[CONTRIBUTION FROM THE CHEMICAL LABORATORY, UNIVERSITY OF WISCONSIN]

The Activity of Hydrogen Chloride in Glacial Acetic Acid and Other Solvents

BY BERNARD O. HESTON AND NORRIS F. HALL

During the past ten years several workers have presented studies of the activity of acids in nonaqueous solvents. Most of these have been concerned with solutions in the alcohols or in mixed solvents. Of those who worked in acid solvents we may mention Hutchison and Chandlee,[1] who studied the activity of sulfuric acid in glacial acetic acid by means of e. m. f. measurements, and Rodebush and Ewart,[2] who have measured the vapor pressure of hydrogen chloride from its solutions in glacial acetic acid. Recently, Eichelberger[3] has made cryoscopic studies in the same solvent.

Solutions of hydrochloric acid in acetic acid are readily studied by means of partial pressure measurements, but accurate results may be obtained only at the higher concentrations. Because of the high partial pressure of hydrogen chloride it is not easy to use the hydrogen electrode in such a system. In this work the chloranil electrode has been substituted for the hydrogen electrode and the activity of hydrochloric acid in glacial acetic acid obtained from measurements of the e. m. f. of cells without liquid junction. The data presented in the following section were obtained from cells of the type.

$$\mathrm{Pt}\left|\begin{array}{l}\text{chloranil (satd.)}\\ \text{hydrochloranil (satd.)}\end{array}\right.\quad \mathrm{HCl}(m)\quad \mathrm{AgCl\ (satd.)}\mid \mathrm{Ag}$$

The difference between the voltages of two such cells at different values of m is a measure of the free energy of dilution of the hydrochloric acid, but if chloranil is reduced at the silver electrode

$$2\mathrm{Ag} + \mathrm{C_6Cl_4O_2} + 2\mathrm{HCl} \longrightarrow 2\mathrm{AgCl} + \mathrm{C_6Cl_4(OH)_2}$$

temporary differences in the hydrochloric acid concentration in different parts of the cell may result. Preliminary tests showed in fact that this side reaction may be made to proceed in the cell for a time with a measurable velocity. The usual way of avoiding such action by separating the chloranil and the silver electrode in opposite arms of an H-cell proved impracticable because of high resistance and the difficulty of equalizing the hydrochloric acid concentration in the two sides. In the cells used, some initial irregularities and part of the difficulty of obtaining reproducible potentials were doubtless due to this side reaction, but almost all e. m. f. measurements in acetic acid are difficult to reproduce within a millivolt. That the measured potentials actually correspond (within the experimental error) to the values of m(HCl) subsequently determined in each case by analysis, and not to some unknown concentration gradient in the cell, is indicated by Fig. 3 as explained below.

Experimental Procedure

Fifty-ml. samples of acetic acid were saturated with chloranil and hydrochloranil so that a considerable excess of both solids was present, a small amount of a stock solution of hydrochloric acid in acetic acid added and the solution thoroughly mixed. This sample was then used to fill two cells each of which contained a silver–silver chloride and a chloranil electrode. These two cells were connected by a broad arm half full of the solution so as to form a sort of H-cell, and permit the two cells to attain the same concentration in case any of the hydrogen chloride escaped while they were being filled. The double cell was then placed in an air thermostat at 25 ± 0.2° and the first readings taken after one hour with subsequent readings at intervals of one hour until two or three readings which checked within one millivolt were obtained. This usually required four or five hours, and readings once obtained did not change over a period of seven or eight hours. It proved impossible to measure the e. m. f. across the H-arm because of the high resistance. The measurements reported are averages of the e. m. f. in the two branches, which usually agreed within 1 millivolt.

After equilibrium was established samples of the solution were run into a measured amount of standard diethylaniline in acetic acid, the weight of the sample obtained and the excess base titrated with standard perchloric acid in acetic acid. The titrations were carried out electrometrically using the chloranil electrode against a saturated calomel half cell with a glass-stoppered salt bridge containing lithium chloride in acetic acid.

Apparatus and Materials.—The silver–silver chloride electrodes were prepared by the method of Carmody[4] and about twenty of these electrodes were used in the course of this work.

The acetic acid was prepared by redistilling C. P. "99.98%" acetic acid obtained from the Niacet Chemical Company. The redistilled acid was of the same quality as that previously used in this Laboratory, the melting point of which was reported as 16.63°. More recent work of Eichelberger and La Mer[5] would indicate that our acid possibly contained about 0.02% water.

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[1] Hutchison and Chandlee, THIS JOURNAL, **53**, 2881 (1931).

[2] Rodebush and Ewart, *ibid.*, **54**, 419 (1932).

[3] Eichelberger, *ibid.*, **56**, 799 (1934).

[4] Carmody, *ibid.*, **51**, 2905 (1929).

[5] Eichelberger and La Mer, *ibid.*, **55**, 3635 (1933).

A stock solution of hydrochloric acid in acetic acid was prepared by passing dry hydrogen chloride gas through the acetic acid. The hydrogen chloride was generated by the action of concd. sulfuric acid on C. P. sodium chloride, passed through a spray trap and a column filled with glass wool and beads moistened with concd. sulfuric acid.

Perchloric acid was made up by diluting a stock solution previously prepared in this Laboratory from the anhydrous acid. It was standardized against weighed samples of diphenylguanidine.

The standard base, diethylaniline, was made by weight to approximately the strength desired and standardized against the perchloric acid.

Experimental Results

Measurements were made at about forty different concentrations extending from $m = 0.000596$ to $m = 0.1365$. From the data obtained we may plot[6] $E^{0\prime}$ against $\sqrt{m}$ and attempt an extrapolation to obtain E^0 in the conventional way. When this is done, however, a curve results which at the lower concentration has not approached nearly enough to a straight line to justify the extrapolation. Due to this fact and to the success of La Mer and Eichelberger in the application of the Gronwall, La Mer and Sandved[7] extension of the Debye and Hückel theory to solutions of sulfuric acid in acetic acid, it was thought advisable to try this method of calculating E^0.

For this purpose the equation

$$-\log f = 19.63\left(\frac{Z^2}{a}\right)\frac{x}{1+x} - 320.7\left(\frac{Z^2}{a}\right)^3 10^3({}^1/{}_2X_3(x) - 2Y_3(x)) - 26{,}200\left(\frac{Z^2}{a}\right)^5 10^5[{}^1/{}_2X_5(x) - 4Y_5(x)]$$

is used where $x = \kappa a$ and for acetic acid $\kappa = 1.1734 \times 10^8\sqrt{c}$. $E^{0\prime}$ has been calculated from concentrations expressed as molality or moles per 1000 g. of solvent. For the calculation of $-\log f$ the concentration is expressed as moles per 1000 cc. of solution. Since the density of the solutions used was unknown, the density of the solvent, 1.043 g./cc., was used to convert from moles per 1000 g. to moles per 1000 cc. A few trial density determinations showed that no significant error was thus introduced.

E^0 is obtained from the equation

$$E^0 = E^{0\prime} - 0.1183 \log f$$

if the proper value of "a" is assumed. If the calculated values of E^0 are plotted against $\sqrt{m}$ for different values of "a," the most probable value of "a" is found to be 11 Å. and for the corresponding value of E^0 to be -1.045 volts.

It is interesting that La Mer and Eichelberger[7] also found a value of 11 Å. for "a" in their studies of sulfuric acid solutions. Using -1.045 volts for E^0 and the calculated values of $E^{0\prime}$ a smooth curve was drawn. Table I shows the values of $E^{0\prime}$ as obtained from this smoothed curve.

TABLE I

$\sqrt{m}$	E_0'	$\sqrt{m}$	E_0'	$\sqrt{m}$	E_0'
0.00	1.045 ($E°$)	0.08	0.932	0.25	0.880
.02	0.992	.10	.922	.30	.871
.04	.964	.15	.9035	.35	.863
.06	.945	.20	.8905	.40	.856

The deviations of the experimental points from the smoothed curve are sufficiently indicated by Fig. 1, in which, to the scale of the diagram, the smoothed curve coincides with the theoretical curve as far as $\sqrt{m} = 0.2$. The broken line represents the theory from this point to higher concentrations. The scattering of the experimental points in the region of higher concentration is believed to be due chiefly to the difficulties encountered in analysis and the transfer of the solutions rather than to irregularities in the determination of the potential of the cells. As usual with measurements in acetic acid the data cannot be significantly related to the Debye-Hückel limiting law.[3]

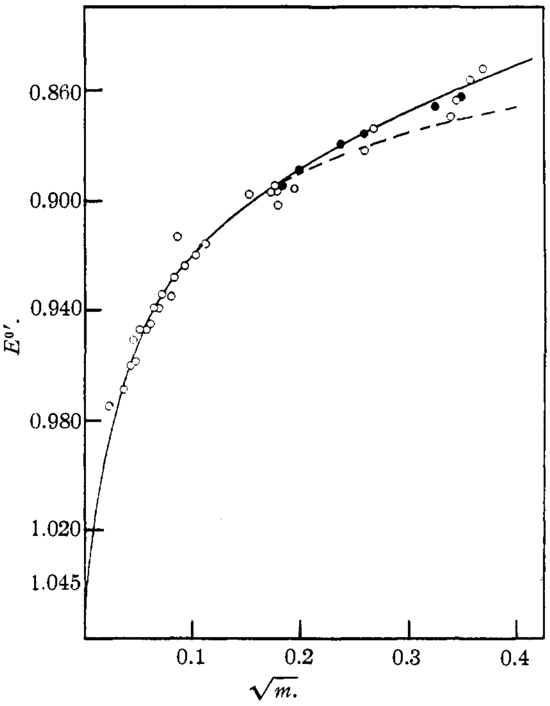

Fig. 1.

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[6] Lewis and Randall, "Thermodynamics," McGraw-Hill Book Co., New York, first edition, p. 334.

[7] La Mer and Eichelberger, THIS JOURNAL, **54**, 2763 (1932); Gronwall, La Mer and Sandved, *Physik. Z.*, **29**, 358 (1928).

It is possible to obtain a partial comparison of these data with the vapor pressure data of Rodebush and Ewart.[2] For a volatile electrolyte the activity coefficient may be calculated from the vapor pressure by the equation[8]

$$p^{1/2}/m = K\gamma$$

where k is an undetermined constant. k is then chosen such that the value of $E^0 - 0.1183 \log \gamma$ as calculated from vapor pressures from the most dilute solution, falls on the e. m. f. curve. With this same value of k other points are calculated. The values of $E^0 - 0.1183 \log \gamma$ obtained in this way (full circles in Fig. 1) are in excellent agreement with the e. m. f. data in so far as the shape of the curve is concerned.

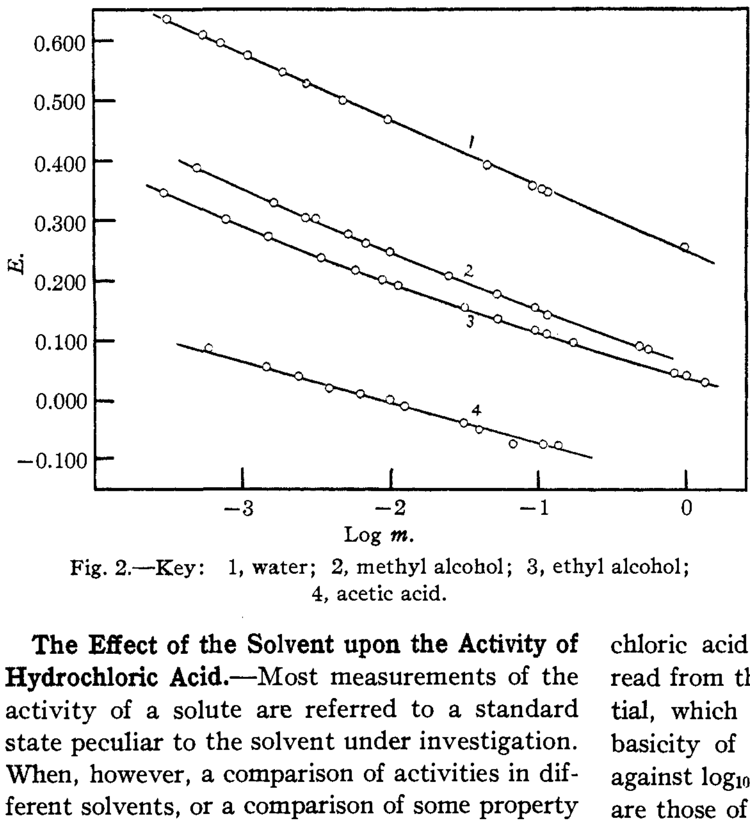

Fig. 2.—Key: 1, water; 2, methyl alcohol; 3, ethyl alcohol; 4, acetic acid.

The Effect of the Solvent upon the Activity of Hydrochloric Acid.—Most measurements of the activity of a solute are referred to a standard state peculiar to the solvent under investigation. When, however, a comparison of activities in different solvents, or a comparison of some property of the solvents, is desired, it becomes advantageous to refer all measurements to the same standard state. Thus in the case of an acid solute the activity of the acid at a given concentration will be an inverse measure of the basic or protophilic character of the solvent, which may be supposed to react as follows

$$\mathrm{HA} + \mathrm{S} = \mathrm{SH}^+ + \mathrm{A}^-$$

and, conversely, at a given value of the activity the concentration of the acid solution is also a measure of the solvent's tendency to behave as a base. The choice of a uniform standard state also makes possible the ready calculation of the free energy of transfer of the solute from a concentration m_1 in one solvent, to a concentration m_2 in another.

For this work the standard state chosen as most convenient is the ideally dilute aqueous solution. This has been suggested by Scatchard[9] and for this particular work presents as usual many advantages over other standard states. Chief among these advantages may be mentioned the host of data for aqueous solutions which are to be found in the literature, the relatively great basicity of water, and the fact that it is a good solvent for salts. The choice of this reference state for the function H_0 by Hammett and his co-workers[10] makes possible an interesting comparison of this significant function with our activity values.

The relative activities of a solute such as hydrogen chloride may be calculated from the e. m. f.'s of two cells without transference but in different solvents (provided the electrodes are the same), in just the same way as in the case of similar cells of two different concentrations in the same solvent. Thus in Fig. 2 are plotted the potentials of the cell $\mathrm{Pt} \mid \mathrm{H_2}\ \mathrm{HCl}\ \mathrm{AgCl}\ (\mathrm{satd.}) \mid \mathrm{Ag}$ over a considerable concentration range in a variety of solvents, and the relative activities of the hydrochloric acid in the different solvents are easily read from the corresponding differences of potential, which also give a measure of the relative basicity of the solvents. The data are plotted against $\log_{10} m$. The values for the aqueous cells are those of Carmody.[11] For methyl alcohol the data of Woolcock and Hartley,[12] for ethyl alcohol those of Nonhebel and Hartley,[13] and for acetic acid our own data (Table I) are used. Since our measurements were made with chloranil rather than with hydrogen electrodes, it was necessary to subtract from our measured values of $-E$ a value for the potential of the chloranil electrode

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(8) Lewis and Randall, Ref. 6, p. 329.

(9) Scatchard, THIS JOURNAL, **47**, 2098 (1925).

(10) Hammett and Paul, *ibid.*, **56**, 827, 830 (1934).

(11) Carmody, *ibid.*, **54**, 188 (1932).

(12) Woolcock and Hartley, *Phil. Mag.*, [7] **5**, 1133 (1928).

(13) Nonhebel and Hartley, *ibid.*, [6] **50**, 729 (1925).

referred to hydrogen. The value chosen was +0.680 volt, for the reasons stated elsewhere.[14]

A more general survey of the properties of these systems is obtained if instead of e. m. f.'s we plot another function of the activity. In Fig. 3 we have plotted[15] log a_2 against log N_2, where N_2 is the mole fraction of hydrogen chloride. The data for aqueous solution have been extended by the inclusion of the vapor pressure data given by Wynne-Jones;[16] those for acetic acid by the addition of the vapor pressure data of Rodebush and Ewart.[2] It is also possible to include the results of the partition measurements of Wynne-Jones for benzene and nitrobenzene.[16]

In this figure the log a_2 values for acetic acid all fall upon the same line whether determined from our own e. m. f. data or from the vapor pressure measurements. This is evidence that the e. m. f. values actually correspond (within the experimental error) to the analytically determined molalities and not to unknown concentrations due to diffusion and to the occurrence of the side reaction. The e. m. f. points are directly determined from the voltage differences between acetic acid and water cells at the same molality, while the vapor pressure points are obtained by a direct comparison of the molalities of solutions of hydrochloric acid of the same vapor pressure in water and in acetic acid without the introduction of any arbitrary constant.

The positions of the curves in the figure indicate the relative basicities of the solvents considered. Solvents which have little tendency to accept protons give curves which lie at high values of log a_2 and the solution obeys Henry's law. At the other extreme one should expect to find such solvents as liquid ammonia in which salt formation is presumably complete below $N_2 = 0.5$. At values of N_2 greater than 0.5 the curve would represent the vapor pressure of hydrogen chloride over the salt or crystalline solvates (now considering the hydrogen chloride as the solvent). Water is apparently an intermediate case.

It is obvious that all these curves must converge to the same point at log $N_2 = 0$. Because of the fact that the partial pressure curve for hydrogen chloride in benzene may be extrapolated in a straight line to give the vapor pressure of pure hydrogen chloride, this point[16] is believed to lie at about log $a_2 = 8.8$. The curve for the aqueous solution is changing in the proper direction at the highest concentration but the other curves do not extend to high enough concentrations to permit extrapolation.

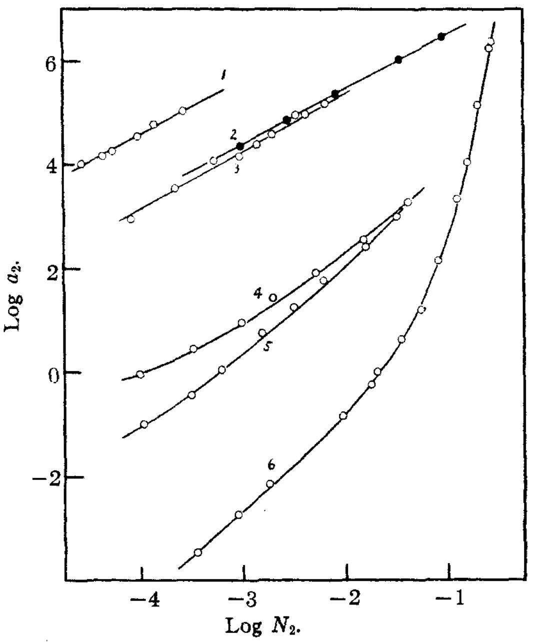

Fig. 3.—Key: 1, benzene; 2, nitrobenzene; 3, acetic acid; 4, ethyl alcohol; 5, methyl alcohol; 6, water.

Such a graph shows the relative basicities of a number of solvents with respect to one acid solute only. A change of solute would undoubtedly produce small changes in the position and character of the curves due to the differences in the properties of the solvates formed. These changes should not be great enough to cause the curves for the alcohols to fall below that for water or for the two alcohol curves to exchange places.

From the graph given it is easy to calculate approximate free energy changes of transfer from one solvent to another at any concentration covered by the graph.

Summary

1. The activity of hydrogen chloride in acetic acid has been determined from e. m. f. measurements in cells without liquid junction.

2. The extended theory of Gronwall, La Mer and Sandved is found to be in good agreement with the experimental data obtained.

3. The present measurements are shown to agree with the vapor pressure measurements of Rodebush and Ewart over the range covered by the latter.

4. The measurements in acetic acid have been

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[14] Heston and Hall, THIS JOURNAL, **55**, 4729 (1933).

[15] Lewis and Randall, Ref. 6, p. 336.

[16] W. F. K. Wynne-Jones, *J. Chem. Soc.*, 1064 (1930).

correlated with other activity measurements of various types for four other solvents and a general discussion has been given of the value of such correlations in comparing the basic strengths of different solvents.

MADISON, WISCONSIN RECEIVED MARCH 15, 1934

[CONTRIBUTION FROM THE JOHN HARRISON LABORATORY OF THE UNIVERSITY OF PENNSYLVANIA]

A Kinetic Study of the Oxidation of Cystine by Iodine[1]

BY KAMENOSUKE SHINOHARA[2] AND MARTIN KILPATRICK

Introduction

Real understanding of the physiological significance of organic sulfur compounds depends largely upon a knowledge of their oxidation processes. Although the study of the oxidation of cysteine (or thiol compounds in general) has been thorough,[3] the oxidation of cystine (or disulfides in general) has received little attention. One of the authors made a brief report[4] on the oxidation of cystine by iodine, following the establishment[5] of the stoichiometric equation

$$\mathrm{RSSR} + 5\mathrm{I}_2 + 6\mathrm{H}_2\mathrm{O} = 2\mathrm{RSO}_3\mathrm{H} + 10\mathrm{HI} \qquad (1)$$

Here it was pointed out that the reaction expressed by equation (1) probably takes place as a series of consecutive bimolecular reactions. In the present paper it will be shown that the rate-determining step in the series of reactions represented in (1) is the reaction between a molecule of cystine and a molecule of iodine or substances whose concentrations are proportional to their concentrations. Moreover, the effect of sodium chloride and of hydrogen and iodide ions will be shown.

Experimental Part

I. Purification of Cystine.—*l*-Cystine from Merck and Company was purified by the method of Gerwe[6] with a slight modification in the final stage. The modification consisted of suspending the *l*-cystine hydrochloride in redistilled water, dissolving it by introducing hydrogen chloride gas, and neutralizing with ammonia gas, upon which an abundant precipitate of *l*-cystine was obtained. The original *l*-cystine was found to contain 4×10^{-3}% iron when analyzed by the method of Lachs and Friedenthal.[7] The iron content was reduced by the purification to 4×10^{-5}%, which is practically the amount found in the hydrochloric acid used for the color test. The specific rotation $[\alpha]_{5461}^{25}$ of a 1% solution of *l*-cystine in *N* hydrochloric acid was −248.5° for the original cystine, and −250.7° for the purified material. The latter value is practically the maximum rotation obtainable for *l*-cystine.[8]

II. Technique.—Since even in acid solution cystine decomposes slowly on standing to produce cysteine,[9] it was necessary to make up a fresh solution of *l*-cystine for each experiment. *l*-Cystine was first dissolved in 2 *M* hydrochloric acid, and water was added until the solution was 0.01 *M* in *l*-cystine and 0.2 *M* in hydrochloric acid. To a 500-cc. flask containing the *l*-cystine and other constituents (excepting iodine) at 25 ± 0.01°, the desired volume of iodine solution in 4% potassium iodide was added, and the flask was filled to the mark with water and shaken. At intervals, 25 cc. of solution was pipetted out and run into a volume of standard sodium thiosulfate solution slightly less than that required for titration of the iodine present. The titration was then completed quickly by addition of sodium thiosulfate solution from the buret.

In order to correct for loss of iodine due to volatilization, reaction with impurities in the acetic acid used for buffers, etc., a blank experiment was run each time. The composition of the blank solution was the same as that of the reaction mixture except that an equivalent amount of *dl*-alanine was substituted for *l*-cystine. The consumption of iodine by cystine was calculated from the difference between the values obtained by titration of the blank and test solutions.

III. Determination of the Extent and Order of the Reaction.—To determine to what extent the stoichiometric equation (1) holds, two experiments were carried out, one in a sodium acetate–acetic acid buffer,[10] one in

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[1] This paper is abstracted from the first part of the thesis presented in April, 1933, by Kamenosuke Shinohara to the Faculty of the Graduate School of the University of Pennsylvania, in partial fulfilment of the requirements for the degree of Doctor of Philosophy.

[2] Robert McNeil Fellow at the Research Institute of the Lankenau Hospital, Philadelphia, Pa. The experimental work was carried out at the Institute.

[3] Ghosh, Raychandhuri and Ganguli, *J. Indian Chem. Soc.*, **9**, 43 (1932); **9**, 53 (1932); Elliot, *Biochem. J.*, **24**, 310 (1930); Elvehjem, *ibid.*, **24**, 415 (1930).

[4] Shinohara, *J. Biol. Chem.*, **97**, X11 (1932).

[5] Shinohara, *ibid.*, **96**, 285 (1932).

[6] Gerwe, *J. Biol. Chem.*, **91**, 57 (1931).

[7] Lachs and Friedenthal, *Biochem. Z.*, **32**, 130 (1911).

[8] Toennies and Lavine, *J. Biol. Chem.*, **89**, 153 (1930).

[9] Shinohara and Kilpatrick, *ibid.*, **105**, 241 (1934).

[10] According to Sano [*Biochem. Z.*, **168**, 29 (1926)] the solubility of *l*-cystine at P_{H} 3.0–6.0 is 4.64×10^{-4} *M* at 25°. However, upon dilution or neutralization of a solution of cystine in acid, precipitation does not occur for some time even though the concentration of cystine exceeds the solubility. Nor does seeding with a crystal of cystine start crystallization. Solutions 5, 8 and 10×10^{-4} *M* in *l*-cystine, in a buffer solution of the composition $\mathrm{CH_3COONa}$ 0.40 *M*, $\mathrm{CH_3COOH}$ 0.40 *M*, KI 0.024 *M*, HCl 0.060 *M*, remained clear for at least one week. A solution originally 12×10^{-4} *M* in *l*-cystine in the above buffer showed a very slight precipitate on the second day, and a solution originally 20×10^{-4} *M*, a precipitate after four hours. That the salt concentration has small effect upon the solubility was demonstrated by experiments in buffer solutions 0.4, 0.2 and 0.02 *M*, respectively, in sodium acetate and acetic acid. An excess of cystine was shaken with the solution at 25 ± 1° for three days, and upon analysis at the end of a week, the solutions were found to contain 5.2, 5.4 and 5.7×10^{-4} moles of cystine per liter, respectively.