REVIEW OF AMERICAN CHEMICAL RESEARCH.

GENERAL AND PHYSICAL CHEMISTRY.

A. A. NOYES, REVIEWER.

A Revision of the Atomic Weight of Cobalt, III. The Analysis of Cobaltous Chloride and Oxide. By THEODORE WILLIAM RICHARDS AND GREGORY PAUL BAXTER. *Proc. Am. Acad.*, 35, 61–72.—Cobaltous chloride was purified by passing air through its ammoniacal solution, precipitating with hydrochloric acid the purpureo-cobaltic chloride formed, recrystallizing this six times from ammonia solution in a platinum dish, precipitating in a Jena flask with hydrochloric acid, and heating the precipitate first in an air-bath at 200°, and then in a current of nitrogen and hydrochloric acid to expel ammonium chloride. The weighed cobaltous chloride was reduced to cobalt by hydrogen, the small quantity of alkaline chloride present in the metal and of ammonium chloride volatilizing during the reduction being determined and corrected for. The atomic weight value so obtained was 59.044, against 58.995 previously found from the analyses of cobaltous bromide (*this Rev.*, 5, 43), but the authors regard the new value as undoubtedly too large, owing to the presence of impurity, probably silica, in the chloride. The authors also attempted to prepare a cobaltous oxide of definite composition by igniting the higher oxide in a nearly complete vacuum of varying pressure, but found by reducing with hydrogen that the atomic weight values varied from 58.93 to 59.07. They conclude that the higher oxide could not be entirely decomposed without causing some reduction of the monoxide to metal cobalt. They regard the value 58.995 obtained from the analysis of the bromide as the most probable one. The article closes with a critical review of previous determinations.

Are Further Experiments Needed for Determining the Atomic Weight of Oxygen ? By EDWARD W. MORLEY. *J. Am. Chem. Soc.*, **22,** 51–61 ; *Science*, **11**, 281–288.—The author recalls the fact that our present value of the atomic ratio between hydrogen and oxygen is based solely on the composition of one compound, water, and on two methods of its determination, its synthesis by weight and the volumetric ratio taken in connection with the densities of the component gases, and suggests that the certainty attaching to the ratio would be greatly increased by a new complete synthesis of water by an independent method, and by further experiments on the density of hydrogen and on the volumetric ratio. He also considers probably practicable an accurate complete synthesis of hydrochloric acid from its elements and a determination of the ratio Na : H by dissolving sodium in water and determining the loss in weight. He raises, however, at the close the question whether the value of the results to science would justify the great labor involved in these investigations, and desires an expression of opinion on this point. In view of the fact that the atomic ratio of hydrogen and oxygen is stoichiometrically one of the most important of all chemical constants, and is likely also to have great significance theoretically, it seems certain to the reviewer that scientists will not rest satisfied until that ratio has been established with the greatest possible certainty and accuracy, and it seems therefore highly desirable that an investigator who possesses the rare manipulative skill, the critical appreciation of sources of error, and the long experience, required for the successful prosecution of such researches, should continue to devote himself to them.

The Nature of the Change from Violet to Green in Solutions of Chromium Salts. By F. P. VENABLE AND F. W. MILLER. *J. Am. Chem. Soc.*, **20,** 484–497.—The authors describe experiments on the fraction (found to be two-thirds) of the total sulphuric acid that is precipitated by barium chloride from green chrome alum solutions, on the amounts of ammonia required to produce a permanent precipitate in the green and violet solutions (found to be equal in the two cases), and on the composition of the gummy precipitate produced by adding alcohol to the green solution. They conclude from their experiments that the hypothesis of Recoura that the green chromium sulphate exists in solution as $Cr_4O(SO_4)_4.SO_4 + H_2SO_4$ is untenable.

The Nature of the Change from Violet to Green in Solutions of Chromium Salts. By W. R. WHITNEY. *J. Am. Chem. Soc.*, **21,** 1075–1084.—In a previous article (see the preceding review) Venable and Miller claim to have refuted the hypothesis of Recoura as to the composition of green chromium sul-

phate. After presenting ten facts previously known as evidence of the presence of free sulphuric acid, the author describes the results of transference experiments of his own, in which the theoretical amount of free acid was actually caused to migrate out from the green solution, while practically no acid was separated by the current from the violet solution. The author also communicates experiments confirming those of Fayre and Valson to the effect that only about one-half of the sulphuric acid in a green solution of chrome alum is precipitated immediately in the cold, and points out that the different results of Venable and Miller are no doubt due to the fact that they waited twenty-four hours before filtering off the barium sulphate. He also points out that their experiments, showing that equal quantities of ammonia are required to barely precipitate the green and violet solutions, do not have the significance attributed to them, since it is well known that alkalies at once convert the violet into the green modification.

The Nature of the Change from Violet to Green in Solutions of Chromium Salts. By F. P. VENABLE. *J. Am. Chem. Soc.*, **22,** 111–114.—After stating the reasons for the contrary opinion maintained in a previous article (see above), and otherwise interpreting some of the results there described, the author accepts the hypothesis of Recoura upheld by Whitney (see the preceding review), that a green solution of chromium sulphate has the composition, $Cr_4O(SO_4)_4.SO_4 + H_2SO_4$.

The Relation of Physical Chemistry to Technical Chemistry. By WILDER D. BANCROFT. *J. Am. Chem. Soc.*, **21,** 1101–1107.— The author emphasizes the importance of a knowledge of the principles of physical chemistry to the manufacturing chemist, and maintains that aside from the value of this knowledge, physico-chemical research is a better training for industrial work than research in synthetic organic chemistry. The reviewer is also of this opinion, especially with reference to this country, where the positions filled by chemists involve, as a rule, the improvement and development of processes rather than the synthetic production of new compounds.

On the Surface-tensions of Mixtures of Sulphuric Acid and Water, and the Molecular Mass of Sulphuric Acid. By C. E. LINEBARGER. *J. Am. Chem. Soc.*, **22,** 5–11.—The author has determined for each 10° between 0° and 50° or 70° the surface-tensions of aqueous sulphuric acid of strengths varying from 2.65 up to 95.0 per cent. acid. He finds that the surface-tension of the strongest acid is greatly increased by the addition of relatively small amounts of water, and is but slightly affected by

rise of temperature. The latter fact is considered to point to a high degree of polymerization of the sulphuric acid molecules.

Seventh Annual Report of the Committee on Atomic Weights. Results Published in 1899. By F. W. CLARKE. *J. Am. Chem. Soc.*, **22**, 70–80.—The average values of the atomic weights published last year and compiled by the author are given below, the values in parentheses being those adopted by the author in his previous report. All values are referred to oxygen taken as 16.000 :

Boron, Gautier	11.016	(10.95)
Nitrogen, Dean	14.031	(14.04)
Calcium, Richards	40.126	(40.07)
Nickel, Richards and Cushman	58.709	(58.69)
Cobalt, Richards and Baxter	58.995	(58.99)
Molybdenum, Vandenberghe	96.069	(95.99)
Tungsten, Hardin	184.0	(184.83)
Cerium, Kölle	138.81	(139.35)
Palladium, Hardin	107.014	(106.36)

On Some Abnormal Freezing-point Lowerings Produced by Chlorides and Bromides of the Alkaline Earths. By HARRY C. JONES AND VICTOR J. CHAMBERS. *Am. Chem. J.*, **23**, 89–105.—The molecular lowerings of the chlorides and bromides of barium, strontium, calcium, and magnesium are all found to have a minimum value of $4.8°-5.1°$ between the concentrations 0.1 and 0.2 molecule per liter. The values increase considerably with increasing concentration (in most cases up to $5.4°-5.8°$ at 0.6 mol per liter), in spite of the fact that the dissociation is continually decreasing, as is shown by the electrical conductivity measurements which are also given by the authors for all eight salts over a wide range of concentration. The authors suggest that the abnormal freezing-point values may arise from the combination of the undissociated salt-molecules with a large number of water molecules, thereby reducing appreciably in the concentrated solutions the amount of solvent. Freezing-point and conductivity determinations are also given for solutions of cadmium chloride and bromide, but in these cases the molecular freezing-point lowering exhibits no minimum, but steadily decreases with rising concentration.

The Melting-point of Chloral Hydrate. By C. G. L. WOLF. *J. Phys. Chem.*, **4**, 21–32.—The author has determined the melting-point of chloral hydrate under different conditions, and concludes that the variations in it found by him and previous investigators are due, not to the existence of an isomeric form, but to the presence of dissociation-products in varying amounts. When a thermometer bulb was coated with chloral hydrate and heated in a vacuum, these conditions being such that the dissociation-

products could rapidly escape as vapors, the substance did not melt below 72°, at which point rapid sublimation took place, however. On the other hand, when it was heated in a closed tube, time being allowed for equilibrium to be reached between the solid, liquid, and vapor, the melting-point was about 47°.

The Relation of the Taste of Acid Salts to Their Degree of Dissociation. By Louis Kahlenberg. *J. Phys. Chem.*, 4, 33-37.— The author finds that the acid sodium salts of oxalic, malonic, succinic, malic, tartaric, fumaric, maleic, and citric acids all taste still noticeably sour in solutions containing one mol in 100–400 liters, varying only between these narrow limits, with the nature of the salt. The variation in sourness is thus very much less than the variation (1 : 70 in the extreme case) in the concentration of the hydrogen ions in the different salt solutions. Moreover, the sourness of acid sodium tartrate and malate solutions, where the dissociation is not more than a few per cent., was found to be nearly equal to that of hydrochloric acid only three times as dilute, confirming the result of Richards with acetic and hydrochloric acids, that the sourness does not depend solely on the concentration of the hydrogen ions. The author concludes that, if his results are to be interpreted by the dissociation theory, the univalent anions (COOH—COO, etc.) must, like the hydrogen ion, be assumed to have a sour taste, an " explanation which is unsatisfactory to say the least." This explanation does not, however, seem a necessary, or even a probable one, to the reviewer, for the intensity of the sour taste probably depends on the amount of some unknown chemical change produced by the acid in the ends of the sensory nerves in a short interval of time. If this change was caused catalytically by the hydrogen ions, one might expect proportionality between its rate and the concentration of those ions. If, however, the hydrogen ions directly enter into the reaction and are consumed by it, their concentration in the layer in immediate contact with the nerve will be rapidly reduced, especially if the acid solution is dilute ; and the hydrogen ions so removed can be replaced, in case the acid is completely dissociated, only by the slow processes of diffusion and convection, while if the acid is more concentrated but only slightly dissociated, an immediate new supply of hydrogen ions is furnished by the instantaneous process of dissociation. Thus, according to this hypothesis, which is *a priori* as reasonable as the other, the degree of sourness of a dilute acid solution will, in general, depend on the total concentration of the acid as well as on the concentration of the hydrogen ions, which is, in fact, the principle which expresses the experimental results thus far reached.

On Thermal and Dynamic Coefficients. By J. E. Trevor. *J. Phys. Chem.*, **3**, 573–576.—The article is a continuation of the general discussion of the partial derivatives of the four fundamental thermodynamic functions (*this Rev.*, **6**, 9).

On the Emission and Absorption of Water Vapor by Colloidal Matter. By P. Duhem. *J. Phys. Chem.*, **4**, 65–122.— The article consists of an extended mathematical discussion of the phenomenon named in the title, when it is affected by hysteresis, and of an application of the conclusions reached to von Bemmelen's observations on the dehydration of gelatinous silica and colloidal ferric hydroxide.

The Melting-point of Formyl Phenyl Acetic Ester. By C. G. L. Wolf. *J. Phys. Chem.*, **4**, 123–129.—The article is largely devoted to a discussion of the observation that the ester, after heating at 100° for some hours (by which it is melted and converted in large part into the enol form), separates crystals of the keto form more quickly if the liquid is first cooled to —80° and then allowed to stand at 20° than if it is cooled directly to 20° and allowed to stand, the difference being attributed to a more rapid reconversion into the keto form, arising from the fact that the temperature in the former case is much further below the equilibrium temperature. Thus fall of temperature may accelerate a reaction.

Freezing-point Curve for Water Containing Hydrochloric Acid and Phenol. By J. A. Emery and F. K. Cameron. *J. Phys. Chem.*, **4**, 130–134.—The authors have determined the freezing-points of water and hydrochloric acid solutions, 0.024– 0.500 normal, when saturated with phenol. The lowerings are found to be the sums of the lowerings caused by the two substances separately. The freezing-point of the saturated aqueous solution of phenol is —1.179°.

Thermal Effects of the Dilution of Some Salts. By F. P. Dunnington and T. Hoggard. *Am. Chem. J.*, **22**, 207–211. —The authors have determined the heat effect resulting when, to nearly saturated solutions of fifteen neutral ammonium, potassium, sodium, lithium, magnesium, calcium, and strontium salts, water is added, one molecule or a few molecules at a time. The heat of dilution was found to be positive for seven of the salts, negative for seven others, and positive for the first additions of water, and then negative for subsequent additions in the case of the remaining salt, calcium nitrate. The values for the first mol of water added are greatest for lithium, calcium, and magnesium chlorides (+678, +508, and +370 cals.).

A Contribution to the Study of Liquid Mixtures of Constant Boiling-point. BY GARNETT RYLAND. *Am. Chem. J.*, 22, 384–396.—The author has fractionated mixtures of eighty pairs of liquids boiling separately less than 40° apart, and finds that 45 of them yield a constant boiling fraction of the same or lower boiling-point than the more volatile constituent, two of them, such a fraction of a higher boiling-point than the less volatile constituent, and one of them, ethyl iodide (b. p. 72°) and benzene (b. p. 79°), a fraction boiling (at 74°–75°) between the boiling-points of the two constituents. This last case, which is rather remarkable from a theoretical standpoint, is further exemplified by an earlier observation of Bauer on a mixture of ethylene and propylene bromides. As this case does not seem to have been hitherto considered theoretically, the reviewer would point out that if the vapor-pressure for any composition of the mixture has a minimum value lying between the two limiting vapor-pressures of the pure liquids, it must necessarily have also a maximum value for some other composition, or *vice versa*, and therefore it must be possible to obtain *two different* mixtures of constant boiling-point, a conclusion which it would be interesting to test experimentally. The author gives the composition, determined by repeated trials of synthetic mixtures, and the boiling-point under atmospheric pressure, of the various constant boiling mixtures ; and compares the observed composition of many of them with that calculated under the assumption that each component volatilizes in gaseous volumes proportional to its own vapor-pressure in the pure state at the boiling-point of the mixture. In almost all cases the so-calculated values are much larger for the less volatile, less concentrated component than the observed ones, a result which is, however, in accord with the law of Raoult, which requires that the vapor-pressure of the smaller component be relatively more reduced than that of the larger component. In the case of the mixtures of methyl alcohol and benzene and of ethyl alcohol and benzene, distillations were made at different pressures (22–77 cms.), the composition of the constant boiling fractions being then found to vary several per cent. In concluding, the author emphasizes the very *common* occurrence of mixtures of constant boiling-point and the consequent frequent impossibility of separating liquids by fractional distillation.

On a Hypothesis to Explain the Partial Non-explosive Combination of Explosive Gases and Gaseous Mixtures. BY W. G. MIXTER. *Am. J. Sci.*, 157, 327–334.—The fact that explosions do not occur by sparking explosive mixtures of gases when their pressure is less than a certain limiting value is attributed to the infrequency of impacts of molecules having a velocity or inter-

nal energy adequate for chemical union, there being more time in a rare gas than in a dense one for a molecule with energy adequate for combination to lose this energy by radiation.

On the Applicability of the Dissociation Theory to the Electrolysis of Aqueous Solutions Containing Two Electrolytes with a Common Ion. By J. G. MacGregor. *Phys. Rev.*, **8**, 129–140.—The author has calculated from the known dissociation-values and transference numbers of the component salts, the transference relations which should prevail in a mixture of two salts with a common ion, and compares the calculated results with those observed by previous investigators. The agreement is only a rough one. The author calculates the dissociation constants of the salts in question at different concentrations with the help of the ordinary mass-action law, disregarding the now well-established fact that they do not conform to that law.

The Conductivity and Dissociation of Some Electrolytes. By William Foster, Jr. *Phys. Rev.*, **8**, 257–281.—The author has determined the electrical conductivity of solutions of ammonium nitrate, magnesium sulphate, magnesium chloride, sodium ammonium hydrogen phosphate, disodium hydrogen phosphate, and potassium dihydrogen phosphate, of sodium hydroxide, and of phosphoric, oxalic, and citric acids, at concentrations varying from 1.0 to 0.0001 equivalents per liter. The temperature coefficients of conductivity and the specific gravities of normal solutions of these substances were also determined. The degrees of dissociation are calculated for all these substances at the various dilutions and compared with Loomis' results by the freezing-point method. The agreement is close except in the cases of the magnesium salts where differences of 10 to 20 per cent. occur, and of sodium hydroxide and potassium dihydrogen phosphate, where the differences are about 5 per cent.

H. M. Goodwin, Reviewer.

The Electrochemical Equivalents of Copper and Silver. By Theodore W. Richards, Edward Collins, and George W. Heimrod. *Proc. Am. Acad.*, **35**, 123–150.—This valuable contribution to electrochemistry should be read by all interested in the determination of the fundamental electrochemical constant. The investigation was undertaken primarily to reconcile, if possible, the discrepancy existing between the atomic weight of copper as determined by electrochemical and by purely chemical methods. A very careful study was first made of the conditions to be observed in a copper voltameter, and the effect of the various sources of error introduced by variable current density, acidity of solution, etc., is minutely discussed. A number of experi-

ments in which two copper voltameters with electrodes of 50 sq. cm. and 25 sq. cm. surface respectively, in series with two silver voltameters, gave the following results for the electrochemical atomic weight of copper :

Uncorrected results of Rayleigh, Gray, Shaw, Vanni, etc	$\begin{cases} 63.30 \text{ to} \\ 63.50 \end{cases}$
Large plates in cupric solutions at 20°............	63.47
" " " " " " 0°............	63.525
Small " " " " " 0°............	63.547
Medium " " cuprous " " 0°............	63.573
" " " " " " 60°............	63.615
Corrected results for cupric solutions............	63.563
At. wt. copper by chemical processes	63.604
Discrepancy	0.041

The experiments designated "in cuprous solutions" were made with copper sulphate which had been, by previous electrolysis with an alternating current, saturated with cuprous salt ; they were carried out to determine the maximum value for the electrochemical atomic weight, since if any cuprous ions carry a part of the current the deposit should be too great. The increase in value with rise of temperature (producing increase in solubility of cuprous salt) proves this to be the case, so that 63.573 can be taken as an extreme upper limit. The discrepancy between the chemical and electrochemical atomic weight is thus seen to be greater than the experimental error of either determination, and the authors therefore sought the cause in the silver voltameter. By a very careful investigation of this they were able to show that the effect well known to previous investigators that the deposit of silver in a silver voltameter depended on whether the solution had been previously used or not, arose from the formation of a substance in the neighborhood of the silver anode, which tended to deposit too much silver if allowed to diffuse to the cathode. The form of silver voltameter was therefore changed, the anode being placed within a porous cell, and this compared with Lord Rayleigh's and Patterson and Guthe's form of voltameter. In all cases the new form of voltameter gave the lowest results. The difference amounted to 0.081 per cent. Correcting the atomic weight of copper referred to the ordinary silver voltameter by this amount the authors find its value must lie between 63.598 and 63.615 as outside limits. The chemical value 63.604 lies between these values, in truth, "a remarkable confirmation of the results of the porous cup voltameter." The mean of the corrected values of Lord Rayleigh and Mrs. Sidgwick, F. and W. Kohlrausch, Kahle, and Patterson and Guthe gives 0.0011173 gram silver per ampere second, or 0.0011172 as the mean of the extreme values. On

this latter basis one gram equivalent of an electrolyte transports 96610 coulombs of electricity instead of 96540, the value usually assumed.   The investigation furnishes probably the most rigid experimental proof of Faraday's law yet published.

The Influence of Temperature, Pressure, Used Solutions, and Size of Anodes on the Deposit of Silver Voltameters. By J. F. MERRILL. *Phys. Rev.*, **10**, 167-175.—Pressures up to 103 atmospheres had no effect on the amount of silver deposited, nor did a variation of temperature from 0° to 90° C. The results with five old solutions compared with freshly prepared solutions showed the deposit to be greater from the former by about 1 part in 11600.   Later, the old solution, after being repeatedly used, gave an increase of as much as one part in 1700.   The solution did not become acid, however, if it had been boiled.   The size of the anodes, except in extreme cases, has no appreciable influence.   The nature of the surface of the cathode seems to have a slight effect.

The Influence of Electrification upon the Surface Tension of Water and Mercury. By ERNEST MERRITT AND SAMUEL J. BARNETT. *Phys. Rev.*, **10**, 65-74.—The authors show by a careful theoretical discussion of the method employed by Barnett for determining the effect of electrification on the surface-tension of water and mercury (*Phys. Rev.*, **6**, 257, 1898) that the effects observed can be wholly accounted for by purely electrostatic effects, and that therefore no effect of pure electrification upon surface-tension has yet been observed.   The electro-capillary phenomena of mercury in contact with electrolytes are, of course, not here considered.

Polarization and Internal Resistance of Copper Voltameter. By B. E. MOORE. *Phys. Rev.*, **10**, 34-52.—By a somewhat elaborate system of circuit breaking and making devices, for details of which reference must be made to the original, the author has measured the total polarization produced by direct currents at the electrodes of a copper voltameter, at times varying from 0.00006 second to 0.045 second after the polarizing electromotive force has been removed.   The results are discussed graphically, and the author concludes that, if the true maximum value of the polarization (P) could be measured, the simple formula $r = \dfrac{E - P}{i}$ would give very nearly the true value of the resistance (r) of electrolytes for all current strengths.

The Electrolytic Reduction of Potassium Chlorate. By ADOLPH L. VOEGE. *J. Phys. Chem.*, **3**, 577-601.—The reduction of this salt was experimentally studied under as varied con-

ditions as possible, and the following conclusions reached: The reduction is greater in acid than in alkali solution, and varies with the strength of acid; it increases greatly with the temperature, and increases also with decreasing current density; this last influence, however, varies with the nature of the metal used for the electrodes, copper, zinc, cadmium, and platinum having been studied.

The Electrochemical Equivalent of Carbon. BY H. C. PEASE. *J. Phys. Chem.*, **4**, 38-41.—The mean of five experiments on the loss of weight of a carbon anode in fused potassium hydroxide in an electrolytic cell with an iron cathode gave 3.32 as the equivalent weight of carbon. Coehn found the value in sulphuric acid to be 3.0. Carbon thus comports itself in these electrolytes as a quadrivalent element.

On the Inversion of the Hepta- and Hexahydrates of Zinc Sulphate in the Clark Cell. BY H. T. BARNES. *J. Phys. Chem.*, **4**, 1-21.—The paper contains the complete results of experiments published by Callendar and Barnes in 1897 (*Proc. Roy. Soc.*, **62**, 150) on the change of hydration of zinc sulphate crystals indicated by the change in temperature coefficient of a Clark cell in the neighborhood of 39° C. The temperature coefficient of cells of various forms was measured from 15° to 50°, and formulæ for the electromotive force deduced for various intervals. The electromotive force curves for cells containing the hepta- and hexahydrates of zinc sulphate cross at 38.78° C., which therefore corresponds to the transition temperature. From solubility experiments the transition temperature is 39.95°. The difference is explained by a secondary influence due to the presence of the mercurous sulphate in the Clark cell.

The Electrolytic Deposition of Brass. BY J. LIVINGSTON R. MORGAN. *J. Am. Chem. Soc.*, **22**, 93-99.—An attempt to explain the process of brass plating on the basis of the osmotic theory of the cell. The explanation does not seem clear to the reviewer, and is not verified by any experimental data.

The Specific Gravity and Electrical Resistance of Metallic Tellurium. BY VICTOR LENHER AND J. LIVINGSTON R. MORGAN. *J. Am. Chem. Soc.*, **22**, 28-31.—The tellurium was prepared by reducing an alkaline solution of the oxide by means of sugar (Lenher). For the specific gravity of the powdered metal 6.1993 was found as a mean of three determinations at 20°. The specific resistance was determined by casting the metal in rods of various lengths (57.5 mm. to 162 mm.) and diameters (2.92 mm. to 8.89 mm.) and measuring their resistance by the usual bridge method. Contact was made by means of mercury.

The values obtained for the specific resistance vary from about 300 to 1150; the mean, about 500, is only about one-fourth of the value found by Matthieson.

ANALYTICAL CHEMISTRY.

ULTIMATE ANALYSIS.

H. P. TALBOT, REVIEWER.

The Iodometric Determination of Small Quantities of Carbon Monoxide. By LEONARD P. KINNICUTT AND GEORGE R. SANFORD. *J. Am. Chem. Soc.*, **22**, 14–18.—The carbon monoxide is passed over iodine pentoxide at a temperature of 150°, and the reduced iodine is collected in a potassium iodide solution and titrated with thiosulphate solution. Accurate determinations were made when the amount of the monoxide in the air was as small as 2.5 parts in 100,000.

The Volumetric Determination of Magnesia. By JAMES OTIS HANDY. *J. Am. Chem. Soc.*, **22**, 31–39.—The precipitate of magnesium ammonium phosphate, thrown down under definite conditions, is collected upon a filter, washed with dilute ammonia, and the filter exposed to the air until the ammonia has volatilized, which is found to be the case when the filter has dried about one-half inch from its circumference. The precipitate and filter are then treated with standard sulphuric acid in excess, and after complete solution the excess is determined by titration with standard sodium hydroxide solution, using methyl orange as an indicator.

Notes : The Retention of Moisture by Asbestos, and The Determination of Graphite by Loss. By GEORGE AUCHY. *J. Am. Chem. Soc.*, **22**, 46–48.

The Determination of Sulphur in Bitumens. By A. C. LANGMUIR. *J. Am. Chem. Soc.*, **22**, 99–102.—The author comments upon the criticisms made by S. F. and H. E. Peckham (*this Rev.*, **5**, 100) upon an article by E. H. Hodgson (*this Rev.*, **5**, 10). He prefers the Eschka method as modified by Heath (*this Rev.*, **4**, 121). While it is undoubtedly true that the sulphur in illuminating gas is often a source of considerable error in sulphur determinations, it seems to the reviewer doubtful whether the data given by von Meyer and quoted by Langmuir are of general application. It is certainly true that Boston gas, for instance, introduces much less sulphuric acid during evaporations, even under unfavorable conditions, than was found by von Meyer in his laboratory.