Energy

Energy is the capacity to do work (w) or to produce heat (q):

$$\Delta U = q + w$$

- we cannot measure the absolute energy, only the change of energy in a system
- work is the energy transfer due to an applied force
- heat is energy transfer due to a difference in temperature (anything other than work)

Systems

Some common terms in thermodynamics:

- System: what we wish to investigate (such as a chemical reaction, for example)
- Surroundings: everything else
- Universe: the sum of the system and its surroundings

Types of Systems

1. **Open:** allows for the transfer of both matter and energy from the system to the surroundings.

2. **Closed:** allows the transfer of energy but not matter from the system to the surroundings.

3. **Isolated:** neither matter nor energy can be transferred from the system to the surroundings.

4. **Adiabatic:** allows the transfer of work to and from the system, but not the transfer of heat or matter (no loss or gain of energy through heat).

1st Law of Thermodynamics

The 1st Law of Thermodynamics states that energy cannot be created nor destroyed, but instead it is conserved through the transfer of heat and work from one form to another. From this, we can derive some important equations:

$$\Delta U_{\text{isolated system}} = 0$$

$$\implies \Delta U_{sys} = -\Delta U_{surr}$$

$$\implies \Delta U_{univ} = \Delta U_{sys} + \Delta U_{surr} = 0$$

In other words, the total change in a systems internal energy is the sum of the energy transferred as heat and work.

$$\Delta U = \Delta E_{sys} = q + w$$

Internal Energy

This internal energy (U) consists of:

- translational energy (motion)
- energy stored in the bonds (rotational, vibrational, electronic)

- energy stored in intermolecular forces

Internal energy is an extensive property.

Sign Convention

TABLE 5.1 The Sign Conventions* for *q*, *w*, and ΔU

q	+	w	=	ΔU
+		+		+
+		−		Depends on the sizes of q and w
−		+		Depends on the sizes of q and w
−		−		−

*For q: + means system *absorbs* heat; - means system *releases* heat.
For w: + means work done *on* system; - means work done *by* system.

Heat

Heat is an extensive property, meaning it depends on the amount of the substance. The quantity of heat depends on:

- change of temperature
- amount of the substance
- nature of the substance

Heat capacity is the amount of heat required to change the temperature of system or substance by 1°. There are a few variations of heat capacity:

1. **Heat Capacity:**
    - units of $J/^\circ C$
    - intensive property
2. **Specific Heat Capacity**
    - denoted by c
    - units of $J/(g \cdot ^\circ C)$
    - extensive property
3. **Molar Heat Capacity**
    - denoted by C
    - units of $J/(mol \cdot ^\circ C)$
    - extensive property

Calorimetry

A calorimeter is an isolated system where no energy or matter is exchanged with surroundings. It is used to measure the change in heat during a reaction.

$$q = mc\Delta T$$

- q is the heat lost or gained
- c is the specific heat capacity of the substance
- m is the mass of the substance (in grams)
- ΔT is the change in the temperature

You can also use moles instead of mass if you replace the specific heat capacity with molar heat capacity:

$$q = nC\Delta T$$

Bomb Calorimetry

In our labs, we use a styrofoam cup and a lid as a calorimeter. More accurate calorimeters can be made using an insulated steel container. Another type of calorimeter is the bomb calorimeter, where a substance is ignited and the change in heat is measured.