DETECTION OF ALCOHOL IN TRANSPARENT SOAPS.

By H. JAY.

It appears that every article manufactured with the aid of alcohol is required on its introduction into France to pay duty on the supposed quantity of this reagent which has been used in its preparation. Certain transparent soaps of German origin are now met with, made, as is alleged, without alcohol, and the author proposes the following process for verifying this statement by ascertaining the presence or absence of alcohol in the manufactured article: 50 grms. of soap are cut into very small pieces and placed in a phial of 200 c.c. capacity; 30 grms. sulphuric acid are then added, and the phial is stoppered and agitated till the soap is entirely dissolved. The phial is then filled up with water, and the fatty acids are allowed to collect and solidify. The supernatant liquid is drawn off, neutralized, and distilled. The first 25 c.c. are collected, filtered, and mixed, according to the process of MM. Riche and Bardsy for the detection of alcohol in commercial methylenes, with $\frac{1}{2}$ c.c. sulphuric acid at 18° B., then with the same volume of permanganate (15 grms. per liter), and allowed to stand for one minute. He then adds 8 drops of sodium hyposulphite at 33° B., and 1 c.c. of a solution of magenta, 1 decigram. per liter. If any alcohol is present there appears within five minutes a distinct violet tinge. The presence of essential oils gives rise to a partial reduction of the permanganate without affecting the conversion of alcohol into aldehyd.

ON THE CALORIFIC POWER OF FUEL, AND ON THOMPSON'S CALORIMETER.

By J. W. THOMAS, F.C.S., F.I.C.

A SIMPLE experiment, capable of yielding results which shall be at least comparative, has long been sought after by large consumers of coal and artificial fuel abroad in order to ascertain the relative calorific power possessed by each description, as it is well known that the proportion of mineral matter and the chemical composition of coal differ widely. The determination of the ash in coal is not a highly scientific operation; hence it is not surprising that foreign merchants should have become alive to the importance of estimating its quantity. While, however, the nature and quantity of the ash can be determined without much difficulty, the determination of the chemical composition of coal entails considerable labor and skill; hence a method giving the calorific power of any fuel in an exact and reliable manner by a simple experiment is a great desideratum. This will become more obvious when one takes into consideration the many qualities and variable characters of the coals yielded by the South Wales and North of England coal fields. Bituminous coals—giving some 65 per cent. of coke—are preferred for some manufacturing purposes and in some markets. Bituminous steam coals, yielding 75 per cent. of coke, are highly prized in others. Semi-bituminous steam coals, yielding 80 to 83 per cent. of coke, are most highly valued, and find the readiest sale abroad; and anthracite steam coal (dry coals), giving from 85 to 88 per cent. of coke (using the term "coke" as equivalent to the non-volatile portion of the coal) is also exported in considerable quantity. Now the estimation of the ash of any of these varieties of coal would afford no evidence as to the class to which that coal belongs, and there is no simple test that will give the calorific power of a coal, and at the same time indicate the degree of bituminous or anthracitic character which it possesses.

In order to obtain such information it is necessary that the percentage of coke be determined together with the sulphur, ash, and water, and these form data which at once show the nature of a fuel and give some indication of its value. To ascertain the quantity of the sulphur, ash, and water with accuracy involves more skill and aptitude than can be bestowed by the non-professional public; the consequence is that experiments entailing less time and precision, like those devised by Berthier and Thompson, have been tried more or less extensively. In France and Italy, Berthier's method—slightly modified in some instances—has been long used. It is as follows:

70 grammes of oxide of lead (litharge) and 10 grammes of oxychloride of lead are employed to afford oxygen for the combustion of 1 gramme of fuel in a crucible. From the weight of the button of lead, and taking 8,080 units as the equivalent of carbon, the total heat-units of the fuel is calculated. This experiment is very imperfect and erroneous upon scientific grounds, since the hydrogen of the fuel is scarcely taken into account at all. In the first place, hydrogen consumes only one quarter as much oxygen as carbon, and, furthermore, two-ninths only of the heating power of hydrogen is used as the multiplying number, viz., 8,080, while the value of hydrogen is 34,462. In other words, one-eighth only of the available hydrogen present in the fuel is shown in the result obtained. Apart from this my experience of the working of Berthier's method has been by no means satisfactory. There is considerable difficulty in obtaining pure litharge, and it is almost impossible to procure a crucible which does not exert a reducing action upon the lead oxide. Some twelve months ago I went out to Italy to test a large number of cargoes of coal with Thompson's calorimeter, and since then this apparatus has superseded Berthier's process, and is likely to come into more general use. Like Berthier's method, Thompson's apparatus is not without its disadvantages, and the purpose of this paper is to set these forth, as well as to suggest a uniform method of working by means of which the great and irreconcilable differences in the results obtained by some chemists might be overcome. It has already been observed that a coal rich in hydrogen shows a low heating power by Berthier's method, and it will become evident on further reflection that the higher the percentage of carbon the greater will be the indicated calorific power. In fact a good sample of anthracite will give higher results than any other class of coal by Berthier's process. With Thompson's calorimeter the reverse is the case, as the whole of the heating power of the hydrogen is taken into account. In short, with careful working, the more bituminous a coal is the more certain is it that its full heating power shall be exerted and recorded, so far as the apparatus is capable of indicating it; for when the result obtained is multiplied by the equivalent of the latent heat of steam the product is always below the theoretical heat units calculated from the chemical composition of the coal by the acid of Favre and Silbermann's figures for carbon and hydrogen. On the other hand, when the heating power of coal low in hydrogen is determined by Thompson's calorimeter, much difficulty is experienced in burning the carbon completely; hence a low result is obtained. From a large number of experiments I have found that when a coal does not yield more than 86 per cent. of coke, it gives its full comparative

heating power, but it is very questionable if equal results will be worked out if the coke exceeds the above amount, although I have met with coals giving 87 per cent. of coke, which were perfectly manageable, though in other cases the coal did not burn completely. It will be noted that the non-volatile residue of anthracite is never as low as 86 per cent., and this, together with the very dry steam coals and bastard anthracite (found over a not inextensive tract of the South Wales Coal field), form a series of coals, alike difficult to burn in Thompson's calorimeter. Considerable experience has shown that in no single instance was the true comparative heating power of anthracite or bastard anthracite indicated. With a view to accelerate the perfect combustion of these coals, sugar, starch, bitumen, and bituminous coals—substances rich in hydrogen—were employed, mixed in varying proportions with the anthracitic coal, but without the anticipated effect. Coke was also treated in a like manner. Without enlarging further upon these futile trials—all carefully and repeatedly verified—the results of my experiments and experience show that for coals of an anthracitic character, yielding more than 87 per cent. of coke, or for coke itself, Thompson's calorimeter is not suited as an indicator of their comparative calorific power, for the simple reason that some of the carbon is so graphitic in its nature that it will not burn perfectly when mixed with nitrate and chlorate of potash. A sample of very pure anthracite used in the experiments referred to, gave 90.4 per cent. of non-volatile residue, and only 0.84 per cent. of ash. This coal was not difficult to experiment with, as combustion started with comparative ease and proceeded quite rapidly enough, but in every instance a portion of the carbon was unconsumed, and consequently instead of about 13° of rise in temperature only 10° were recorded.

Since the calorific power of a coal is determined by the number of degrees Fahrenheit which a given quantity of water is raised in temperature by a known weight of fuel, it follows that every care should be taken that the experiment be performed under similar atmospheric conditions. The oscillation of barometric pressure does not appear to affect the working, but the temperature of the room in which the work was done, and especially that of the water, are most important considerations. It has been observed by some who have used this apparatus—and I have frequently noticed it myself—that the lower the temperature of the water is under which the fuel is burnt the higher is the result found. This has been explained on the assumption that the colder the water used, the greater is the difference between the temperature of the room and that of the water; hence it would be expedient that in all cases when such experiments are made the same difference of temperature between the air in the room and the water employed should always exist. For example, if the temperature of the room were 70°, and the water at 60°, then the same coal would give a like result with the water at 40° and the room at 50°. This has been regarded as the more evident, because the gases passing through the water escape under favorable conditions of working at the same temperature as the water, and are perfectly deprived of any heat in excess of that possessed by the water. Under these circumstances it would seem only reasonable that this assumption should be correct. It was, however, found after a large number of experiments upon the same sample of coal that this was not the case. 30 grammes of coal which raises the temperature of the water 13.4°, when the water at starting was 60° and the room at 70°, gives 13.7° rise of temperature with the water at 40° and the room at 50°. Conversely, when the water is at 70° and the room at 80°, a lower result is obtained. The explanation appears to be this: The gas which escapes from the water was not in existence in the gaseous form previous to the experiment, and the heat communicated to the gas being a definite quantity it follows that the more the gas is cooled the greater the proportion of chemical energy in the shape of heat will be utilized and recorded as calorific power.

In order, therefore, to make the experiment more simple and workable at all temperatures, a sample of coal was selected, which should be perfectly manageable and readily consumed. Appended is an analysis of the coal employed (from Elbow Vale, Monmouthshire):

	Composition per cent.
Carbon	88.33
Hydrogen	5.18
Oxygen	3.28
Nitrogen	0.55
Sulphur	0.70
Ash	1.26
Water (moisture)	0.80
	100.00

In the following experiments the standard temperature of the water was taken as 60° F., and as the coal gave 13.4° of rise of temperature, 67° F. was selected as the standard room temperature. The reason for this room temperature is obvious, for, whatever heating effect the higher temperature of the room may have upon the water in the cylinder during the time occupied by the first half of the experiment, would be compensated for by the loss sustained during the second half of the experiment, when the temperature of the water exceeded that of the room. The mean of numerous trials gave 13.4° F. rise of temperature, equal to 14.74 lb. of water per lb. of coal. When the water was at 50° and the room at 57°, the mean of several experiments gave 13.5° rise of temperature. When the water was 40° at starting and the room at 47°, 13.65° was the average rise of temperature. Trials were made at intermediate temperatures, and the results always showed that higher figures were recorded when the water was coldest. With a view of getting uniformity in the results it was thought well to make experiments, in order to find out what temperature the room should be at, so that this coal might give the same result with the water at 50°, 40°, or at intermediate temperatures. Without going much into detail, it was found that when the temperature of the room was at 40° and that of the water 40°, and the experiment was rapidly and carefully performed, 13.4° rise of temperature was given; but this result could be obtained without special effort when the room was 42° and the water 40° at starting. It is evident that the cooling effect of the air in the room upon the water cylinder is very appreciable when the water has reached 13° above that of the room. When the water was at 50° and the room at 55°, the coal gave 13.4° rise with ease and certainty, and it would not be out of place to remark here that with those coals which burn well in Thompson's calorimeter, the results of several trials are remarkably uniform when properly performed. With the water at 70° and the room at 80°, a like result was worked out. Experiments at intermediate temperatures were also carried out (see table in sequel). It is true that the whole difference of temperature we are dealing with in making these corrections is only 0.25, but 0.2 in the result, when

multiplied by 537 to bring it into calories, as is done by the authorities in Italy, makes more than 100 heat units—a serious difference when 5d. per ton fine is attached to every 100 calories lower than the number guaranteed.

Taking the latent heat of steam as 537° C., and multiplying this number by 14.74, the evaporative power of the coal used in these experiments, its equivalent in calories is 7,915. From the analysis of this coal, disregarding the nitrogen and deducting an equivalent of hydrogen for the oxygen present, the total heat units given by Favre and Silbermann's figures for carbon (8,080) and hydrogen (34,462) will be 8,746. It will be seen, therefore, that the calorific power, as determined by Thompson's apparatus, gives a much lower result when multiplied by 537 than the heat units calculated from the chemical composition of the coal. When I used Thompson's apparatus in the chemical laboratory at Turin to determine the evaporative power of various cargoes of South Wales coal, it was agreed by mutual consent that the temperature of the water at starting should be 39° F. (the temperature at which the heat unit was determined). The temperature of the room was about 60°, but this varied, as the weather was somewhat severe and changeable. Under these conditions, with the water at 39° and room 60°, the coal which gives 14.74 lb. of water per lb. of coal, will give as high as 15.88 lb. of water per lb. of coal. This result multiplied by 537=8,496 calories, approaching much more nearly to the theoretic value. This method of working is still practiced abroad, but experience has shown that very widely differing results follow when working in this manner, especially if the temperature of the room is changeable, as it naturally is where ash determinations and other chemical work is proceeding simultaneously. The time the experiment lasts, taking the reading on a quickly rising thermometer and other considerations, render the experiments anything but trustworthy when 0.2 of a degree makes a difference of more than 100 calories. In the instructions supplied with Thompson's calorimeter nothing is said as to the temperature of the room in which the experiment is performed, but simply that the water shall be at 60° F. If, with the water at 60°, a room were at 50°, as it often is in winter, a good coal would give 14 lb. of water per lb. of coal as the evaporative power; but if in summer, the room were at 75° and the water at 60°, the same coal would give 15 lb. of water per lb. of coal. If further evidence were needed of the effect of temperature a consideration of the experiments already referred to will show how necessary it is that some general rule shall be adopted. Considerable stress is laid (in the instructions) upon the quantity of oxygen mixture used being determined by rough experiments. This I have found leads to erroneous conclusions unless a number of experiments are tried in the calorimeter, as it often happens that the quantity which appears to be best adapted is not that which yields a trustworthy result. There are many samples of South Wales coal, 30 grains of which will require 10 parts of oxygen mixture in order to burn completely, but since a little oxygen is lost in drying and grinding, and few samples of chlorate are free from chloride, it is not safe to use less than 11 parts of oxygen mixture, but this amount is sufficient in all cases, and never need be exceeded. I have made numerous experiments with various coals (anthracite, steam, semi-bituminous, and bituminous, including a specimen of the ten yard coal of Derbyshire), and find that with 11 parts of chlorate and nitrate of potash, they are all perfectly manageable and yield the best results. It is quite clear that the excess of chlorate is decomposed in all instances, and the latent heat of the oxygen evolved, but those coals which are best to experiment with did not yield results that differed when the quantity of oxygen mixture was reduced to nearly the limit required for combustion of the coal. Under these circumstances, therefore, the constant use of 11 parts of oxygen mixture—a suitable quantity for all coals exported—would enable operators to obtain similar figures, and make the test uniform in different hands.

The following is a brief outline of the method of procedure recommended: Sample the coal until an average portion passes through a sieve having 64 meshes to the square inch. Take about 300 grains (20 grammes) of this and run through a brass wire gauze having 4,600 meshes to the square inch, taking care that the whole sample selected is thus treated. One part of nitrate of potash and 3 parts of chlorate of potash (dry) are separately ground in a mortar, and repeatedly sifted through another wire gauze sieve, having 1,000 meshes to the square inch, in order that the oxygen mixture shall not be ground to an impalpable powder, as this is very undesirable. It absorbs moisture rapidly, and interferes with the regularity of the combustion when very fine. 330 grains of the powder are weighed out (after drying), and intimately incorporated with 30 grains of coal—better with a spatula than by rubbing in a mortar—and then introduced into a copper cylinder (3½ inches long by ¼ inch wide, made from a copper tube), and pressed down in small portions by a test-tube with such firmness as is required by the nature of the coal, not tapped on the bottom, since the rougher portions of the oxygen mixture rise to the surface. As the temperature of a room is almost invariably much higher than the water supply, a little hot water is added to that placed in the glass cylinder, until the difference of temperature between the water and the room is about the mark indicated in the following table:

Room at	The water should be
80° F.	70° F.
72	64
67	60
60	54
55	50
50	46
42	40

Say, for example, the room was at 57° and the water placed in the cylinder was at 46°: add a little hot water and stir with the thermometer until it assumes 52°. By the time the excess of water has been removed with a pipette until it is exactly level with the mark, and all is ready, the temperature will rise nearly 0.5°. Let the thermometer be immersed in the water at least three minutes before reading. The fuse should be placed in the mixture, and everything at hand before reading and removing the thermometer. After igniting the fuse and immersing the copper cylinder in the water, the apparatus should be kept in the best position for the gases to be evolved all around the cylinder, and the rate of combustion noted. Some coals are very unmanageable without practice, and samples of "patent fuel" are sometimes met with, containing unreasonable proportions of pitch, which require some caution in working and very close packing, inasmuch as small explosions occur during which a little of the fuel escapes combustion.

In order that the experiment shall succeed well, experience has shown that the nature of the fuse employed has much to do with it. Plaited or woven wick is not adapted, and will

fail absolutely with dry coals, unless it is made very free burning. In this case not less than three-quarters of an inch in length is necessary, and the weight of such is very appreciable. I always use Oxford cotton, and thoroughly soak it in a moderately strong solution of nitrate of potash. When dry it should burn a little too fast. The cotton is rubbed between two pieces of cloth until it burns just freely enough; then four cotton strands are taken, twisted together, and cut into lengths of $\frac{3}{4}$ inch and thoroughly dried. Open out the fuse at the lower end when placing it in the mixture so as to expose as much surface as possible in order to get a quick start, but carefully avoid pressing the material, and use a wire to fill up close to the fuse. A slow start often spoils the experiment, through the upper end of the cylinder becoming nearly filled up with potassic chloride, etc.

By paying attention to such details, and following the method recommended, the apparatus yields very satisfactory results with bituminous and semi-bituminous coals.—*Chemical News*.

EXPLOSION AS AN UNKNOWN FIRE HAZARD.

WORDS pass along with meanings which are simple conventionalities, marking current opinions, knowledge, fancies, and misjudgments. They attain to new accretions of import as knowledge advances or opinions change, and they are applied now to one set of ideas, now to another. Hence there is nothing truer than the saying, "definitions are never complete." The term explosion in its original introduction denoted the making of a noise; it grew to comprehend the idea of force accompanied with violent outburst; it is advancing to a stage in which it implies combustion as associated with destruction, yet somewhat distinct from the abstract idea of the resolution of any form of matter into its elementary constituents. The term, however, as yet takes in the idea of combustion as a decomposition in but a very limited degree, and it may be said to be wavering at the line between expansion and dissociation.

Strictly, in insurance, fire and explosion are different phenomena. A policy insuring against fire-loss does not insure against loss by explosion. It thereby enforces a distinction which exists, or did exist, in the popular mind; and fire, in an insurance sense, as distinct from explosion, was accurately defined by Justice McIlvaine, of the Supreme Court of Ohio (1872), in the case of the Union Insurance Company vs. Forte, *i. e.*, an explosion was a remote cause of loss and not the proximate cause, when the fire was a burning of a gas jet which did not destroy, though the explosion caused by the burning gas-jet did destroy. Earlier than this decision, however (in 1852), Justice Cushing, of the Supreme Court of Massachusetts, in *Scripture vs. Lowell Mutual Fire Insurance Company*, somewhat anticipated later definition, and pronounced for the liability of the underwriter where all damage by the explosion involves the ignition and burning of the agent of explosion. That is, for example, the insurer is liable for damage caused by an explosion from gunpowder, but not for an explosion from steam. The Massachusetts Judge did not conceive any distinction as to fire-loss between the instantaneous burning of a barrel of gunpowder and the slower burning of a barrel of sulphur, and insurance fire-loss is not to be interpreted legally by thermo-dynamics nor thermo-chemistry. While the legal principles are as yet unsettled, the tenor of current decisions may be summed up as follows: If explosion cause fire, and fire cause loss, it is a loss by fire as proximate cause; and if fire cause explosion, and explosion cause loss, it is a loss by fire as efficient cause. Smoke, an imperfect combustion, damages, in an insurance sense, as well as flame, which is perfect combustion; and where there is concurrence of expanding air with expanding combustion, the law settles on the basis of a common account. It's all "heat as a mode of motion."

Explosions are the resultants of elemental gases, vaporization, combination, contact of different substances, as well as of the specifically named explosives. With new processes in manufacture, involving chemical and mechanical transformations, and other uses of new substances and new uses of old substances, explosions increase. The flour-dust of the miller, the starch-dust of the confectioner, increase in fineness and quantity, and they explode; so does the hop-dust of the brewer. In 1844, for the first time, Professors Faraday and Lyell, employed by the British government, discovered that explosion in bituminous coal mines was the quickening of the comparatively slow burning of the "fire-damp" by the almost instantaneous combustion of the fine coal-dust present in the mines. The flyings of the cotton mill do not explode, but flame passes through them with a rapidity almost instantaneous, yet not sufficient to exert the pressure which explodes; the dust of the wood planer and sawer only as yet makes sudden puffs without detonating force. Naphtha vapor and benzine vapor are getting into all places. One of the latest introductions is naphtha extracting oil from linseed, and then volatilized by steam superheated to 400° F. This combination reminds us, as to effectiveness, of the combination at the recent Kansas City fire, when cans of gunpowder and barrels of coal oil both went up together.

But it is the unsuspected causes of explosion which make the great trouble, and prominent among these is conflagration as itself the cause of explosion, and such explosion may develop gases which are non-supporters of combustion as well as those which are inflammable. You throw table salt down a blazing chimney to set free the flame-suppressing hydrochloric acid, you discharge a loaded gun up a blazing chimney to put out the fire by another agency; still the salt, with certain combinations, may be explosive, a resinous vapor may be combustible in a hydrochloric atmosphere, and gunpowder isn't harmless when thrown upon a blaze—in fact, our common fire-extinguisher, water, has its explosive incidences as liquid as well as vapor.

Gases explosive in association may be set free by the temperature of a burning building and get together. In respect to the old conundrum, "Will saltpetre explode?" Mr. A. A. Hayes, Prof. Silliman, and Dr. Hare's views were, as to the explosions in the New York fire of 1845, that in a closed building having niter in one part and shellac or other resinous material in another, the gaseous oxygen generated from the niter and the carburated hydrogen from the resins mingling by degrees would at length constitute an explosive mixture. A brief consideration of specific explosives uniting may serve to illustrate this phase of the subject.

Though the explosion of gunpowder is the result of a chemical change whereby carbonic acid gas at high tension is evolved (due to the saltpetre and the charcoal), the effect and rapidity of action are greatly promoted by the addition of sulphur. On the contrary, dynamite, now so important, and various similar explosives, are but mixtures of nitro-glycerine with earthy substances, in order to diminish and make more manageable the development of the rending force of the base. The explosive power of any substance is the

pressure it exerts on all parts of the space containing it at the instant of explosion, and is measured by comparing the heat disengaged with the volume of gas emitted, and with the rapidity of chemical action. In the case of gunpowder, the proper manipulation and division of the grains is important, because favoring rapid deflagration; but in a purely chemical explosion, each separate molecule is an explosive, and the reaction passes from the interior of one to the interior of another, suddenly driving the atoms much further apart than their naturally infinitesimal vibrations.

Purely chemical explosives like nitro-glycerine, gun-cotton, the picrites, and the fulminates, present a terrible danger from the unknown mode of the new union of atoms, and reaction of the particles within themselves, in spontaneous explosions happening in irregular manner. Some curious circumstances attend the manufacture and use of gun-cotton,* nitro-glycerine, and dynamite. Baron von Link, in his system of the artillery use of gun-cotton, diminishes the danger of sudden explosion by twisting the prepared cotton into cords or weaving it into cloth, thereby securing a more uniform density. Mr. Abel's mode of making gun-cotton, which explosive is now used more than any other by the British government, includes drying the damp prepared cotton upon hot plates, *freely open to the air*. If ignited by a flame, however, in an unconfined place, gun-cotton only burns with a strong blaze, but if *confined* where the temperature reaches 340° F., it explodes with terrific violence. Somewhat similar is the action of nitro-glycerine and dynamite, which simply *burn* if ignited in the open air, while the same substance will *explode* through a very slight concussion or by the application of the electric spark; a red-hot iron, also, if applied, will explode them when a flame will not. With care, nitro-glycerine can be kept many years without deterioration; and it has been heated in a sand-bath to 80° C. for a whole day without explosion or alteration. One curious experiment is deserving of mention: If a broad-headed nail be partly driven into pine wood, and then some pieces of dynamite placed on the head of the nail, the latter may be struck hard blows with a wooden mallet without exploding the dynamite *so long as the nail will continue to enter the wood*.

Taking gunpowder as the unit, picrate of potash (picric acid and potassium) has five times more force, gun-cotton seven and a half times, and nitro-glycerine ten times more force. There are others still more powerful, but less known and used, and some explosives are quite uncontrollable and useless.

But the particular object of these remarks is to refer to articles of merchandise non-explosive under general conditions, but so in particular circumstances, as the two fire-extinguishers, water and salt, are explosive under given conditions. The memorable fire which, in July, 1850, destroyed three hundred buildings in Philadelphia, upon Delaware avenue, Water, Front, and Vine streets, was largely extended by explosions of possibly concealed or unknown materials, the presence of the generally recognized explosives being denied by the owners of the properties.

"The germ of the first knowledge of an explosive was probably the accidental discovery, ages ago, of the deflagrating property of the natural saltpetre *when in contact with the incandescent charcoal*."† Although much manipulation is deemed necessary to form the close mechanical mixture of the materials of gunpowder, it has never been proved that such intimate previous union is necessary to precede the chemical reaction causing explosion; indeed, some explosions in powder works, before the mixture of the materials, or just at its commencement, seem to point to the contrary. It is also certain that in the manufacture of gunpowder the usual nitrate of potassium (saltpetre) can be replaced by the nitrates of soda, baryta, and ammonia, also by the chloride of potassium; charcoal by sawdust, tan, resin, and starch; and though a substitute for sulphur is not easily found, the latter, or a similar substance, is not an absolute necessity in the composition of gunpowder.‡

The generally received theory of the chemical action which makes gunpowder explosive is that it is due to the superior affinity of the oxygen of the niter (KNO₃) for the carbon of the charcoal, and the production of carbonic acid gas (CO₂) and carbonic oxide (CO) suddenly and in great volume. The latter extinguishes flame as well as the former, unless its own flammability is supported by the oxygen of the atmosphere until the degree of oxygenation CO is reached. Considering that water (H₂O) is composed of two volumes of hydrogen and one of oxygen, and that under an enormously high temperature and the excessive affinity of oxygen gas for potassium or sodium (freed from nitrate union), dissociation of the water may be possible, aided by its being in the form of spray and steam, we would hesitate to deny that an explosive union of suitable crude salts could occur during the burning of a building containing them when water for extinguishment was put on. Any one who has seen the brilliance with which potassium and sodium burn upon water can easily imagine how such strong affinity of oxygen for these substances might aid in severing its union in water in their presence and under extraordinary heat. It might be safe so say that the presence of water under very high temperature may be as aidful to form an explosive among such salts as have been named, as sulphur is for the rapid combustion of gunpowder.

In the review for August, 1863 (Saltpetre Deflagrations in Burning Buildings and Vessels—Water as an Explosive Agency), it was shown that Mr. Boyden's experiments in 1861–62 proved that explosions would occur when water was put upon niter heated alone, and stronger explosion from niter, drywood, and sulphur; also explosion when melted niter was poured on water. The following points we reproduce for comparison: If common salt be heated separately to a bright heat, and water at 150° F. poured on it, an explosion will occur. Niter mixed with common salt, placed upon burning charcoal, and water added, produce a stronger explosion than salt alone. Heating caustic potash to a white heat, and adding *warm or hot water*, produces explosion. At a Boston fire small explosions were observed upon water touching culinary salt highly heated. Anthracite coal and niter heated in a crucible exploded when *sea water* was poured on them.

The production of explosion by the putting of water on nitrate of potassium and chloride of sodium arises from the union, at high temperature, of the oxygen of the water with the potash and soda. Of the three liberated gases, hydrogen only is inflammable, and the other two suffocative of flame; but together the nitrogen and chlorine are not to be under-

valued, for chloride of nitrogen is ranked as the most terrible and unmanageable of all explosives. Chlorine is a great water separator, but in the present case its affinity for hydrogen would result in hydrochloric acid, a fire extinguisher.

What happens in chemical experiment may be developed on a large scale in burning grocery, drug, or drysalter's stores, when great quantities of materials, such as just mentioned, including common salt, almost always present, are heated most intensely, and then subjected to the action of water in heavy dashes, or in form of spray or steam.

Picric acid, the nature of which we have several times previously mentioned, and which explodes at 600° F. (only 28° above gunpowder), may also be an element in such explosions during fires. Its salts form, in combinations, various powerful explosives, much exceeding gunpowder in force; and they have been used to a considerable extent in Europe. Picric acid, now much employed by manufacturers and dyers for obtaining a yellow color, is always kept in store largely by drysalter and druggists, and generally by dyers, but in smaller quantity.

In a very destructive fire which occurred in Liverpool, Eng., in October, 1874, involving the loss of several "fire-proof" stores, repeated explosions of the vapor of turpentine rent ponderous brick arched vaults, and exposed to the flames stocks of cotton, etc., in the stories above. This conflagration was started by the carelessness of an *employé* in snuffing a tallow candle with his fingers and throwing the burning snuff into the open bung-hole of a sample barrel of turpentine, of which liquid there were many hundreds of barrels on storage in the buildings. Turpentine vapor united with chlorine gas may not produce explosion, but by spreading flames almost instantly throughout the burning buildings, such burnings have practically equaled, if not excelled, explosions, which may sometimes be fire-extinguishers. In such cases detonation may be prevented by there being ample space to receive the suddenly ignited vapor, lessening the tension of it, but carrying the flames much more rapidly than otherwise to inflammable materials at great distance.

If disastrous results have arisen from the vapor of turpentine as a fire spreader in vaults without windows, it is possible that if a quantity of hot water were suddenly converted into steam in closely confined spaces, effects of pressure might be observed, less destructive perhaps, but resembling those which other explosives might produce. If the immense temperature attained in some conflagrations be considered—sufficient to melt iron and vitrify brick—it is possible to conceive of water as being instantly converted into steam. Even a very small quantity of water thus expanded could produce most disastrous results. While such formation of steam, if it happened, would certainly extinguish most flames in direct contact, the general phenomena shown would be explosive.

A curious circumstance occurred at the Broad street (N. Y.) fire in 1845, previously mentioned. The fire extended through to Broadway, and almost to Bowling Green. A shock like a dull explosion was heard, and by many this was attributed to the effects of gunpowder and saltpetre. Several firemen were, at the moment of the shock, on the roof of the burning building, when the whole roof was suddenly raised and then let down into the street, carrying the men with it uninjured. One of the firemen described the sensation "as if the roof had been first hoisted up and then squashed down." *Query*: Was this like the common lifting and falling back of the loose lid of a tea-kettle containing boiling water? Was it from steam—at a low pressure perhaps—seeking vent through the roof in like manner to the raising of the kettle-lid? Without dilating on this part of the subject, we mention it as a possible cause of minor explosions—doubtless to become better known in future. It may even be that explosions happening from steam acting in close spaces may have been attributed to gunpowder, or to niter and other salts, separately, but suddenly caused to combine in chemical reaction.—*American Exchange and Review*.

CARBON.—SYMBOL C.—COMBINING WEIGHT 12.

By T. A. POOLEY, B.Sc., F.C.S.

THIS element, which next deserves our attention, is one of great importance and wide distribution; it occurs in nature in both the free and the combined states, and the number of compounds which it forms with other elements is very large. Unlike the previous elementary bodies we have studied, carbon is only known to us in the solid form when free, although many of its combinations are gaseous at the ordinary temperature and pressure. Carbon is known to exist in several different physical states, thus illustrating what chemists call *allotropism*, which means that substances of identical chemical composition sometimes possess altogether different outward and physical appearances. Thus the three states in which pure carbon exists, viz., diamond, graphite, or plumbago, and charcoal are as different as possible, and yet chemically they are all exactly the same substance. The diamond is the purest carbon, and occurs in the crystalline form known as a regular octahedron; the diamond is one of the hardest substances known, and is therefore, utilized for cutting glass; it has also a very high specific gravity, namely, 3.5, which means that it is three and a half times heavier than water, and it is far heavier than any of the other allotropic modifications of carbon. Graphite or plumbago, the second form in which carbon occurs, is widely distributed in nature, and the finer qualities are known as black lead, although no lead enters into their composition, as they are composed of carbon almost as pure as the diamond; the specific gravity of graphite is only 2.3. Charcoal, the third allotropic modification of carbon, is by far the most common, and is formed by the natural or artificial disintegration of organic matters by heat; we thus have formed wood charcoal, animal charcoal, lamp-black, and coke, all produced by artificial means, and we may also class with these coal, which is a natural product, and which contains from 85 to 95 per cent. of pure carbon.

Wood charcoal is made by heating wood in closed vessels or in large masses, when all the hydrogen, oxygen, and nitrogen are expelled in the gaseous state, and the carbon is left mixed with the mineral constituents of the wood; this form of carbon is very porous and light, and is used in a number of industrial processes.

Animal charcoal, as its name implies, is the carbonaceous residue left on heating any animal matter in a retort; and contains, in addition to the carbon, a large proportion of phosphates and other mineral salts, which, however, can be extracted by dilute acids. Animal charcoal possesses to a remarkable degree the property of removing color from solutions of animal and vegetable substances, and it is used for this purpose to a large extent by sugar refiners, who thus

* The purest gun-cotton may be regarded as a cellulose, in which three atoms of hydrogen are replaced by three molecules of peroxide of nitrogen.

† Encyclopædia Britannica, new edition, viii, p. 806.

‡ Vide Abel's Experiments in Gunpowder, as detailed in Phil. Trans. Roy. Soc., 1874.—*Vide also Bull. Soc. d'Encouragement*, Nov., 1880, p. 633, *Sur les Explosives*.