The Stabilities of Metal Chelate Compounds formed by Some Heterocyclic Acids. Part I. Studies in Aqueous Solution.

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The stability constants have been measured for some chelated compounds with metals formed by picolinic, 6-methylpicolinic, and quinoline-2- and -8-carboxylic acid in aqueous solution. The figures obtained by potentiometric titration cannot be considered as accurate but conform with the results which might be expected for these compounds. All are weak chelating agents. The complexes with *iso*quinoline-1-carboxylic acid are too insoluble for measurements to be made.

A NUMBER of reagents containing a carboxylic acid group together with a heterocyclic nitrogen atom are used in quantitative analysis, the most important being quinoline-2-carboxylic (quinaldinic), quinoline-8-carboxylic, and picolinic acid. Both the first two have been suggested as gravimetric reagents (see, *e.g.*, Flagg, " Organic Reagents in Gravimetric and Volumetric Analysis," Interscience Publ. Corp., 1948), usually for copper; for this metal Duval (*Analyt. Chim. Acta*, 1951, **6**, 47), using the thermobalance, has investigated the two systems, together with that for zinc and quinoline-8-carboxylic acid. In addition, quinaldinic acid (Majumdar and Sen, *J. Indian Chem. Soc.*, 1950, **27**, 245) and more particularly picolinic acid (*idem, Analyt. Chim. Acta*, 1953, **8**, 369) have been suggested as colorimetric reagents for iron. Stock (*Sborník Mezinárod. polarog. sjezdu u Praze*, 1st Congr., 1951, Part I, p. 371) has studied the behaviour of quinoline-2- and -8-carboxylic acid on polarographic reduction, comparing them with 8-hydroxyquinoline; he has investigated the possibility of amperometric titration in the cases of copper and zinc.

It is assumed that all these reagents act by forming chelated rings with metal ions, the ring being six-membered in the case of quinoline-8-carboxylic acid and five-membered in the other two cases. With these last two the ring is the same as that formed by 8-hydroxyquinoline; as, however, the dissociation constant of the carboxylic acid group is less by several powers of 10 than that of the hydroxyl group it might be expected that correspondingly less stable complexes would be formed. Many measurements of stability constants have been made on the compounds formed by 8-hydroxyquinoline and metal ions (see, *e.g.*, Martell and Calvin, " Chemistry of the Metal Chelate Compounds," Prentice Hall, 1952), but almost none on compounds such as those listed previously and containing a carboxylic acid group. In view of this it was decided to study compounds of this type much more systematically.

The most important measurements of stability are those of Wenger, Monnier, and Epars (*Helv. Chim. Acta*, 1952, **35**, 396) in which the overall stability constant of the unstable red species formed between ferrous iron and quinaldinic acid was evaluated spectrophotometrically as $2{\cdot}6 \times 10^{-8}$, and of Majumdar and Sen (*Analyt. Chim. Acta*, *loc. cit.*) in which the red, soluble complex formed by ferrous iron and picolinic acid was similarly found to have an overall stability constant varying between $2{\cdot}2 \times 10^{-5}$ and $6{\cdot}4 \times 10^{-5}$ over the temperature range studied.

The compounds studied in our first investigations were picolinic, 6-methylpicolinic, quinaldinic, *iso*quinoline-1-carboxylic, and quinoline-8-carboxylic acid. All were sufficiently soluble in water for dissociation constants to be obtained therein, and all formed compounds with a wide range of metals. Some of these derivatives, as is well known, are insoluble in water; the stabilities of those which proved sufficiently soluble were determined by potentiometric titration and form the basis of this paper. A rather shorter series of metals was studied in 50% (v/v) aqueous dioxan and these results will be published separately.

EXPERIMENTAL

Materials.—Picolinic acid was prepared by oxidation of α-picoline by alkaline permanganate (*Org. Synth.*, 1940, **20**, 79), 6-methylpicolinic acid by Black, Depp, and Corson's method (*J. Org. Chem.*, 1949, **14**, 147), and quinoline-8-carboxylic acid according to Campbell, Kerwin, La Forge, and Campbell (*J. Amer. Chem. Soc.*, 1946, **68**, 1844). *iso*Quinoline-1-carboxylic acid was prepared by Padbury and Lindwall's method (*ibid.*, 1945, **67**, 1269), but recrystallised from hot water (Solomon, *J.*, 1947, 129); it had m. p. 168—169°, in agreement with Solomon (168—171°) but not with Padbury and Lindwall (161°). The 8-hydroxyquinoline was of B.D.H. " AnalaR " grade, and the quinaldinic acid was obtained from Messrs. Light. All these reagents were recrystallised from suitable solvents, either to constant m. p. or until only one spot was obtained on a chromatogram. Different batches of material were checked as giving identical dissociation constants.

Metals were used in the form of their perchlorates (G. Frederick Smith Co.).

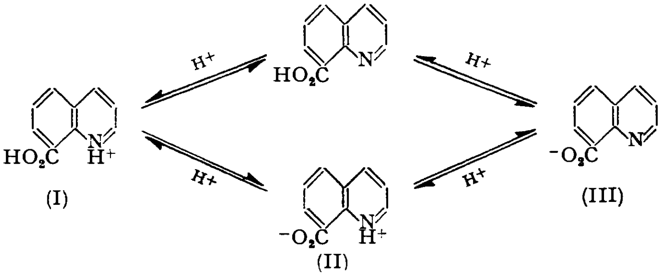
Potentiometric Titrations.—Titrations were carried out in a glass vessel of about 100 ml. capacity, provided with a Perspex cover, which was drilled with holes to admit glass and calomel electrodes, the elongated tip of a 10-ml. burette, and a jet which projected under the surface of the solution in the vessel and through which was admitted nitrogen, freed from oxygen and carbon dioxide by Fieser's solution; the nitrogen served both to stir the solution and to preserve an inert atmosphere. The vessel was kept in a thermostat at 25° ± 0·1°.

For the determination of stoicheiometric dissociation constants a measured volume of an aqueous solution of the acid was placed in the vessel, and 5 ml. of approx. 0·1N-perchloric acid were added, followed by 50 ml. of air-free water. The final solution was about 10^{-3}M with respect to the organic acid. For the determination of stability constants the water was replaced by a solution of metal perchlorate, such that the final molar concentration of the bivalent metal was half that of the organic acid.

In either case the pH of the solution was found by means of a Marconi pH meter. Carbonate-free approximately 0·1N-sodium hydroxide was then added in small portions, and the pH taken after each addition and the attainment of equilibrium. Before and after each titration the pH meter was calibrated with standard buffer solutions. The ionic strength of the solutions during titration was of the order of 0·02 and varied only slightly.

Spectrophotometric Measurements.—Absorption spectra were measured with a Unicam S.P. 500 spectrophotometer with minimum slit widths and matched 1-cm. silica cells. Hydrochloric acid was used to obtain very low pH's, acetate buffers for intermediate ones, and sodium hydroxide for higher values. The ionic strength was made constant at 0·5M with sodium chloride. The temperature was 18° ± 2°.

Results.—The acids under consideration are all dibasic, the equilibria which depend on pH being, *e.g.*:



For the present work it is unimportant whether form (II) is the neutral molecule, the zwitterion, or a mixture of both, but the matter will be considered in a subsequent paper. Here it will simply be assumed that the equilibria are governed by two dissociation constants. Preliminary calculations show that these in all cases are sufficiently far apart for no overlapping to occur. Hence a separate evaluation of each may be made.

Values obtained are given in Table 1. All are stoicheiometric, strictly valid only for the

TABLE 1.

Acid	First dissn. const.	Second dissn. const.	Ref.	Acid	First dissn. const.	Second dissn. const.	Ref.
Quinoline-8-carboxylic ...	2·0	6·71		Picolinic	1·5	5·43	
Quinoline-2-carboxylic ...	1·9	4·92			1·60	5·40	3
		5·0	1		1·08	5·32	4
		4·92	2			5·49	2
*iso*Quinoline-1-carboxylic	not detd.	4·97		6-Methylpicolinic	0·9	5·77	
				8-Hydroxyquinoline (for comparison)	5·00	9·70	

(1) Wenger, Monnier, and Epars (*loc. cit.*); stoicheiometric constant by pH titration at 20° and ionic strength M/30. (2) Holmes and Reed, unpublished work: stoicheiometric constants obtained spectrophotometrically at 18° ± 2° and ionic strength 0·5M. (3) Jellinek and Urwin (*J. Phys. Chem.*, 1954, **58**, 548); spectrophotometrically at 20° ± 2° and ionic strength 0·01. (4) Evans, Herington, and Kynaston (*Trans. Faraday Soc.*, 1953, **49**, 1284); apparent constants determined spectrophotometrically at 25° ± 0·1° and ionic strength 0·030M.

Stability Data for Chelated Complexes.—Concn. of NaOH used, 0·1034N; concn. of $HClO_4$ used, 0·0933N (except where otherwise stated).

I. *Quinoline-*8*-carboxylic acid.*

A. Copper: 0·0671N-NaOH; 0·1049N-$HClO_4$. 0·0235 g. of acid. Initial vol. 65·0 ml.
NaOH, 7·360 ml.; pH 3·20; $\bar{n}$ 0·3; log 1/[L⁻] 5·3.

B. Nickel: 0·0229 g. of acid. Initial vol. 69·2 ml.

NaOH, ml.	4·583	4·720	4·861	5·024	5·164	5·310	5·451	5·548	5·660
pH	4·53	4·95	5·22	5·51	5·80	6·08	6·38	6·63	7·09
$\bar{n}$	0·13	0·32	0·56	0·74	0·89	1·07	1·23	1·31	1·35
log 1/[L⁻]	4·93	4·60	4·40	4·18	3·97	3·81	3·66	3·54	3·27

C. Cobalt: 0·0229 g. of acid. Initial vol. 68·5 ml.

NaOH, ml.	4·605	4·735	4·869	5·031	5·154	5·313	5·470	5·633
pH	5·26	5·62	5·89	6·16	6·35	6·59	6·86	7·27
$\bar{n}$	0·07	0·22	0·36	0·48	0·57	0·68	0·82	0·92
log 1/[L⁻]	4·23	3·92	3·79	3·53	3·42	3·29	3·20	3·10

D. Zinc: 0·0229 g. of acid. Initial vol. 66·8 ml.

NaOH, ml.	5·010	5·240	5·519
pH	6·36	6·70	7·16
$\bar{n}$	0·23	0·30	0·40
log 1/[L⁻]	3·30	3·11	2·95

E. Lead: 0·0229 g. of acid. Initial vol. 66·6 ml.

NaOH, ml.	4·820	4·979	5·241
pH	6·20	6·40	6·73
$\bar{n}$	0·01	0·11	0·24
log 1/[L⁻]	3·43	3·24	3·07

TABLE 1. (*Continued.*)

F. Manganese : 0·0229 g. of acid. Initial vol. 66·9 ml.
NaOH, 5·341 ml.; pH 6·84; $\bar{n}$ 0·22; log 1/[L$^-$] 3·03.

G. Mercury : 0·0671N-NaOH; 0·1049N-HClO$_4$. 0·0294 g. of acid. Initial vol. 65·0 ml.
NaOH, 10·107 ml.; pH 7·62; $\bar{n}$ 0·08; log 1/[L$^-$] 2·71.

II. *Quinaldinic acid.*

A. Copper : A precipitate appears at too low a pH for values to be obtained.

B. Nickel : 0·0199 g. of acid. Initial vol. 68·1 ml.

NaOH, ml.	4·140	4·520	4·666	4·861	5·050	5·381
pH	3·15	3·36	3·46	3·65	3·88	4·59
$\bar{n}$	0·23	0·58	0·71	0·91	1·05	1·38
log 1/[L$^-$]	4·65	4·53	4·48	4·38	4·21	3·80

C. Cobalt : 0·0199 G. of acid. Initial vol. 68·5 ml.

NaOH, ml.	4·080	4·530	4·750	4·934	5·110	5·295	5·412
pH	3·19	3·52	3·72	3·99	4·27	4·70	5·23
$\bar{n}$	0·10	0·41	0·61	0·79	0·95	1·11	1·30
log 1/[L$^-$]...........	4·57	4·33	4·19	3·99	3·82	3·58	3·17

D. Zinc : A precipitate forms at too low a pH.

E. Lead : 0·0671N-NaOH; 0·1049N-HClO$_4$. 0·0347 g. of acid. Initial vol. 65·0 ml.
NaOH, 8·039 ml.; pH 3·86; $\bar{n}$ 0·08; log 1/[L$^-$] 3·68.

F. Manganese : 0·1122N-NaOH; 0·0149N-HClO$_4$. 0·0333 g. of acid. Initial vol. 65·0 ml.

NaOH, ml.	4·825	5·328	5·720
pH	3·82	4·43	4·81
$\bar{n}$	0·12	0·34	0·60
log 1/[L$^-$]...........	3·72	3·25	3·08

G. Mercury : A precipitate is formed at too low a pH.

III. iso*Quinoline*-1-*carboxylic acid.*

All the metals considered are precipitated at too low a pH for measurements to be made in aqueous solution.

IV. *Picolinic acid.*

A. Copper : 0·1122N-NaOH; 0·1049N-HClO$_4$. 0·01690 g. of acid. Initial vol. 65·0 ml.

NaOH, ml.	1·950	2·100	3·160	4·200	4·560	4·923	5·202	5·560	5·688
pH	2·28	2·30	2·44	2·66	2·77	2·92	3·09	3·50	3·84
$\bar{n}$	0·74	0·83	1·20	1·45	1·55	1·65	1·69	1·76	1·81
log 1/[L$^-$]	6·13	6·13	6·13	6·07	6·03	5·98	5·86	5·57	5·29

B. Nickel : 0·0174 g. of acid. Initial vol. 69·4 ml.

NaOH, ml.	2·331	3·370	4·270	4·526	4·742	4·910	5·082	5·535
pH	2·45	2·60	2·87	2·96	3·05	3·15	3·25	3·73
$\bar{n}$	0·58	0·98	1·09	1·19	1·28	1·34	1·43	1·66
log 1/[L$^-$]	5·89	5·85	5·64	5·59	5·55	5·48	5·45	5·21

C. Cobalt : 0·0174 g. of acid. Initial vol. 68·9 ml.

NaOH, ml.	3·832	4·249	4·565	4·935	5·235	5·515
pH	2·77	2·90	3·05	3·27	3·56	4·13
$\bar{n}$	0·77	0·91	0·99	1·17	1·35	1·55
log 1/[L$^-$]	5·60	5·50	5·39	5·26	5·08	4·68

V. 6-*Methylpicolinic acid.*

A. Copper : 0·0205 g. of acid. Initial vol. 69·6 ml.

NaOH, ml. ...	2·258	3·396	3·804	4·106	4·640	4·971	5·282	5·392	5·571
pH	2·42	2·61	2·71	2·80	3·01	3·24	3·59	3·80	4·31
$\bar{n}$	0·63	0·97	1·06	1·12	1·29	1·39	1·49	1·53	1·70
log 1/[L$^-$]	6·26	6·20	6·14	6·07	5·96	5·93	5·53	5·37	5·06

B. Nickel : 0·0205 g. of acid. Initial vol. 69·4 ml.

NaOH, ml. ...	4·240	4·491	4·645	4·825	4·998	5·150	5·275	5·350	5·510	5·660
pH	3·23	3·42	3·57	3·79	4·03	4·27	4·40	4·65	5·08	5·77
$\bar{n}$	0·23	0·39	0·51	0·67	0·87	1·02	1·16	1·27	1·50	1·14
log 1/[L$^-$]	5·34	5·19	5·07	4·90	4·75	4·59	4·54	4·31	4·12	3·42

C. Cobalt : 0·0205 g. of acid. Initial vol. 68·2 ml.

NaOH, ml. ...	4·231	4·370	4·591	4·682	4·770	4·860	4·950	5·051	5·147	5·492	5·572
pH	3·35	3·51	3·82	3·98	4·13	4·28	4·43	4·58	4·75	5·52	5·85
$\bar{n}$	0·06	0·11	0·26	0·32	0·46	0·56	0·67	0·79	0·92	1·30	1·11
log 1/[L$^-$]	5·17	5·03	4·75	4·61	4·51	4·39	4·28	4·18	4·07	3·64	3·33

conditions under which they have been obtained. For all the acids considered the first dissociation constant has a pK of about 2, and at this value small errors in pH measurement can cause appreciable error in the value obtained. In the determination of stability constants, values of this order are not important to the calculation where chelation is at pH's of 4 or more. But since the compounds under consideration are chelated with some metals at pH's much less than this, it was necessary to use an accurate value, and this was obtained by spectrophotometric means. With quinoline-8-carboxylic acid, however, chelation occurred at higher values of pH, so the value recorded is that obtained by potentiometric titration.

When chelation occurs at very low pH, values of $\bar{n}$ (see *J.*, 1954, 2904) obtained may be appreciably in error, because of the importance of the hydrogen-ion concentration term in the expression for $\bar{n}$ at such pH's. Consequently $\bar{n}$ values obtained at pH <*ca.* 2·5 have not generally been quoted. Similarly, figures at high values of $\bar{n}$ are often not given, either because precipitation occurs or because the graph indicates that no further chelation takes place.

No further metals were used with either picolinic acid or 6-methylpicolinic acid.

DISCUSSION

It can be seen from the tables of $\bar{n}$ values that the complexes are formed at low pH's and, as already pointed out above, this may tend to a lack of symmetry in the lower part of the formation curve. For this reason, or because precipitation makes it possible to obtain only a short portion of the formation curve, formation constants cannot be evaluated in all cases and those quoted are not all of the same accuracy.

In Bjerrum's original work it is shown that the values of log $1/[\mathrm{L}^-]$ at $\bar{n} = \frac{1}{2}$ and $\bar{n} = 1\frac{1}{2}$ may be taken as the first and second formation constants if the spreading factor is $\geqslant 10$. If this condition is not obeyed, as is the case with most of the complexes under consideration, then the values of K_1 and K_2 may be calculated from

$$\frac{\bar{n}}{(\bar{n}-1)[\mathrm{L}]} = \frac{(2-\bar{n})[\mathrm{L}]}{(\bar{n}-1)} K_1K_2 - K_1$$

by the method of least squares, as has been stated by Irving and Rossotti (*J.*, 1953, 3397). To obtain both K_1 and K_2 by this method, however, a reasonable portion of the formation curve must be obtained and this has not been possible in every case. Because of this, and because it was thought that the figures are not of sufficient accuracy to merit the more extensive treatment, values of log K_1, log K_2, and log $K_{\mathrm{av.}}$ are simply values of pL for which $\bar{n} = 0·5$, 1·5, and 1·0, respectively. In some instances it has not been possible to obtain the upper portion of the formation curve and in such cases the value of log K_2 has not been estimated from those of log $K_{\mathrm{av.}}$ and log K_1 since with most of the compounds under present consideration it is the lower portion of the curve which is most susceptible to error. The figures obtained are given in Table 2. Nevertheless, information of some relevance is given both by the figures and by the shape of the curves, the drop in pL with higher values of $\bar{n}$ in the case of 6-methylpicolinic acid being noteworthy as evidence of the instability of the bis-complex.

TABLE 2.

	Copper			Nickel			Cobalt		
Acid	log K_1	log K_2	log $K_{\mathrm{av.}}$	log K_1	log K_2	log $K_{\mathrm{av.}}$	log K_1	log K_2	log $K_{\mathrm{av.}}$
Picolinic	—	6·0	6·1	5·9	5·4	5·7	6·0	4·8	5·4
6-Methylpicolinic	6·3	5·6	6·1	5·1	4·1	4·6	4·5	3·3	3·9
Quinoline-2-carboxylic	—	—	—	4·6	—	4·3	4·3	—	3·8
Quinoline-8-carboxylic	—	—	—	4·4	—	3·9	3·5	—	—

The first dissociation exponents of the parent acids are in the order quinoline-8-carboxylic = quinoline-2-carboxylic > picolinic > 6-methylpicolinic acid; and the second dissociation exponents in the order quinoline-8-carboxylic > 6-methylpicolinic > picolinic > quinoline-2-carboxylic acid. From these, it would be expected that quinoline-8-carboxylic acid would form the most stable complexes of the series since it has the highest pK in both cases. That it does not do so could be explained in terms either of steric hindrance or of the existence of the 6-membered chelate ring (known to be in general less stable than a 5-membered ring), or of both. Very probably both effects are present but the limited range of data does not allow a distinction to be made.

Although values for log K_1 are not available in every case, the trend of log K_1 and log $K_{\text{av.}}$ is the same for both nickel and cobalt, and is : picolinic > 6-methylpicolinic > quinoline-2-carboxylic > quinoline-8-carboxylic acid. This does not follow the order of either dissociation constant. Irving and Rossotti (*J.*, 1954, 2910), when considering complexes formed by 8-hydroxyquinoline and similar compounds, showed that the second pK for the acid and log K_1 for the complex should be related, but that steric hindrance or the effect of the first acid dissociation constant might upset this relationship. With our compounds it is obvious that a simple relation between the second dissociation constant and log K_1 or log $K_{\text{av.}}$ does not exist. The effect of the first pK would be to make complexes of quinoline-2-carboxylic acid more stable with respect to picolinic acid, and those of 6-methylpicolinic acid less stable. In the latter case, however, the complexes are formed at a higher pH than corresponding ones derived from picolinic acid, and the effect of variation in the first dissociation constant in calculations of $\bar{n}$ is negligible at $\bar{n}$ values other than the lowest, a change in pK of as much as 0·6 producing little change with $\bar{n}$ values of 0·5 or more. With picolinic acid, higher $\bar{n}$ values must be attained before such changes produce negligible results, but at values of $\bar{n} = 1$ the two systems may fairly be compared for nickel and cobalt. When this is done the lower stabilities of the 6-methylpicolinic acid complexes leave no doubt that steric hindrance is operative.

In the case of quinoline-2-carboxylic acid the presence of the benzene ring in the 5 : 6-position must lead to the same effect.

All the data are in agreement with the Irving and Williams's order of stabilities (*J.*, 1953, 3192). In particular, the order is maintained with quinoline-2-carboxylic acid although the pH's of incipient precipitation of the complexes of this acid are anomalous, the nickel complex in particular being precipitated out of the expected order.

Fuller curves obtained from titration in a dioxan–water medium, to be reported later, lend support to the deductions made here.

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