Chapter 14 - Heat

The temperature of an object is a measure of the amount of motion the molecules of that object have; the higher the temperature, the larger the amount of motion, and the higher the internal kinetic energy of the molecules.

Heat $^{(Q)}$ is the transfer of internal energy from one object (hot) to another (cold).

An example is putting ice cubes in a drink: the hot drink transfers its internal energy to the cooler ice cubes. As a result, the temperature of the hot object goes down, and the temperature of the cooler object (may) go up (if there is no phase change)

The amount that the temperature goes up (or down) in response to gaining (or losing) heat energy is:

$$Q = mc\Delta T \qquad \text{solid/liquids}$$

$$= nC_V\Delta T \qquad \text{gases}$$

where m is the object's mass and c is the specific heat, different for different kinds of materials.

In general, it's not too convenient to talk about the <u>mass</u> of a gas, so it can be easier to use that second expression in terms of number of moles, n. In this case, c_V is the <u>molar specific heat</u>.

For a monatomic gas (e.g. Ar, Ne, etc)

$$c_V = \frac{3}{2}R$$

For a diatomic gas (eg. O_2, N_2, etc)

$$c_V = \frac{5}{2}R$$

This has to do with the number of different ways the molecules can move, or their number of "degrees of freedom".

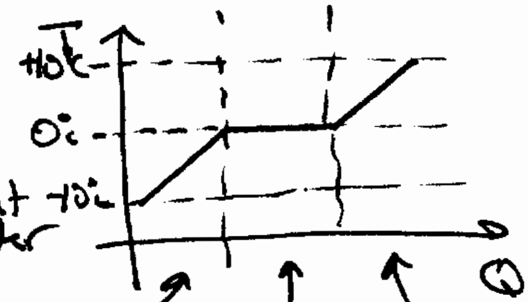
It can also be possible to add heat energy to an object without raising its temperature if it undergoes a <u>phase change</u>, like ~~s~~ solid $\leftrightarrows$ liquid or liquid $\leftrightarrows$ gas. In this case, the heat energy goes into breaking the intermolecular bonds rather than raising the temp. The amount of energy required is

$$Q_{\substack{\text{phase}\\\text{change}}} = mL$$

where L is the <u>latent heat</u>, and depends upon the material and the phase change.

An example is heating ice from -10°C, say, to +10°C.

The total heat required is:



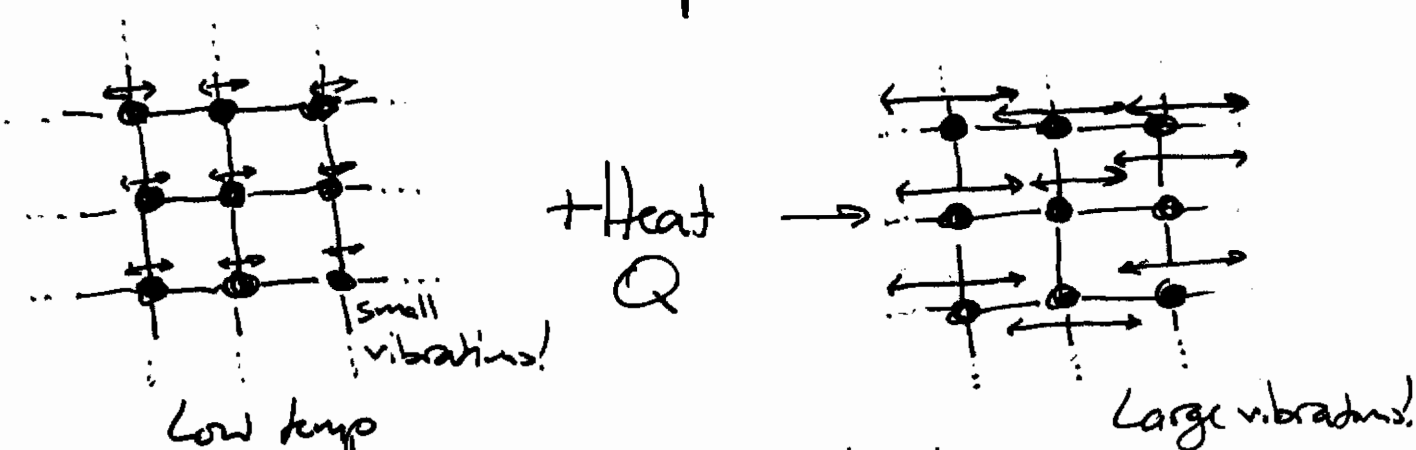
$$Q_{net} = Q_{\substack{Heat\\ice}} + Q_{\substack{melt\\ice}} + Q_{\substack{Heat\\water}}$$

$$= m_{ice} c_{ice} (T_{melt} - T_0) + m_{ice} L_f + m_{ice} c_{H_2O} (T_f - T_{melt})$$

Heat can be transferred between two objects using three different mechanisms, distinguished by what happens to the molecules of the transmitting material:

<u>Conduction</u>: Heat flow through a solid, like a metal.

The heat energy results in the increase in molecular motion of the molecules about a fixed position.

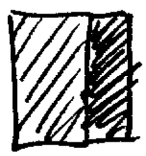


The rate at which heat is conducted through a material is:

$$H = \frac{Q}{t} = kA\frac{(T_H - T_c)}{L}$$

where K is a material-specific thermal conductivity, A is the cross-sectional area of the conductor, L is its length, and T_H and T_C are the temps of the hot end and cold end, resp.

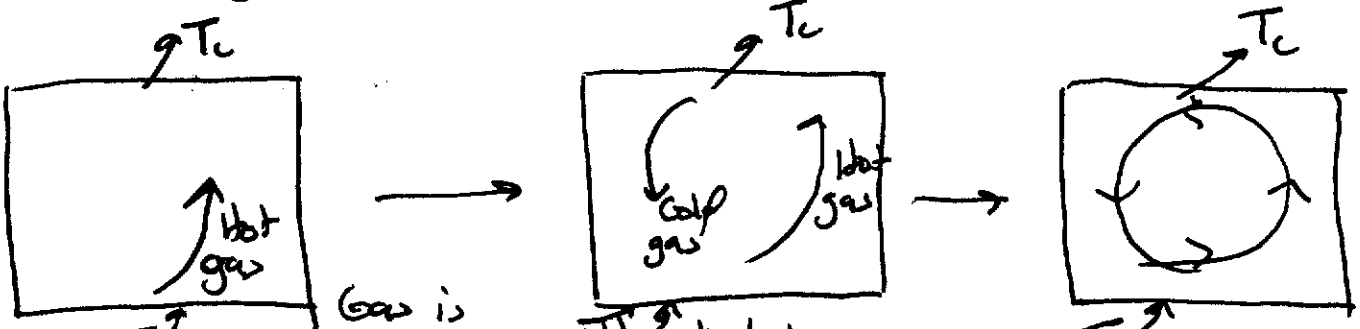
If you have multiple materials in contact, the rates of heat flow through each must be equal; this is like continuity in fluid flow. Anything that goes in has got to come out!



$$T_H \rightarrow H_1 \rightarrow H_2 \rightarrow T_C$$

$$H_1 = H_2$$

<u>Convection</u> - This is where the molecules of the medium actually move! This generally occurs in a fluid, like a liquid or gas.



Gas is heated + starts to rise due to Archimedes!

The hot gas displaces colder gas, causing it to sink

The cooler gas is heated, causing it to rise, and starting a <u>convection cycle</u>

Radiation - All objects emit electromagnetic radiation (i.e. light) as a result of their temperature. You have seen how when you heat an object up, it starts to glow. This is a result of the amount of EM radiation emitted being dependent upon the temperature. As you increase the temp, the amount and rate of energy emitted increases, which we see as a glow!

The rate of energy emission (i.e. the power) is related to the temp by

$$P = \epsilon \sigma A T^4 = \frac{Q}{t}$$

σ is the Stefan-Boltzmann constant, $\sigma = 5.67 \times 10^{-8}\ W/m^2 K^4$.

A is the surface area of the object; the larger it is, the more energy it can emit

T is the object's temperature

ϵ is a material-specific emissivity - some materials radiate better than others. A perfect radiator/absorber ($\epsilon = 1$) is called a Blackbody.

In general, an object will also absorb energy from radiation from its surroundings at a rate

$$P_{absorb} = \epsilon \sigma A T_0^4 \qquad T_0 = \text{surrounding temp}$$

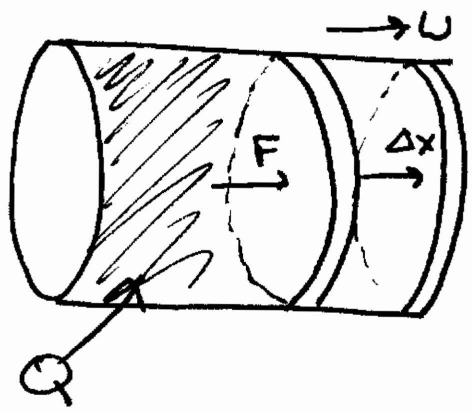
The net rate of heat loss is the difference between emission and absorption:

$$P_{net} = P_{emit} - P_{absorb} = \epsilon \sigma A (T^4 - T_0^4)$$

Chapter 15 - Thermodynamics

Thermodynamics is the study of how you can use Heat to do Work. You can use this to build engines or refrigerators!

For example, a gas trapped in a piston can be heated, causing the gas to expand. The gas is exerting a force on the piston (F), causing it to move (Δx); this is work!



$$W = F \cdot \Delta x$$

In general, if you add heat to a gas, you will increase its internal energy (temperature):

$$\uparrow Q \rightarrow \uparrow U$$

If the gas expands, it does work and loses energy

$$\uparrow W \rightarrow \downarrow U \qquad \text{(Work done by gas)}$$

Hence, the total change in internal energy is

$$\Delta U = Q - W \qquad \text{(Work done by gas)}$$

$$= Q + W \qquad \text{(Work done ON gas)}$$

First Law of Thermodynamics!

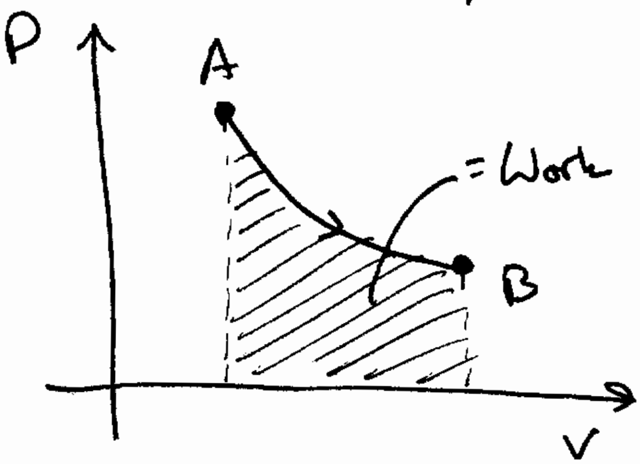
If the process is at <u>constant pressure</u>,

$$W = P\Delta V \quad \text{(Work done BY gas)}$$

$$= -P\Delta V \quad \text{(Work done ON gas)}$$

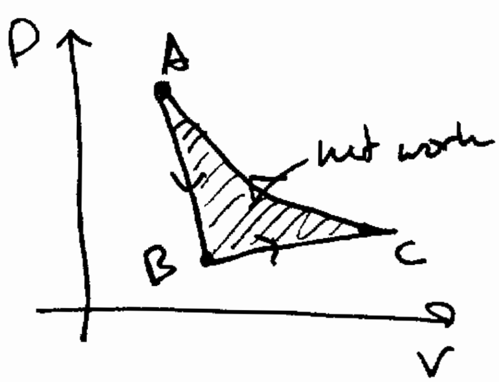
The positive and negative signs are an endless source of confusion; on your exam, do whatever the formula sheet says to do if there is any confusion! I am going to insist that we are very lenient about this.

You can draw thermodynamic processes on a P-V plot:



The area under a P-V curve is equal to the work done!

If you construct an engine that performs some cycle, that will look like:



When the engine completes a cycle, it has returned to its original state; in other words

$$\Delta U_{A \to A} = 0$$

From the first law,

$$\Delta U = Q_{net} - W_{net} = 0 \qquad \text{(Work done BY gas)}$$

$$\Rightarrow Q_{net} = W_{net}$$

The net heat in is

$$Q_{net} = Q_{Hot} - Q_{cold}$$

or

$$W_{net} = Q_{net} = Q_{Hot} - Q_{cold}$$

Hence, the efficiency, or the amount of work you get out compared to the amount of heat energy you put in is

$$e = \frac{W_{net}}{Q_H} = \frac{Q_H - Q_c}{Q_H}$$

$$= 1 - \frac{Q_c}{Q_H}$$

Remember, in general efficiency MUST be less than one!