# Bond Enthalpy & Enthalpy of Reaction

**Unit:** Thermodynamics

### Skills:

* Utilize Latent Heat and Heat Capacity equations
* calculate ΔHrxn from bond enthalpies
* calculate ΔHrxn from  data

### Notes:

Last time, we determined a method for measuring the heat energy absorbed or given off by a system using calorimetry.

We were introduced to heat (q) equations to measure energy flow between the reaction and the water of the calorimeter:

1. **Latent Heat:** The amount of heat energy required to push a substance through a phase change (overcome intermolecular forces)

q = m∆H

∆H can be:

∆Hf - heat of fusion (energy required to melt 1 g of ice)

= 334 J/g or 79.9 cal/g

∆Hv - heat of vaporization (energy required to boil 1 g of water)

= 2260 J/g or 540 cal/g

1. **Heat Capacity:** The amount of energy required to increase the temperature of 1 g of a substance 1°C

q = mCp∆T

C = the specific heat capacity of that substance

Cice = 2.05 J/g°C or 0.5 cal/g°C

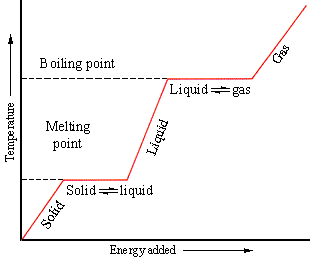
Cwater = 4.18 J/g°C or 1 cal/g°C

Csteam = 2.01 J/g°C or 0.48 cal/g°C



So, if you have to change the temperature of a substance AND go through a phase change, you would need to use the latent heat equation AND the heat capacity equation. In other words, each section of a heating/cooling curve has a “unique” equation to describe the energy needed to go through it.

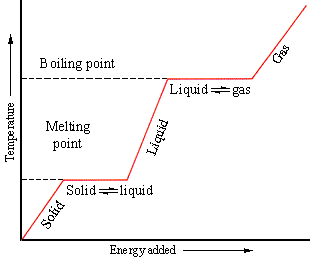
Ex: calculate the amount of energy required to increase the temperature of ice from -10°C to 35°C.



Notice the slopes on the heating cooling curve.

* What do they represent?
* How is the added thermal energy being used?
* Which equation describes this thermal energy flow?

Now investigate the plateaus on the heating cooling curve.

* What do they represent?
* How is the added thermal energy being used?
* Which equation describes this thermal energy flow?

We can also describe the energy released or absorbed during chemical reactions by describing the energy held within the chemical bonds of the molecules/compounds.

chemical bond: when two atoms are joined through the sharing of electrons. (We can say this because no bond is 100% ionic—even CsF has a small amount of covalent character.)

Any chemical bond that actually forms is *always* more stable than the corresponding unbonded atoms. If the bond were less stable than the unbonded atoms, the bond would fail to form.

* Where does this stability come from?

enthalpy: stored (potential) chemical energy in the form of heat. In thermodynamics, we use the variable *H* to represent enthalpy.

state function: a function (variable) whose value depends only on the current state of the system, and whose value is independent of the pathway. Enthalpy is an example of a state function.

In thermodynamics, variables representing state functions are represented by capital letters. (G, H, S)

## Bond Dissociation Energies

“Bond energy” is the energy required to break a chemical bond. Generally, bond dissociation energies are for homolytic dissociation (*i.e.,* the electrons are equally split between the two atoms). Heterolytic bond dissociation energies are always higher.



Here are some common heterolytic bond dissociation energies:

|  |  |
| --- | --- |
| Bond | Bond Dissociation Energy |
| C−C | 346 |
| C=C | 602 |
| C≡C | 835 |
| C−H | 411 |
| C−O | 358 |
| C=O | 799 |
| O=O | 494 |
| O−H | 459 |

You can use bond dissociation energies to determine the enthalpy change of a chemical reaction. For example, consider the combustion of methane:

CH4 + 2 O2 → CO2 + 2 H2O

This actually means:

 + 2  →  + 2 

break 4 C−H bonds + 2 O=O bonds → form 2 C=O bonds + 4 O−H bonds

By convention, energy is positive if we put it in (endothermic), and negative if it comes out (exothermic). This means the enthalpy needed to break a bond is positive, and the enthalpy released by forming a bond is negative.

Adding the energies of all of the bonds broken and formed in the equation gives the following:



## Enthalpy of Formation

The enthalpy of formation () of a compound is the Δ*H*rxn for the formation of a compound from its elements in their natural state. For Δ*Hf* values measured/calculated at “standard state” (usually defined to be 1 atm and 25°C), we add a superscript °.

The enthalpy of formation of any pure element in its natural state is defined to be zero.

For example:

2 Al (s) + 3 O2 (g) → Al2O3 (s) 

Because  for Al (s) is 0, and  for O2 (g) is also 0, this means

 for Al2O3 (s) = 

## Calculating Δ*H*rxn from Enthalpy of Formation Data

Rather than calculate  values from bond energies, it is much more convenient to look them up in a table. Thus, for the reaction:

CH4 (g) + 2 O2 (g) → CO2 (g) + 2 H2O (g)

we could simply look up the  values for each of the compounds:

|  |  |
| --- | --- |
| Compound |  |
| CH4 (g) | −74.8 |
| O2 (g) | 0 |
| CO2 (g) | −393.5 |
| H2O (g) | −241.8 |

Then:

