ORGANIC AND BIOLOGICAL CHEMISTRY

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Infrared Spectra of Aqueous Solutions. II. Iminodiacetic Acid, N-Hydroxyethyliminodiacetic Acid and Nitrilotriacetic Acid[1]

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Infrared spectra of aqueous solutions of iminodiacetic acid, N-hydroxyethyliminodiacetic acid and of nitrilotriacetic acid are reported at various *p*H values. On the basis of the antisymmetric stretching bands of the ionized and un-ionized carboxyl groups, the structures of the free ligands are inferred and are compared with the results of potentiometric studies. The C–O stretching bands for the hydroxyethyl derivative are also described.

Introduction

As an approach to the study of the structures of metal chelates of aminopolycarboxylic acids in aqueous solution, it was decided to study the infrared spectra of three typical ligands as a function of *p*H, and to compare the results with the corresponding potentiometric titration curves. The infrared data are of considerable interest in themselves because of the structural information that such data supply about the ligands in aqueous solution, as well as providing further examples of the use of a new experimental technique. The resulting infrared spectra, and any structural information obtained from them, would be useful in following metal chelate formation in aqueous solution with various metal ions, as has recently been done in the case of metal chelate compounds of α-amino acids.[3]

Although potentiometric measurements in aqueous solution have provided valuable information on solution equilibria, such measurements have the disadvantage of not giving information on the structures of the reacting species. Infrared measurements on aqueous solutions, taken as a function of *p*H, or of hydrogen ion concentration, can in favorable cases provide the type of microscopic information needed for a more precise understanding of the aqueous neutralization reactions of complex organic acids.

Experimental

Compounds.—The disodium salt of iminodiacetic acid was purchased from the Eastman Organic Chemicals Co. N-Hydroxyethyliminodiacetic acid and nitrilotriacetic acid were obtained through the courtesy of the Dow Chemical Co., Midland, Mich.

Spectral Measurements.—Infrared spectra were measured with a Perkin–Elmer model 21 infrared spectrophotometer equipped with sodium chloride optics. The samples were aqueous solutions of the ligands in 0.10-ml. cells of 0.19-mm. thickness having barium fluoride windows. All aqueous solutions were prepared with 99.5% D_2O as solvent, purchased from New England Nuclear Corp., Boston, Mass. The concentrations of the ligand solutions prepared for measurements were in the range of 5~15 weight per cent. of solute. The spectra of the solids were measured by means of the potassium bromide pellet technique.

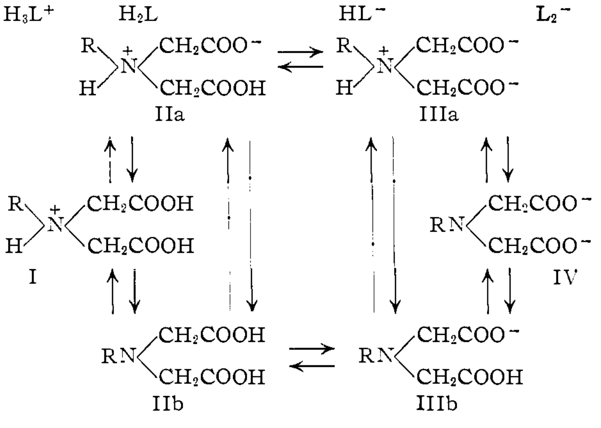
***p*H Measurements.**—The *p*H values of the solutions were controlled by the addition of small amounts of hydrochloric acid or metallic sodium, and were measured with a Beckman model GS *p*H meter fitted with extension glass and calomel electrodes, and calibrated with standard Beckman buffer solutions. The *p*H values listed were obtained from ordinary aqueous solutions which had the same molar concentration of solute as the 99.5% D_2O solutions used for infrared measurements. Thus the *p*H values of the solutions used for spectral measurements were slightly different, but well within the tolerances required for meaningful interpretation of the results and for comparison with previously published potentiometric data.

Results and Discussion

Iminodiacetic Acid (IMDA).—There are six possible species of iminodiacetic acid in aqueous solution as is shown in Scheme I (R = H). In order to distinguish between these species and to follow their interconversion, the infrared spectra of the solutions were measured as a function of *p*H. The results are illustrated in Fig. 1.

It is well known that the antisymmetric stretching band of the un-ionized carboxyl group occurs at 1730~1710 cm.$^{-1}$, whereas that of the ionized anion is found at 1620~1580 cm.$^{-1}$. By the use of this criterion together with the results of potentiometric titration[4] ($pK_2 \cong 2.54$, $pK_3 \cong 9.12$), the predominant species present at various *p*H values can be deduced, as is shown in Table I.

Scheme I.—Solution equilibria of IMDA and HIMDA (R = H, IMDA; R = CH_2CH_2OH, HIMDA).



It is interesting that the observed changes in the spectra (Fig. 1) give direct proof of the nature of the ions in solution. The completely protonated form I dissociates at low *p*H to give the dipolar ion IIa. The next step is the complete loss of protons from the carboxyl groups to give an anion IIIa in

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[1] This work was supported by a research grant, A-3095, from the National Institute of Arthritis and Metabolic Diseases, Public Health Service.

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[3] K. Nakamoto, Y. Morimoto and A. E. Martell, *J. Am. Chem. Soc.*, **83**, 4528 (1961).

[4] S. Chaberek and A. E. Martell, *ibid.*, **74**, 5052 (1952).

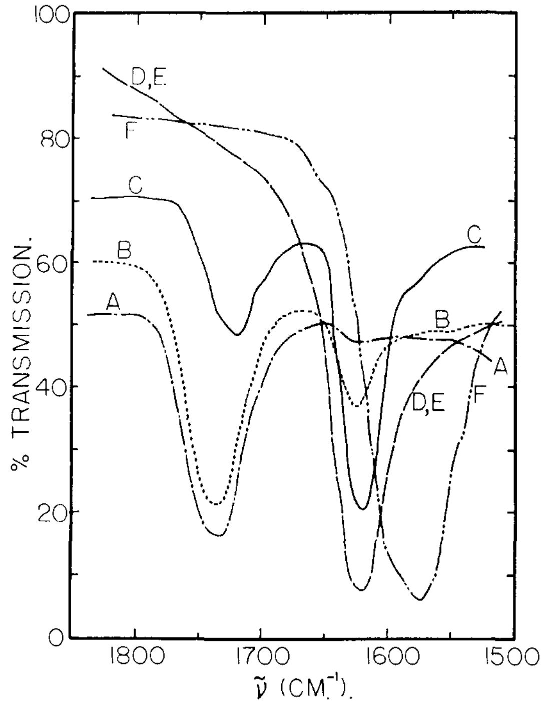


Fig. 1.—Infrared spectra of IMDA in D_2O solutions (concn. 6%): A, *p*H 0.4; B, *p*H 1.1; C, *p*H 2.2; D, *p*H 4.1; E, *p*H 7.5; F, *p*H 10.9.

which the remaining proton must be on the nitrogen atom. At high *p*H values, the band at 1619 cm.$^{-1}$ disappears and a new band appears as 1575 cm.$^{-1}$. Since IV is the only possible species under these conditions, the observed shift of the COO^- band to a lower frequency must be due to the loss of formal positive charge on the nitrogen atom of IIIa. Thus the frequencies of IIIa and IV are different, although both of them are due to the ionized COO^- groups.

Table I

Antisymmetric Carboxyl(ate) Absorption Bands of IMDA as a Function of *p*H

*p*H	Observed bands, cm.$^{-1}$ COOH	COO$^-$		Predominant species
0.4	1733(s)	....	....	I
2.2	1721(m)	1619(s)	....	IIa
4.1	....	1619(vs)	....	IIIa
10.9	....	....	1575(vs)	IV

The observed appearance and disappearance of the COOH and COO^- bands, respectively, is the first direct evidence for the structures of these aqueous amino acid species, which has been deduced previously by arguments based on the relative basicities of the carboxylate anion and of the amino group. These neutralization reactions are dramatically illustrated by the diminution and growth of the absorption bands.

Nitrilotriacetic Acid (NTA).—As is shown in Scheme II, there are eight possible species in aqueous solutions of nitrilotriacetic acid. Table II lists the antisymmetric carboxyl(ate) stretching frequencies at various *p*H values.

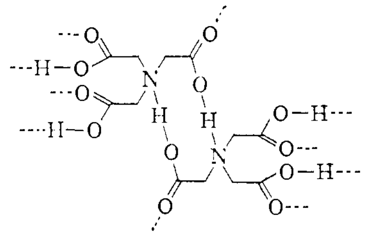
Table II

Antisymmetric Carboxyl(ate) Absorption Bands of NTA as a Function of *p*H

*p*H	Observed bands, cm.$^{-1}$ COOH	COO$^-$		Predominant species
2.3	1720(vw)	1623(m)	.....	VIIa
5.9	......	1625(vs)	.....	VIIIa
7.9	......	1610(vs)	.....	VIIIa
10.2	......	.....	1584(vs)	IX

It is known from potentiometric data[5] that the *pK* values of H_3L, H_3L^- and HL^{2-} are 1.89, 2.49 and 9.73, respectively. A comparison of these values with the data in Table II indicates that VIIa and VIIIa are predominant between *p*H 2.3 and 6. At high *p*H values, the COO^- band shifts to a lower frequency as was shown above for IMDA. This is again due to the loss of a positive charge on the nitrogen atom in going from VIIIa to IX. The neutral species, H_3L, is only slightly soluble in aqueous solution. Therefore it was not possible to distinguish between VIa and VIb from aqueous infrared spectra. It is noted, however, that the crystalline form of H_3L exhibits only one band at 1730 cm.$^{-1}$ favoring a structure in which none of the carboxyl groups are ionized.

This result is somewhat unexpected, in view of the differences in basicities of the carboxylate and amino groups. The fact that the frequency of the carboxyl band of nitrilotriacetic acid solid is, if anything, slightly higher than that of the same compound in solution may be due to less coördination of the carbonyl groups through hydrogen bonding in the solid state than takes place in solution through coördination with water molecules. A slight increment of positive charge on the nitrogen atom resulting from hydrogen bonding to a carboxyl group in the solid compound would also tend to shift the carboxyl carbonyl absorption in the observed direction. A possible arrangement of functional groups in the solid state is indicated by the structure



One would expect the protonated carboxyl groups, two per molecule, to be hydrogen bonded in an intermolecular fashion with other carboxyl groups as are most organic acids. The observation of only a single absorption band is explainable if it is assumed that the influence on the carboxyl frequency of hydrogen bonding to nitrogen would be approximately the same as hydrogen bonding to oxygen. The above structure is one of the many possible arrangements in the solid state compatible with the experimental data.

An attempt also was made to prepare H_4L^+ ion in concentrated hydrochloric acid. However, the

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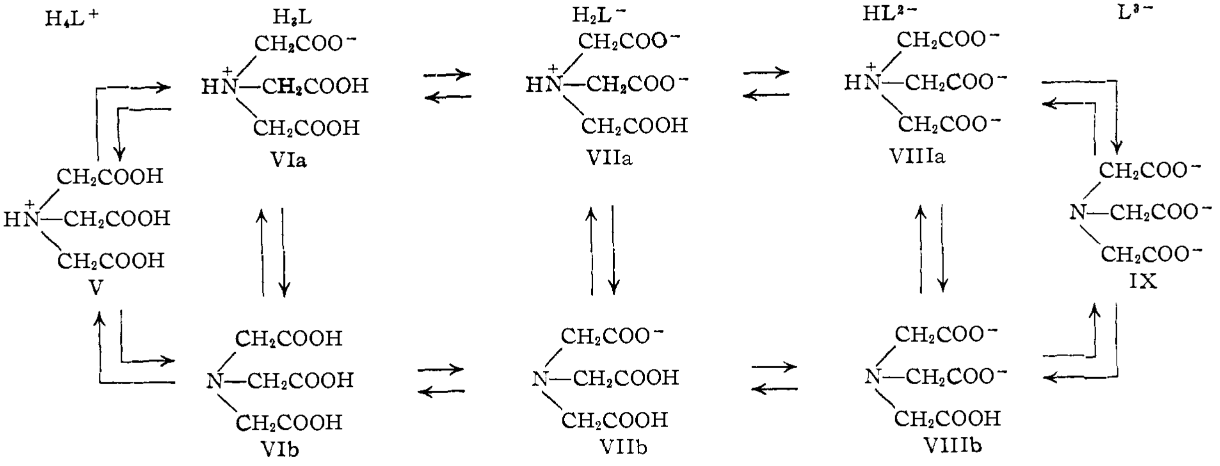
[5] G. Schwarzenbach and W. Biedermann, *Helv. Chim. Acta*, **31**, 331 (1945).

Scheme II.—The solution equilibria of NTA.



spectrum obtained was identical to that of H_3L in the crystalline state. Perhaps extremely strong acids such as concentrated sulfuric acid may be necessary to form such a cation in sufficiently high concentration to make infrared measurements possible.

N-Hydroxyethyliminodiacetic Acid (HIMDA).—The ligand N-hydroxyethyliminodiacetic acid is interesting as a well-known analog of both IMDA and NTA. In most of its coördination compounds, including proton coördination, the hydroxyethyl group functions as a neutral group in aqueous solutions. In some complexes, however, it seems to lose a proton to form the very basic alkoxide anion which contributes very strongly to the stability.

If one neglects possible slight ionization of the hydroxyethyl group, there are six possible species in aqueous solution, as is shown in scheme I (R = CH_2CH_2OH). Table III and Fig. 2 (A,B) show the results obtained at various *p*H values.

TABLE III

ANTISYMMETRIC CARBOXYL(ATE) ABSORPTION BANDS OF HIMDA AS A FUNCTION OF *p*H

*p*H	Observed bands, cm.$^{-1}$ COOH	COO$^-$	COO$^-$	Predominant species
0.2	1722(s)	.....	.....	I
0.5	1728(s)	1629(m)	.....	I, IIa
1.9	1728(m)	1628(vs)	.....	IIa, IIIa
2.4	1727(w)	1627(vs)	.....	IIa, IIIa
4.0	.....	1628(vs)	.....	IIIa
8.7	.....	1625(s)	1583(s)	IIIa, IV
10.1	.....	1625(w)	1583(vs)	(IIIa), IV
12.4	.....	.....	1582(vs)	IV

At extremely low pH (0.2), only one band appears at 1722 cm.$^{-1}$. This band is evidently due to the un-ionized COOH groups of I. At *p*H 0.5, however, bands appear at 1728 and 1629 cm.$^{-1}$, which must be due to un-ionized and ionized carboxyl groups, respectively. At this slightly higher *p*H, it is probable that partial conversion of I to II takes place. The appearance of the second band, therefore, is definite evidence for a species which has an ionized COO^- group. This conclusion is consistent with the previously accepted structure of this ligand in aqueous solution, which was based on the relative acidities of the COOH and R_3NH^+ groups.

At *p*H 1.9, the band at 1728 cm.$^{-1}$ is very much weakened whereas the band at 1629 cm.$^{-1}$ is very much strengthened. According to the results of potentiometric studies,[6] the value of pK_1 of this

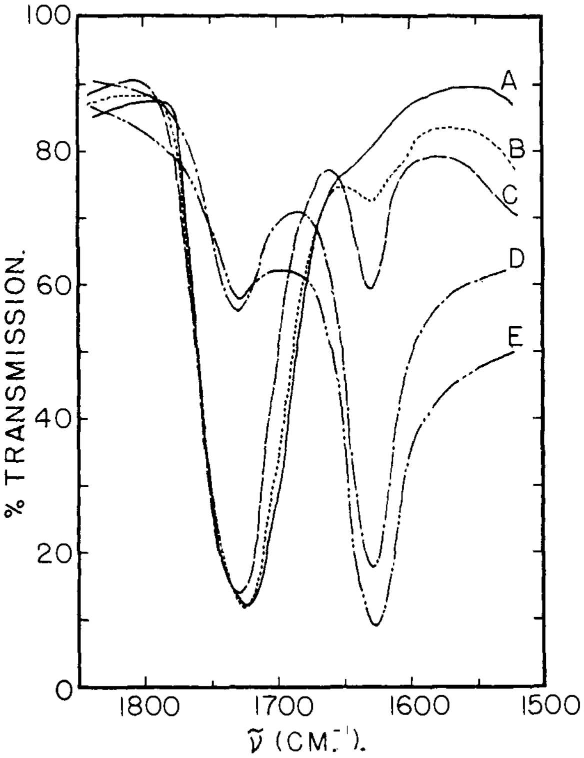


Fig. 2A.—Infrared spectra of HIMDA in D_2O solutions (concn. 5%): A, *p*H 0.2; B, *p*H 0.3; C, *p*H 0.5; D, *p*H 1.9; E, *p*H 2.4.

compound is 1.96. Therefore, at pH 1.9, about the same amount of H_2L and HL^- should be present. If so, the results of these infrared studies is in accord with the conclusion that IIa and IIIa represent the structure of H_2L and HL^-, respectively.

At pH 4.0, the band at 1728 cm.$^{-1}$ disappears entirely, showing a complete conversion of IIa to IIIa. At the *p*H 8.7, which is the value of pK_2 of

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(6) S. Chaberek, Jr., R. Courtney and A. E. Martell, *J. Am. Chem. Soc.*, **74**, 5057 (1952).

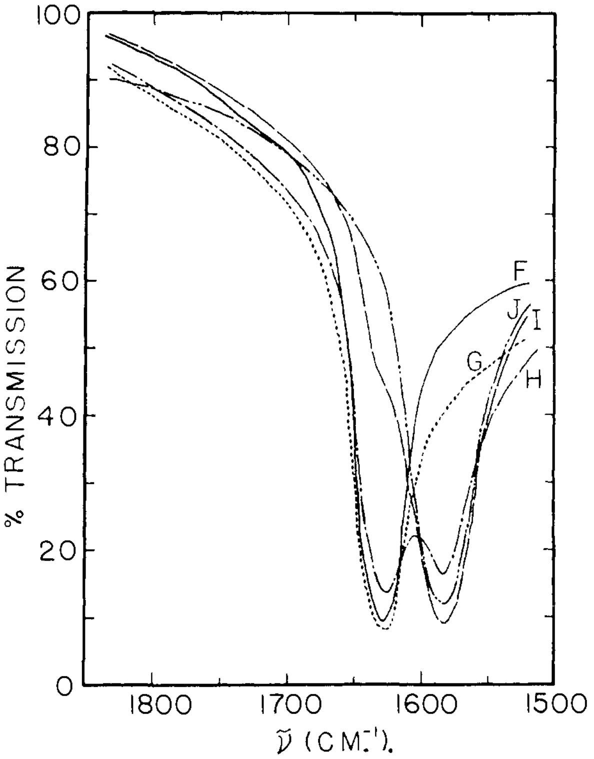


Fig. 2B.—Infrared spectra of HIMDA in D_2O solutions (concn. 5%): F, *p*H 4.0; G, *p*H 6.5; H, *p*H 8.6; I, *p*H 10.0; J, *p*H 12.4.

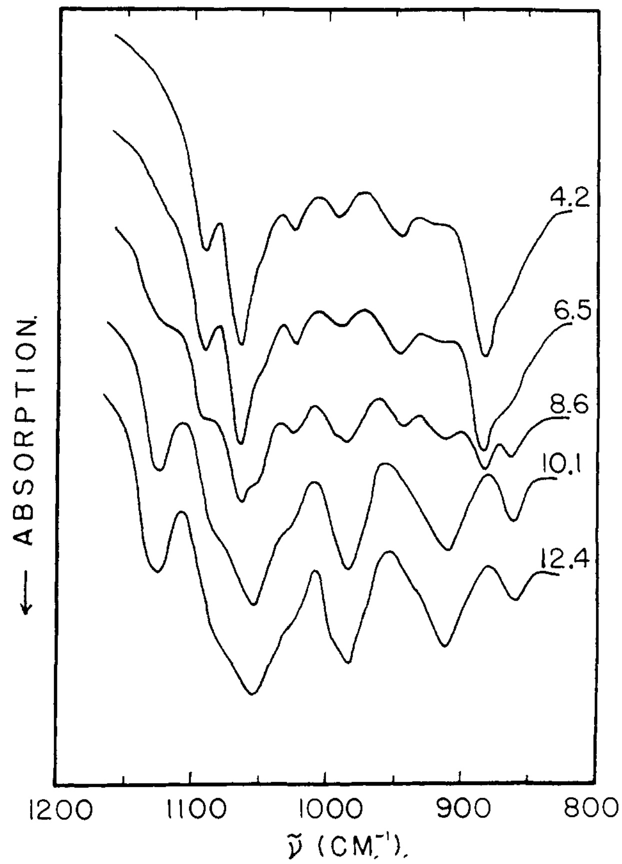
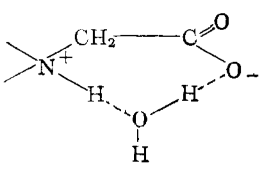


Fig. 3.—Infrared spectra of HIMDA in the low frequency region (15% HIMDA in D_2O solution; *p*H values as indicated).

HIMDA, two bands are observed at 1625 and 1583 cm.$^{-1}$. The band at 1625 cm.$^{-1}$ is accordingly attributed to IIIa, whereas the band at 1583 cm.$^{-1}$ must be due to IV, which differs from I, IIa and IIIa in the loss of the positive charge on the nitrogen atom. Such a shift of the COO^- band to a lower frequency is always observed when the proton of the $HN^+\equiv$ group is ionized. This effect is probably due to the loss of the inductive effect of the $HN^+\equiv$ group which occurs in ionization. Above *p*H 10, the band at 1625 cm.$^{-1}$ disappears, and only the band at 1583 cm.$^{-1}$ is observed, indicating the complete conversion of IIIa to IV. An alternative interpretation of the reason for this frequency shift is the interaction of the carboxylate and $\equiv NH^+$ groups through hydrogen bonding. Since any significant influence of the nitrogen atom on the carboxylate group through direct hydrogen bonding would be sterically impossible, such hydrogen bonding, if it exists, must take place through a solvent molecule, as indicated by the structure



In order to obtain evidence for the nature of the CH_2OH group at high *p*H, the infrared spectra were obtained in the range between 1200 and 800 cm.$^{-1}$, where the C–O stretching bands probably occur. As is seen in Fig. 3, the bands at 1125, 990, 913 and 860 cm.$^{-1}$ become stronger whereas the bands at 1091, 1025, 950 and 885 cm.$^{-1}$ become weaker as the *p*H increases. Since strong coupling is expected between the C–O, C–C and C–N stretching vibrations,[7] it is not possible to assign these bands empirically. Therefore they are attributed to the absorption of the N–C–C–O group as a whole. The observed *p*H dependence of the spectra cannot be attributed to the ionization of the CH_2OH group since the *p*H of the solution is too low to ionize the group. On the other hand, these observations may be explained by the loss of a positive charge on the nitrogen in the ionization of IIIa to IV. The resulting change in the inductive effect of the nitrogen atom would be expected to influence the vibrational frequencies of the N–C–C–O group as well as that of the COO^- group described above. Here again an alternate interpretation for the observed frequency shift exists in the possibility of intramolecular hydrogen bonding through one (or more) coördinated solvent molecules.

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(7) A. V. Stuart and G. B. B. M. Sutherland, *J. Chem. Phys.*, **24**, 599 (1956).