WJEC (Wales) Physics A-level

Topic 3.4: Thermal Physics

Notes

Internal Energy

Imagine a cup of coffee placed on a table, then let's say we slide the cup across the table. What is the energy of the cup of coffee? It turns out that this question isn't specific enough; we could have said that the coffee has energy due to its high temperature, or we could have argued that the coffee has kinetic energy seeing as it is sliding across the table.

Both of these points are true, but they are considering **different types of energy**. The coffee is made up of a collection of molecules which have kinetic and potential energy, this is what is called **internal energy**.

The internal energy of an object is the **sum of the kinetic and potential energies** of all of its particles.

The internal energy of an object depends on the **state** of the object. This just means that it can depend on the thermodynamical quantities of the object such as **temperature, pressure and volume**.

Internal Energy of an Ideal Gas

One key assumption about an ideal gas is that the particles don't interact and so **only have kinetic energy**, this means that the **internal energy of an ideal gas is the sum of the kinetic energy of the particles.**

The kinetic energy of one mole of an ideal gas is:

$$KE\ of\ one\ mole = \frac{3}{2}RT$$

If the gas contains n moles of particles then the total kinetic energy **and thus internal energy** shall be:

$$U = \frac{3}{2}nRT$$

This shows that the **internal energy of an ideal gas is proportional to and only depends on temperature.** The lower the temperature the less internal energy the gas will have. The smallest temperature possible is called **absolute zero;** it is the starting point of the Kelvin temperature scale at 0K, (or -273.15°C). **An object with a temperature of absolute zero has the minimum internal energy.**

In a real gas the distance between the particles will affect their potential energy, so the internal energy would depend on volume in addition to temperature.

The First Law of Thermodynamics

The internal energy of an object can be changed in two ways, either:

Heat is transferred to/from the object: In our example with the cup of coffee, over time the temperature of the coffee will drop as it transfers heat to the surrounding air. This will cause a drop in the internal energy of the coffee.

Work is done on/by the object: A gas which is being compressed will experience an increase in internal energy.

The First Law of Thermodynamics is an equation which relates **changes in internal energy** to the heat transferred and work done:

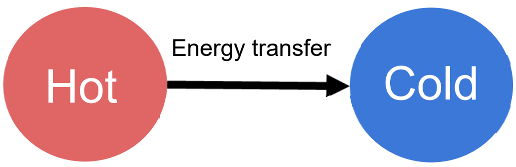
$$\Delta U = Q - W$$

In this equation, ΔU is the **change in internal energy**. Q is the **transfer of heat to the body**, and W is the **work done by the body** on the surroundings. Note that Q and W **can both either be positive or negative.** If Q is negative the internal energy is reduced, so Q is negative when the body **loses** heat to the surroundings. If W is negative, seeing as it already has a negative sign in the first law equation ΔU will be positive, so W is negative when the surroundings do work **on the object**.

The First Law summarises two important points. One is that **heat and work are both forms of energy.** The second is that **energy is conserved.**

Heat Transfer

The heat transfer between two objects depends on their temperature; **heat is only ever transferred from a hotter body to a colder body:**



Generally speaking, heat is transferred at a faster rate when the **difference in temperature** between the objects is larger. This means that an ice cube placed in a room at 30°C will absorb heat faster than if it were placed in a room at 20°C.

The transfer of heat will become gradually slower until eventually the temperature of both objects is the same. **Two objects with the same temperature are in thermal equilibrium,** where heat is no longer transferred between them.

Note that sometimes the word **system** is used instead of object. For example, the air in a room and a cup of coffee are both thermodynamic systems.

For most systems there is a simple relationship that describes how the temperature of the object changes when heat is added. This is shown below:

$$\Delta Q = mC\Delta T$$

In this equation, ΔQ is the amount of heat that is supplied to the system, m is the mass of the system, C is the **specific heat capacity** and ΔT is the **temperature change** of the system due to the added heat.

The **specific heat capacity is defined to be the amount of energy needed to raise the temperature of 1KG of the substance by 1K. It has units of** $J\,KG^{-1}K^{-1}$. This is basically a measure of how resistant the substance is to having its temperature changed. For example, water has a specific heat capacity of $4200\,J\,KG^{-1}K^{-1}$ whereas iron has $450\,J\,KG^{-1}K^{-1}$ meaning that **water takes much more energy to heat up than iron.** This is essentially why water is used in central heating systems, **seeing as a small amount of water can hold a lot of energy.**

Example question:

How much energy does a kettle need to supply to boil 1.2KG of water?

Let's assume the water starts off at room temperature at 20°. To boil the water must reach 100°, so $\Delta T = 80K$. Since we are finding the **difference** in temperature it doesn't matter if we use Degrees or Kelvin as our temperature unit.

Using the fact that for water $C = 4200\,J\,KG^{-1}K^{-1}$, we find $\Delta Q = 1.2 * 4200 * 80 = 403200\,J$

Most kettles have a power of around 3000 Watts, meaning this will take the kettle about 2 minutes to boil the water.

Energy Transfer as Work

From the First Law of Thermodynamics, the internal energy of a system will change if the system does work.

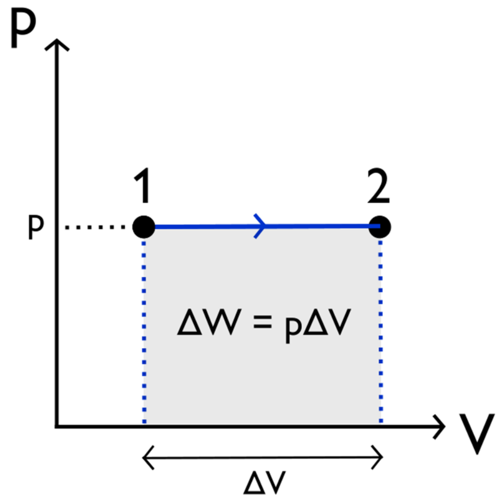
How the system actually does this work depends on what the system is, but for an **ideal gas work is done by the gas if it expands**. In order for the gas to expand, the pressure of the gas has to exert a force on the container walls causing them to move outwards.

Similarly to how pushing an object along a rough surface requires work from the person pushing, the gas must work to expand.

For a constant force F which moves an object through a distance Δx, the work done is $W = F\Delta x$. Similarly if pressure is constant, and the volume changes by ΔV then the **work done by the gas is** $W = p\Delta V$.

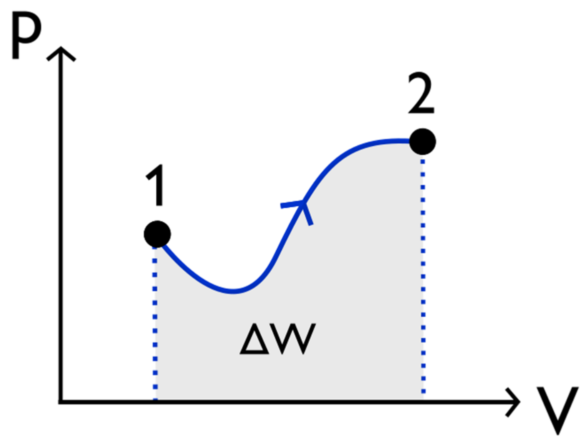
If the gas is compressed instead, then ΔV **shall be negative** meaning the work done by the gas is negative. This just means that the surroundings have done work on the gas, increasing its internal energy.

A graph of the pressure against volume for a gas undergoing expansion at constant pressure is shown here:



The gas expands from state 1 to state 2 whilst at a constant pressure. This plot helps to show that the **work done on the gas is the area under the graph.**

The plot below shows a situation where the **pressure isn't constant**. However, the **area under a pressure-volume curve is always the work done by the gas.** Importantly we cannot use the formula $\Delta W = p\Delta V$ if the pressure changes.



It is relatively easy to do work on a gas by changing its volume, but what about for solids and liquids? These states of matter are much harder to compress, so their **volume is approximately constant.** This means that generally **work is not done on solids or liquids**, and so the only way to change their internal energy is via transfer of heat. This gives the first law of thermodynamics for solids and liquids as:

$$\Delta U = \Delta Q$$

Where here we have set $\Delta W = 0$.

The internal energy of a perfect gas is due to the kinetic energy of the particles it contains, but how do liquids and solids hold their internal energy?

For solids this depends on the material, but an important example is for metals. Metals contain atoms which are more or less fixed in their position, along with a **sea of free electrons.** These free electrons hold most of the metals' internal energy because they are able to move rapidly with large kinetic energies. The atoms can vibrate around their fixed positions but this corresponds to a small amount of energy relative to the electrons.

For this reason, **metals are very good conductors of heat.** They generally have very **small specific heat capacities**, as adding a small amount of heat to a metal will cause a large increase in temperature.

Solids which do not contain free electrons are called **insulators**. These are much **worse conductors of heat than metals,** meaning they have **larger heat capacities.** Their internal energy is due to the vibration of atoms.

Similarly, most liquids do not contain free electrons and so are generally not great conductors of heat. Unlike the atoms in solids, the particles in liquids aren't fixed in position meaning they are free to move around and carry kinetic energy. This means liquids generally have higher heat capacities than insulating solids.