water, and the other into the pound of oil. The two liquids will of course receive heat, each from its own piece of copper, and they will therefore rise in temperature. Let the rise in temperature be carefully noted by means of two identical thermometers, one immersed in the water and the other in the oil. As soon as the mercury in the two thermometers has ceased to rise, we may assume that the pieces of copper have parted with their surplus heat, but it will be found that the temperature of the water is $68\frac{1}{2}°$, while that of the oil is nearly 92°. Here then we have a pound of copper at 212°, which is only capable of heating a pound of water, having the original temperature of 55°, up to $68\frac{1}{2}°$. In other words, while the copper has lost $212° - 68\frac{1}{2}° = 143\frac{1}{2}°$ the water has gained only $68\frac{1}{2}° - 55° = 13\frac{1}{2}°$. While in the case of the oil, the copper has lost $212° - 92° = 120$, and the oil has gained $92° - 55° = 37°$. Now the amount of heat lost by the copper in each case is of course exactly equal to that gained by the water in the one instance, and by the oil in the other; therefore it is evident that it takes less heat to raise the temperature of the oil by 37° than it does to raise that of the water by $13\frac{1}{2}°$, while the same quantity which suffices to produce this latter effect upon the water is competent to raise the temperature of the copper $143\frac{1}{2}°$.

These figures show conclusively how very differently the temperatures of different bodies are affected by different quantities of heat.

Specific Heat.—The amount of heat which a body of unit mass absorbs in order that its temperature may be raised by one degree; or, *vice versâ*, the amount of heat which the body parts with while its temperature is lowered one degree, is called its Capacity for Heat.

To compare this quantity for different bodies we must first fix upon some unit of quantity of heat. The unit generally adopted in Great Britain is the quantity of heat required to raise one pound of pure water from the temperature of 39° to 40°; or, *vice versâ*, the quantity of heat

parted with by the water in cooling from 40° to 39°. It is necessary thus to specify the temperature, because water, and indeed most bodies, have different capacities for heat at different temperatures. This quantity of heat is called the British Thermal Unit. The capacities for heat of other bodies are designated numerically, by comparing the quantities of heat necessary to raise their temperatures by one degree with unity.

The ratio of the quantity of heat required to raise the temperature of a body one degree, to the quantity required to raise an equal weight of water one degree, is called the Specific Heat of the body. Thus, if it take half the quantity of heat to raise one pound of ice from 20° to 21° that it does to raise a like quantity of water from 39° to 40°, then the specific heat of ice is said to be $\frac{1}{2}$ or ·5.

It is often necessary in questions connected with the steam engine to know how much matter at one temperature it will take, in order to raise a certain quantity of matter of another temperature to some third temperature. These questions are easily solved in the following manner:

Let M be the mass of one of the bodies.
„ M′ be the mass of the other body.
„ T° be the temperature of the body of mass M.
„ T′° be the temperature of the body of mass M′.
„ S be the specific heat of the body of mass M.
„ S′ be the specific heat of the body of mass M′.

When the bodies are mixed together the hotter of them will lose heat to the colder, till at last they attain some common temperature; and the quantity of heat lost by the one substance must be exactly equal to the quantity gained by the other, since the total quantity of heat remains unchanged. Let the body of mass M be the hotter of the two, and let the common temperature which both attain when mixed be called $\theta°$. Then, the quantity of heat lost by one pound of the hotter body in cooling from T° to $\theta° = S\,(T° - \theta°)$, and consequently the quantity lost by M pounds, is M.S ($T° - \theta°$). Similarly, the quantity gained by the other body is M′.S′ ($\theta° - T'°$), and since these two quantities are equal, we obtain the equation

$$M.S. (T° - \theta°) = M'.S'. (\theta° - T'°).$$

This equation is only true provided the whole effect of heat upon bodies is the changing of their temperatures; but it is known that this is not the only effect. We shall afterwards see that a large quantity of heat may be added to bodies without changing their temperatures in the least, and that its effect is in these cases to change the state of constitution of the body; as, for instance, when ice of 32° is changed into water of 32°, or water of 212° changed into steam of 212°. For such cases, therefore, as the mixing of ice and water together, the above equation does not hold good. The equation is also only true on the supposition that the specific heat of bodies is the same at all temperatures. This also is not, strictly speaking, true, but for ordinary purposes the error thus introduced may be neglected.

Mechanical Equivalent of Heat.

Having thus examined the question of the measurement of heat, we are now in a position to revert to the subject of the equivalence of heat and energy. What we want to know is, firstly, how much heat can be got by the doing of a certain quantity of work. The converse question, viz., how much work can be got out of a certain quantity of heat, is of a more complicated character, and its discussion must be postponed till the following chapter.

The first question was accurately settled experimentally by Dr. Joule, of Manchester. He conducted an immense number of experiments on the friction of various solids and liquids, and on the compression of gases. His experiments on the friction of fluids were carried out in the following way. The work was done by causing a known weight to

descend through a given distance; the weight during its descent caused, by means of a suitable mechanism, a paddle to revolve inside a closed vessel filled with the liquid to be experimented upon. This paddle, by agitating the contents of the vessel, produced friction between the particles of the liquid, the walls of the vessel and the paddle, which friction would of course be converted into heat, and would raise the temperature of the vessel and its contents. Careful allowance was made for the resistance caused by the friction of all mechanism exterior to the vessel, and also for all the heat which escaped into the sides of the vessel or into the air. The temperature of the liquid was carefully noted, first before the experiment commenced, and next after the weight had descended through a given distance. The rise of temperature multiplied by the mass of liquid, multiplied by its specific heat, gave, after making all allowances, the quantity of heat generated by the descent of the weight.

The result was that a quantity of work represented by 772 foot-pounds is capable, when all converted into heat, of raising the temperature of a pound of pure water, weighed in vacuo, and having the temperature of 50°, through one degree Fahrenheit.

In other words: *The British Thermal Unit is equivalent to 772 foot-pounds of work.* This number is called the mechanical equivalent of heat.

This result has been fully confirmed by numerous other experiments, made on various substances and in various ways, and it constitutes by far the most important practical discovery which has yet been made in the science of heat. Another way of putting the above result is this. If a pound of water be allowed to fall in vacuo down a height of 772 feet, and if all the heat generated by its impact at the end of its descent be collected into the pound of water, its temperature will be raised one degree.

The equivalent of the units of mechanical work in thermal units can now be readily expressed. For example, one foot-pound of work is equivalent to $\frac{1}{772}$nd part of a thermal unit. One horse-power exerted for a minute, or 33,000 foot-pounds, is equivalent to 42·74 thermal units.

It might at first be supposed that if by doing 772 foot-pounds of work on a pound of water, we thereby raise its temperature one degree, the converse of this must also be true, viz. that by cooling a pound of water by one degree we should thereby be enabled with a suitable apparatus to do 772 foot-pounds of work. It will be seen hereafter that this is not possible; but before this question can be thoroughly understood we must examine into the effect of heat upon gases and water, as it is in general through the medium of these substances that heat is converted into mechanical effect.

The Effect of the Application of Heat to Gas.

The effect of heat upon water has more to do with the subject-matter of this book than has its effect upon gases; but as steam is an imperfect gas, and as the laws which govern the behaviour of gases are much simpler than those for water and steam, we will commence with the subject of gases.

Gases differ from solids and liquids in that they have no tendency to keep to any fixed form and volume. A small portion of gas if introduced into a closed empty vessel will at once expand, so as to fill the whole of it, and will press with a certain equal pressure against every equal portion of the containing surface of the vessel. If by any means the vessel be enlarged, the gas will expand still further, so as to fill it completely as before, but its density or weight per unit of volume will be less, and the pressure which it exerts against the sides of the vessel will also be less.

Take for instance a cylinder closed permanently at one end, and containing a movable piston, by means of which the volume of the portion below the piston can be changed

at pleasure. Let the area of the horizontal section of the cylinder be one square foot, and let the piston, which is supposed to be without weight, be placed at a height of one foot above the bottom of the cylinder, so that the space beneath it is filled with one cubic foot of gas or air at the ordinary temperature and pressure (say 14·7 lbs. per square inch) of the atmosphere. The volume of the air is now one cubic foot; and its pressure is 14·7 lbs. per square inch, or 2116·8 lbs. per square foot. Next, let the piston be raised by hand to a new position, two feet from the bottom of the cylinder, the temperature of the inclosed air being maintained constant. The volume of the air is now doubled or two cubic feet. If we had a proper instrument for measuring pressures, we should find that the pressure per square foot is only half what it was before, or 1058·4 per square inch, and as the same weight of air now occupies twice the original bulk, its weight per cubic foot is halved If instead of raising the piston we had weighted it, or pushed it down to within half a foot of the bottom of the cylinder, and if we had taken care to keep the temperature of the inclosed air constantly the same, we should thus have halved the volume, which would now be half a cubic foot; and doubled the pressure to 4233·6 lbs. per square foot, and also doubled the density or weight per cubic foot.

These facts are expressed generally by Boyle's law of the pressure and volume of gases, which is as follows:

The volume of a portion of gas varies inversely as the pressure, so long as the temperature is constant.

This law may be expressed in other words as follows:

The product of the pressure multiplied by the volume of a portion of gas is a constant quantity, so long as the temperature is constant.

If we remember, that in the above experiment the weight of the gas per cubic foot diminished in proportion as the volume increased, we may express the law otherwise thus:

The pressure of a gas is proportional to its density.

If P represent the pressure in any units, say in pounds, per square foot, and V represent the volume in cubic feet, and C is a constant quantity, to be determined experimentally for one particular case, then the algebraical expression for Boyle's law is

$$P \times V = C \quad \text{or} \quad P = \frac{C}{V}.$$

Graphic Representation of Boyle's Law.

The law may be exhibited graphically in the following manner. Take a line OV, fig. 11, along which volumes are measured to any scale; and a line OP at right-angles to OV, along which pressures are measured, likewise to scale. Now, reverting to the former experiment, measure off along OV a distance O*a* to scale, representing the original volume of the air beneath the piston, viz. one cubic foot, and draw *a*A at right-angles to O*a* so as to represent to scale the original pressure,

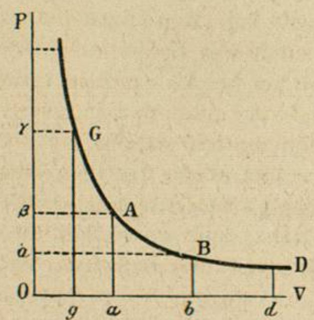

Fig. 11.

viz. 2116·8 lbs. per square foot, then the product of this volume and pressure is represented by the area of the rectangle O*a*A*β*, which is the constant quantity in the above equation. Next draw O*b* to represent the volume in the second stage of the experiment, viz. two cubic feet, and *b*B to represent the corresponding pressure, viz. 1058·4 lbs. per square foot. Then the product of these two quantities is represented by the rectangle O*b*B*α* which is equal to the original rectangle O*a*A*β* because its base O*b* = 2O*a*, while its height *b*B = ½*a*A. Next, for the third stage, measure O*g* equal to half a cubic foot, and *g*G equal to twice the original pressure, or 4233·6 lbs. per square foot. Then the product of this pressure and volume is represented by the rectangle O*g*G*γ*, which likewise is equal to the original rectangle. In a similar way we can obtain the point D such that the rectangle O*d* × *d*D equals the original rectangle, and similarly any other number of points. Now, a curve drawn through the angles G A B D &c. of any number of rectangles of equal area, arranged as in the figure, is called a rectangular hyperbola; and not merely for these points, but also for every other point along the curve, the condition holds good, viz. that the rectangle formed by drawing perpendiculars to the lines OV and OP is equal to the original rectangle.

If we designate all lines measured parallel to OV by the symbol v, and all corresponding lines measured parallel to OP by the symbol p, then it is evident that the curve GABD is expressed by the equation $pv=c$, in which c represents the area of the original rectangle. This equation corresponds exactly, in form, with the equation used to express Boyle's law. The lines OV and OP are called respectively the axes of volume and of pressure; they are also what are called the asymptotes of the hyperbola GABD. The lines drawn from any point on the curve, perpendicular to the axes, are called the ordinates of the point.

Isothermal Lines.

We see, therefore, that the volume and pressure of a portion of gas, when expanding or being compressed, the temperature remaining always the same, may be represented graphically by the ordinates of an hyperbola. Such an hyperbola as has just been drawn is called an *isothermal line of expansion or compression of a gas*, or briefly an *isothermal*. This term, which is derived from two Greek words signifying *equal* and *heat*, signifies that the temperature is constant throughout the expansion or compression represented by the line.

The system of representing graphically the varying pressures and volumes of a portion of gas, by means of a line should be carefully marked, for it is, as we shall afterwards see, the basis of the graphic representation of the working of all gas or steam engines, called indicator diagrams. The method may also be used in many theoretical investigations connected with the expansion of gases and steam under different circumstances, in place of complicated algebraical expressions, or also to supplement and explain these latter.

Effect of varying the temperature of a gas.—Let us now revert to the cylinder of the previous experiment, the piston being, as before, one foot above the bottom, and the air beneath it being at the pressure of the atmosphere, but at the temperature of melting ice. Let us now heat the air by any means so as to bring it up to the temperature of boiling water. During the rise of temperature the piston will gradually ascend, and when the temperature of 212° has been attained, it will have risen to ·366 feet or 4·39 inches above its original position, and will remain there so long as the temperature is maintained at 212°, and the other circumstances continue unchanged.

If the piston had not been allowed to rise, but the air beneath it had been heated as before up to 212°, its pressure would then have increased from 1 to 1·3654 atmospheres—i.e. from 14·7 to 20·08 lbs. per square inch. During the first of these operations the air is said to be heated at constant pressure, and during the second at constant volume.

The fact that air, and indeed all gases, increase in bulk when heated from freezing to boiling point by a fixed fraction (which is nearly the same for every gas) of their original volumes at freezing-point, was first discovered in France by Charles. Hence the numerical statement of the amount of the expansion is usually called Charles' law. The dilatation of gases was afterwards investigated more completely by Dalton in this country, and by Gay Lussac in France; hence

the statement is also often called the law of Dalton and of Gay Lussac. Though both these philosophers agreed as to the total increase in the bulk of a gas when raised in temperature from 32° to 212°, there was, nevertheless, an important difference between them in investigating the increase for each individual degree between these two temperatures.

Gay Lussac's view was that the increase for each degree is a certain fixed quantity, which quantity is a definite fraction of the volume at any temperature. Thus, if V_0 be the volume of the gas at 0°, and V the volume at any other temperature $t°$, and if a be the fraction of V_0 by which the volume is increased for each degree of rise of temperature (commonly called the coefficient of expansion), then, according to Gay Lussac,

$$V = V_0 + t.a.V_0 = V_0(1 + ta).$$

The total increase in volume between 32° and 212° is, as stated above, in the ratio of 1 to 1·3654, and the increase per degree is therefore ·3654 ÷ 180 = ·00203. If the original volume be taken at 0°, the corresponding increase per degree = ·00217.

In a similar manner, if the volume of a gas be kept constant and it be heated from 32° to 212°, the pressure will be increased in the ratio of 1 to 1·3654, and the increase per degree of temperature will be as before, ·00203. This experimental result can also be deduced by the aid of Boyle's law from the known increase of volume when the temperature is raised and the pressure kept constant. For, let the pressure be P and the original volume at 32° be V. When heated to 212° at constant pressure the volume becomes V × 1·3654. Let the gas now be compressed at the constant temperature 212° back to its original volume, and the pressure by Boyle's law will become $P\frac{V \times 1{\cdot}3654}{V}$

$= P \times 1{\cdot}3654.$

According to Dalton, however, if a be the coefficient of expansion for a rise of one degree from 32° to 33°, so that the volume at 33° $=V+aV$; then at 34° the volume becomes $(V+aV)+a\ (V+aV) = V\ (1+a)^2$, and at 35° it is $V\ (1+a)^3$ and so on. In other words, a gas at a given temperature, say 50°, for a rise of one degree, say to 51°, increases in bulk by a fixed fraction of its volume at 50° and not of its volume at 32°—as stated by Gay Lussac—and so on for any other temperature. It will be seen that there is a most important difference between the two laws, and many of the theoretical calculations relating to steam would be largely modified by the adoption of Dalton's instead of Gay Lussac's statement. It might have been thought that the question of accuracy between the two could be easily settled by experimental investigation, and the subject has naturally received much attention at the hands of Regnault, Stewart, and others. It is, however, not possible absolutely to prove the truth of either law experimentally, because, as we have seen, no accurate experimental method exists of measuring temperature quantitatively. The experiments above referred to are, however, generally accepted as proving that Gay Lussac's statement is the nearer to truth of the two, and, in conformity with this generally received opinion, his law will be made use of throughout the remainder of this book.

It should here be stated that both Boyle's and Gay Lussac's laws are only true for perfect gases. For actual gases Regnault has found :

1. That the product $P \times V$ is not quite constant, especially for those gases which can be liquefied.

2. That the coefficients of expansion of air and other simple and compound gases are not quite identical.

3. That the coefficients of expansion of all gases, with the exception of hydrogen, increase somewhat as their density increases.

4. The coefficients of expansion of all gases become more nearly equal to each other as their pressures diminish.

The Air Thermometer.

One of the most important practical deductions from the above laws is the construction of the air thermometer. There are several reasons why air, or any other gas, is a better substance for measuring temperature than any liquid or solid. Foremost among these is the fact, that no two different liquids or solids will agree when used for measuring temperatures other than the two originally selected temperatures, which are invariably those of freezing and boiling water. On the other hand, air and all gases, when at the same pressures, will expand by almost exactly identical amounts when subjected to the same temperatures. Another reason for preferring gases to other substances for thermometrical purposes is that the specific heats of gases remain the same at all temperatures.

Without going into the question of air thermometers in actual use, a very simple theoretically possible form of such an instrument will now be described for the purpose of illustrating what is meant by *absolute temperature.*

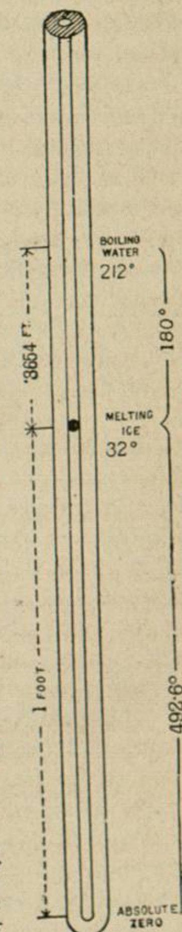

Fig. 12.

Take a tube with a uniform bore, and inclose in it a portion of air or other gas, in such a way that, at the temperature of melting ice, the air will occupy a space of one foot measured from the bottom of the tube and will be separated from the external atmosphere by an air-tight piston, which, however, is free to move up and down the bore of the tube (see fig. 12). The pressure on the inclosed gas is to be maintained constant. If, now, the tube be exposed to the steam of boiling water, the inclosed air will expand, causing the piston to rise to a point 1·3654 feet above the end of the tube. A mark at this point will indicate the temperature of 212°. If we divide the space between this and freezing-point, viz. ·3654 feet, into 180 equal divisions, each of these will, if Gay Lussac's law be true, represent one degree Fahrenheit. Hence we see that for each degree of rise of temperature the volume of the gas inclosed in the thermometer increases by a fraction of its original volume at $32° = \frac{\cdot 3654}{180} = \cdot 00203$.

Absolute Temperature.

If we choose we can now extend these subdivisions below freezing and above boiling point as far as we like. As the space between freezing and boiling points, viz. ·3654 feet, contains 180 divisions, the space below freezing-point to the bottom of the tube, viz. one foot, will contain 492·6 divisions or degrees. In other words, the bottom of the tube is $492·6° - 32° = 460·6°$, or in round numbers, 461° below zero. As we shall see hereafter, great use is made in thermodynamical calculations of the above fact.

This point has been called the *absolute zero of temperature*, and the first conclusion which would be deduced from the foregoing would be, that by depriving a portion of gas of all its heat we should thereby reduce its bulk to nothing. It is needless to say that all reasoning as to the condition of a gas when deprived of all its heat is mere speculation, as we have no experimental knowledge of the subject whatever. Dalton's law, it will be observed, would lead to quite other conclusions. It is, however, very convenient in calculations respecting gases to measure temperature from the absolute zero, instead of from the zeros of any of the scales in common use. Temperature thus measured is called *absolute temperature*. To convert temperature measured on Fahrenheit's scale to absolute temperature, we have only to add 461° to the reading. Thus the absolute temperature of boiling water would be $461° + 212° = 673°$.

Combination of Boyle's and Gay Lussac's laws.—Let V_0 be the volume of a portion of gas at the pressure p and the temperature 0°.

Let V be the volume at any other temperature $t°$, the pressure remaining p.

Then by Gay Lussac's law $V = V_0(1 + at) \therefore V_0 = \frac{V}{1 + at}$.

Now by Boyle's law the volume V', at any pressure p', and temperature t', multiplied by its pressure p', equals the volume at pressure p, and same temperature t', multiplied by its pressure p,

$$\text{or,} \quad V'.p' = V_0(1 + at').p = V.\frac{1 + at'}{1 + at}.p.$$

Now by reference to the description of the air thermometer and the definition of the absolute zero, it will be seen that $1 + at'$ and $1 + at$ bear the same relations to t' and t that the absolute temperatures do to the temperatures on the Fahrenheit scale; therefore writing τ' and τ for the absolute temperatures instead of $1 + at'$ and $1 + at$, we get the very simple equation

$$\frac{V'p'}{\tau'} = \frac{Vp}{\tau},$$

or in other words, *the product of the volume and pressure of a portion of gas is proportional to the absolute temperature.*

The above equation will be found of the greatest use in solving all questions as to the varying volumes, pressures and temperatures of a portion of gas when its condition at any one temperature is given.

The same result might have been arrived at in a simpler way by the mere inspection of the air thermometer, for this latter is nothing more nor less than the record of a series of experiments on the varying volume of a portion of gas when the pressure remains constant and the temperature changes. The varying volume is in fact the exact measure of the varying temperature. Hence the product of the pressure and volume is exactly proportional to the length of the column of gas—i.e. to the absolute temperature. Now there is no reason why any one pressure should be chosen rather than any other, hence the above statement is perfectly general.

The specific heat of gases.—The specific heat of a gas, is, in accordance with the definition on page 39, the ratio of the amount of heat required to raise a given weight of it one degree in temperature, to the amount required to raise the same weight of water one degree. In heating a gas, it is possible to do so when the volume is kept constant, or the pressure constant, or partly in the one way and partly in the other. The first two cases are the most important. It will be seen that, if the mechanical theory of heat be true, it will take more heat to raise the gas one degree in temperature when the pressure is kept constant while the volume is variable than in the reverse case. For, take a cylinder closed at one end, having a section of one square foot, and confine a cubic foot of gas in it, by means of a piston free to move. The cubic foot of gas has then to sustain the pressure due to the weight of the atmosphere plus the weight of the piston. When, therefore, the inclosed gas is heated and expands, the whole weight of the atmosphere and piston is raised through a certain space, and work is done. Consequently, the heat supplied to the gas is partly expended in raising its temperature, and partly in doing external work. Now, in the case of heating at constant volume, the weight of the atmosphere is not raised at all, and no external work is done, and therefore less heat is required in this than in the former case.

The ratio which these two specific heats of a gas bear to each other can be easily ascertained. Reverting to the cylinder, let the cubic foot of gas be at the temperature of 32°, and let it be heated till its volume is doubled. To do this the temperature must be raised, according to Gay Lussac's law, by 492·6°.

Now, according to Regnault's experiments the specific heat of gas—say air—is, when heated at constant pressure, 0·2375, therefore the quantity of heat in thermal units required to effect the above operation is equal to the original weight of the cubic foot of air at the atmospheric pressure of 14·7 lbs. per square inch, and temperature of 32°, multiplied by the rise in temperature, multiplied by the specific heat. Now the weight of the cubic foot of air, under the above circumstances, is ·0807 lb. The rise of temperature is 492·6° and the specific heat is 0·2375 ; therefore, the quantity of heat required is,

$$\cdot 0807 \times 492\cdot 6 \times 0\cdot 2375 = 9\cdot 422 \text{ thermal units.}$$

Now the external work done during the process is equal to the weight of the atmosphere, viz. 14·7 lbs. per square inch, resting on the surface of the piston, viz. 144 square inches, multiplied by the height through which it is raised viz. one foot.

$$= 14\cdot 7 \times 144 \times 1 = 2116\cdot 8 \text{ foot-pounds.}$$

Now 2116·8 foot-pounds is equivalent to $\frac{2116\cdot 8}{772} = 2\cdot 74$ thermal units. Therefore, of the above quantity of thermal units 2·74 have been expended in doing external work, and

$$9\cdot 422 - 2\cdot 74 = 6\cdot 682 \text{ thermal units}$$

represent the quantity of heat expended in raising the temperature of the air, provided that no other effect has been produced by the heat.

Now, before we can establish a ratio from the above data between the two specific heats, we must ascertain whether heat has been expended in any other way than in doing external work, and in raising the temperature of the mass of gas. It will be noticed, that the gas, at the end of the above experiment, is in a different molecular condition to what it was at the commencement, the particles composing the gas being much further apart. It might be thought that part of the heat was expended in effecting this molecular change. Whatever may be the case with other substances, it has been experimentally proved by Joule that no heat is expended in this way in the case of the permanent gases. The experiment consisted in warming a mass of gas in a closed vessel, and then, by opening a cock, allowing the confined and heated gas to expand into another vessel, in communication with the first, and which was in a condition of vacuum. In this case the gas evidently expanded without doing any work; therefore, no heat was consumed in this way. At the end of the experiment the temperature was found to be unchanged, showing that no heat had been expended in altering the distance apart of the particles.

Reverting now to the original experiment, we see that the only effects produced when the gas was heated at constant pressure were:—1. The doing of external work; 2. The raising the temperature of the mass of the gas; and 3. The further separation of the particles of the gas. We have seen that the last effect required no expenditure of heat at all, and consequently we may be sure that the above-named quantity of 6·682 thermal units represents the amount of heat expended in raising the temperature of the gas.

If, now, the mass of gas be heated at constant volume instead of at constant pressure, no external work is done, and no separation of the particles takes place, and the whole heat is expended in raising the temperature of the gas. The quantity of heat required for this purpose is, as we have seen, only 6·682 thermal units. The specific heat of the gas at constant volume is, therefore, less than the specific heat at constant pressure in the ratio of 6·682 to 9·422, and consequently the numerical value of the specific heat at constant volume is

$$0{\cdot}237 \times \frac{6{\cdot}682}{9{\cdot}422} = 0{\cdot}167.$$

Our knowledge of the specific heats of various gases is derived chiefly from Regnault's experiments, which were all made upon gases at constant pressure. The experimental determination of the specific heat at constant volume is a very difficult operation, hence it is usually derived theoretically from the specific heat at constant pressure in the manner explained above. Though the true value of the ratio of the two specific heats has not been confirmed by direct experiment, still it has been assumed in calculations, such as the theoretical determination of the velocity of sound in air; and as the calculated velocity agrees practically with the result of experiment, we may assume that the value for the ratio has been indirectly confirmed by experiment.

The Effect of the Application of Heat to Water.

Having thus briefly considered the effect of heat when applied to gases, we must now consider its effect on water. This latter is the more complicated subject of the two, but is at the same time of the greater importance in the study of the steam engine.

It will be convenient to commence with the solid form of water or ice at the temperature of 0° F. In this condition the specific heat is about 0·5, that is, it takes half as much heat to raise a pound of ice one degree in temperature that it does to raise the same weight of water by the same amount. To raise the temperature of the ice up to 32° requires then $32 \times 0{\cdot}5 = 16$ units of heat. At this temperature its volume compared with water at its maximum density is as 1·0908 to 1.

If we continue to heat the pound of ice at 32° it will begin to melt, but the temperature will remain stationary