Some employ zinc to form the alloy, but aluminum is preferable for the following reasons not stated in the text-books:

First.—Zinc melts and volatilizes at a comparatively low temperature, before the silicon has a chance to become properly alloyed. It is also difficult to combine the reaction-product with molten zinc, since the latter is heavier and causes the former to float.

Second.—The copious white fumes and blue flame evolved by zinc, even though the crucible be covered and the oxidation minimized, cause large amounts of "zinc wool" to condense on the interior of furnace and crucible, thus interfering with the operation.

Third.—The zinc-silicon alloy is hard and resists solution with hot or cold hydrochloric acid, while the aluminum alloy is readily dissolved.

Fourth.—Zinc tends to produce needle-shaped crystals of silicon, which are not so striking as the graphitoidal spangles obtained from the aluminum alloy.

Graphitoidal silicon resists oxidation before the blowpipe. Under the microscope, by reflected light, the faces of incomplete octahedra are visible in the small irregularly shaped masses. The yield is rarely more than ten per cent. by weight of the silica employed.

There are indications that the element silicon itself, heretofore classed among laboratory curiosities, may become commercially important.

On account of its high electrical resistance, its use in the form of compact rods has already been suggested for electric heating apparatus.

THE COMPARATIVE VALUE OF CERTAIN REAGENTS FOR REMOVING LIME AND MAGNESIA FROM NATURAL WATERS FOR INDUSTRIAL USES.[1]

By Martin L. Griffin.

IT is only within the past few years, fifteen to twenty possibly, that the purification of water mechanically and chemically has claimed the serious attention of engineers and chemists. Everything is done on such a vast scale to-day and competition is

[1] Read at the meeting of the New York Section of the American Chemical Society, May 5. 1899.

so great that things which were counted of little consequence a few years ago may determine success or failure now. Water contained the same impurities a thousand years ago as it does to-day, but they did not play the same rôle. We all remember when every house had a cistern to catch the pure rain-water from the clouds, when industries sought the source of pure water for their uses. It is the common practice of cities to-day to do this, though there are conspicuous exceptions where extensive filtration and purifying plants are being constructed. Formerly paper-mills were in the habit of digging wells or arching over large areas in the beds of streams and canals in order to obtain a pure water supply. Sponges were used as a filtering medium. Fifteen years ago the company which I continue to serve purified its water for making chemical fiber by pumping the water through a mass of excelsior, which at frequent intervals gave way and passed into the pulp to the confusion of the superintendent, who at first concluded that he had a " bad cook."

These inadequate and unscientific practices are fast giving way to scientific and efficient processes. In 1883 and 1884 J. W. Hyatt and Isaiah Smith Hyatt brought out the modern system of filtration and were the first to claim the use of a coagulant for the precipitation of organic matter in water. One would think that soluble organic matter as it naturally occurs in surface waters would be the most difficult of removal, but with the modern filter using a coagulant this is now accomplished beautifully, leaving nothing to be desired. If we could precipitate and filter out the lime, magnesia, and sulphuric acid as completely as we can remove organic matter, there would be little work remaining for the chemist to investigate in this direction. This, no doubt, has been a favorite theme for speculation with many chemists ; however, our inability to do this completely should not prevent putting forth our best endeavors to render a valuable service.

In 1882, Dr. C. B. Dudley obtained a patent for a process using caustic soda for purifying water, claiming as a result that "the lime and magnesia separate as carbonates and the iron and alumina probably as hydrated oxides, etc." This statement assumes that the lime and magnesia exist as bicarbonates only and leaves the subject vague and indefinite.

In 1883 the same author obtained a patent for the use of soda-lime for similar purposes, stating the result of the whole operation to be, '' that by far the largest part of the mineral impurities existing in the water previous to treatment are removed.''

Doubtless the failure to accomplish all that was expected of this process brought out another claim by Dr. Dudley the next year, which is in effect an amendment to the soda-lime process. In this he adds the use of barium hydroxide, and states that "the result of the whole operation will be that nearly all the mineral impurities occurring in the water, including the sulphuric acid, will be removed from it.''

At the first general meeting of this society held in Newport, R. I., in 1890, and which I had the pleasure of attending, Dr. Chas. A. Doremus presented a short paper calling attention to the use of fluoride, particularly sodium fluoride for softening hard waters. At that time the most important statement made was that '' the precipitation of magnesium is especially thorough and noteworthy.''

In 1893 the same author directed the attention of chemists again to the same subject, perhaps laying more stress on the ability of sodium fluoride to prevent incrustation in feed waters. This statement also appears : '' When sodium fluoride is added to waters containing even a moderate proportion of calcium and magnesium salts, a precipitate forms at once.'' In his patent covering the use of this chemical process, he says : '' Since the character of the precipitate produced by a fluoride may be such that it is difficult for a filter to separate it from the water, it may prove advisable to use with the fluoride some chemical capable of producing with the foreign substance in the water a bulky precipitate which shall act mechanically in removing any precipitate of fluoride also produced. Since magnesium fluoride is precipitated in a more flocculent form than calcium fluoride, water containing much magnesium salts may be treated effectually with the fluoride alone, etc. Since caustic soda causes a bulky precipitate when added to waters containing calcium salts, it may be employed with a fluoride for the above-mentioned purpose.''

He also says that he has found the use of carbonate and caustic soda with a fluoride more serviceable with certain waters

than when the fluoride alone was used, and that "highly carbonated waters especially seemed to require the use of some caustic alkali." I call attention to these statements, as there seems to be a lack of harmony with the claims made for fluoride. We shall also see that caustic soda in waters containing only calcium sulphate, unless in unusually large amounts, produces no precipitate whatever, and that waters highly charged with carbonates are best suited for the fluoride treatment because of the easy reaction. It was not until I received the January number of the Journal this year, containing an article on "Sodium Aluminate as a Means for the Removal of Lime and Suspended Matter from Waters for use in Boilers," by Prof. Mabery and Edwin B. Baltzley, that I determined to make a systematic quantitative investigation into the value of these and other reagents for removing lime and magnesia from natural waters.

Their method is based on certain reactions which the authors give, all depending on the presence of calcium or magnesium bicarbonate in the waters at the beginning and giving calcium and magnesium carbonates as a final result of the reaction, including the precipitation of the alumina as hydrate.

In regard to waters which contain no carbonates, as for instance, a certain well in Ashtabula. Ohio, cited by the authors, their explanation does not apply. But this is not important. We know that the results depend upon the relative affinities of the bases and acids entering into the reactions and the insolubility of the resulting calcium and magnesium salts.

The authors state that they have made a test comparing the use of different proportions of caustic soda, with aluminate, with the result that "the removal of lime was not much more than half the quantity taken out by aluminate, and the magnesia was not affected." This is certainly a result which chemists could not reasonably expect. The use of sodium aluminate for such purposes is only another way of introducing the strong base sodium, since nothing can be claimed for the precipitated alumina except for its mechanical action. We would therefore look for the same kind of an action whether we were to use sodium hydroxide or aluminate, the free base, however, acting more promptly and with greater intensity.

Also in the case of waters containing only calcium and

magnesium sulphates we would have expected better results in the removal of the magnesium than the calcium, since magnesium hydroxide is far more insoluble than calcium hydroxide.

In the case of one sample the authors say that " air was blown through the solution for a few minutes." No explanation is given for this. Out of some seven samples of water treated, over ninety per cent. of the calcium and magnesium were reported removed from six of them, while the remaining sample, though only an ordinary river-water, gave " unsatisfactory results." It seems that some explanation should have been given for this. The practicability of the process is beyond the scope of this paper.

The reagents selected for my investigation were sodium hydroxide, sodium phosphate, sodium fluoride, sodium aluminate, and barium hydroxide. It will be observed that the primary intent in this work is to introduce a strong, free base, throwing down the insoluble carbonates and hydroxides, of which iron, aluminum, and magnesium are quite insoluble, while that of calcium is easily soluble. We might pause if it were not for this solubility of calcium hydroxide, but unfortunately again the calcium salts are the principal mineral impurities in natural waters, so we pursue our search for a reagent possessing a very strong base and an acid having a greater affinity for the bases calcium and magnesium, and whose salts are insoluble, and we include sodium phosphate and fluoride.

With regard to sodium aluminate, this may be classed with the free base inasmuch as the alumina acts as a very weak acid. I have included it in the test in consequence of the claims made for it by the authors. The principal precipitates resulting from the use of these five reagents are calcium and magnesium carbonates, phosphates, fluorides and hydroxide, iron and aluminum hydroxides, and barium sulphate. With the exception of magnesium phosphate and calcium hydroxide, they are among the most insoluble precipitates we have, but the greater solubility of freshly made precipitates and the greater inactivity of reagents in weak solutions are obstacles to success. The relative affinities also play au important rôle.

The experiments were conducted upon a half liter of water usually, and the results are given in fractions of a gram per

liter.   Those who desire to read the results in grains per United States gallon (231 cu. in.), may use the factor 58.38, and for the Imperial gallon (277¼ cubic inches) the factor 70.   The half liter of the water at the ordinary atmospheric temperature was transferred to a flask of double the capacity, when a slight excess of the reagent to be employed was introduced.   The whole was shaken, corked, and allowed to stand on an average of one or two days.   The water was then filtered through a close, heavy paper S. S. 598 and the lime and magnesia in solution determined.   The reagents employed were normal solutions of sodium hydroxide and fluoride and tenth-normal sodium phosphate and barium hydroxide.   The aluminate solution was prepared from a pure aluminum sulphate containing 56.68 per cent. anhydrous sulphate, ten grams to the liter, to which sodium hydroxide was added to alkaline reaction.

In order to obtain some idea of what to expect I began the work with simple solutions of calcium and magnesium.   The first was a solution of calcium sulphate containing 0.072 gram calcium oxide, and there remained after treatment with

	Hours.	Calcium oxide. Gram.
Aluminate solution	46	0.060
Sodium hydroxide	76	No precipitate
Barium hydroxide	40	0.060
Sodium phosphate	48	0.032
Sodium fluoride	28	0.014

Here we observe that all reagents producing the hydroxides are unsatisfactory, while the best results are obtained with the fluoride.

Next a solution of magnesium sulphate containing 0.061 gram magnesium oxide was taken, and there remained in solution after treatment with

	Hours.	Magnesium oxide. Gram.
Aluminate solution	25	0.0115
Sodium hydroxide	27	0.0349
Barium hydroxide	21	0.0183
Sodium phosphate	44	No precipitate
Sodium fluoride	26	" "

In this case we see that the reagents producing hydroxides are much more successful.

A solution of magnesium chloride containing 0.0558 gram magnesium oxide contained after treatment with

	Hours.	Magnesium oxide. Gram.
Aluminate solution...........	17	0.0082
Sodium hydroxide............	17	0.018
Barium hydroxide............	12	No precipitate
Sodium phosphate............	120	" "
Sodium fluoride	28	" "

The latter was concentrated by boiling to one-fifth the original bulk and still no precipitate appeared, showing that no reaction could take place. In this instance only the aluminate and caustic gave results, the former doing particularly well.

We now pass to some natural waters. A sample from the Hudson River at Mechanicsville contains

	Gram.		Gram.
Calcium carbonate.....	0.0234 }	equivalent calcium oxide	0.0176
" sulphate......	0.0110 }		
Magnesium carbonate .	0.0102	" magnesium "	0.0050

and there remained after treatment with

	Hours.	Calcium oxide. Gram.	Magnesium oxide. Gram.
Aluminate solution......	46	0.0180	0.0048
Sodium hydroxide.......	42	0.0160	0.0043
Barium hydroxide.......	21	0.0176	0.0046
Sodium phosphate.......	24	No precipitate.	
Sodium fluoride..........	24	" "	

In this case none of the reagents were able to effect any appreciable reduction in the amounts of lime and magnesia.

The next is a sample from the middle of Lake Champlain opposite Ticonderoga, and contains somewhat more lime and magnesia, though it is not a hard water. It contains

	Gram.		Gram.
Calcium carbonate....	0.0341 }	equivalent calcium oxide	0.0255
" sulphate.....	0.0156 }		
Magnesium carbonate	0.0147	" magnesium "	0.0070

and there remained after treatment with

	Hours.	Calcium oxide. Gram.	Magnesium oxide. Gram.
Aluminate solution	47	0.0256	0.0068
Sodium hydroxide.......	25	0.0256	0.0064
Barium hydroxide	25	0.0260	0.0060
Sodium fluoride..........	47	0.0250	0.0057

We draw the same conclusions here as in the case of the Hudson River water.

The next is a sample of tap-water from the city of Niagara Falls and is the Niagara River water.   It contains

	Gram.		Gram.
Calcium carbonate···· 0.0332 }	sulphate····· 0.0640 }	equivalent calcium oxide 0.0448	
Magnesium carbonate 0.0272	··	magnesium "	0.0129

Calcium carbonate···· 0.0332 ⎫ equivalent calcium oxide 0.0448
"   sulphate····· 0.0640 ⎭
Magnesium carbonate 0.0272   ··   magnesium "   0.0129

and there remained in solution after treatment with

	Hours.	Calcium oxide. Gram.	Magnesium oxide. Gram.
Aluminate solution······	45	0.0260	0.0118
Sodium hydroxide·······	46	0.0100	0.0057
Barium hydroxide ·······	45	0.0430	0.0100
Sodium fluoride ·········	46	0.0190	0.0130

We observe that sodium hydroxide gives the best results with fluoride second.

The next sample is from a driven well in Mechanicsville which I will call B.   It contains

Calcium carbonate···· 0.0620 ⎫ equivalent calcium oxide 0.0760
"   sulphate····· 0.1003 ⎭
Magnesium carbonate· 0.0552   "   magnesium "   0.0263

and there remained in solution after treatment with

	Hours.	Calcium oxide. Gram.	Magnesium oxide. Gram.
Aluminate solution······	18	0.0650	0.0237
Sodium hydroxide·······	17	0.0160	0.0147
Barium hydroxide ·······	24	0.0260	0.0150
Sodium phosphate·······	17	0.0650	0.0240
Sodium fluoride ·········	46	0.0270	0.0234

Sodium and barium hydroxide give the best results with this water.

The following is also from a local well which I will call D. It contains

Calcium carbonate···· 0.0398 ⎫ equivalent calcium oxide 0.0940
"   sulphate····· 0.1760 ⎭
Magnesium carbonate· 0.0703   "   magnesium "   0.0330

There remained in solution after treatment with

	Hours.	Calcium oxide. Gram.	Magnesium oxide. Gram.
Aluminate solution 24		0.080	0.0236
Sodium hydroxide......... 27		0.022	0.0133
Barium hydroxide 24		0.063	0.0029
Sodium phosphate........ 17		0.068	0.0273
Sodium fluoride 71		0.035	0.0338

Again we observe that the best results are obtained with sodium hydroxide, with sodium fluoride second in reference to lime, and barium hydroxide second when we include magnesia.

The next sample is from the Chazy Marble Lime Co.'s quarries in northern New York, and is used in their boilers. It contains

$$\left.\begin{array}{l}\text{Calcium carbonate....} \quad 0.1100 \\ \text{\quad '' \quad sulphate.....} \quad 0.0570\end{array}\right\} \text{equivalent calcium oxide } 0.0850$$

Magnesium carbonate 0.0323   ''   magnesium ''   0.0154

There remained in solution after treatment with

	Hours.	Calcium oxide. Gram.	Magnesium oxide. Gram.
Aluminate solution 63		0.0670	0.0136
Sodium hydroxide....... 40		0.0150	0.0111
Barium hydroxide 41		0.0240	0.0115
Sodium fluoride 40		0.0170	0.0115

In this case sodium hydroxide and sodium fluoride gave nearly equally good results.

The following sample is from a well fourteen feet in solid rock near Lake Champlain in Chazy, N. Y.   It contains

$$\left.\begin{array}{l}\text{Calcium carbonate....} \quad 0.1818 \\ \text{\quad '' \quad sulphate.....} \quad 0.0148\end{array}\right\} \text{equivalent calcium oxide } 0.1140$$

Magnesium carbonate· 0.0520   ''   magnesium ''   0.0248

and contained in solution after treatment with

	Hours.	Calcium oxide. Gram.	Magnesium oxide. Gram.
Aluminate solution 88		0.0720	0.0208
Sodium hydroxide....... 42		0.0140	0.0186
Barium hydroxide 40		0.0300	0.0146
Sodium fluoride 42		0.0220	0.0172

Three out of the four reagents gave very satisfactory results.

The next is a sample of mine water from the Coxe Iron Mfg. Co., at Drifton, Pa., and contains

	Gram.			Gram.
Ferrous sulphate	0.0440	equivalent ferrous oxide		0.0208
Calcium sulphate	0.0534	" calcium "		0.0220
Magnesium sulphate	0.1554	" magnesium "		0.0518
Free sulphuric acid	0.1450			

and there remained in solution after treatment with

	Hours.	Ferrous oxide. Gram.	Calcium oxide. Gram.	Magnesium oxide. Gram.
Aluminate solution	71	trace	0.0200	0.0516
Sodium hydroxide	16	none	0.0220	none
Barium hydroxide	16	"	0.0220	0.0064
Sodium fluoride	112	No trace of any precipitate		

showing that this reagent was of no use. The sulphuric acid was not determined. We notice that sodium and barium hydroxide gave very satisfactory results. Barium hydroxide would doubtless be preferred, as it would largely remove the sulphuric acid.

The following sample is from the mouth of Onondaga Creek, Syracuse, N. Y., and contains

	Gram.			
Total solids	1.6200			
Sodium chloride	0.6550			
Magnesium chloride	0.0254 ⎫	equivalent magnesium oxide		0.0298
" carbonate	0.0493 ⎭			
Calcium sulphate	0.2176 ⎫	" calcium	"	0.3480
" carbonate	0.4614 ⎭			

and contained in solution after treatment with

	Hours.	Calcium oxide. Gram.	Magnesium oxide. Gram.
Aluminate solution	98	0.0840	0.0014
Sodium hydroxide	17	0.1280	0.0043
Barium hydroxide	136	0.0260	0.0028
Sodium fluoride	117	0.0420	0.0288

In this case barium hydroxide gave the best results, with sodium fluoride second, in reference to lime.

The last sample is from the magnesian lime quarries of the Snow Flake Lime Co., Bowling Green, Ohio, and contains

	Gram.			Gram.
Calcium carbonate	0.0870	} equivalent calcium oxide		0.2080
" sulphate	0.3868			
Magnesium carbonate·	0.1662	" magnesium "		0.0792

and contained in solution after treatment with

		Calcium oxide.	Magnesium oxide.
	Hours.	Gram.	Gram.
Aluminate solution······	63	0.0080	none
Sodium hydroxide·····	16	0.0090	0.0018
Barium hydroxide ·······	16	0.0680	0.0064
Sodium fluoride ·········	21	0.0250	0.0597

In this case we obtain the best results with the aluminate, but the treatment with sodium hydroxide follows so closely that practically there is little difference. The use of barium hydroxide and sodium fluoride are not so satisfactory.

In this work I have endeavored to get samples of characteristic and representative natural waters, such as are commonly met with in the industries, and think the ground is fairly well covered, so that the deductions to be made shall be true and useful.

We note first the limit to which we are able to purify waters in this way, and within what limits we are able to effect a reduction in the amount of earthy substances, by the use of any of the above reagents. The Hudson River contains 0.0176 gram calcium oxide and 0.005 gram magnesium oxide. These amounts we cannot reduce. Lake Champlain, containing 0.0255 gram calcium oxide and 0.007 gram magnesium oxide, could not be reduced, but increasing amounts above these figures show a reduction by one or more of the reagents. I conclude, therefore, that waters containing from 0.020 to 0.025 gram calcium oxide and from 0.005 to 0.007 gram magnesium oxide cannot be purified appreciably, although we observe that frequently the amounts of these impurities can be reduced below these limits in the case of waters highly charged, which is a favorable feature of the work. As regards the removal of calcium carbonate, the most effectual reagents are sodium hydroxide and sodium fluoride, and in some cases we observe that the aluminate solution gave particularly good results as in the last sample, but I will refer to this again. The best results show a reduction in the amounts of calcium oxide of 0.008 to 0.010 gram. It appears also that waters containing a large propor-

tion of calcium sulphate with carbonate can be successfully treated with these reagents as in the case of limestone quarry waters. Barium hydroxide did not prove as effective as I anticipated in the removal of lime.

As regards the removal of magnesium salts, we observe again that sodium hydroxide averages best, with barium hydroxide following very closely, while aluminate solution takes third place, and sodium fluoride produces practically no results on any of the samples. Sodium phosphate was so unpromising from the start, and the removal of the phosphoric acid from the water after treatment added so much to the labor of the work, that I abandoned its use.

We cannot conclude, however, that sodium hydroxide is the best reagent for all waters. Sodium fluoride is much to be preferred in waters containing calcium sulphate and chloride, and is of no value in the removal of magnesium salts.

Barium hydroxide is most serviceable in the treatment of acid mine waters and gave very satisfactory results in the case of Onondaga Creek water, the character of which may be seen from the analysis.

Referring back to the references in the early part of this paper, we are able to place definite limits to Dr. Dudley's processes and to know how far they may be useful. Dr. Doremus' claims for sodium fluoride for the precipitation of calcium salts are fairly well established, but for the precipitation of magnesium salts it is a failure, and for the removal of iron and magnesium very questionable. In my experiments with sodium aluminate I endeavored to follow the suggestions of Prof. Mabery, using somewhat less than the theoretical amount of the combined sodium oxide (Na_2O) in the aluminate necessary to replace the bases lime and magnesia, although the reason for doing this is not plain to me. In every case the action of the aluminate was comparatively slow, and in the case of the lime salts did not give satisfactory results or meet the claims of its authors. It proved more effectual in the removal of magnesium salts, which was to be expected.

I desire to refer to the composition of the aluminate reagent before the subject is left. I prepared this reagent as Prof. Mabery directs, from sodium hydroxide and aluminum sulphate,

	Originally.		Aluminate.		N,NaOH.		$\frac{N}{10}$Ba(OH)$_2$.		N,NaFl.		$\frac{N}{10}$Na$_2$HPO$_4$.	
	CaO.	MgO.	CaO.	MgO.	CaO.	MgO.	CaO.	MgO.	CaO.	MgO.	CaO.	MgO.
CaSO$_4$ solution	0.0720		0.0580	0.0115	0.0720	0.0349	0.0600	0.0183	0.0140		0.0320	0.0610
MgSO$_4$ solution		0.0610		0.0082		0.0180		0.0558		0.0558		0.0558
MgCl$_2$ solution		0.0558		0.0048		0.0043		0.0046		0.0050		0.0050
Hudson River	0.0176	0.0050	0.0180	0.0048	0.0160	0.0043	0.0176	0.0046	0.0176	0.0050	0.0176	0.0176
Lake Champlain	0.0255	0.0070	0.0256	0.0068	0.0256	0.0064	0.0260	0.0060	0.0250	0.0057		
Niagara River	0.0448	0.0129	0.0260	0.0118	0.0100	0.0057	0.0430	0.0100	0.0130	0.0130		
Well B	0.0760	0.0263	0.0650	0.0237	0.0160	0.0147	0.0260	0.0150	0.0190	0.0234		0.0650
Well D	0.0940	0.0334	0.0800	0.0236	0.0220	0.0133	0.0630	0.0029	0.0350	0.0338		0.0273
Chazy Quarry	0.0850	0.0150	0.0670	0.0136	0.0150	0.0111	0.0240	0.0115	0.0115	0.0680		0.0240
Chazy Well	0.1140	0.0248	0.0720	0.0208	0.0140	0.0186	0.0300	0.0146	0.0170	0.0172		
Coxe Mine	0.0220	0.0518	0.0200	0.0516	0.0220	0.0000	0.0220	0.0064	0.0220	0.0518		
Onondaga Creek	0.3480	0.0298	0.0840	0.0014	0.1280	0.0043	0.0260	0.0028	0.0420	0.0288		
Bowling Green	0.2080	0.0792	0.0080	0.0000	0.0090	0.0018	0.0680	0.0064	0.0250	0.0597		

from which it will be seen that the available sodium oxide in combination with the alumina is but a small proportion of the composition of the reagent, there being three inert molecules of sodium sulphate to one of sodium oxide beside the alumina.   I have no doubt this exerted some influence on the sample of water to which the reagent was added in causing a more complete precipitate of lime and magnesia in some cases.

It would be better to prepare this reagent from pure alumina and sodium hydroxide. As it is, the results are comparable with those reported by Prof. Mabery.

If this reagent could be made as effective in simple practice as the authors claim, leaving practically no alumina in the filtered water, its discovery would prove to be of immense value at once.

In conclusion I would say that there is little doubt but that satisfactory reagents and means will soon be discovered whereby all objectionable waters may be successfully treated at comparatively small expense.

Later I may have something further to contribute to this end.

REVIEW.

THE RETORT COKE OVEN AND THE CHEMISTRY OF ITS BY-PRODUCTS.

There is no innovation into the realm of metallurgical science, these last days of the nineteenth century, which will affect more widely and will, one might say, so revolutionize the whole system of iron- and steel-making in this country as will the introduction of the retort oven for the production of coke, with the simultaneous recovery of the by-products, tar, ammonia, benzene, cyanides and illuminating and fuel gas.

Think of the devastation in the beehive coke oven districts for the year 1897.   One has only to take the government statistics of the coke manufactured in this country for 1897 and put opposite it the coke to be made from the same coal in retort ovens of the Semet-Solvay type together with the long list of valuable by-products obtainable, to see the vast amount of valuable material destroyed annually by the present system of coking coal in beehive ovens, and to realize the extent to which humanity has been deprived of what rightly belongs to it.   The figures below will illustrate what was actually obtained from the beehive coke ovens in the United States from 22,140,000 tons of