

[CONTRIBUTION FROM THE MASSACHUSETTS INSTITUTE OF TECHNOLOGY.]

REVIEW OF AMERICAN CHEMICAL RESEARCH.

VOL. VI. NO. I.

ARTHUR A. NOYES, Editor.

REVIEWERS: Analytical Chemistry, H. P. Talbot and W. H. Walker; Biological Chemistry, A. G. Woodman; Carbohydrates, G. W. Rolfe; General Chemistry, A. A. Noyes; Geological and Mineralogical Chemistry, W. O. Crosby and M. L. Fuller; Inorganic Chemistry, Henry Fay; Metallurgical Chemistry and Assaying, H. O. Hofman; Organic Chemistry, J. F. Norris; Physical Chemistry, H. M. Goodwin; Sanitary Chemistry, E. H. Richards; Industrial Chemistry, A. H. Gill and F. H. Thorp.

GENERAL AND PHYSICAL CHEMISTRY.

A. A. NOYES, REVIEWER.

The Development and Application of a General Equation for Free Energy and Physico-Chemical Equilibrium. By GILBERT NEWTON LEWIS. *Proc. Am. Acad.*, 35, 1-38.—The title well indicates the general scope of the article, of which a detailed review is impracticable. The following laws and phenomena are considered by the author from the point of view of his general free energy equation: law of mass-action; law of constant distribution-coefficients; change of equilibrium with temperature including van't Hoff's equation; van der Waals' equation; heats of vaporization and vapor-pressure curves; specific heats of liquids and vapors; osmotic pressure of concentrated solutions; distribution of a solute between two solvents; single potential differences and electromotive force of concentration cells and of those composed of the same metal in contact with the same electrolyte dissolved in two different solvents. The mathematical treatment of the author is unusually clear and explicit.

Contribution to Our Knowledge of Aqueous Solutions of Double Salts.—*II. Chlorides.* By HARRY C. JONES AND KENJIRO OTA. *Am. Chem. J.*, 22, 5-14.—*III. Chlorides and Bromides.* By HARRY C. JONES AND NICHOLAS KNIGHT. *Am. Chem. J.*, 22, 110-141.—The electrical conductivity of aqueous solutions, varying in concentration from 0.5-1.0 mol. per liter to a very high dilution, of a considerable number of double chlorides and bromides, was measured and compared with the sum of the conductivities of the separate components of

the double salt, which were in many cases measured by the authors, but in others were taken from Kohlrausch. The following table presents the results in a greatly condensed form. Under *v* is given the volume in liters in which one gram-molecular weight of the double salt is dissolved, and under "per cent. diff." is given the percentage difference (unfortunately not calculated by the authors) between the sum of the conductivities of the separate component salts and the conductivity of the double salt. The four salts in the first horizontal series were investigated by Jones and Ota, the remainder by Jones and Knight.

2KCl.ZnCl ₂ .		KCl.CdCl ₂ .		NaCl.AlCl ₃ .		2NH ₄ Cl.HgCl ₂ .	
<i>v</i>	Per cent. diff.	<i>v</i>	Per cent. diff.	<i>v</i>	Per cent. diff.	<i>v</i>	Per cent. diff.
1	35.8	2	26.4	2.88	13.1	5	12.1
40	2.6	32	9.7	23.04	5.2	40	3.6
2000	0.6	1024	0.1	368.6	0.2	80	1.1
2NaCl.ZnCl ₂ .		NH ₄ Cl.MgCl ₂ .		KCl.Mg ₂ .Cl ₂ .		SrCl ₂ .2CdCl ₂ .	
1.14	47.5	1.92	13.1	1.95	17.3	2.18	34.2
24.16	9.4	30.72	5.0	31.2	12.5	34.9	13.6
773.	1.8	983.	2.4	998.	5.2	1116.	5.8
2NaBr. ₃ CdBr ₂ .		3NH ₄ Br.ZnBr ₂ .		BaBr ₂ .CdBr ₂ .		KBr.CdBr.	
4.88	30.6	1.6	23.2	1.97	31.8
39.1	14.6	25.9	7.8	31.5	9.9	30	13.7
1563.	6.3	1039.	2.4	1249.	3.7	1200	5.7
KCl.NH ₄ Cl.		KCl.2NH ₄ Cl.					
1	11.9	1	14.1				
32	3.8	32	3.7				
640	--0.4	640	--0.8				

The mixtures of potassium and ammonium chlorides were investigated to determine the effect on one another of two salts which do not form a double salt. In normal solutions of these salts it will be seen that the reduction of the dissociation in the mixture causes a difference in conductivity of 12-14 per cent. The very considerable magnitude of this effect makes somewhat doubtful the general conclusion of the authors that the (other) "double chlorides and bromides exist, to a considerable extent, in the more concentrated solutions, and are completely broken down by water only at very great dilutions." Their conclusion would seem to be warranted in the case of the zinc and cadmium double salts, but not in the case of those of aluminum, mercury, and magnesium. The freezing-point lowering caused by seven of the above double salts was also determined at various dilutions. From the fact that some of these, as well as some of the component salts, exhibit a *minimum* molecular lowering between 0.1 and 0.5 molecular, the authors conclude that the

freezing-point method is not suitable for the investigation of double salts. It would seem to the reviewer, however, that, in spite of the very considerable deviations from theory which no doubt exist, much could be learned from a direct experimental comparison of the component salts and their mixtures, and it is to be hoped that the authors will continue their investigation in this direction.

The Hydrolysis of Stannic Chloride. BY WILLIAM FOSTER, JR. *Phys. Rev.*, 9, 41-56.—The author has determined the electrical conductivity of both concentrated and dilute aqueous stannic chloride solutions, and the increase in their conductivity with the time. He finds that very concentrated solutions (5.0-2.5 equivalents per liter) exhibit no time-change; that solutions of medium concentrations increase slowly in conductivity, while dilute ones (0.5-0.001 normal) change so rapidly that the initial velocity is not measurable. The maximum values reached by the dilute solutions are identical with that shown by hydrochloric acid of the same concentration. These results are substantially the same as those obtained by Goodwin (*this Rev.*, 3, 22) in his similar investigation on ferric chloride. The assumption of complete hydrolysis in dilute solutions is in complete accord with the freezing-point determinations of Loomis, the transference experiments of Hittorf, and the thermochemical measurements of Thomsen.

The Electrical Resistance of Lead Amalgams at Low Temperatures. BY GEORGE W. GRESSMAN. *Phys. Rev.*, 9, 20-29.—The results are presented graphically. As the resistance suddenly decreases very markedly when the amalgam begins to freeze, the freezing-points can be readily determined. These are, for a moderate proportion of lead, *higher* than that of pure mercury; for example, the latter freezes at -38.8° , while a 4.2 per cent. amalgam freezes at -37.7° , a 7.1 per cent. one at -37.0° , and a 11.2 per cent. one at -30.1° .

Molecular Weights of Liquids, III. BY CLARENCE L. SPEYERS. *J. Am. Chem. Soc.*, 21, 282-287.—The author shows graphically that, according to the experiments of Lehfeldt, in the case of mixtures of carbon tetrachloride with benzene and of the same with toluene, the partial vapor-pressure of each constituent is nearly proportional to its percentage molecular concentration, even between the extreme limits of 0 and 100 per cent., but that this is not true of mixtures of ethyl alcohol with benzene or with toluene. He concludes from these facts that the molecules of the constituents of the first-named mixtures are not, and that those of the last-named mixtures are, polymerized.

On Van't Hoff's Equation and the Molecular Weights of Liquids. BY C. L. SPEYERS. *J. Am. Chem. Soc.*, **21**, 725-732.

The Inversion of Sugar by Salts. BY L. KAHLENBERG, D. J. DAVIS, AND R. E. FOWLER. *J. Am. Chem. Soc.*, **21**, 1-23.—The effect of most of the salts investigated by the authors has been determined previously by Walker and Aston (*J. Chem. Soc.*, **67**, 576) or by Long (*J. Am. Chem. Soc.*, **18**, 120, 693). The original features of the investigation are the application of the freezing-point lowering as a means of determining the extent of the inversion, and the measurement of the inverting effect of certain colored salts (copper sulphate and chloride and nickel sulphate) by that means. The authors' theoretical considerations in regard to hydrolysis are misleading, this phenomenon depending on the degrees of dissociation of water, of the salt itself, and of the acid and base composing it, and not at all, as the authors state, on the "ionizing tendency" of the separate ions. It may be further pointed out that the inverting effect of a salt of a weak base and the number of hydrogen ions its solution contains are a measure of its hydrolysis only in the case where the acid is completely dissociated; and, therefore, the fact that sulphates of the heavier metals invert less readily than the chlorides, does not necessarily signify a less degree of hydrolysis in the former salts, since much undissociated sulphuric acid may be present.

Derivatives and Atomic Mass of Palladium. BY WILLET LEPLEY HARDIN. *J. Am. Chem. Soc.*, **21**, 943-955.—Diphenyl-, dipyridine-, and diquinoline-pallad-diammonium chlorides and bromides, and di-piperidine-pallad-diammonium chloride were prepared by the addition of the bases to an acid solution of palladium chloride or bromide. The composition of the phenyl derivative is $\text{Pd}(\text{NH}_2\text{C}_6\text{H}_4)_2$; and the others are analogous, e.g. $\text{Pd}(\text{C}_6\text{H}_4\text{NCl})_2$. The names assigned to the latter are, therefore, irrational. To determine the atomic weights of the metal, the author has reduced diphenyl-pallad-diammonium chloride and bromide and ammonium palladium bromide by hydrogen at a red heat. As a mean of the three series of closely checking determinations he gets 107.014 which is much higher than the values adopted by Clarke and Richards (106.36 and 106.5).

The Conversion of Ammonium Thiocyanate into Thiourea and of Thiourea into Thiocyanate. BY JOHN WADDELL. *J. Phys. Chem.*, **2**, 525-535.—The author shows that these two reactions reach the same state of equilibrium, the mixture then consisting of 21.2 per cent. of thiourea at 152° - 153° and 20.9 per cent. at 160° - 161° ; and he finds that the velocity of each of them

corresponds as closely as the rather large experimental errors permit, with the requirements of the mass-action law for a reversible monomolecular reaction. The rate of the reactions is about twice as great at 160° as it is at 152°.

Pressure-temperature Diagrams for Binary Systems. BY WILDER D. BANCROFT. *J. Phys. Chem.*, 3, 1-11.

Reversible Reactions. BY JOHN WADDELL. *J. Phys. Chem.*, 3, 41-45.—The author shows, as has been previously done in detail by Ostwald (*Lehrbuch*, II, 2, 251, 257), that the velocity of a reversible first order reaction can be expressed by the equation $\frac{dx}{dt} = (k_1 + k_2)(\xi - x)$, in which k_1 and k_2 are the velocity constants of the opposing reactions and ξ is the value of x after equilibrium is reached; but that the corresponding simple expression, $\frac{dx}{dt} = (\xi - x)^2$, is not applicable to reversible second order reactions. An omission made by Ostwald in deriving the formula that does hold true in the latter case is pointed out.

Demonstration of the Phase Rule. BY PAUL SAUREL. *J. Phys. Chem.*, 3, 69-71; 137-143.

Dissociation Studies, I. BY WILDER D. BANCROFT. *J. Phys. Chem.*, 3, 72-94.—A graphical discussion of the equilibrium of two component systems.

The Electromotive Forces of Concentration Cells. BY J. E. TREVOR. *J. Phys. Chem.*, 3, 95-105.—The author shows the effect of the heat of dilution (which is ordinarily neglected) on the general equations expressing the relation between electromotive force, concentration, and temperature.

Correction. BY WILDER D. BANCROFT. *J. Phys. Chem.*, 3, 107.—This refers to an error in the calculation of the electromotive force between amalgams in an article by Cady (*J. Phys. Chem.*, 2, 554).

The Composition of Mixed Vapors, I. BY HECTOR R. CARVETH. *J. Phys. Chem.*, 3, 193-213.

On Maxwell's Theorem. BY PAUL SAUREL. *J. Phys. Chem.*, 3, 214-216.

Ternary Mixtures, IV. BY WILDER D. BANCROFT. *J. Phys. Chem.*, 3, 217-231.

Solid Solutions. BY HAMILTON P. CADY. *J. Phys. Chem.*, 3, 127-136.—The author states that naphthalene forms two series

of solid solutions both with monochloroacetic acid and with glycolic acid, but hardly any evidence is presented to show that the supposed solid solutions are not simple mixtures of the two separate solid phases.

The Equilibrium of Stereoisomers, III. BY WILDER D. BANCROFT. *J. Phys. Chem.*, 3, 144-155.

On the Molecular Weight of Sulphur in Carbon Bisulphide Solutions. BY H. T. BARNES. *J. Phys. Chem.*, 3, 156-159.—The author calculates a value for the molecular weight from the density (*sic*) of the solutions.

Freezing-point in Ternary Mixtures. BY JOHN WADDELL. *J. Phys. Chem.*, 3, 160-164.

Electrolytic Dissociation and Toxic Effect. BY J. F. CLARK. *J. Phys. Chem.*, 3, 263-316; *Botan. Gaz.*, 28, 289.—The author has determined the minimum concentration of a number of chemical substances that suffices (1) to retard or make irregular the mycelial development and prevent the fruiting of the spores of five different forms of mold, (2) to inhibit entirely their development, and (3) to kill them. The second of these effects is most accurately determinable and is the one mainly discussed. The results in the following table show the number of mols of each substance required to produce an inhibiting effect equal to that of one mol of mercuric chloride when dissolved in the same quantity of water, namely, 7300 liters.

Potassium bichromate	1.1	Trichloroacetic acid	321.
Silver nitrate	1.3	Hydrogen peroxide	375.
Potassium chromate	1.4	Ferrous sulphate	411.
Formaldehyde	5.	Copper sulphate	468.
Hydrocyanic acid	11.	Copper nitrate	479.
Cadmium nitrate	22.	Nitric acid	503.
Potassium cyanide	91.	Potassium hydroxide	593.
Nickel sulphate	120.	Sodium salicylate	650.
Ammonium hydroxide	182.	Sulphuric acid	732.
Cobaltous sulphate	206.	Hydrochloric acid	821.
Monochloroacetic acid	207.	Zinc sulphate	2150.
Dichloroacetic acid	229.	Potassium iodide	8775.
Acetic acid	296.	Ethyl alcohol	13164.

Various other neutral salts were also partially investigated; namely, potassium chloride, bromide, nitrate, and sulphate, and the potassium salts of the three chloroacetic acids. All of these salts were only very slightly toxic, having from $\frac{1}{11}$ to less than $\frac{1}{32}$ of the effect of hydrochloric acid. A consideration of the data from a theoretical standpoint leads the author to conclude that

the hypothesis that "the ionized part of a dissolved substance and that only is chemically active" is in disagreement with his own physiological results and also those of others. The main reason for this conclusion is that the partially dissociated acids like acetic and the chloracetic acids have nevertheless a much more powerful toxic effect than hydrochloric acid and than their neutral salts which contain the same ions in much larger quantity. Without wishing to defend the almost obviously untenable assumption that undissociated molecules never enter into chemical reactions, the reviewer would point out that one important fact has been generally overlooked in the interpretation of the physiological results with reference to the dissociation theory; and that is the well-known *catalytic* effect of hydrogen ions. Thus it is by no means justifiable to assume that because the NO_3^- ion has no appreciable toxic influence in a potassium nitrate solution, that the same is true of it in a nitric acid solution where hydrogen ions are simultaneously present. On the contrary, the fact that nitric is more toxic than hydrochloric acid would seem to be due more probably to the oxidizing effect of the NO_3^- ion in *acid* solution than to the undissociated HNO_3 molecules.

Some Boiling-point Curves. BY J. K. HAYWOOD. *J. Phys. Chem.*, 3, 317-327.—The author has determined the boiling-points at nearly constant pressure of alcohol-water, alcohol-ether, alcohol-carbontetrachloride, chloroform-carbontetrachloride, chloroform-methylalcohol, acetone-water, acetone-methylalcohol, acetone-ether, and acetone-carbontetrachloride mixtures. The results are presented graphically as well as in a tabular form.

Some Boiling-point Curves, II. BY J. K. HAYWOOD. *J. Am. Chem. Soc.*, 21, 994-1001.—In this paper the boiling-points of mixtures of benzene with chloroform, with carbon tetrachloride, with ether, with acetone, and with methyl alcohol, and of methyl alcohol with ethyl alcohol, with carbon tetrachloride, and with ether are given, and the corresponding boiling-point curves are shown. Of these curves those for benzene with methyl and ethyl alcohols, and for methyl alcohol with carbon tetrachloride, alone show the existence of a minimum boiling-point. No very definite conclusions are drawn from the results.

The Deliquescence of Potassium Nitrate, Sodium Nitrate, and Ammonium Nitrate. BY F. L. KORTRIGHT. *J. Phys. Chem.*, 3, 328-333.—These three salts were placed on watch-glasses in desiccators over mixtures of sulphuric acid and water of known concentrations and vapor-pressures, gradually increasing the amount of water in the acid until the salt gained in

weight. The minimum pressures of water vapor causing deliquescence were thus found to be at 20° between 8.5 and 10.8 mm. for ammonium nitrate, between 12.3 and 13.5 mm. for sodium nitrate, and between 15.5 and 16.5 mm. for potassium nitrate. The vapor-pressure of pure water at 20° is 17.4 mm. The author has determined by extrapolation from existing data the vapor-pressures of saturated solutions of the salts and find these in agreement with the deliquescence pressures, as theory requires.

On the Stability of Equilibrium of a One Component System.

BY PAUL SAUREL. *J. Phys. Chem.*, 3, 334-336.

Notes on the Energy Theory. BY J. E. TREVOR. *J. Phys. Chem.*, 3, 339-348.—The author differentiates the energy-changes that may take place in any system into (mechanical) work terms and into what he calls work-equivalent terms; the former are the product of forces into changes of space (distance, surface, or volume), and the latter, including all other energy-changes, are the product of potentials into changes of "quantity-coordinates."

Minimum Boiling-points and Vapor Compositions. BY J. H. PETTIT. *J. Phys. Chem.*, 3, 349-363.—The author has determined the boiling-points of various mixtures of methyl alcohol and ether, methyl alcohol and chloroform, and methyl alcohol and acetone, and of the four pure substances. The last two mixtures have a minimum boiling-point when the concentration of the methyl alcohol is about ten per cent. The purpose of the work was to test the "theorem" of Bancroft that all miscible liquids whose temperature-vapor pressure curves intersect form mixtures with a maximum or minimum boiling-point at some concentration. Besides his own experiments the author takes into consideration those of others, and concludes that the theorem is fully supported by the facts.

Dissociation Pressure before H. Sainte-Claire Deville. BY P. DUHEM. *J. Phys. Chem.*, 3, 364-378.—The author points out that in 1837, a French scientist, Georges Aimé, published a thesis in which he clearly stated the principle of constant dissociation-pressure, and communicated experiments confirming it, on calcium carbonate and hydrochloric acid, and sodium carbonate and various organic acids, in which he showed that these reactions ceased when the partial pressure of the carbon dioxide reached a definite value which was independent of the amounts of carbonate and acid brought together, of the capacity of the closed vessel, and of the presence or absence of air in the vessel.

On the Paraanisaldoximes. BY H. R. CARVETH. *J. Phys. Chem.*, 3, 437-451.—This investigation is along the lines of previous ones from the same laboratory. (See *this Rev.*, 4, 97.)

On the Relation between Pressure and Evaporation. BY EDWIN H. HALL. *J. Phys. Chem.*, 3, 452-456.—The author states that Nernst in deriving the equation for isothermal chemical equilibrium neglects the work done on the condensed substances that may be involved; and by making this correction the author derives a relation between temperature, total pressure, and vapor-density in the case of a pure solid or liquid. The reviewer would point out, however, that if the pressure is so high as to make the mentioned correction appreciable, the deviations from the simple gas laws would be so also; and the derived relation would have a different form.

Thermal Coefficients. BY J. E. TREVOR. *J. Phys. Chem.*, 3, 523-547.—A complete mathematical treatment of the fundamental relations between pressure, volume, temperature, entropy, energy, free energy, the heat function, and the thermodynamical potential.

Hydrates in Solution. BY WILDER D. BANCROFT. *J. Phys. Chem.*, 3, 551-554.—Nernst showed that, in the case of a hydrate of the formula, $A \cdot nH_2O$, dissociating into its components A and n molecules H_2O , the mass-action law requires, since for not too concentrated solutions the concentration of the water is constant, that the concentrations of the hydrate and anhydrous substance be proportional to each other, or that the proportion of hydrated substance be independent of the concentration. This last statement, which is obviously to be understood to mean that the ratio of hydrate to the sum of hydrate and anhydrous substance is constant, is misinterpreted by the author so as to signify that the ratio of hydrate to the total amount of substance present, no matter what its molecular condition may be, is constant; and it is correctly shown by him that in this sense the statement is not applicable to hydrated salts which at the same time undergo electrolytic dissociation; for in that case the amount of undissociated anhydrous salt is, it is well-known, not proportional to the total amount of salt present. In regard to the equilibrium of hydrated salts in solution, the reviewer would add that it is just as probable that the ions are hydrated as that the undissociated molecules are.

The Specific Heat of Metals at Low Temperatures. BY C. C. TROWBRIDGE. *Science*, 8, 6-11. The author has determined by the method of mixtures the mean specific heat of copper, iron, and aluminum between $+13$ and -181.4° , the temper-

ature of boiling oxygen obtained from liquid air, and finds the values 0.0868, 0.0914, and 0.1833 respectively, which are less than the mean specific heats between 23° and 100° by 7.6, 21.3, and 15.7 per cent., respectively.

H. M. GOODWIN, REVIEWER.

Note on the Coefficient of Expansion of Ice. BY EDWARD L. NICHOLS. *Phys. Rev.*, 8, 184-186.—The author gives the result of a direct determination of the coefficient of expansion of artificial ice cut in the form of a bar 45 cm. long, 12 cm. wide, and 5 cm. thick. The measurements were made on a dividing-engine, during a period when the temperature varied from -3° to -17°. The value 0.0000540 ± 0.0000020 was obtained for a mean temperature interval of -8° to -12°. This value agrees well with that obtained by a similar method by Struve (0.0000530), and that obtained by Plücker and Geissler (0.0000528) by the dilatometer method. The results of Brunner and of Marchand are much lower: the differences may be due to the specimen of ice employed.

The Electrical Conductivity of Non-Aqueous Solutions. BY AZARIAH T. LINCOLN. *J. Phys. Chem.*, 3, 456-494.—This article forms a continuation of a paper by Kahlenberg and the author on the Dissociating Power of Solvents (*this Rev.*, 5, 47). Additional qualitative determinations of the solubility of eighteen inorganic salts, mostly chlorides, were first made in twenty-nine organic solvents. Quantitative conductivity measurements were then made of solutions of ferric chloride in various organic solvents, as well as of several solutions of bismuth, antimony, aluminum chloride, and silver nitrate. The work and conclusions of other investigators are discussed and recapitulated in connection with the results obtained and the general conclusion is drawn that the dissociation theory and all other proposed theories fail completely to explain the results thus far obtained for non-aqueous conducting solutions. No rigid classification of solvents with reference to the conductivity of their solutions is made, although in general the Thomson-Nernst dielectric constant criterion is found to hold true.

The Influence of Substituents on the Electrical Conductivity of Benzoic Acid. BY ALFRED TINGLE. *J. Am. Chem. Soc.*, 21, 792-804.—The electrical conductivity of the alkali salts of the ortho and para halogen-substituted benzoic acids was determined at 25° C. for dilutions of 64 and 128 liters. The difference between the molecular conductivity of the same salts of two metals was found to vary widely with the acid with which the metals were in combination. The results are, however, consid-

ered only preliminary. If further measurements are to embrace but two dilutions, it would be much better to choose 37 or 64 liters and 1024 liters for the reasons pointed out by Bredig, and then compare the extrapolated μ_{∞} values.

Differences of Potential between Metals and Non-Aqueous Solutions of Their Salts. BY LOUIS KAHLENBERG. *J. Phys. Chem.*, 3, 379-404.—The potential differences between silver and decinormal solutions of silver nitrate in various organic solvents were measured against a normal electrode by the usual compensation method with capillary electrometer. Also those between a number of other metals and decinormal solutions of such of their salts as were sufficiently soluble in organic solvents were determined. In all eighty-nine combinations including some twenty or more solvents were investigated. The conclusion drawn by Jones from a similar investigation with aqueous, methyl and ethyl alcoholic and acetone solutions, namely, that the electrolytic solution tension of a metal is a function of the solvent, is further verified by the more extensive data thus obtained. From the values of the potential difference of different metals in the various solvents, it is probable that the order of the metals in the voltaic series also varies with the solvent. Values of P , the electrolytic solution pressure, could not be computed for lack of data on the dissociation of the dissolved salts. On account of the fact that the electrolytic solution pressure varies with the solvent or mixture of solvents, the author seems to question the definiteness and value of the concept of the quantity as such. This seems to the reviewer unnecessary, for the same change of numerical value is observed in the solubility of all substances in different solvents; thus, ordinary solubility (that is, the solution pressure of the solute) and electrolytic solution pressure of a metal, are analogous quantities and are completely determined for a given substance by temperature, pressure, and the nature of the solvent.

The Specific Heat of Solutions which Are Not Electrolytes. BY WM. FRANCIS MAGIE. *Phys. Rev.*, 9, 65-85.—The author deduces (under certain assumptions), the relation that the specific heats of solvent and solute are constant at all concentrations. In somewhat different form the same result has also been recently published by Schiller. To test this relation, the specific heat of various solutions, cane-sugar, dextrose, mannite, and urea in water, as well as of solutions in alcohol and aniline was determined by a modification of Pfaundler's method. The relation was found to be verified in many cases. Alcohol, either as solute or solvent, formed an exception. The data are as yet insufficient to warrant the general conclusion that the molecular heat of a dissolved substance is independent of the solvent.

On the Spectra of Krypton. BY C. RUNGE. *Astro-phys. J.*, 10, 73.—The krypton was prepared as described by Ramsay and Travers from liquid air, and contained some argon. The gas emitted two spectra, one with Leyden jar or spark gap in the secondary circuit, the other without Leyden jar and spark gap, as is the case with argon. The lines due to krypton were measured, argon lines being eliminated by comparison with a photograph of an argon spectrum. It is interesting to note that the author identifies the lines of Ramsay's metargon with the carbon lines, as was first pointed out by Schuster. He suggests that they may be due to a compound of argon and carbon which is not decomposed by sparking in an atmosphere of oxygen.

INORGANIC CHEMISTRY.

HENRY FAY, REVIEWER.

A Contribution to Our Knowledge of Dicarboxyl Cuprous Chloride. BY WILLIAM APP JONES. *Am. Chem. J.*, 22, 287-311.—Dicarboxyl cuprous chloride, prepared by passing carbon monoxide into cuprous chloride moistened with hydrochloric acid, was obtained in white crystalline plates which decomposed rapidly on drying. The compound was analyzed by decomposing it by heat, measuring the carbon monoxide liberated, and weighing the water; also, by determining the increase in weight of cuprous chloride when carbon monoxide is passed over it. The analyses show that the compound has the composition represented by the formula $\text{Cu}_2\text{Cl}_2 \cdot 2\text{CO} \cdot 4\text{H}_2\text{O}$. The compound is decomposed by diminished pressure and by increase of temperature. The gas given off does not show any marked increase in chemical activity. Oxygen and chlorine cause only very slight decomposition, and hydrogen and nitrogen are without action. The general properties of dicarboxyl cuprous chloride are similar to those of other inorganic compounds containing carbon monoxide. It is to be considered as a true chemical compound.

The Occlusion of Hydrogen by Metallic Cobalt and Other Metals. BY GREGORY PAUL BAXTER. *Am. Chem. J.*, 22, 351-364.—The amounts of hydrogen occluded by metallic cobalt, nickel, copper, and silver were determined by igniting the reduced metals in a current of air and weighing the water formed. The volume of occluded hydrogen depends upon the amount of surface presented, the purity of the metal, and the temperature of reduction. With cobalt the metal in the form of ingots occludes less hydrogen than that deposited in the form of foil by the electric current. The fine powder obtained from the reduction of the oxide occludes hydrogen in amounts varying with the purity