174. *Metal Complexes of* N-(2-*Pyridylmethyl*)*iminodiacetic Acid.*

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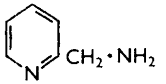
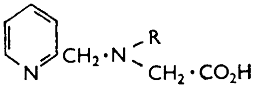
N-(2-Pyridylmethyl)iminodiacetic acid has been prepared and the stoichieometric stability constants of its proton and metal complexes have been measured potentiometrically at 20° in a medium of constant ionic strength $\mu = 0{\cdot}1$M.

The stability order is $Li^+ < Ba^{2+} < Sr^{2+} < Tl^+ < Mg^{2+} < Ca^{2+} < Ag^+ < Mn^{2+} < Fe^{2+} < Cd^{2+} < Pb^{2+} < Zn^{2+}$ for 1 : 1 complexes. However, 1 : 2 complexes are formed with Mn^{2+}, Co^{2+}, Cu^{2+}, and Cd^{2+}. The magnitudes of the stability constants are only slightly less than those for corresponding complexes of the much more basic aliphatic analogue ethylenediamine-*NN*-diacetic acid; possible reasons for this are discussed.

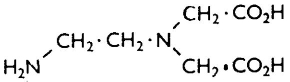
CONSIDERED as a donor atom, heterocyclic nitrogen is less basic than aliphatic nitrogen towards hydrogen ions and it might be expected to form less stable complexes with metals. However, the possibility of additional bonding through the π-orbitals of the heterocyclic ring and suitable orbitals of the metal ion might well reverse the picture. To examine this possibility we have prepared *N*-(2-pyridylmethyl)iminodiacetic acid (II; R = $CH_2{\cdot}CO_2H$) and compared its ability to form proton and metal complexes with that of the aliphatic analogue ethylenediamine-*NN*-diacetic acid (III).



(I; B)



(II: R = H; HY
R = $CH_2{\cdot}CO_2H$; H_2L)



(III; H_2Z)

Although Chaberek and Martell have referred to the preparation of *N*-(2-pyridylmethyl)aminoacetic acid (II; R = H) no details were given[1] and the stabilities of few of its metal complexes appear to have been studied.

The dibasic acid (II; R = $CH_2{\cdot}CO_2H$) was prepared by condensation of chloroacetic acid in alkali with 2-aminomethylpyridine (I), itself prepared by the reduction of pyridine-2-aldoxime. The stoicheiometric dissociation constants defined by $K_1 = [H^+][HL^-]/[H_2L]$ and $K_2 = [H^+][L^{2-}]/[HL^-]$ were obtained by potentiometric titration by procedures already described,[2] and are summarised in Table 1 together with values for ethylenediamine-*NN*-diacetic acid (III) and the conjugate acids of *N*-(2-pyridylmethvl)aminoacetic acid (II; R = H) and 2-aminomethylpyridine (I).

TABLE 1.

Stoicheiometric acid dissociation constants valid for $\mu = 0{\cdot}1$M.

Compound		pK_1	pK_2	pK_3	Ref.
(I)	H_2B^{2+}	1·78	8·18	—	*a*
		3·1	8·51	—	*b*
(II: R = H)	H_3Y^{2+}	2·19	2·46	8·18	*c*
(II: R = $CH_2{\cdot}CO_2H$)	H_2L	2·85 ± 0·02	8·25 ± 0·01	—	*d*
(III)	H_2Z	5·58	11·05	—	*e*

(*a*) Lane and Thompson, *J. Amer. Chem. Soc.*, 1960, **82**, 4179; 25°; 50% dioxan. (*b*) Goldberg and Fernelius, *J. Phys. Chem.*, 1959, **63**, 1245; 30°; $\mu \longrightarrow 0$. (*c*) La Coste (ref. 1); 25°; $\mu = 0{\cdot}1$M. (*d*) Present work; 20°; $\mu = 0{\cdot}1$ (KNO_3). (*e*) Schwarzenbach, Anderegg, Schneider, and Senn, *Helv. Chim. Acta*, 1955, **38**, 1147; 20°; $\mu = 0{\cdot}1$ (KCl).

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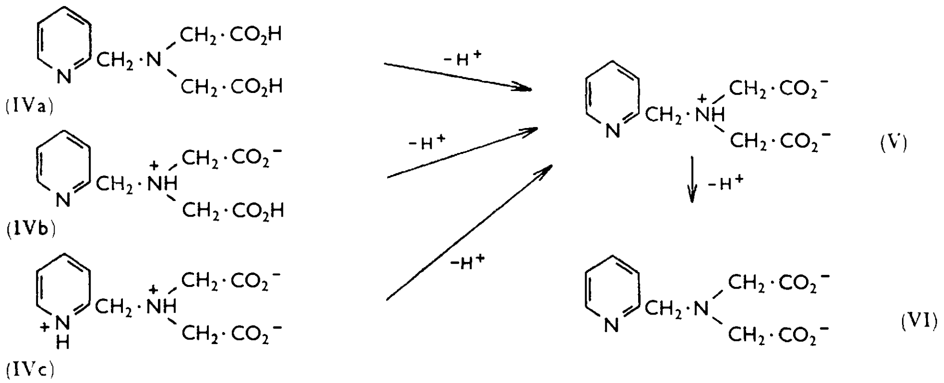
[1] La Coste, Doct. Dissertation, Clarke University, 1957, quoted by Chaberek and Martell in " Organic Sequestering Agents," Prentice Hall, New York, 1960.

[2] Irving and da Silva, *J.*, 1963, 448.

The value found for pK_2 = 8·25 is much the same as those found [3] for other derivatives of iminodiacetic acid and corresponds to the dissociation of a proton from the ammonium ion (IV); the value pK_2 = 8·18 for the dissociation of the proton from HB^+ is directly comparable. On the other hand, it is not immediately obvious whether the value pK_1 = 2·85 refers to the dissociation of a proton from one of the two equivalent carboxyl groups (IVa), or from the remaining carboxyl group or the pyridinium ion of the zwitterion (IVb) or (IVc). There can be no doubt that the value pK_1 = 1·78 relates to the dissociation of a proton from the pyridinium ion of the acid H_2B^{2+} conjugate to (I). The measurements refer to 50% dioxan and it is known that an increase of ~1 pK unit in the dissociation



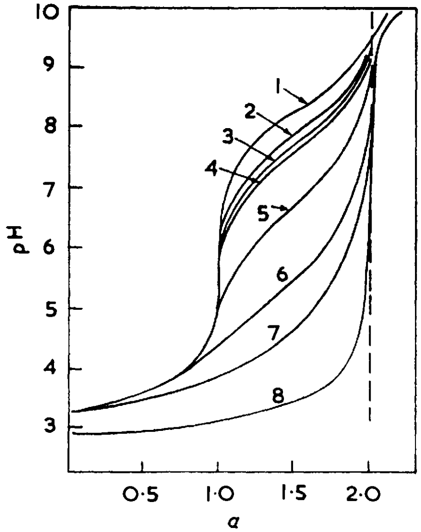


Fig. 1.

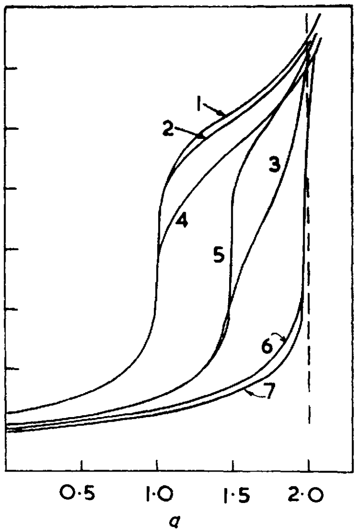


Fig. 2.

Fig. 1. Titration curves for *N*-(2-pyridylmethyl)iminodiacetic acid and a variety of ions.

1, Acid. 2, Ba^{2+}. 3, Sr^{2+}. 4, Mg^{2+}. 5, Ca^{2+}. 6, Ag^+. 7, Mn^{2+}. 8, Pb^{2+}. The curve for Pb^{2+} is almost coincident with that for Co^{2+}, and that for Mg^{2+} coincides with the curve for Tl^+.

Fig. 2. Titration curves for *N*-(2-pyridylmethyl)iminodiacetic acid (1) and (2) Li^+ (10 : 1), (3) Cd^{2+} (1 : 2), (4) Tl^+, (5) Co^{2+} (1 : 2), (6) Fe^{2+} (1 : 1), and Cd^{2+} (1 : 1).

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[3] *Chem. Soc. Special Publ.* No. 6, " Stability Constants. Part I. Organic Ligands," London, 1957.

TABLE 2.

Titration of *N*-(2-pyridylmethyl)iminodiacetic acid at 20°.

(1) $C_L = 1{\cdot}000 \times 10^{-3}$M; $C_M = 0{\cdot}010$M; $C_B = 0{\cdot}1008$M-tetramethylammonium nitrate.

	pH							
a	K^+ *	Mg^{++}	Ca^{++}	Sr^{++}	Ba^{++}	Ag^+	Tl^+	Pb^{++}
0	3·258							2·840
0·102	3·301							
0·203	3·352							2·883
0·305	3·441							
0·406	3·480							2·936
0·508	3·561							
0·609	3·664							2·992
0·711	3·800							
0·812	4·000							3·061
0·914	4·382							
1·015	6·560	5·684	5·040	5·890	6·021	4·450	5·692	3·146
1·117	7·371	6·541	5·660	6·740	6·904	4·721	6·584	3·195
1·218	7·699	6·901	6·010	7·093	7·257	4·953	6·949	3·250
1·320	7·930	7·168	6·282	7·349	7·510	5·162	7·208	3·311
1·421	8·120	7·397	6·525	7·569	7·715	5·371	7·436	3·388
1·523	8·299	7·611	6·769	7·767	7·910	5·596	7·650	3·477
1·524	8·476	7·830	7·022	7·974	8·104	5·850	7·869	3·590
1·726	8·672	8·070	7·310	8·199	8·319	6·171	8·100	3·747
1·827	8·907	8·350	7·659	8·469	8·570	6·120	8·386	3·990
1·929	9·211	8·740	8·161	8·841	8·925	7·421	8·752	4·546
2·030	9·614	9·373	9·209	9·425	9·459	9·195	9·364	9·020
2·132	9·965	9·861	9·850	9·891	9·905	9·892	9·862	9·748

* No metal added, 10 ml. of water being used in place to make up the volume to 120 ml. The results are used to calculate the acid dissociation constants of the ligand acid.

(2) $C_L = 9{\cdot}75 \times 10^{-4}$M; $C_M = 0{\cdot}010$M; $C_B = 0{\cdot}1008$M.

	pH			
a	Li^+ †	Co^{2+}	Cd^{2+}	Zn^{2+}
0	—	2·847	2·921	2·829
0·208	—	2·854	2·969	2·873
0·417	—	2·947	3·020	2·925
0·625	—	3·005	3·080	2·985
0·833	—	3·075	3·151	3·052
1·042	6·550	3·160	3·238	3·139
1·146	7·304	3·211	3·290	3·189
1·250	7·622	3·269	3·348	3·244
1·354	7·846	3·331	3·415	3·310
1·458	8·032	3·410	3·495	3·388
1·563	8·209	3·503	3·592	3·480
1·667	8·381	3·627	3·720	3·601
1·771	8·570	3·790	3·900	3·767
1·875	8·790	4·068	4·195	4·038
1·979	9·077	4·867	5·050	4·805
2·083	9·448	9·100	9·270	8·436

† $C_M = 0{\cdot}100$M. Volumes corrected to give $\mu = 0{\cdot}1$M.

(3) $C_L = 1{\cdot}000 \times 10^{-3}$M; $C_M = 0{\cdot}005$M; $C_B = 0{\cdot}1008$M.

	pH		
a	Cd^{2+}	Co^{2+}	Cu^{2+}
1·000	3·503	3·498	3·496
1·500	5·158	6·579	6·675
1·600	5·870	7·460	7·530
1·650	6·121	7·670	7·729
1·700	6·369	7·881	7·880
1·800	6·896	8·250	8·140
1·850	7·217	8·459	8·254
1·900	7·639	8·685	8·361
1·950	8·302	8·962	8·468
2·000	9·060	9·270	8·570

(4) $C_L = 1{\cdot}000 \times 10^{-3}$M; $C_B = 9{\cdot}524 \times 10^{-2}$M.

	pH		
a	Mn^{2+}	Mn^{2+}	Fe^{2+}
1·047	3·932	—	3·350
1·142	4·063	—	3·407
1·237	4·211	—	3·470
1·332	4·381	—	3·540
1·428	4·571	—	3·620
1·523	4·796	6·740	3·720
1·570	—	7·095	—
1·618	5·059	7·351	3·849
1·665	—	7·562	—
1·713	5·391	7·755	4·012
1·760	—	7·934	—
1·808	5·839	8·121	4·267
1·855	—	8·330	—
1·904	6·591	8·564	4·720
1·951	—	8·860	—
2·000	8·720	—	5·650
C_M	0·010	0·005	0·010

constant occurs for pyridine itself when the volume concentration of dioxan is reduced from 50% to zero.[4,5] On this basis a value for $pK_1 \approx 2{\cdot}8$ appears to correspond to the dissociation of the pyridinium proton in aqueous solution. On the other hand, values for pK_1 in the range 2—3 are common for derivatives of iminodiacetic acid for which the dissociation of a carboxylate proton is the only possibility.[3] If the value $pK_1 = 2{\cdot}19$ corresponds to the dissociation of the carboxylic hydrogen from H_3Y^{2+} the sequence (IVc) $\longrightarrow$ (V) $\longrightarrow$ (VI) seems most likely, though it is not improbable that the measured pK values refer to complex equilibria involving proton losses of comparable energy requirements.

Figs. 1 and 2 show typical titration curves for the variation of pH with the degree of neutralisation, a, in systems containing known total amounts, C_L and C_M, of ligand and metal, respectively, in a medium of constant ionic strength $\mu = 0{\cdot}1\text{M-KNO}_3$ at 20°. Metal stability constants defined by $K_{ML} = [ML]/[M][L]$ and $\beta_2 = [ML_2]/[M][L]^2$ (where charges are omitted in the interests of generality) were calculated as described previously.[6]

EXPERIMENTAL

All metal salts were of " AnalaR " grade. A laboratory sample of lithium bromide was repeatedly recrystallised from water, and the concentration of its solutions was determined by Volhard's method. Distilled water was passed through a mixed-bed resin, and its conductivity checked before use.

Preparation of N-(2-*Pyridylmethyl*)*iminodiacetic Acid* (II; R = $CH_2{\cdot}CO_2H$).—Pyridine-2-aldoxime [5 g.; m. p. 113—114° (from benzene)] in ethanol (75 ml.) was reduced with zinc dust (40 g.) and glacial acetic acid (40 ml.) added in portions during 8 hr. After basification the amine was removed by steam-distillation and its dihydrochloride recrystallised from alcohol (m. p. 223—225°; yield 5·5 g., 75%). This salt (5·4 g., 0·03 mole) was neutralised with 5N-sodium hydroxide, mixed with monochloroacetic acid (6·0 g., 0·06 mol. + excess) previously neutralised with 5N-sodium hydroxide, and heated to the b. p. while 5N-alkali (12 ml., 0·06 mol.) was added at such a rate that the pH was maintained between 8 and 10 (45 min.). The solution was boiled then for a further 10 min., chilled in ice-salt, and made strongly acid with hydrochloric acid. No complexone separated at this stage (A), so a portion of the solution was treated with an excess of copper sulphate and cautiously neutralised. The insoluble copper complex was collected, washed, suspended in boiling water, and treated with hydrogen sulphide. After separation of copper sulphide the filtrate was concentrated under reduced pressure, crystals of the desired acid (II; R = $CH_2{\cdot}CO_2H$) being obtained; these could be used as seed at stage A in later work. N-(2-*Pyridylmethyl*)*iminodiacetic acid* was then obtained as needles which, recrystallised from aqueous ethanol, had m. p. 174—175° (30%) [Found: C, 53·6; H, 5·3; N, 12·2%; M (by titration), 224. $C_{10}H_{12}N_2O_4$ requires C, 53·6; H, 5·4; N, 12·5%; M, 224·2).

Titration Procedure.—Details of apparatus have been given previously.[2] Mixtures containing 100 ml. of the ligand acid (of concentration C_L), 10 ml. of metallic salt (concentration C_M), and 10 ml. of 1·200M-potassium nitrate were titrated with carbonate-free potassium hydroxide (concentration C_B). The pH of the mixtures for various degrees of neutralisation, a (moles of alkali added per mole of ligand), are recorded in Table 2.

RESULTS AND DISCUSSION

Table 3 presents the results of the present measurements together with values for the stability constants of metal complexes of the analogous ligands *N*-(2-pyridylmethyl)aminoacetic acid (II; R = H) and ethylenediamine-*NN*-diacetic acid (III).

A comparison of results for acids (II; R = H and $CH_2{\cdot}CO_2H$) shows the expected increases in stabilities caused by the additional chelate ring provided by the latter acid and establishes that these ligands are tridentate and tetradentate, respectively, in their 1 : 1 complexes. Only with the transition metals was there tendency to take up a second

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[4] Irving and Rossotti, *Acta Chem. Scand.*, 1956, **10**, 72.

[5] Rossotti (Mrs. H. S.), D.Phil. Thesis, Oxford, 1954.

[6] Irving and Skelton, *J.*, 1958, 3540.

molecule of *N*-2-pyridylmethyliminodiacetic acid. Comparison of this ligand (II; R = $CH_2 \cdot CO_2H$) with the analogue (III) in which the major change is the replacement of the pyridine nucleus by an aliphatic secondary amine group reveals unexpected features. Despite the greater basicity (ΔpK is ~2·8 log units for both pK_1 and pK_2) of the aliphatic ligand the stabilities of its metal complexes are not so much larger and some values (*e.g.*, with Ca^{2+}, Sr^{2+}, and Ba^{2+}) are actually lower. It is evident that some new factor must be taken into account.

TABLE 3.

Stability constants of metal complexes. Data valid for μ = 0·1M.

	(II; R = H) *	(II; R = $CH_2 \cdot CO_2H$) †			(III) †	
Cation	log K_{ML}	log K_{ML}	log K_{ML_2}	log K_{MHL}	log K_{ML}	log K_{ML_2}
Mg^{2+}		3·90 ± 0·01			4·53	
Ca^{2+}		4·92 ± 0·02			4·63	
Sr^{2+}		3·65 ± 0·01			3·55	
Ba^{2+}		3·40 ± 0·01			3·19	
Li^{+}		1·71 ± 0·01				
Na^{+}		0·85 ± 0·03				
Tl^{+}		3·84 ± 0·01				
Ag^{+}		6·09 ± 0·01				
Mn^{2+}	4·2	6·97 ± 0·01	3·63 ± 0·01		7·71	3·70
Fe^{2+}		8·94 ± 0·01		3·84	9·81	3·86
Co^{2+}	8·1	10·16 ± 0·05	3·18 ± 0·05	4·95	11·78	4·13
Ni^{2+}	10·2	§	—	5·80	13·73	5·91
Cu^{2+}	11·8	§	(3·5)		15·90	
Zn^{2+}	7·6	10·87 ± 0·03	—		11·93	
Cd^{2+}	6·1	9·45 ± 0·02	5·29 ± 0·01		10·58	6·01
Pb^{2+}		10·31 ± 0·04	—		12·22	2·90

* Chaberek and Martell, see ref. 1; 25°; μ = 0·1. Values of K_{ML_2} not quoted. † Present work: 20°; μ = 0·1 (KNO_3). ‡ Schwarzenbach, Anderegg, Schneider, and Senn, *Helv. Chim. Acta*, 1955, **38**, 1147; 20°; μ = 0·1 (KCl) except Cd^{2+}, Pb^{2+} (μ = 0·1M-KNO_3). § Stabilities of ML too high for measurements by the present technique; value in parentheses is approximate.

Now pyridine is a π-acceptor and since some of the metals studied can donate suitable *d*-electrons to strengthen the initial co-ordinate bond by d_π-bonding it seemed possible that their interaction with the aromatic ring system of acid (II) might provide the explanation for the enhanced stability. This effect could certainly be expected with Ag^+, and to examine the influence of possible π-bonding we have plotted the stability constants of several complexes of this ion against the corresponding values for complexes of Tl^+ where, in consequence of the different electronic structure (. . . $5d^{10}6s^2$), donor π-bonding cannot take place (Fig. 3).

A line of approximately unit slope and intercept 0·5 log unit is obtained but the position of *N*-(2-pyridylmethyl)iminodiacetic acid (no. 10) is clearly anomalous and points to an abnormally stable silver complex. With some reservations, which will be commented upon below, we attribute this enhanced stability to π-bonding.

The same method can be extended to other pairs of ions and Fig. 4 shows a comparison between lead(II), where the presence of the inert pair again inhibits d-π bonding, and Cd^{2+} where, although suitable *d*-orbitals are provided for π-bonding, the higher positive charge renders them less available than in the isoelectronic Ag^+.[7] Fig. 4 indeed shows a small stabilisation for *N*-(2-pyridylmethyl)iminodiacetic acid (no. 12), but the effect is no longer so obvious. Plots (not reproduced) of results for Co^{2+} against Pb^{2+}, and for Zn^{2+} against Pb^{2+}, again show small stabilisations of complexes of the heterocyclic ligand with elements capable of donor π-bonding. It is significant that the biggest anomaly revealed by Fig. 4 is for ligand no. 16, *N*-(2-mercaptoethyl)iminodiacetic acid, $HS \cdot C_2H_4 \cdot N(CH_2 \cdot CO_2H_2)$, where donor π-bonding involving the sulphur atom could be of importance.

It must be emphasised that with the alkaline-earth metals the effects are less pronounced than they appear at first sight. This is because ethylenediamine-*NN*-diacetic acid is itself not a powerful ligand towards ions of rare-gas structure and probably also towards ions with the terminal shells $(n-1)d^{10}ns^2$. This is shown in Fig. 5 which displays the reasonably linear relation between the stability constants of calcium-ligand complexes and the pK values for the corresponding proton–ligand complexes. It will be seen that ethylenediamine-*NN*-diacetic acid is hardly better than an ordinary tridentate ligand when its high proton affinity is taken into consideration. This agrees with the known result that nitrogen-donor ligands of low basicity show little tendency (compared with oxygen) to co-ordinate to alkali-metal and alkaline-earth ions. In view of this circumstance we shall compare the behaviour of *N*-(2-pyridylmethyl)iminodiacetic acid, (H_2L), with that of a hypothetical ideal ligand, H_2X, of the same basicity and chelating potentialities. The stability constant for such an ideal ligand and calcium ions is interpolated from Fig. 5 which summarises data for ligands of the type $X{\cdot}N(CH_2{\cdot}CO_2H)_2$ as a linear plot of the equation $\log K_{ML} = a \times pK + b$. Apart from *N*-(2-pyridylmethyl)-iminodiacetic acid (no. 12) the two ligands of exceptional stability are nos. 13 (N-*o*-carboxyphenyliminodiacetic acid) and (no. 14) (nitrilotriacetic acid), both of which contain carboxy-groups suitably located to form strong chelate rings. If these two ligands are excluded

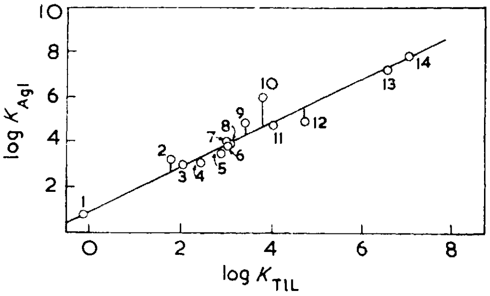

FIG. 3.

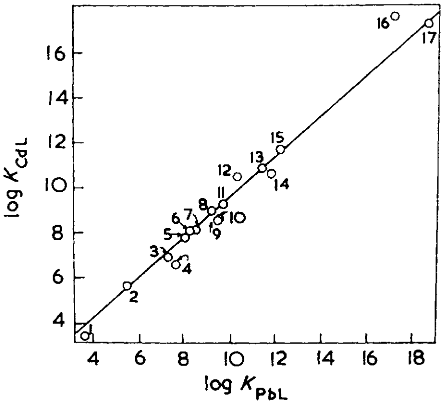

FIG. 4.

FIG. 3. A comparison of the stabilities of complexes of silver(I) and thallium(I) with a series of ligands. Data for 20° and $\mu = 0{\cdot}1$M are from ref. 7 and the present work supplemented by unpublished results for ligands no. 2—9 and 11—13. (1) Acetic acid; (2) iminodiacetic acid; (3) ethylenediaminetetra-acetic acid (MHY); (4) propane-1,2-diaminetetra-acetic acid (MHY); (5) *N*-*o*-carboxyphenyliminodiacetic acid; (6) 2-hydroxycyclohexyliminodiacetic acid; (7) methyliminodiacetic acid; (8) furfuryliminodiacetic acid; (9) cyclohexyliminodiacetic acid; (10) *N*-(2-pyridylmethyl)iminodiacetic acid; (11) *N*-(tetrahydropyran-2-ylmethyl)iminodiacetic acid; (12) nitrilotriacetic acid; (13) ethylenediaminetetra-acetic acid; (14) propane-1,2-diaminetetra-acetic acid.

FIG. 4. A comparison of the stabilities of complexes of cadmium with those of lead. Data for 20° and $\mu = 0{\cdot}1$M are from ref. 7 and present work supplemented by unpublished measurements for ligands no. 4 and 11.
(1) Phenyliminodiacetic acid; (2) di(carboxymethyl)aminoethyltrimethylammonium; (3) 2-ethoxycarbonyliminoethyliminodiacetic acid; (4) iminodiacetic acid; (5) methyliminodiacetic acid; (6) 3,3-dimethylbutyliminodiacetic acid; (7) carbamoylmethyliminodiacetic acid; (8) *N*-(2-methylthioethyl)iminodiacetic acid; (9) 2-methoxyethyliminodiacetic acid; (10) 2-hydroxyethyliminodiacetic acid; (11) 2-hydroxycyclohexyliminodiacetic acid; (12) *N*-(2-pyridylmethyl)iminodiacetic acid; (13) 2-carboxyethyliminodiacetic acid; (14) nitrilotriacetic acid; (15) ethylenediamine-*NN*-diacetic acid; (16) *N*-(2-mercaptoethyl)iminodiacetic acid; (17) ethylenediaminetetra-acetic acid.

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[7] Irving, *Chem. Soc. Special Publ.*, No. 13, p. 26, London, 1959.

from the relationship, the stability constant for the calcium complex of the ideal or " theoretical " ligand of pK 8·25 would be log K_{CaX} = 4·3, definitely lower than the value observed for *N*-(2-pyridylmethyl)iminodiacetic acid (II). Fig. 6 reproduces this straight-line plot for calcium together with similar lines for complexes of magnesium, strontium, barium, and zinc. From these plots were obtained the interpolated values for the " theoretical " stabilities summarised in Table 4. The additional stabilisation is greatest for zinc, and with the alkaline-earth metals it appears to decrease with increasing ionic radius.

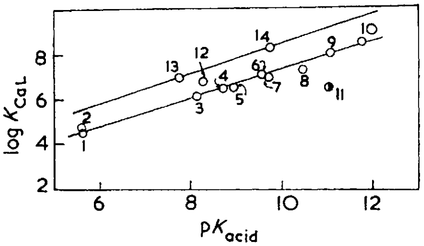


FIG. 5.

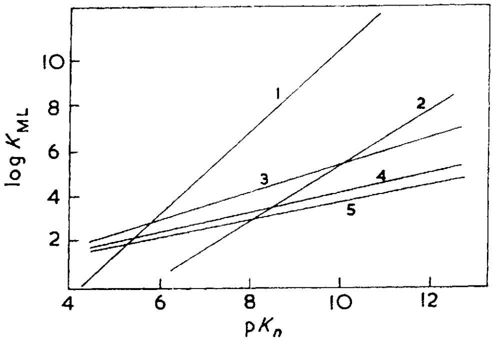


FIG. 6.

FIG. 5. The linear relationship between the stabilities of calcium and proton complexes for a number of tetradentate ligands. (1) Di(carboxymethyl)aminoethylammonium cation; (2) *o*-methoxyphenyliminodiacetic acid; (3) taurine-*NN*-diacetic acid; (4) 2-hydroxyethyliminodiacetic acid; (5) 2-methoxyethyliminodiacetic acid; (6) 2-hydroxycyclohexyliminodiacetic acid; (7) 2-carboxyethyliminodiacetic acid; (8) 2-phosphonoethyliminodiacetic acid; (9) *o*-hydroxyphenyliminodiacetic acid; (10) 2-hydroxybenzyliminodiacetic acid; (11) ethylenediamine-*NN*-diacetic acid; (12) *N*-(2-pyridylmethyl)iminodiacetic acid; (13) *o*-carboxyphenyliminodiacetic acid; (14) nitrilotriacetic acid.

FIG. 6. Linear plots of the relationship, log $K_{\mathrm{ML}} = a \times \mathrm{p}K_{\mathrm{HL}} + b$, for complexes of (1) zinc, (2) magnesium, (3) calcium, (4) strontium, and (5) barium.

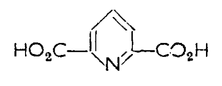
TABLE 4.

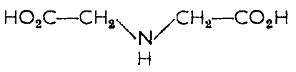
A comparison of the stabilities of complexes of metals with *N*-(2-pyridylmethyl)iminodiacetic acid (II; H_2L) with those of the " ideal " ligand (H_2X) of the same basicity.

Ion	log K_{ML}	log K_{MX}	log $K_{\mathrm{ML}}/K_{\mathrm{MX}}$	r (Å)
Zn^{2+}	10·87	7·3	3·6	0·74
Mg^{2+}	3·90	3·1	0·8	0·65
Ca^{2+}	4·92	4·3	0·6	0·99
Sr^{2+}	3·65	3·2	0·5	1·13
Ba^{2+}	3·40	3·0	0·4	1·35

Anderegg has recently investigated three ligands in which a pyridine nitrogen acts as a donor atom, *viz.*, picolinic acid, 6-methylpicolinic acid, and pyridine-2,6-dicarboxylic acid (VII), and finds them to form complexes of remarkably high stability. In discussing whether or not this could be correlated with the possibility of donor π-bonding from the metal to the aromatic ring of the ligand, Anderegg compared the stabilities of complexes of the acid (VII) with those of the aliphatic analogue iminodiacetic acid (VIII) (Im) and plotted the difference in free energy as measured by ΔK (= log $K_{\mathrm{M(VII)}}$ − log $K_{\mathrm{M(VIII)}}$) against the ionic radius (circles in Fig. 7). The stabilisation is seen to be greatest for the

(VII)





$HO_2C-CH_2-NH-CH_2-CO_2H$

(VIII)

alkaline-earth metals, not the transition metals where *d*–π bonding might have been expected. There is also an increasing preference for the aromatic ligand as the ionic radius is increased: this is attributed to the greater separation between the nitrogen atom and the carboxylic-oxygen atom in the pyridine derivative.[8] Anderegg's plot in Fig. 5 of ref. 8 is slightly misleading owing to confusion between the ionic radii of cobalt, zinc, and copper; the error does not affect the argument and has been corrected in Fig. 7. On the same graph (Fig. 7, crosses) we have plotted the differences in stability between metal complexes of the analogous ligands (II; $R = CH_2{\cdot}CO_2H$) and (III) as a function of the ionic radii of the metals. The similarity in the trend of the differences is remarkable and although the stabilisation is now smaller the greatest effects are still shown by the alkaline-earth metals. We must, however, point out that this method of comparison is not really sound because it fails to take into account differences in the basicity of the ligands and the different effects that these will have on different metal complexes. To illustrate this point consider Fig. 6. If complexes of a (hypothetical) ligand of pK 7 were compared with those of a ligand of pK 9, the stability order would change from Zn > Ca > Sr > Ba > Mg to Zn > Ca > Mg > Sr > Ba. Moreover, since the slopes of the plots increase in the order Ba < Sr < Ca < Mg < Zn any difference between, *e.g.*, complexes of magnesium and strontium will change, not only in sign, but also in magnitude. From Table 4 it would appear that the " stabilisation " is certainly a function of radius but, as pointed out before, the preference for the aromatic ligand *decreases* as the ionic radius (*r*) increases. Whatever the explanation may be there seems no doubt that the alkaline earths co-ordinate more strongly to pyridine-nitrogen than would be expected from their behaviour towards aliphatic nitrogen. It is worth recalling that the alkaline earths form quite stable complexes with the bidentate ligand 8-hydroxyquinoline and its derivatives.

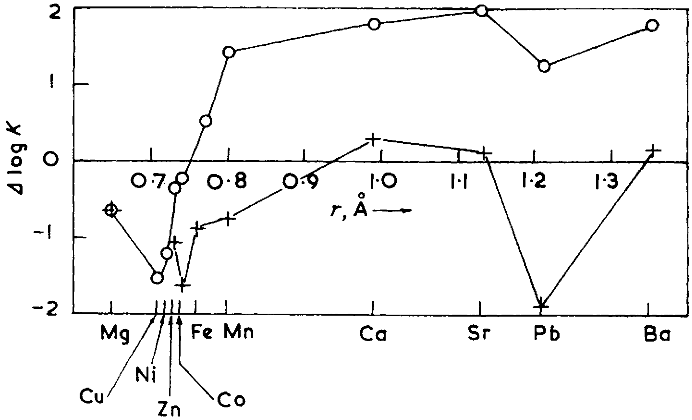


FIG. 7. Stabilisation, ΔK, of 1 : 1 complexes of ligands containing a pyridine-nitrogen atom in comparison with analogues with aliphatic nitrogen atom. Circles, log $K_{M(VII)}$ − log $K_{M(VIII)}$, after Anderegg (ref. 8). Crosses, log K(Metal–N-2-pyridylmethyliminodiacetic acid) − log K(Metal–methylenediamine-NN-diacetic acid). Since Co^{2+} and Zn^{2+} have identical radii the point for Zn^{2+} is displaced 0·01 Å to the left.

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[8] Anderegg, *Helv. Chim. Acta*, 1960, **43**, 414.