Section 1.1.2 – Heat Transfer

Heat

Count Rumford stated in the 18th century, that **heat was a form of energy**. He further stated that:

"Hot objects had heat energy, the same way that a moving object has kinetic energy or an elevated object has gravitational potential energy."

Recall that energy and work are measured in the SI unit of the **Joule (J).**
1 Joule is the energy expended when a force of 1 Newton is applied through a distance of 1 metre.

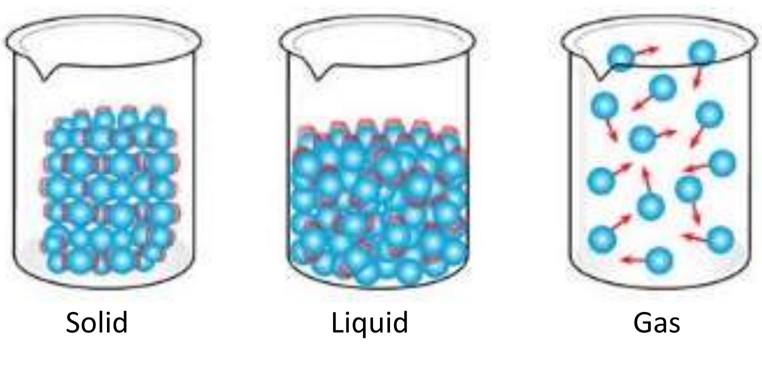
NB: 1 Joule of heat energy would raise the temperature of 1 gram of dry air by 1 °C

Heating refers to a particular type of energy transfer which occurs in response to a temperature difference. The particular form of energy involved is called **internal energy.** During heat transfer between two objects, the internal energy of the hotter object decreases, while the internal energy of the colder object increases. When the energy is in transit it is often referred to as heat.

The Kinetic Theory of Matter

In order to understand more complex problems associated with heat and energy transfer within a system, one needs to first understand the **kinetic theory of matter**.

The kinetic theory of matter states that "***all matter is made of small particles that are in random motion and that have space between them***". This means that no matter what phase matter is in, it is made of separate, moving particles.



In solids, particles are closely packed in a regular arrangement. The particles vibrate about a fixed position.

In liquids, particles are closely packed in a random arrangement. The particles can move through the liquid but they cling together.

In gases, the particles are far apart. There motion is random and independent of the other particles. Particles have much more kinetic energy and therefore move quicker.

The kinetic theory of matter can explain several properties associated with each state:

Solids

Have a fixed shape and cannot flow	The particles cannot move from place to place
They cannot be compressed or squashed	The particles are close together and have no space to move into

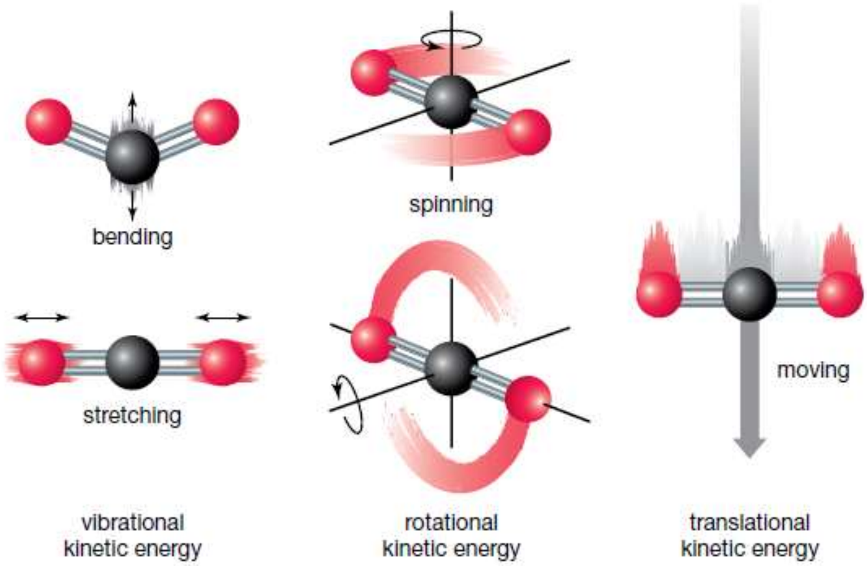
Liquids

They flow and take the shape of their container	The particles can move around each other
They cannot be compressed or squashed	The particles are close together and have no space to move into

Gases

They flow and completely fill their container	The particles can move quickly in all directions
They can be compressed or squashed	The particles are far apart and have space to move into

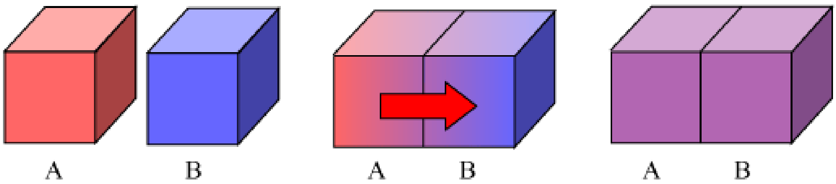
Consider the numerous different forms of molecular kinetic energy available to gases.



Temperature is the measure of the **average translational kinetic energy** of particles. The other modes of kinetic energy contribute to the internal energy but not the temperature. Temperature provides a measure of the "**hotness**" of an object.

Thermal Equilibrium

Energy is always transferred **from a region of high temperature to a region of lower temperature**. This process occurs until both regions eventually reach the same common temperature. At which point a state of **thermal equilibrium** is reached.



(a) **(b)** **(c)**

(a) Two bodies A and B are of different temperatures, the temperature of A is higher than that of B. (b) When they are in contact, heat is transferred from A to B. (c) Heat transfer will stop when both A and B reach an equilibrium temperature.

Example 1

When ice blocks are added to a room temperature drink, energy from the drink is actually transferred to the ice blocks. In doing so the drink is cooled. Eventually the ice blocks melt and the combined solution reaches a constant temperature. At which point the system is said to be in thermal equilibrium.

Example 2

When a red hot horseshoe is quenched in a bucket of water, some of the kinetic energy from the horseshoe is transferred to the water. Eventually the translational kinetic energy of both the horseshoe and the water become the same. At which point the system is said to be in thermal equilibrium.

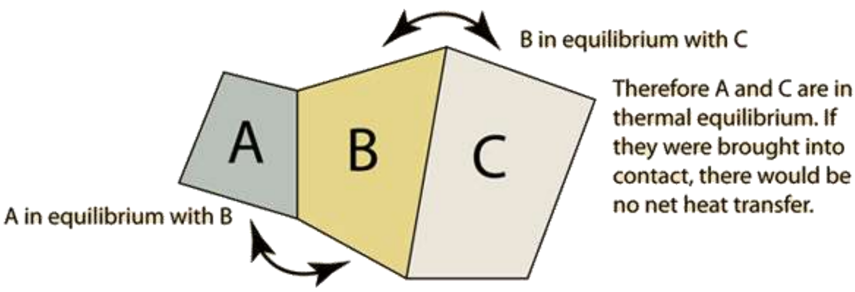
NB: the final temperature at thermal equilibrium depends upon the initial temperature of the two items, the amount of each material and the properties of the materials from which both items are made.

The Laws of Thermodynamics

"***Thermodynamics is a branch of physics concerned with heat and temperature and their relation to energy and work***."

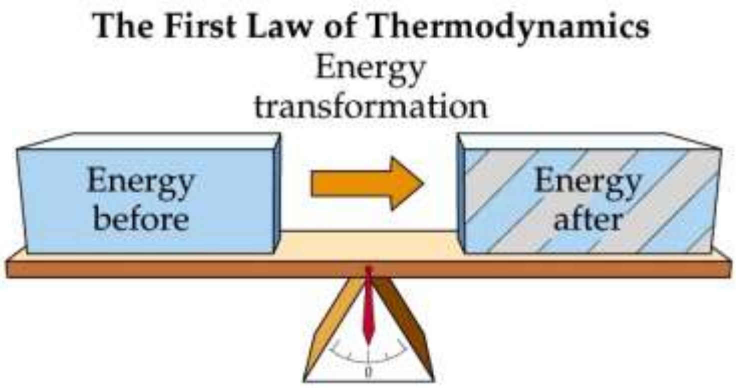
Historically there were three laws of thermodynamics established during the 19th century. During the 20th century an additional law, associated with thermal equilibrium was also established. This additional law was fundamental to the pre-existing other three laws. Accordingly it was called the "Zeroth Law of Thermodynamics"

Zeroth Law of Thermodynamics



The Zeroth Law of Thermodynamics states that "***if two bodies are each in thermal equilibrium with some third body, then they are also in equilibrium with each other***". Accordingly all three bodies are all the same temperature.

First Law of Thermodynamics



The First Law of Thermodynamics states that "***heat is a form of energy, and heat energy cannot be created or destroyed. It can, however, be transferred from one location to another and converted to and from other forms of energy.***"

The following equation is used to quantify the change of internal energy of a system. It takes into consideration any heat that has either been absorbed or released. It also incorporates if work is done by the system on the surroundings or if work is done on the system by the surroundings.

$$\Delta U = Q - W$$

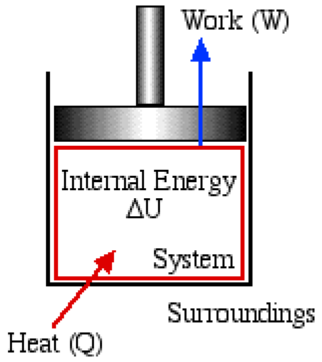
Where ΔU represents the change in internal energy of a system
Q represents the heat added **to** the system
W represents the work done **by** the system

NB: The sign allocated to the heat (Q) and the work (W) is very important.

Positive Q (Q > 0) represents heat transferred **into the system**
Negative Q (Q < 0) represents heat transferred **out of the system**

Positive W (W > 0) represents the system does **work on the surroundings**
Negative W (W < 0) represents the surroundings doing **work on the system**

Example 3



Consider the piston system show left of page.

It consists of a sample of gas that is contained within the walls of the cylinder and the moveable piston head.

Scenario 1 If the piston head is fixed and the cylinder heated
Heat has been transferred into the system (Q > 0)
No work is done by the system on the suroundings and vice versa (W = 0)

∴ There is a net gain in internal energy (ΔU ↑) because ΔU = Q – W

Scenario 2 If the piston head is fixed and the cylinder cooled
Heat has been transferred out of the system (Q < 0)
No work is done by the system on the suroundings and vice versa (W = 0)

∴ There is a net reduction in internal energy (ΔU ↓) because ΔU = Q – W

Scenario 3 If the piston head is free to move and the cylinder heated
Heat has been transferred into the system (Q > 0)
Work is done by the system on the suroundings (W > 0)

∴ There is a no net change in internal energy (ΔU remains constant) because ΔU = Q – W

Example 4

Consider our system to be a warm cup of coffee

How can a system have heat added to it?

- add some extra hot water into the cup
- place the cup in front of a heater
- Place the cup in a microwave

How can a system have work done on it?

- stir the cup with a spoon (aggressively)
- place a lid on the cup and shake it

By adding heat to the system or by doing work on the system the **internal energy of the system is increased**.

How can a system have heat removed from it?

- add ice cubes into the cup
- blow across the top of the cup

How can a system do work on its surroundings?

- the coffee "pops" the lid off the cup

By removing heat from the system or by the system doing work on its surroundings the **internal energy of the system is reduced**.

Example 5

45 Joules of heat energy is added to a system. The system then does 16 Joules of work. How much did the internal energy of the system change?

ΔU = ?
Q = +45 J (+ve as heat is transferred **into the system**)
W = +16 J (+ve as work is **done by** the system on the surroundings)

$\Delta U = Q - W$
= 45 -16
= **<u>29 J</u>** The system increased its internal energy by 29 Joules

Example 6

James sits on top of a helium filled balloon. In doing so the balloon receives 200 Joules of heat energy. The pressure placed upon the balloon does a further 50 Joules of work on the helium filled balloon. How much did the internal energy of the system change?

ΔU = ?
Q = +200 J (+ve as heat is transferred **into the system**)
W = -50 J (-ve as work is **done on** the system by the surroundings)

$\Delta U = Q - W$
= 200 – (-50)
= **<u>250 J</u>** The system increased its internal energy by 250 Joules

NB: As the internal energy increased in both examples, so too did the temperature of both systems.