ELECTRICAL SECTION.

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The Electrolysis of Water.*

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The title of this discussion may be a misnomer, because, as is well known, the purest water which can be obtained is practically a non-conductor of electricity, and therefore cannot be electrolysed. In elementary lectures on chemistry, perhaps not more than a quarter century ago, the professor explained, when demonstrating the decomposition of water by electricity, that some sulphuric acid was added to the water so as to make it conducting, and then the current flowed and water was decomposed,—but the sulphuric acid had no part or function in that decomposition. The explanation may probably still be regarded as about half true; at least, the final result of passing the current is to give us the constituents of water in the proportion in which they exist in water, and so the sum total of all the actions taking place, simple or as complex as we wish to imagine them, is *the decomposition of water*.

We may conveniently divide our subject into its experimental, its theoretical and its applied phases. The experimental part will deal with the conditions under which water is decomposed, and the facts concerning the electrolytic phenomena; the theoretical part will concern itself with the why and wherefore of these phenomena, and the question of the theoretical connection between the work done in electrolysis and the electric energy consumed; the applied section will treat of the commercial applications of these reactions, the types of apparatus used and the uses of the gases produced.

*Review to date.—J. W. R.

I.

The phenomena of the electrolysis of aqueous solutions which yield oxygen and hydrogen, have been studied for a century. The materials which are dissolved in or mixed with pure water in order that the current may be transmitted are usually sulphuric acid, caustic soda or potash, sulphate of sodium or potassium or carbonate of sodium or potassium. Besides these, many oxygen acids and salts, particularly oxygen salts of the alkaline metals, exercise the desired function, and give the desired gaseous products when worked between proper limits of temperature, concentration of solution and current density; outside of these limits, results different from the normal or obtained.

CONSTANTS OF DECOMPOSITION.

The following constants are those most generally accepted:

WATER:

Molecular weight.....	18
Specific gravity.....	1
1 liter weighs.....	1000 grams.
1 coulomb decomposes.....	0.093926 m. grams
1 ampere-hour decomposes.....	0.3351 grams.

OXYGEN:

Atomic weight.....	16.
Molecular weight, O ²	32.
Specific gravity (Air = 1).....	1.10563
Specific gravity (hydrogen = 1).....	16.
1 liter at 0 deg. C. and 760 m. m. pressure weighs.....	1.4303 grams.
1 gram occupies.....	699 cub. centimeters.
1 coulomb liberates 0.0829 m.grams =	0.058 cub. centimeters.
1 ampere-hour liberates 0.298 gms. =	207.2 cub. centimeters

HYDROGEN:

Atomic weight.....	1
Molecular weight.....	2
Specific gravity (Air = 1).....	0.0696
1 liter at 0 deg. C. and 760 m. m. pressure weighs.....	0.09 grams.
1 gram occupies.....	11.11 liters.
1 coulomb sets free 0.01036 m. gms =	0.1150 cub. centimeters.
1 ampere-hour sets free 0.0373 gms. =	414.4 cub. centimeters.

MIXED GASES (detonating gas).

Specific gravity (Air = 1)..... =	0.4150
1 liter at 0 deg. C. and 760 m. m. pressure weighs.....	0.563 gram.
1 gram equals.....	1.865 liters.
1 Coulomb sets free 0.0933 m. gms. =	0.1725 cub. centimeters.
1 ampere-hour sets free 0.3353 gms. =	621.6 cub. centimeters.

PHENOMENA DURING ELECTROLYSIS.

It is to be assumed that the electrodes used are such as will be permanent, not being attacked by the electrolyte or the products of its decomposition. Such electrodes are not easy to find, because at the anode the tendency is strongly towards oxidation, at the cathode strongly towards reduction. Even carbon (graphite) used as anode seems to suffer some oxidation, platinum is fairly fixed and gold seems to be practically unattacked. A lead anode is at first superficially oxidized to PbO^2 , and having reached the limit of its oxidation remains afterwards unaffected. As cathode, gold is recommended strongly by Mr. Carl Hering as being unaffected; platinum slowly blackens, evidently from the superficial formation of a combination with hydrogen and its subsequent decomposition, leaving spongy platinum; lead cathodes blacken and disintegrate superficially, powdered lead falling to the bottom. The action is evidently analogous to that of platinum, but more active. Iron makes a satisfactory cathode in commercial work.

The electrolytes used in the laboratory are usually dilute sulphuric acid or caustic soda. In the case of dilute acid it is imperative to use as anode a material not attacked by the acid alone, and not forming soluble or insoluble sulphate by the action of the current; the cathode may be any material not attacked by the acid alone, and not occluding, dissolving or combining too actively with hydrogen. In the case of caustic alkali, the anode must be insoluble in caustic alkali alone, and not forming soluble or insoluble oxide or hydrate under the action of the current; the cathode may be any material not attacked by caustic alkali alone, strong or dilute, and not forming easily alloys with sodium or the alkaline metals or occluding or dissolving hydrogen gas. In case the cathode can unite easily with sodium or an alkaline metal, we find that it disintegrates rapidly, evidently because of the successive formation of the

alloy and its subsequent decomposition by the water of the electrolyte.

The electrolyte is warmed by the current passing through it, the heating effect being proportional to its resistance as an electric conductor and to the square of the current passing (Ohm's laws as applied to any conductor). The resistance of the electrolytic cell can be calculated from its dimensions and the known resistivity of the solution used, and thus the energy absorbed for any known current, and converted into heat, can be calculated. Or, the cell can be placed in a calorimeter, and the heating effect of the current measured directly; or, the cell itself may be regarded as a calorimeter, the weights and specific heats of electrolyte, electrodes and thermometer obtained, and from the rate at which its temperature begins to rise immediately on starting the current, the number of calories liberated in it per unit of time can be reckoned up. This latter method, though little used, is probably the simplest, easiest and most accurate of the methods proposed.

Energy disappears in the cell in addition to that which re-appears as heating effect such as is calculated or observed above, and the amount thus disappearing is measured by the difference between the total energy expended by the current and the amount of energy appearing as heat. For every watt of electric energy used, the energy expended may be taken as 0.2385 gram calories. If the electrical energy were all converted into heat, the calorimetric effect which would be measured would agree with the above calculation, *i. e.*, 0.2385 gram calories for every watt-second. But investigation and measurement will show the calorimetric effect to correspond to only a fraction of the energy of the current, generally about one-half. A large part of the energy of the current has therefore disappeared in some other function, and that function is the chemical work of liberating the hydrogen and oxygen from water. We are on safe ground when we say that the electric energy thus disappearing in chemical work is exactly equivalent in amount (measured either in thermal or mechanical units of work) to the chemical work done.

The amount of work necessary to split up nine grams of water into eight grams of oxygen gas and one gram of hydrogen gas is, as nearly as is at present known, represented by

34,500 gram calories. An ampere-hour liberates 0.0373 grams of hydrogen, and therefore does an amount of chemical work represented by $0.0373 \times 34,500 = 1287$ gram calories. But, a watt-hour of electric energy represents only $0.2385 \times 60 \times 60 = 859$ gram calories, and therefore we need at least $1287 \div 859 = 1.5$ watt-hours to liberate 0.0373 grams of hydrogen. An ampere-hour of current flow can only represent 1.5 watt-hours of electric energy when it falls in potential 1.5 volts; and therefore we have the necessity that a continuous drop of potential of 1.5 volts must be provided for in order to decompose water, that is, for the chemical work of decomposition.

The work absorbed in the decomposition may also be observed in other ways. If the size of the electrodes be taken, their distance apart, and the conductivity of the acid solution used be known, the heating effect of the current in overcoming the ohmic resistance of the bath can be calculated by the well-known formula $\text{Watts} = RA^2$, and thus the heating effect calculated instead of being directly determined calorimetrically. Agreements will be found to be good, if carefully determined in both ways. The total energy of the current employed being $(V \times A)$ Watts, the difference between $(V \times A)$ and RA^2 is the work done in chemical decomposition. This will be found, if carefully determined, to agree closely with the chemical work done, viz., to amount to 34,500 gram calories per gram of hydrogen liberated, or 2.61 gram calories per cubic centimeter of hydrogen gas, or 1287 gram calories per ampere-hour used.

The rate at which decomposition takes place in any given apparatus depends, of course, on the number of amperes passing through it, but this, in its turn, depends on the voltage maintained across the terminals of the bath, although not in a direct proportion. If a voltage of 1.5 is applied, it will be found that practically no current flows, evidently because that much potential is required to start electrolysis and to provide the energy of decomposition. When a slightly higher voltage is applied current begins to flow, but no visible evolution of gas occurs, with ordinary-sized electrodes, until the solution becomes saturated with hydrogen and oxygen gases. With two to three volts liberal evolution of gases usually results.

When the apparatus is placed under pressure more amperes will pass through for the same applied voltage, or the same

number of amperes can be passed through at a smaller applied voltage. If the latter condition prevails, it will be found that the output of gas per ampere-hour is decreased, and also that the calorimetric heating effect in the cell decreases. The cause of these phenomena is that under pressure considerably more hydrogen and oxygen dissolve in the electrolyte, resulting on the one hand in increasing the amount of re-combination of the hydrogen with dissolved oxygen at the cathode and of oxygen with dissolved hydrogen at the anode, and on the other hand in increasing the conductance of the solution and so diminishing its resistance and the consequent heating effect. Whether the amount of gases evolved per watt-hour is increased or diminished I do not know; this field is ripe for considerable experimentation, since no accurate information on the subject has as yet been published.

The amounts of gases liberated and collected in well-designed apparatus are always less than the theoretical amounts, the loss in oxygen being usually greater than that of hydrogen. Losses of both gases caused by re-combination to form water diminish the amounts of each collected, but diminish them in the same proportion as they are being evolved; that is, for example, if recombination is five per cent., then one-twentieth of each gas is being lost. Temporary loss of hydrogen may occur at the beginning of electrolysis by its absorption by the cathode, which, if of palladium, for instance, could absorb 700 times its volume of hydrogen before becoming saturated; but, such absorption has a limit, and after the cathode has absorbed all the hydrogen it can the action must cease, for good, and thenceforth the normal amount of hydrogen be liberated and no more be lost by cathodic absorption. At the anode, oxygen may similarly be lost at first by superficial oxidation of the anode, but this action usually ceases as soon as a continuous coating of oxide is formed. A lead anode, for instance, in dilute sulphuric acid, is quickly coated with brown per-oxide, PbO_2 , with consequent disappearance of oxygen, but the oxidation soon ceases, and no more is thereafter so lost.

Another cause of disappearance of oxygen is the formation of hydrogen peroxide or of persulphuric acid or persulphates in the electrolyte. The amounts to which these may form depends primarily on the strength of acid used, temperature of

the electrolyte and the current density used. Elbs and Schönherr* determined the loss of oxygen in the formation of persulphuric acid, at ordinary temperatures, to be:

Specific gravity of sulphuric acid used	Percentage loss of oxygen using current densities per sq. meter of		
	5 amp.	50 amp.	100 amp.
I.15			7.0
I.20		4.4	20.9
I.25		29.3	43.5
I.30	1.8	47.2	51.6
I.35	3.9	60.5	71.3
I.40	23.0	67.7	75.7

In commercial practice the strength of acid used and current density are so adjusted as to fall well within the minimum values of above table, but in laboratory tests these conditions may easily be exceeded, and large loss of energy ensue. It should be observed, however, that when the persulphuric acid in solution reaches the cathode, by circulation or diffusion, it is easily reduced by hydrogen or combines with hydrogen to form ordinary sulphuric acid, thus causing an amount of hydrogen to disappear equivalent to the amount of oxygen absorbed or lost in the formation of the persulphuric acid. Variations in temperature have a very great effect on the possible loss of oxygen by formation of persulphuric acid. Under conditions where 32.2 per cent. of the oxygen was so lost, at 27°C, the raising of the temperature to 60°C. reduced this loss to 0.59 per cent. Running the electrolytic cell warm is one of the best means for reducing this source of loss of oxygen, and of hydrogen also.

Loss of oxygen, very similar in its mechanism to loss by formation of persulphuric acid, may result under special conditions from anodic formation of hydrogen peroxide. The condition favorable for its formation, however, is the use of acid over 60 per cent. strong, and this is rarely or never used in commercial practice. If it were, important losses of oxygen, and also of hydrogen, would occur in the formation and subsequent reduction of the peroxide, in a manner exactly analogous to the formation and reduction of persulphuric acid.

A decreased output of oxygen, measured by volume, occurs

*Zeitschrift für Elektrochemie 1, 417,468; 2, 162,245 (1895).

most frequently from the formation and liberation of ozone, every two volumes of which is formed from three volumes of oxygen gas, so that for every liter of ozone formed the volume of oxygen appears to diminish half a liter. Conditions favoring the formation of ozone are: increased current density, increased concentration of acid up to fifty per cent., low temperature; working with low-current density, weak acid, and warm solution, the amount of ozone formed is very small. For many practical purposes, the formation of ozone does not impair but rather increases the efficiency of the oxygen gas with which it is mixed, which thus gains in intensity of action more than it loses from reduction of volume.

II.

Concerning the *modus operandi* of the evolution of hydrogen and oxygen gases from dilute sulphuric acid or caustic soda solutions, taken as the most important examples, there are many theories. The oldest is, that the acid or salt makes the water conducting, and it (the water) is then decomposed by the current. Another view is that the acid or salt only is decomposed by the current, and that the water present is attacked by the products of this decomposition and thus chemically decomposed. A third view is that the *mixture* of acid and water, or salt and water, conducts the current, and that this *mixture* is the thing which is decomposed, or resolved by the current, into hydrogen, oxygen and acid or salt. These three views may be expressed by the reactions:

The first theory is untenable, because it is no less logical to speak of the water making sulphuric acid conducting than it is to say that acid makes the water conducting. Pure water is practically a non-conductor, yet a small amount of sulphuric acid mixed with it gives a solution which conducts excellently; pure anhydrous sulphuric acid is practically a non-conductor, yet a small amount of water mixed with it gives a solution which conducts excellently. The only logical and proper conclusion is that the conductivity is a function of the mixing, and that it is the mixture which must be regarded as the conductor of electricity. An analogous case would be the fusibility of silicates of lime. In a blast furnace, pure silica is practically infusible; so is pure lime. Mix them, and the mixture fuses easily in the furnace. No metallurgist says that the lime makes the silica a fusible substance, or that the silica makes the lime a fusible substance; but it is rightly said that the mixture or combination of the two is a fusible substance, and that the fusibility is a property of the mixture only. Likewise to borrow the logical process, the conductivity of aqueous sulphuric acid is a property inherent in and characteristic of the *mixture* and not of either ingredient.

The second theory regards the acid or salt constituent as being the bearer of the electric current, *i. e.*, the conductive constituent of the electrolyte, and as being the constituent primarily decomposed by the current. The first part of this statement is untenable for exactly the same reason as the first theory, which has just been explained at length. If it were allowable or necessary to regard the current as being carried through the electrolyte by the acid or salt constituent alone, then we could acknowledge the force of the corollary, that it alone was primarily decomposed by the current, and that the evolution of hydrogen and oxygen gases was really secondary. But, since the first statement is untenable, the second loses its element of necessity, and must rest upon other grounds of proof or support, if they can be found. The current not being carried by either constituent of the electrolyte alone, but by the electrolyte as a whole, the assumption of the primary and

secondary actions which this second theory requires becomes superfluous and is in fact illogical. Why should the current, after arriving at the electrodes through the medium of the mixed electrolyte, and not through either constituent of it alone, end by decomposing the acid into two constituents which then react upon the water to liberate its constituents? If the current is able to decompose the acid, and the constituents of the acid are able to decompose the water, why not assert at once that the current decomposes the water at the electrodes, and cut out the illogical and unnecessary complication which the theory under discussion involves?

The third theory stands upon the observed fact that only the mixture of acid, or salt, and water conducts; that the *mixture* is, therefore, the conducting body and not either constituent of it; that the current is carried from one electrode to the other as a purely physical operation by the electrolyte or mixture *as a whole*, and not by either of the constituents which united to form it; that, in fact, the constituents of the electrolyte have lost their intrinsic individuality on becoming mixed, having combined with each other or mutually dissolved each other in such manner as to constitute a resultant body or material which contains neither of its ingredients *per se*, but is a compound material which must be considered a unit for the purpose of electric conductivity. This body conducting the current as a whole, from one electrode to the other, the electrochemical action, the chemical decomposition by the electric influence, takes place only at the surface of the electrodes, where the current passes from electrolyte to electrode, and there decomposes the compound body, the electrolyte. In the case of dilute sulphuric acid, the electrolyte contains hydrogen, sulphur and oxygen, in amounts corresponding to a mixture of H^2So^4 and H^2O in some proportion. Calling the electrolyte $\text{H}^2\text{SO}^{4+x} \cdot \text{H}^2\text{O}$, we can write this empirically as

and the electrochemical action is expressable by the formula

or, perhaps more logically, for the purposes of comparison,

and in caustic soda solution, $\text{NaOH} \cdot x\text{H}_2\text{O}$,

On these principles, neither the water nor the acid or salt conduct the current through the bath, but the *mixture* of the two; also, neither the water nor the acid or salt is decomposed, but the *mixture*,—the action of the current being to deprive it (the mixture) of hydrogen and oxygen. These views express more exactly the phenomena, agree more closely with the logic of the situation and the observed facts, and contain no unfounded or unreasonable assumptions.

III.

In Engelhardt's monograph on "Die Elektrolyse des Wassers," he classifies the commercial apparatus used under three heads, according to whether they are used for obtaining oxygen alone, or both oxygen and hydrogen mixed together, or both oxygen and hydrogen separately led off. This classification is, however, a matter of mechanical detail alone, with which we will not concern ourselves.

The real commercial classification is into apparatus with diaphragms and without diaphragms to separate the gaseous products from each other, and further into apparatus with unit couples in multiple and those in series. The multiple or series types have been used according to whether it was desired to work with low or high voltage. Where direct-current electric lighting circuits, of say 110 volts, are available, it is very convenient to have series apparatus adapted to work on those circuits; even when dynamos are installed, it is well to consider that lighting generators of standard voltage are to be found more plentifully on the market, and usually sell at lower prices, than specially-constructed low-voltage machines.

Without further classification we will take up the description of the most successful types of commercial apparatus.

EARLIER FORMS OF APPARATUS.

D'Arsonval, in 1885, was perhaps the first to instal a plant for furnishing oxygen electrically in the laboratory. He used 30

per cent. caustic soda solution as electrolyte, cylindrical sheet-iron electrodes, a current density of two amperes per square decimeter, and enclosed the anode in a woolen bag, to serve as a diaphragm. Only the oxygen was saved. The apparatus used sixty amperes, furnished some 100 to 150 liters of oxygen daily, and was in use several years.

Latchinoff used an asbestos cloth partition, ten per cent. caustic soda solution, iron electrodes, 3.5 amperes per square d.m., and 2.5 volts working tension; or with a five to fifteen per cent. sulphuric acid solution used leads anodes and carbon cathodes. In his first apparatus, the units were all in parallel, but afterwards he used series electrodes, the one side of an electrode acting as an anode and the other as a cathode; a series of forty was used on a normal lighting circuit, with current density of ten amperes per square d.m., and parchment partitions between the electrodes to separate the gases. *Latchinoff* was also the first to carry out the decomposition under pressure, using a strong iron vessel as electrolyser, and by an ingenious system

of floating valves keeping the pressure of the two gases equal in the apparatus. Fig. 1 shows this apparatus, the action of which will be evident from a short inspection.

Dr. O. Schmidt, in 1899, perfected this series type of apparatus, using bi-polar electrodes, and constructed very practical and workable apparatus in the shape of a filter press, adapted to work directly on 110 volt direct-current lighting circuits. (Fig. 2.)

Each frame in the press contains an iron electrode, which acts as a double-pole (bi-polar) electrode, sheets of asbestos cloth held between the frames acting as partitions, reinforced with rubber on the edges for making tight joints. The electrolyte is ten per cent. solution of potassium carbonate, fill-

Fig. 1.

ed into the apparatus through the stand-pipe on the right, which communicates with all the compartments through holes in the frames similar to the usual filter-press construction. The gases evolved escape by similar passages into the cylinders on the left end, where they separate from the electrolyte and pass upwards, while the electrolyte, dragged by the gas bubbles, flows downwards back into the apparatus, thus maintaining an efficient circulation. With forty plates, about 2.5 volts is absorbed in each cell, using a current density of about two amperes per square decimeter. The oxygen obtained is 97 per

Fig. 2.

cent. pure, which is raised to 99.985, if desired for medical purposes, by passing through a tube filled with platinized asbestos heated to 100 deg. C. Five sizes of apparatus are made, adapted for 15, 30, 60, 100 or 150 amperes, and each may use 25, 40 or 80 frames, for working on 65, 110 or 220 volt circuits. The prices vary from \$450 down to \$130 per kilowatt of power capacity, according to the size of the machine, and it has been shown that plants in Europe can make good profits when selling compressed oxygen gas at \$1 and hydrogen at \$0.31 per cubic meter. The Machinen Fabrik Oerlikon in Zurich make

these apparatus, and nearly a score of plants thus equipped are in operation in Europe.

Schoop, in 1900, devised an apparatus with non-conducting and non-porous partitions, which has gone into considerable commercial use. Fig. 3 shows the section of the apparatus, where *a a* are the tubular electrodes of sheet hard-lead, enclosed by glass or clay suspended tubes *c*, which are perforated at their lower ends; the electrode surface is further increased by fine hard-lead or iron wires hung inside the tubular electrodes, the latter being perforated above the level of the electrolyte in order to let the internally-generated gas escape. Each cylinder contains two anodes and two cathodes. When alkaline electrolyte is

Fig. 3.

used and iron electrodes, the working voltage is 2.25; when sulphuric acid of density 1.235 is used, with hard-lead electrodes, the working voltage is 3.6 to 3.9. The output is given as 68 litres of oxygen and 136 litres of hydrogen per electrical horsepower hour. Fig. 4 shows an experimental plant on this system installed at the Cologne Accumulator works of G. Hagen at Kalk on the Rhine. Only one size of apparatus has so far been made, suitable for 175 amperes and costing \$69.25 per unit.

Fig. 4.

Fig. 6.

Garuti, in 1892, introduced a new electrolytic principle into these apparatus for the decomposition of water. He uses a

Fig. 5.

nearly complete metallic partition between the electrodes, and avoids the evolution of gases on this partition by keeping the working voltage between the electrodes below three volts. It will be easily understood (now that Garuti has thought of utilizing the principle), that the metallic partition can only act as an intermediate or bi-polar electrode by virtue of the current entering and leaving it; but this would make two decompositions between the original electrodes, necessitating an absorption of $2 \times 1.5 = 3$ volts in decomposition. As long as the working voltage is kept below 3, the partition must act merely as a partition, the same as a non-conducting partition. The advantage gained is in the simplicity and economy of making the partitions of sheet metal instead of burnt clay, rubber, glass, etc.

Garuti devised many modifications in the de-

tails of his cells, of which Fig 5 may represent the most recent. The original forms made of sheet lead (using dilute sulphuric acid electrolyte) got out of shape too easily, and were replaced by sheet-iron apparatus, using caustic soda solution. The electrodes are only twelve millimeters from each other, and separated by a sheet-iron partition with small perforations in it, the latter allowing free passage of current but being too small to allow any gas bubbles to pass. The alternate compartments are connected with oxygen and hydrogen mains, in which are enlargements for collection of spray and moisture, which runs back into the cell. Current densities of two to three amperes per square d.m. are possible with a working voltage between 2.45 and 3, using caustic soda solution of 21° Baumé. The size cell shown in Fig. 6 is intended to take 400 amperes, and to require one kilowatt of power. A 100-horse-power plant set up at Rome cost \$14,000, when transformers and gas compressing plant were added in, \$20,000. Nearly a dozen plants are now working on this system in Europe.

Siemens Bros. & Co. and Obach devised the apparatus shown in Fig. 7, the principle being similar to that of Garuti. The cast-iron vessel *a* is surrounded by heat-retaining material, in order that the temperature of the cell may be automatically raised and thus its running resistance lowered. A cylindrical iron anode *f* is separated from the encircling cathode *g* by a cylinder of wire netting *c*, held in place by the porcelain block *k*. The electrolyte is diluted caustic soda; the gases escape above from the spaces *n* and *m*. The whole apparatus is set on insulating porcelain feet. The normal type of apparatus is built to take 750 amperes at 3 volts drop of potential, and furnishing eleven cubic meters of oxygen and twenty-two cubic meters of hydrogen per twenty-four hours, using up 162 kilowatt-hours. The cost of this unit is given as \$187.50, amounting to \$83.25 per kilowatt of current employed.

Schukert & Co. are selling an apparatus which cannot be properly described because it has not been patented, and certain details are kept secret. It is, however, a sort of combination of the ideas of Garuti and Siemens Bros. and Obach. Units taking 200 and 600 amperes have been built, operating at 2.9 volts, and costing only \$62.50 each, or \$36.00 per kilowatt used. If these figures are correct, it is the cheapest commercial apparatus yet

devised for this purpose; but unfortunately, satisfactory confirmation of these details is lacking.

Laboratory uses for apparatus decomposing water are probably confined to furnishing oxygen and hydrogen, for which any of the above types erected on a small scale, are suitable, or for measuring the electric current. If the volume of gas is found, which is evolved in a given time, the instrument integrates the amperes of current passing; that is, measures the coulombs which have passed. When thus used it has been called a *voltameter*, but the name is a misnomer, and has been abandoned for the much more appropriate term of the oxy-

Fig. 7.

hydrogen *coulometer* (coulomb-meter). The laboratory apparatus devised by Hofman is shown in its most recent form in Fig. 8, where the external reservoir enables one to read the volume of either gas at atmospheric pressure. Fig. 9 shows a coulomb-meter for measuring heavier currents, devised by Minet, in which the gases evolved collect in the reser-

voir V and by the pressure which they exert give corresponding indications on the manometer H. The instrument is

calibrated by using a constant current passing through a standard ammeter in series, and thus its co-efficient found, the indications being proportional to a constant times the indicated pressure.

The commercial applications of the electrolysis of water are numerous. The mixed gases, or both gases, are useful for obtaining high temperatures, and are used in melting platinum and similar metals, repairing of water-tube boilers, melting together plates of glass, melting out frozen tap-holes of iron blast-furnaces; the Drummond lime light is another application; quite recently blasting cartridges have been devised for dangerously gassy mines, in

which water is decomposed by an electric current and the highly compressed mixture ignited afterwards by an electric spark. Hydrogen alone is useful for filling balloons, several European governments having electrolytic plants for filling their military balloons; also for soldering purposes, particularly the autogenous soldering of lead and aluminum. The absence of arsenic makes the electrolytic product particularly desirable, in comparison with the hydrogen made by using zinc and sulphuric acid; also the cost is only about one-quarter as great. Hydrogen has been proposed for lighting railway cars, etc., using very small Welshbach mantles, with resultant advantages in sanitary respects, smaller danger of explosion and smaller piping. Within a distance from the point of production of 180 miles, it has been calculated that the compressed hydrogen is cheaper for this purpose than calcium carbide and acetylene. The uses of oxygen in chemistry, metallurgy, dentistry, medicine, etc., are too well known to need repetition.

In concluding we would refer to a few of the places where further information about the subject treated of in this lecture may be obtained. In the scientific treatment of the subject, we may refer to:

Dr. H. Daneell: Handbuch der Elektrochemie; Spezielle Elektrochemie. Lieferung 1. Wilhelm Knapp. Halle. 1903.

S. L. Bigelow: On the passage of a direct current through an electrolytic cell. *Journal of Physical Chemistry*, December, 1902.

Fig. 9.

C. J. Reed: Gas polarization in lead accumulators. *Journal of Physical Chemistry*, January, 1901.

J. W. Richards and W. S. Landis: The Electrolysis of Water. *Transactions of the American Electrochemical Society*, Vol. iii (1903), Vol. iv (1903).

The commercial side of the question has been treated by:

M. N. Schoop: Die Industrielle Elektrolyse des Wassers. 1901.

Engelhardt: The Electrolysis of Water. Translated by J. W. Richards. Chemical Publishing Company, Easton, Pa. 1904.

Various writers: *Electrochemical Industry*, 1903 and 1904. Particularly December, 1903, page 579.