesin dragging an organelle)

Electrochemical work in the form of electrochemiluminescence
Demo: Pickle Power!

Another unresolved problem for your generation to solve: why on earth does it glow?
State Functions

A state function depends only on the present state of the system and is independent on how the system got there.

Energy is a state function, but heat and work are not.

\[ \Delta E = E_{\text{final}} - E_{\text{initial}} \]

Other state functions are pressure, volume and temperature.

\[ \Delta P = P_{\text{final}} - P_{\text{initial}} \quad \Delta V = V_{\text{final}} - V_{\text{initial}} \quad \Delta T = T_{\text{final}} - T_{\text{initial}} \]
Heat and Work: Not State Functions

Heat and work applied to the system depend on the process or path taken. They cannot be extracted from the initial and final state of the system.

path 1
$q_1 + w_1$

path 2
$q_2 + w_2$

$\Delta E = q + w$

$E_i$

$E_f$
Thermodynamic Quantities

Thermodynamic quantities always have a number and a sign. The convention is: *if energy flows into the system, it has a positive sign because the system’s energy is increasing.*

![Diagram showing energy flow into and out of a system with signs: ΔE < 0 for energy flow out, ΔE > 0 for energy flow in.]

<table>
<thead>
<tr>
<th>Process</th>
<th>Sign</th>
</tr>
</thead>
<tbody>
<tr>
<td>Work done by the system on the surroundings</td>
<td>−</td>
</tr>
<tr>
<td>Work done on the system by the surroundings</td>
<td>+</td>
</tr>
<tr>
<td>Heat absorbed by the system from the surroundings (endothermic process)</td>
<td>+</td>
</tr>
<tr>
<td>Heat absorbed by the surroundings from the system (exothermic process)</td>
<td>−</td>
</tr>
</tbody>
</table>
Work can be done by a gas undergoing expansion (- $w$), or on a gas undergoing compression (+ $w$).

Mechanical work = force x distance

$$w = F \times \Delta h$$

$$w = P \times A \times \Delta h$$

$$w = - P \Delta V$$

1 L × atm = 101.3 J

Work can be done by a gas undergoing expansion (- $w$), or on a gas undergoing compression (+ $w$).

"$\Delta w = w_{\text{final}} - w_{\text{initial}}$" – is a meaningless equation
Making Snow

\[ \Delta E = q + w \]

\[ q \sim 0 \]

\[ w < 0, \Delta E < 0 \]
(work by the system on the surroundings)

\[ \Delta E = C \times \Delta T < 0 \]
(C = heat capacity > 0)

\[ \Delta T < 0 \]
WATER → SNOW!
The volume of an ideal gas is decreased from 5.0 L to 5.0 mL at a constant pressure of 2.0 atm. Calculate the work (in J) associated with this process. What sign does it have and why?

1. Use the equation for PV work at constant pressure.

For a gas at constant pressure, \( w = -P\Delta V \).

\[ P = 2.0 \text{ atm} \quad \Delta V = 0.0050 \text{ L} - 5.0 \text{ L} = -4.995 \text{ L} \]

\[ w = - (2.0 \text{ atm}) \times (-5.0 \text{ L}) = 10 \text{ L} \times \text{ atm} \]

2. Convert to Joules.

\[ 10 \text{ L} \times \text{ atm} \times \frac{101.3 \text{ J}}{1 \text{ L} \times \text{ atm}} = 1013 \text{ J} = 1.0 \text{ kJ} \]

Work is positive because it is done on the system and therefore it increases the energy of the system.
A balloon is expanded by heating the air inside it. In the final stages of the expansion, the volume of the balloon changes from $4.00 \times 10^6$ L to $4.50 \times 10^6$ L by the addition of $1.3 \times 10^8$ J of energy as heat. Assuming that the balloon expands against a constant pressure of 1.0 atm, calculate $\Delta E$ (J).

1. Write down the equation for $\Delta E$.

   $\Delta E = q + w$

2. Find $q$ and determine its sign.

   $q = 1.3 \times 10^8$ J

   The sign is positive because heat flows into the system
3. Find \( w \) and determine its sign.

\[
\begin{align*}
w &= -P \Delta V \\
\Delta V &= 4.50 \times 10^6 \text{ L} - 4.00 \times 10^6 \text{ L} = 5.00 \times 10^5 \text{ L}
\end{align*}
\]

\[
w = -1.0 \text{ atm} \times 5.00 \times 10^5 \text{ L} = -5.0 \times 10^5 \text{ L} \cdot \text{atm}
\]

\[
-5.0 \times 10^5 \text{ L} \times \text{atm} \times \frac{101.3 \text{ J}}{1 \text{ L atm}} = -5.1 \times 10^7 \text{ J}
\]

The sign is negative because the work is done by the system on the surroundings (air around the balloon).

4. Add \( q \) and \( w \) to get \( \Delta E \).

\[
\Delta E = q + w
\]

\[
\Delta E = 1.3 \times 10^8 \text{ J} - 5.1 \times 10^7 \text{ J} = 8.0 \times 10^7 \text{ J}
\]

\( \Delta E \) is positive, meaning that more energy was put into the system as heat than was expended as work.
Constant P vs Constant V Processes

If a process (a chemical reaction) occurs at a constant volume, then $\Delta V = 0$. Therefore:

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

The change in internal energy is equal to the heat exchange when $V$ is constant (hence the subscript "V")

This makes internal energy the most useful state function for the constant volume processes
6.4 – Enthalpy

Enthalpy ($H$) is defined as:

\[ H = E + P \times V \]

H is a state function because E, P, V are state functions as well.

**But why is it useful?**

For a process at constant pressure and involving only PV work:

\[ \Delta H = \Delta E + P \Delta V \]

By definition, \( \Delta E = q_P + w = q_P - P \Delta V \)

(where \( q_P \) = heat flow at constant pressure, hence the subscript)

Therefore, \( \Delta H = \Delta E + P \Delta V = q_P - P \Delta V + P \Delta V = q_P \)

\[ \Delta H = q_P \]

This makes **enthalpy** the most useful state function for the **constant pressure** processes (e.g., reaction in an open beaker)
Enthalpy of Reaction

For chemical reactions studied at a constant pressure, $\Delta H$ is also called the heat of reaction.

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

If sign of $\Delta H_{\text{rxn}}$ is positive: reaction is endothermic

If sign of $\Delta H_{\text{rxn}}$ is negative: reaction is exothermic
Thermochemical Equations

Thermochemical equations show the enthalpy changes as well as the mass relationships in a reaction.

\[
\text{CH}_4 (\text{g}) + 2 \text{O}_2 (\text{g}) \rightarrow \text{CO}_2 (\text{g}) + 2 \text{H}_2\text{O (gas)} \\
\Delta H_{\text{rxn}} = -802.4 \text{ kJ/mol}
\]

"mol" refers to "moles of reactions as written" (not moles of any specific reactant!)

The physical state of the substances must be explicitly indicated, because the total enthalpy change depends on the physical states of the products and reactants.

\[
\text{CH}_4 (\text{g}) + 2 \text{O}_2 (\text{g}) \rightarrow \text{CO}_2 (\text{g}) + 2 \text{H}_2\text{O (liquid)} \\
\Delta H_{\text{rxn}} = -890.4 \text{ kJ/mol} \text{ – different from above}
\]
Thermochemical Equations

By conservation of energy, the reverse equation is associated with a change of sign for $\Delta H_{rxn}$.

$$\text{CH}_4 \text{ (g)} + 2 \text{ O}_2 \text{ (g)} \rightarrow \text{CO}_2 \text{ (g)} + 2 \text{ H}_2\text{O (gas)}$$
$$\Delta H_{rxn} = -802.4 \text{ kJ/mol}$$

$$\text{CO}_2 \text{ (g)} + 2 \text{ H}_2\text{O (gas)} \rightarrow \text{CH}_4 \text{ (g)} + 2 \text{ O}_2 \text{ (g)}$$
$$\Delta H_{rxn} = +802.4 \text{ kJ/mol}$$

If you multiply both sides of the equation by a factor $n$, then $\Delta H_{rxn}$ must change by the same factor $n$.

$$2 \text{ CH}_4 \text{ (g)} + 4 \text{ O}_2 \text{ (g)} \rightarrow 2 \text{ CO}_2 \text{ (g)} + 4 \text{ H}_2\text{O (gas)}$$
$$\Delta H_{rxn} = 2 \times (-802.4 \text{ kJ/mol}) = -1605 \text{ kJ/mol}$$
The Hindenburg exploded in 1927 during a landing attempt. It contained $2.0 \times 10^8$ L of hydrogen at 1.0 atm and 25°C. Assuming that all of the hydrogen reacted at constant P, how much heat was produced by the explosion? The $\Delta H_{\text{rxn}}$ for this reaction is $-572$ kJ/mol ($= -286$ kJ per each mol of $\text{H}_2$ burned):

$$2\text{H}_2 (g) + \text{O}_2(g) \rightarrow 2\text{H}_2\text{O} (l)$$

1. Find the number of moles of $\text{H}_2$ that burned.

$$n = \frac{PV}{RT} = \frac{1.0 \text{ atm} \times 2.0 \times 10^8 \text{ L}}{0.08206 \text{ L atm / K mol} \times 298 \text{ K}} = 8.2 \times 10^6 \text{ mol H}_2$$

2. Calculate the amount of heat released.

$$q_P = \Delta H = 8.2 \times 10^6 \text{ mol H}_2 \times \frac{-286 \text{ kJ}}{1 \text{ mol H}_2} = -2.3 \times 10^9 \text{ kJ}$$

The sign of $\Delta H$ is negative: this reaction is \textit{exothermic}. 
Enthalpy and Energy

Consider the following reaction at constant pressure and 25 °C:

$$2 \text{Na (s)} + 2 \text{H}_2\text{O (l)} \rightarrow 2 \text{NaOH (aq)} + \text{H}_2 (g)$$

The change in internal energy:

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

$$\Delta H_{rxn} = -367.5 \text{ kJ (per each mol of } \text{H}_2 \text{ generated)}$$

The change in internal energy:

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

$$1 \text{ L} \times \text{atm} = 101.3 \text{ J}$$

$$\Delta E = -367.5 \text{ kJ/mol} - (1 \text{ atm} \times 24.5 \text{ L}) =$$

$$= -367.5 \text{ kJ/mol} - 2.5 \text{ kJ/mol} = -370.0 \text{ kJ/mol}$$

Internal energy change is *almost the same* as enthalpy change, because the $(PV)$ work performed is very small.
Gaseous carbonyl sulfide reacts according to the following thermochemical equation. When 405 g of COS is reacted, calculate the sign and magnitude (kJ) of the heat produced.

\[ 2\text{COS}(g) + 3\text{O}_2(g) \rightarrow 2\text{SO}_2(g) + 2\text{CO}_2(g) \quad \Delta H_{rxn} = -1092.0 \text{ kJ/mol} \]

1. Calculate the number of moles of COS consumed:

\[ n_{\text{COS}} = \frac{405 \text{ g}}{60.08 \text{ g/mol}} = 6.74 \text{ mol} \]

2. Calculate the heat produced in the reaction:

2 moles of reacted COS produce \( q = -1092.0 \text{ kJ} \) per above equation. Therefore, 6.74 mol should produce:

\[ q = - 6.74 \text{ mol COS} \times 1092.0 \text{ kJ/(2 mol COS)} \]

\[ = -3680 \text{ kJ} = -3.68 \times 10^3 \text{ kJ} \] (report in scientific format so that a correct number of significant digits can be included)
6.5 - Calorimetry

The **specific heat capacity** \((C_m)\) or simply **specific heat** of a substance is the amount of heat \((q)\) required to raise the temperature of 1 g of the substance by 1°C (= by 1 K).

The **heat capacity** \((C)\) of a substance is the amount of heat \((q)\) required to raise the temperature of a given quantity \((mass)\) of the substance by one 1°C (= by 1 K).

They are related: \[ C = mass \times C_m \]

They are useful for calculating heat \((q)\) absorbed or released by the substance:

\[ q = mass \times C_m \times \Delta T = C \times \Delta T \]

where \(\Delta T = T_{final} - T_{initial}\)

<table>
<thead>
<tr>
<th>Substance</th>
<th>Specific Heat (J/g · °C)</th>
</tr>
</thead>
<tbody>
<tr>
<td>Al</td>
<td>0.900</td>
</tr>
<tr>
<td>Au</td>
<td>0.129</td>
</tr>
<tr>
<td>C (graphite)</td>
<td>0.720</td>
</tr>
<tr>
<td>C (diamond)</td>
<td>0.502</td>
</tr>
<tr>
<td>Cu</td>
<td>0.385</td>
</tr>
<tr>
<td>Fe</td>
<td>0.444</td>
</tr>
<tr>
<td>Hg</td>
<td>0.139</td>
</tr>
<tr>
<td>H₂O</td>
<td>4.184</td>
</tr>
<tr>
<td>C₂H₅OH (ethanol)</td>
<td>2.46</td>
</tr>
</tbody>
</table>
Sample Problem

Calculate the heat capacity of 1.000 kg of liquid water and the amount of energy needed to warm water up by 10.00 °C

1. Calculate the heat capacity of 1 kg of water:

\[ C_m \text{ (water)} = 4.184 \text{ J} / (\text{g} \times \text{K}) \]
\[ C \text{ (1 kg water)} = 4.184 \text{ J} / (\text{g} \times \text{K}) \times 1000 \text{ g} = 4184 \text{ J} / \text{K} \]

2. Calculate the heat required to produce this temperature change

\[ q = C \times \Delta T = 4184 \text{ J} / \text{K} \times 10 \text{ K} = 41.84 \text{ kJ} \]
**Constant-Volume Calorimetry**

No heat enters or leaves!

\[
q_{\text{sys}} = q_{\text{water}} + q_{\text{bomb}} + q_{\text{rxn}}
\]

\[
q_{\text{sys}} = 0
\]

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

\[
q_{\text{water}} = \text{mass} \times C_m \times \Delta T
\]

\[
q_{\text{bomb}} = C_{\text{bomb}} \times \Delta T
\]

Draw back:
Reaction at constant \( V \)

\[
\Delta H \neq q_{\text{rxn}}
\]

However:

\[
\Delta H \sim q_{\text{rxn}}
\]
Constant-Pressure Calorimetry

No heat enters or leaves!

\[ q_{sys} = q_{water} + q_{cal} + q_{rxn} \]

\[ q_{sys} = 0 \]

\[ q_{rxn} = -(q_{water} + q_{cal}) \]

\[ q_{water} = mass \times C_m \times \Delta T \]

\[ q_{cal} = C_{cal} \times \Delta T \]

(usually quite small)

Reaction at constant \( P \)

\[ \Delta H = q_{rxn} \]
Hess’s Law

Change in enthalpy from reactants to products is independent of the reaction pathway taken.
Hess’s Law

Consider the oxidation of nitrogen to nitrogen dioxide. There are several ways to get from the reactants to the final product:

Path 1: direct

\[ \text{N}_2 (g) + 2\text{O}_2 (g) \rightarrow 2\text{NO}_2 (g) \quad \Delta H_1 = 68 \text{ kJ} \]

Path 2: indirect

\[ \text{N}_2 (g) + \text{O}_2 (g) \rightarrow 2\text{NO} (g) \quad \Delta H_2 = 180 \text{ kJ} \]
\[ 2\text{NO} (g) + \text{O}_2 (g) \rightarrow 2\text{NO}_2 (g) \quad \Delta H_3 = -112 \text{ kJ} \]

Net:

\[ \text{N}_2 (g) + 2\text{O}_2 (g) \rightarrow 2\text{NO}_2 (g) \quad \Delta H_2 + \Delta H_3 = 68 \text{ kJ} \]
Hess’s Law

The same reaction expressed with potential energy diagrams.

Hess’s law is frequently applied to calculate the enthalpy changes for unknown processes. For example, if $\Delta H_1$ and $\Delta H_2$ are known then $\Delta H_3$ can be calculated.
Sample Problem

Make use of the table of enthalpies of combustion to calculate \( \Delta H \) for the following reaction:

\[ \text{C}_4\text{H}_4 \text{ (g)} + 2\text{H}_2 \text{ (g)} \rightarrow \text{C}_4\text{H}_8 \text{ (g)} \]

| Combustion enthalpies | C\(_4\)H\(_4\) | -2341 kJ/mol | C\(_4\)H\(_8\) | -2755 kJ/mol | H\(_2\) | -286 kJ/mol |

1. Write down the sub-reactions and their enthalpies (kJ/mol).

By definition, \textit{combustion reaction} has O\(_2\) as a reactant and CO\(_2\) and H\(_2\)O as products

\[ \text{C}_4\text{H}_4 \text{ (g)} + 5\text{O}_2 \text{ (g)} \rightarrow 4\text{CO}_2 \text{ (g)} + 2\text{H}_2\text{O} \text{ (l)} \quad \Delta H_1 = -2341 \]

\[ \text{C}_4\text{H}_8 \text{ (g)} + 6\text{O}_2 \text{ (g)} \rightarrow 4\text{CO}_2 \text{ (g)} + 4\text{H}_2\text{O} \text{ (l)} \quad \Delta H_2 = -2755 \]

\[ \text{H}_2 \text{ (g)} + \frac{1}{2} \text{O}_2 \text{ (g)} \rightarrow \text{H}_2\text{O} \text{ (l)} \quad \Delta H_3 = -286 \]
2. Reorder the sub-reactions by bringing the reactants of the main equation to the left and the main products to the right. Multiply the stoichiometric coefficients and the corresponding enthalpies by integer factors if needed. The goal is to add the sub-equations and get the equation we want.

In this case, we are inverting the second equation, and multiplying the third equation by 2.

\[
\begin{align*}
\text{C}_4\text{H}_4 \text{ (g)} + 5\text{O}_2 \text{ (g)} & \rightarrow 4\text{CO}_2 \text{ (g)} + 2\text{H}_2\text{O} \text{ (l)} & \Delta H_1 = -2341 \\
4\text{CO}_2 \text{ (g)} + 4\text{H}_2\text{O} \text{ (l)} & \rightarrow \text{C}_4\text{H}_8 \text{ (g)} + 6\text{O}_2 \text{ (g)} & \Delta H_2 = -(2755) \\
2\text{H}_2 \text{ (g)} + \text{O}_2 \text{ (g)} & \rightarrow 2\text{H}_2\text{O} \text{ (l)} & \Delta H_3 = 2 \times (-286) \\
\hline
\text{C}_4\text{H}_4 \text{ (g)} + 2\text{H}_2 \text{ (g)} & \rightarrow \text{C}_4\text{H}_8 \text{ (g)} & \Delta H = \Delta H_1 + \Delta H_2 + \Delta H_3 \\
& & = -158 \text{ kJ/mol}
\end{align*}
\]

3. Add the equations
Sample Problem

Make use of this table of reaction enthalpies to calculate $\Delta H$ for the following reaction:

$$\text{C}_6\text{H}_4(\text{OH})_2(\text{aq}) + \text{H}_2\text{O}_2(\text{aq}) \rightarrow \text{C}_6\text{H}_4\text{O}_2(\text{aq}) + 2\text{H}_2\text{O}(\text{l})$$

1. Reorder the sub-reactions with reactants to the left and products to the right, and correct for stoichiometries.

<table>
<thead>
<tr>
<th>Reaction</th>
<th>$\Delta H$ (kJ)</th>
</tr>
</thead>
<tbody>
<tr>
<td>$\text{C}_6\text{H}_4(\text{OH})_2(\text{aq}) \rightarrow \text{C}_6\text{H}_4\text{O}_2(\text{aq}) + \text{H}_2(\text{g})$</td>
<td>$+177.4$</td>
</tr>
<tr>
<td>$\text{H}_2(\text{g}) + \text{O}_2(\text{g}) \rightarrow \text{H}_2\text{O}_2(\text{aq})$</td>
<td>$-191.2$</td>
</tr>
<tr>
<td>$\text{H}_2(\text{g}) + \frac{1}{2}\text{O}_2(\text{g}) \rightarrow \text{H}_2\text{O}(\text{g})$</td>
<td>$-241.8$</td>
</tr>
<tr>
<td>$\text{H}_2\text{O}(\text{g}) \rightarrow \text{H}_2\text{O}(\text{l})$</td>
<td>$-43.8$</td>
</tr>
</tbody>
</table>

Total $\Delta H = +177.4 - 191.2 - 241.8 - 43.8 = -210.4$ kJ
Hess’s Law

\[ \Delta H_1 = 177.4 \text{ kJ} \]

\[ \Delta H_2 = 191.2 \text{ kJ} \]

\[ \Delta H_3 = -483.6 \text{ kJ} \]

\[ \Delta H_4 = -87.6 \text{ kJ} \]

\[ \Delta H_{\text{reaction}} \text{ is independent of the pathway taken} \]
2. Calculate the total enthalpy by adding up the sub-reactions.

\[ \Delta H = \Delta H_1 + \Delta H_2 + \Delta H_3 + \Delta H_4 = \]

\[ = 177.4 + 191.2 - 483.6 - 87.6 = -202.6 \text{ kJ} \]

The total reaction is exothermic
6.7 – Enthalpy of Dissolution

The enthalpy of solution ($\Delta H_{\text{solv}}$) is the heat generated or absorbed when a certain amount of solute dissolves in a certain amount of solvent.

$$\Delta H_{\text{soln}} = H_{\text{mixed solution}} - H_{\text{separated solute and solvent}}$$

Negative $\Delta H_{\text{solv}}$ – the solution becomes heated upon dissolution of a compound ($H_2SO_4$ is a good example)

<table>
<thead>
<tr>
<th>Compound</th>
<th>$\Delta H_{\text{solv}}$ (kJ/mol)</th>
</tr>
</thead>
<tbody>
<tr>
<td>LiCl</td>
<td>$-37.1$</td>
</tr>
<tr>
<td>CaCl$_2$</td>
<td>$-82.8$</td>
</tr>
<tr>
<td>NaCl</td>
<td>$4.0$</td>
</tr>
<tr>
<td>KCl</td>
<td>$17.2$</td>
</tr>
<tr>
<td>NH$_4$Cl</td>
<td>$15.2$</td>
</tr>
<tr>
<td>NH$_4$NO$_3$</td>
<td>$26.2$</td>
</tr>
</tbody>
</table>

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

Which substance(s) could be used for a cold pack?
Dissolution and Hess’s Law

Lattice energy: energy needed to convert an ionic solid to gas phase ions

Hydration energy: energy released when a gas-phase ion is placed in water

\[ \Delta H_{\text{soln}} = \text{Step 1} + \text{Step 2} = 788 - 784 = 4 \text{ kJ/mol} \]
6.6 - Standard Enthalpy of Formation

\[ \Delta H_f^o = \text{change in enthalpy occurring during the formation of one mole of a compound from its elements with all substances in their standard states. [kJ/mol]} \]

\[ \text{element 1 + element 2 + ...} \rightarrow \text{compound} \]

\[ \text{Ag (s) + 1/2 Cl}_2 \text{ (g)} \rightarrow \text{AgCl (s)} \]

\[ \Delta H_f^o = -127.0 \text{ kJ} \]
Standard States

For an element:

The most stable state in which the element exists under conditions of $P = 1 \text{ atm}$ and $T = 25 \degree \text{C}$

For a compound:

Gaseous substance: $P = 1 \text{ atm}$ and $T = 25 \degree \text{C}$

Liquid or solid substance: pure liquid or solid

Substance in solution: concentration of exactly 1 M.
Standard States

If there more than one form of the element under the standard conditions, use the most stable form.

- Oxygen (g) can exist as O₂ or as O₃ (ozone) at 1 atm and 298 K.
  
  Most stable form = O₂ (g) \[ \Delta H_f^o = 0 \text{ kJ/mol} \]
  
  Less stable form = O₃ (g) \[ \Delta H_f^o = +142.7 \text{ kJ/mol} \]

- Carbon (s) can exist as graphite or as diamond at 1 atm and 298 K.

  Most stable form = graphite \[ \Delta H_f^o = 0 \text{ kJ/mol} \]

  Less stable form = diamond \[ \Delta H_f^o = +1.88 \text{ kJ/mol} \]
Gray tin and white tin are two solid forms of tin. The denser white metallic form is the most stable phase above 13°C, and the powdery gray form is more stable below 13°C.

Darn cold Russian winters… Darn thermo-chemistry….
The standard enthalpy of reaction ($\Delta H_{\text{rxn}}^0$) is the enthalpy of a reaction carried out at 1 atm and 298 K.

It can be calculated from the standard enthalpies of formation using Hess’s law:

$$ aA + bB \rightarrow cC + dD $$

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

$$ \Delta H_{\text{rxn}}^0 = \Sigma n\Delta H_f^0 (\text{products}) - \Sigma m\Delta H_f^0 (\text{reactants}) $$
Calculate the standard enthalpy of formation of CS$_2$ (l) given that:

$$\text{C(graphite)} + \text{O}_2 \ (g) \rightarrow \text{CO}_2 \ (g) \quad \Delta H_{\text{rxn}}^0 = -393.5 \ \text{kJ/mol}$$

$$\text{S(rhombic)} + \text{O}_2 \ (g) \rightarrow \text{SO}_2 \ (g) \quad \Delta H_{\text{rxn}}^0 = -296.1 \ \text{kJ/mol}$$

$$\text{CS}_2(l) + 3\text{O}_2 \ (g) \rightarrow \text{CO}_2 \ (g) + 2\text{SO}_2 \ (g) \quad \Delta H_{\text{rxn}}^0 = -1072 \ \text{kJ/mol}$$

1. Write the balanced standard formation reaction for CS$_2$

$$\text{C(graphite)} + 2\text{S(rhombic)} \rightarrow \text{CS}_2 \ (l)$$

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

$$\text{C(graphite)} + \text{O}_2 \ (g) \rightarrow \text{CO}_2 \ (g) \quad \Delta H_{\text{rxn}}^0 = -393.5 \ \text{kJ/mol}$$

$$2\text{S(rhombic)} + 2\text{O}_2 \ (g) \rightarrow 2\text{SO}_2 \ (g) \quad \Delta H_{\text{rxn}}^0 = -296.1 \ \text{kJ/mol} \times 2$$

$$\text{CO}_2(g) + 2\text{SO}_2 \ (g) \rightarrow \text{CS}_2 \ (l) + 3\text{O}_2 \ (g) \quad \Delta H_{\text{rxn}}^0 = +1072 \ \text{kJ/mol}$$

$$\text{C(graphite)} + 2\text{S(rhombic)} \rightarrow \text{CS}_2 \ (l)$$

$$\Delta H_{\text{rxn}}^0 = -393.5 + (2 \times -296.1) + 1072 = 86.3 \ \text{kJ/mol}$$
Benzene (C₆H₆) burns in air to produce carbon dioxide and liquid water. How much heat is released per mole of benzene combusted? The standard enthalpies of formation of benzene, CO₂, and H₂O are 49.04, - 393.5, and – 187.6 kJ/mol, respectively.

\[
2\text{C}_6\text{H}_6 (l) + 15\text{O}_2 (g) \rightarrow 12\text{CO}_2 (g) + 6\text{H}_2\text{O} (l)
\]

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

\[
\Delta H^0_{\text{rxn}} = [12\Delta H^0_f (\text{CO}_2) + 6\Delta H^0_f (\text{H}_2\text{O})] - [2\Delta H^0_f (\text{C}_6\text{H}_6)]
\]

\[
\Delta H^0_{\text{rxn}} = \begin{bmatrix} 12\times(-393.5) + 6\times(-187.6) \\ - 2\times49.04 \end{bmatrix} = -5946 \text{ kJ}
\]

Each reaction consumes 2 moles of benzene, therefore each mole of benzene will produce:

\[
\frac{-5946 \text{ kJ}}{2 \text{ mol}} = -2973 \text{ kJ/mol C}_6\text{H}_6
\]