# Heat & Thermochemistry

**Unit 8A:** Thermodynamics

### Knowledge/Understanding:

* heat as energy
* principles of heat flow
* state functions/variables
* calorimetry

### Skills:

* solve calorimetry (specific heat) problems

### Notes:

Heat, just like electromagnetic waves, is a form of energy. During chemical reactions, energy is typically transferred in the form of heat.

system: the region being considered in a problem.

surroundings: everything that is outside of the system.

E.g., if a chemical reaction takes place in a 10 L tank, the system is everything inside the tank, and the surroundings are the physical walls of the tank and everything outside of them.

We will use the variable *q* to represent the heat that is being transferred during a chemical reaction.

### Heat vs Temperature

Heat and temperature are often used interchangeably in everyday conversation by muggles, which is inaccurate.

Heat: The ***total*** amount of thermal (kinetic) energy held by the

system. Units = J, kJ, or cal.

Temperature: The ***average*** thermal (kinetic) energy held by the

molecules of the system. Units = ⁰C or K

So, when we start talking about the “heat of formation” during the breaking/formation of bonds, we are talking about a summative quantity of energy (J or cal).

### Sign of Heat Flow

Energy transfer is always in relation to the system, meaning the chemical reaction that is taking place, or more specifically, the molecules involved in the reaction.

Endothermic: heat energy is taken in by the system (feels cold to you,

the surroundings)

Exothermic: heat energy is released by the system (feels hot to you,

the surroundings)

“Positive” heat (+q) means heat flows in to the reaction. Because the heat energy is coming into of the reaction, it is taken from the surroundings’ molecules, reducing their kinetic energy. *I.e., positive heat of reaction means heat is being pulled out of the surroundings, and the temperature goes* ***down.*** (the system would feel cold because it is taking in energy from your hand).

“Negative” heat (-q) means heat flows out of the reaction. Because the heat energy is coming out of the reaction, it is absorbed by the surroundings’ molecules, adding to their kinetic energy. *I.e., negative heat of reaction means heat is being released into the surroundings, and the temperature goes* ***up****.* (the system would feel warm because it is giving energy to your hand).

|  |  |  |  |  |
| --- | --- | --- | --- | --- |
| **Heat Flow** | **Sign of *q*** | **Type of Reaction** | **Products** | **Surroundings** |
| from surroundings to system | + (positive) | endothermic | gain heat | lose heat (get colder) |
| from system to surroundings | − (negative) | exothermic | lose heat | gain heat (get hotter) |

## So what can we learn from heat/temperature?

Now that we can relate the concepts of heat energy and temperature, we can quantify any changes in energy that occur within a system by measuring changes in temperature.

* Measuring the change in temperature of a solution during dissolving or other physical changes should tell us whether or not and how much energy was required/released for that process to occur.
* Measuring the temperature change of a system due to a chemical reaction should tell us the amount of energy given off or absorbed during the reaction.

In order to quantify heat energy values, we must be able to correlate the temperature (T) and heat energy (q) of a system/surroundings.

*q* = *m* *Cp* Δ*T*

*q* = heat (J) *m* = mass (g)

*Cp*= heat capacity 

Δ*T*= temperature change (°C)

**Specific Heat Capacity**

specific heat capacity: a measure of the amount of heat energy required to raise the temperature of 1 gram of a substance 1ºC.

*Cp*: specific heat capacity measured at constant pressure. For gases, this means the measurement was taken allowing the gas to expand as it was heated.

*Cv*: specific heat capacity measured at constant volume. For gases, this means the measurement was made in a sealed container, allowing the pressure to rise as the gas was heated.

For solids and liquids, *Cp* ≈ *Cv*. For gases, *Cp* > *Cv* (always). For ideal gases, *Cp* − *Cv* = R (the gas constant).

When there is a choice, *Cp* is more commonly used than *Cv* because it is easier to measure.

*q* = *m* *Cp* Δ*T*  \*

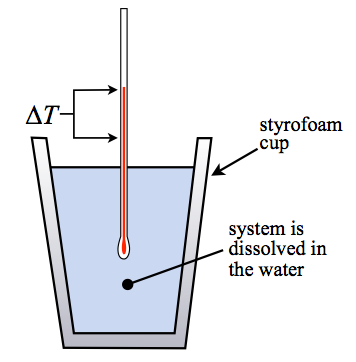
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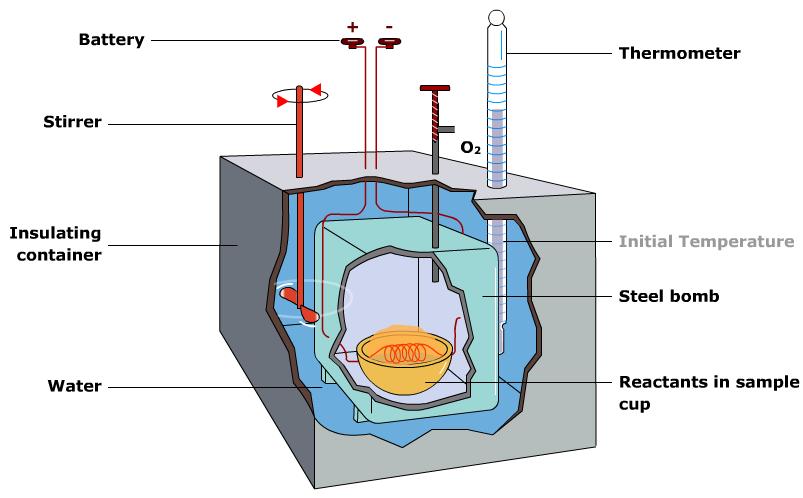
## \*This is the equation provided on the AP testCalorimetry

calorimetry: the measurement of heat flow

calorimeter: an insulated container for performing reactions that involve measuring heat.

coffee cup calorimeter: a calorimeter that is only an insulated container—often containing your system (reactants) dissolved directly in water. Can be made from a styrofoam coffee cup.

bomb calorimeter: a calorimeter for measuring the heat produced by a chemical reaction. The calorimeter contains a mass of water. The heat from the reaction makes the temperature of the water change.



In a calorimeter, the heat of reaction is equal and opposite to the heat of the calorimeter. This is because any heat coming *out of* the reaction is going *into* the calorimeter system. (Assuming no energy loss from the closed system)

*q*rxn = *q*sys =-(qcal + qwater)

Most calorimeters are calibrated with a published heat capacity for the calorimeter Ccal, measured in , to account for energy lost to the calorimeter due to imperfect insulation. This means:

*q*cal = *C*cal Δ*T*

**Example Calorimetry Problem: Melting ice**

Determine an experimental value for the heat of fusion of ice (heat energy required to melt 1 gram of ice).

1. Assemble a coffee cup calorimeter.
2. Add 100.0g of di-water to your calorimeter. Record the initial temperature. (Cwater = 4.18 j/gºC or 1 cal/ºC)
3. Mass out ~3 cubes of ice and place it into the calorimeter containing 100.0 g of water. Insert a temperature probe and record the lowest temperature obtained once all of the ice has melted.
4. What is the experimental value for the specific heat of fusion of ice?

Qcalorimeter = mwater•Cwater•ΔT

Qice = mice•ΔHfusion-ice

1. Draw an energy flow diagram for your calorimeter experiment.

The basis for the solution to this problem is the recognition that the quantity of energy lost by the water when cooling is equal and opposite to the quantity of energy required to melt the ice. In equation form, this could be stated as

**qice = -qwater/calorimeter**

* The positive qice indicates that the system is gaining energy
* The negative qwater/calorimeter indicates the surroundings are losing energy

Actual heat of fusion (∆H) for ice is 79.9 cal/ºC or 334 J/gºC, how close were you?

What are some assumptions/possible sources of error?

* Did you have a closed system?
* Was the water the only thing losing energy in the calorimeter?
  + qsurroundings = qcup + qwater