RECOGNISING ACHIEVEMENT

ADVANCED GCE UNIT **2865/01**

PHYSICS B (ADVANCING PHYSICS)

Advances in Physics

INSERT

THURSDAY 21 JUNE 2007

Afternoon

Time: 1 hour 30 minutes

INSTRUCTIONS TO CANDIDATES

- This insert contains the article required to answer the questions in Section A.

This insert consists of **7** printed pages and **1** blank page.

Thermodynamics and Time's Arrow

Heat and Caloric

We take the principle of conservation of energy so much for granted that it is hard to imagine the long struggle needed to get to our current state of understanding. Originally, work – defined in terms of a force and the distance which it moves – was thought to be quite different from heat. Heat was thought to be an invisible fluid called 'caloric' which was contained within matter. Mass and charge are conserved in physical processes – in the same way, caloric was assumed to be conserved. When you burnt wood, the caloric escaped in the flames; when you filed metal, the escaping caloric made the metal and file hot; when you put a cold spoon into a hot cup of tea, caloric flowed from the hot liquid into the spoon.

That this model was flawed was shown by the American physicist Benjamin Thompson, who observed that making holes in metal with a blunt drill released far more caloric than using a sharp one. In fact, a very blunt drill can release as much caloric as you wish by drilling for a longer time, which leads to the unreasonable conclusion that materials contain an infinite amount of caloric. Thompson rightly concluded that heat was somehow connected with motion. This was borne out by James Joule's experiments, which revealed that, when work is done to heat 1 kg of water, every 4.2 kJ results in a temperature rise of 1 °C.

It is clear that there really is a conserved quantity, but this is energy, not caloric. Heat flow is just the thermal transfer of energy, as work is the mechanical transfer of energy. This became the First Law of Thermodynamics – the internal energy of a system changes when heat moves in or out of it, or when work is done on it or by it, and the change in internal energy can be simply calculated by doing the necessary bookkeeping on the thermal and mechanical energy transfers. With this better understanding of the relationship between heat and work, the stage was set for a clearer understanding of the steam engines that had been a major factor in the Industrial Revolution.

Efficiency and temperature

Early engineers such as James Watt made many improvements to steam engines to improve their efficiency. Unfortunately, it became clear that there was an upper limit to the efficiency of any engine. Any heat engine involves the transfer of heat from a **source** at a high temperature (T_{hot}) to a **sink** at a low temperature (T_{cold}), doing useful work in the process. For the early steam engines, the higher temperature was about 373 K. The lower temperature was the temperature of the surroundings.

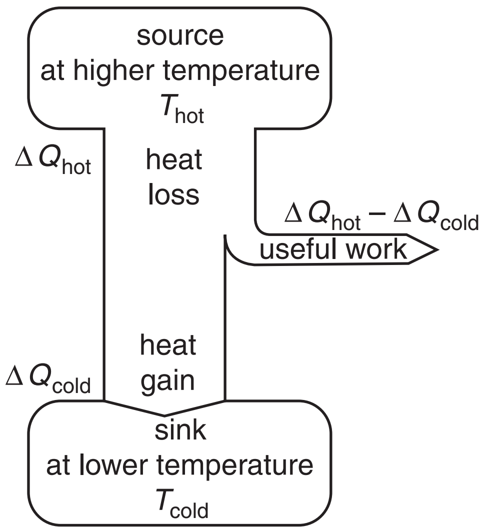


Fig. 1 A heat engine

The French engineer Sadi Carnot found that the efficiency of any heat engine has a maximum theoretical value, dependent only on the two working absolute temperatures, given by

$$\text{maximum theoretical efficiency} = 1 - \frac{T_{\text{cold}}}{T_{\text{hot}}}.$$

This has the clear implication that the maximum theoretical efficiency of early steam engines was bound to be less than 30%. In fact, in the eighteenth century, actual efficiencies were much less. Modern power stations, working with steam at high pressures and at temperatures of several hundred degrees celsius, have reached efficiencies of 40%. The explanation for this relationship involves the introduction of a new concept: **entropy**.

Let me count the ways

Entropy is a measure of the disorder present in a system. As an example, consider a box containing equal numbers of two different types of gas molecule, such as oxygen and nitrogen.

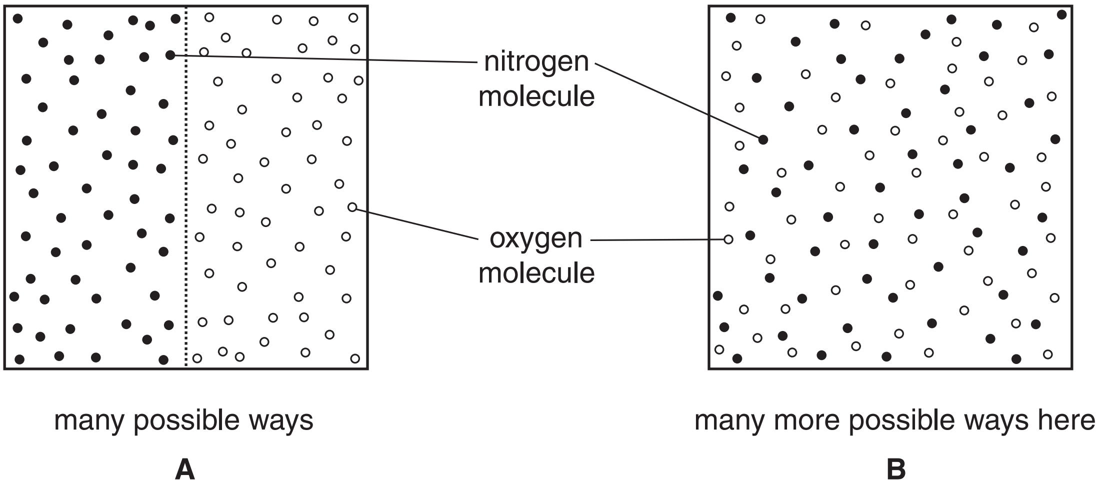


Fig. 2

Suppose that all the oxygen molecules are found in one side of the box, and all the nitrogen molecules in the other, as shown in Fig. 2A. There are obviously very many ways in which the different molecules can be placed to satisfy that condition. However, if you allow each molecule to be anywhere in the box (Fig. 2B), each molecule can be arranged in twice as many ways, so there are very many more possible ways of arranging all the molecules. This is why molecules, moving randomly, are very much more likely to produce a disordered arrangement than an orderly one.

The amount of disorder in a system is measured by a quantity called **entropy**. The more ways that a system can be arranged in, the greater its entropy, so we can say that random processes are likely to increase the entropy. Allowing each molecule to occupy any part of the container in this example greatly increases the entropy.

This example is an illustration of an important principle, the Second Law of Thermodynamics: the total entropy of a object and its surroundings cannot decrease in any process. It must either increase or remain the same, and it can remain the same only in the special circumstance that the process is perfectly reversible.

Energy and Entropy

The Second Law does not apply merely to the way in which matter is organised. It applies also to energy. Einstein showed that a useful model of a solid was to think of it as consisting of a set of separate atoms, each independently oscillating on the 'springs' of the bonds connecting it to the neighbouring atoms in the lattice. Each atom, oscillating on these bonds, can have only certain values of energy, corresponding to equally-spaced energy levels.

Fig. 3 shows two solids, each of 60 atoms, at different temperatures. The hotter has more internal energy than the cooler. Each has its own distribution of quanta of energy, as shown in Fig. 3, where each small black dot represents one quantum of energy. In neither region do all the vibrating atoms have the same amount of energy. As quanta of energy move around at random, some atoms 'get lucky' and so have more quanta of energy while the majority, as with people buying lottery tickets, get none. The hot solid contains more quanta of energy, distributed among its vibrating atoms, than the cold one.

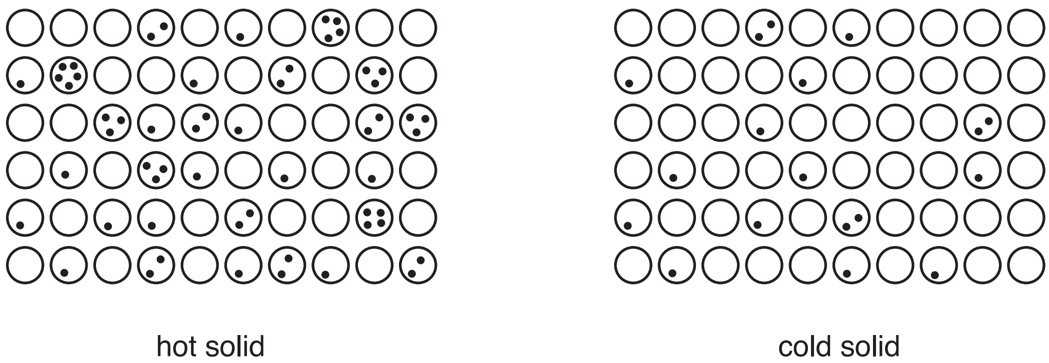


Fig. 3

There are more ways of arranging a larger number of quanta of energy than a smaller number. When the numbers of atoms at each energy level in these two solids are compared, the exponential nature of the distribution of numbers of atoms with different energies can be clearly seen. This is shown in Fig. 4, where each circle represents one atom.

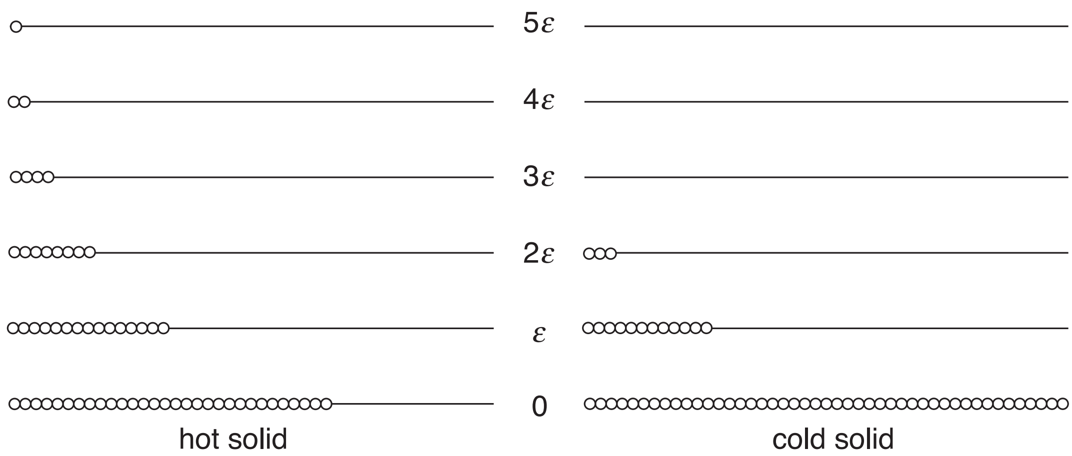


Fig. 4

The Boltzmann factor for the energy distribution is the ratio $\dfrac{\text{number at any level}}{\text{number at the level below}} = e^{-\frac{\varepsilon}{kT}}$, where ε is the energy jump between any two adjacent energy levels. The Boltzmann factor is smaller at lower temperatures, as can be seen by comparing the fractions on each level in Fig. 4.

When two solids at different temperatures act as the source and sink of a heat engine, the random movement of energy quanta results in the energy being distributed more evenly between the two of them. This results in a greater number of ways of arranging the quanta and so is associated with a gain in entropy. This is essentially similar to the mixing of gas molecules in Fig. 2.

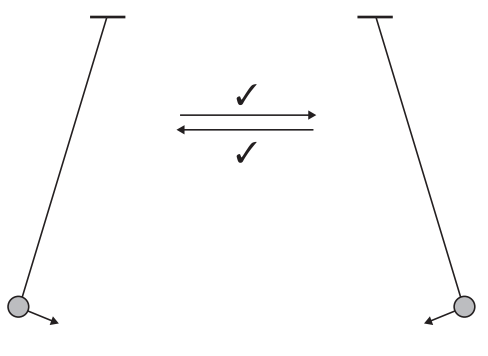
The entropy change ΔS when heat ΔQ flows reversibly into or out of any region at an absolute temperature T is given by

$$\Delta S = \frac{\Delta Q}{T}.$$

So what happens when heat flows from a source to a sink? The source has a entropy loss, as there are now fewer ways in which the energy can be distributed in it, given by $\Delta S_{\text{source}} = \dfrac{\Delta Q_{\text{hot}}}{T_{\text{hot}}}$. However, the sink has an entropy gain, given by $\Delta S_{\text{sink}} = \dfrac{\Delta Q_{\text{cold}}}{T_{\text{cold}}}$, which is larger, so the overall entropy change is an increase, as required by the Second Law. If T_{hot} and T_{cold} are sufficiently different, ΔQ_{cold} can be quite a bit smaller than ΔQ_{hot}, allowing the heat engine to 'steal' the energy difference as work, providing that the entropy gained by the sink is at least as large as the entropy lost in the source.

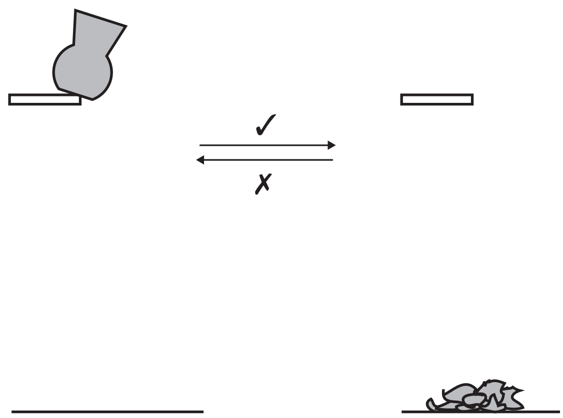
The Arrow of Time

Some changes in physics are perfectly reversible in time, so that a video recording of an event would look much the same if it were played backwards. An example would be a video of several oscillations of a pendulum, providing that no significant amounts of energy had been lost to friction (Fig. 5a). The same could not be said of a fragile object falling off a shelf and smashing on the floor (Fig. 5b).



time-reversible

Fig. 5a



not time-reversible

Fig. 5b

You would know that a video showing the broken pieces of a vase recombining and jumping up onto a shelf must have been played backwards.

The swinging pendulum and falling vase both conserve energy and momentum. This means that they could be 'played in reverse' and still obey those conservation laws. The Second Law of Thermodynamics, that entropy cannot decrease, applies a much tighter constraint. The entropy change in the time-reversible process is zero, so that it can proceed in either direction without contravening the Second Law. This is not true of the other process. The entropy of the vase fragments together with the dispersed energy resulting from the fall is far greater than the entropy of the vase upon the shelf. Only a perfectly reversible process, for which there is no entropy change, can look the same if played backwards. For most processes, there is only one way that they can proceed, and that is to produce greater disorder, and an increase in entropy. This consequence of the Second Law – everything will get more disordered as time goes on – gives an absolute direction to the way in which processes can go: it is the Arrow of Time.

The Beginning and End of the Universe

Astronomical evidence suggests that the Universe started in a hot, dense state about 13.7 billion years ago and expanded rapidly, cooling as it did so. Fluctuations in density in the early Universe resulted in matter – principally hydrogen and helium atoms – condensing under gravity into clouds which were to form the galaxies that we now observe. Within these gas clouds, stars formed as the compressed gases reached high enough temperatures (tens of millions of K) and pressures (about 10^{15} Pa), where nuclear fusion could take place.

Within the cores of stars, smaller nuclei fuse into larger ones, releasing energy. Eventually, in the smallest stars, fusion stops and they shrink into white dwarfs, emitting planetary nebulae. In the cores of the largest stars, the nuclei with the strongest possible binding energies are produced, stopping nuclear fusion. These stars then explode as supernovae. The products of dead stars may be recycled as new stars form, but the process cannot continue for ever. Once nuclei with the strongest binding energies have been produced, no further fusion is possible, and no more stars can be formed.

As the Universe has continued to expand up to the present time, the wavelength of the radiation permeating the whole of space, initially at about 3000 K, has stretched a thousand-fold. The photon energies of this background now correspond to the microwave region of the spectrum, and to a temperature of about 3 K. As space continues to expand, the overall entropy of the Universe continues to increase. As a consequence, all the energy in the Universe is becoming more and more randomly spread.

It seems inevitable that the Universe will eventually consist of cold matter and low energy photons. The energy released during the lifetime of all the stars will be spread evenly throughout the entire Universe, raising its temperature a very small amount above the current 2.7 K. This scenario – not the only one, as the presence of much unobservable dark matter in the Universe will radically affect its expansion – has been referred to as the Heat Death of the Universe.

END OF ARTICLE

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