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Studies on the Sulfur-containing Chelating Agents. XXXIX.[1] Determination of Stability Constants of Some Metal Complexes of N-Phenyl- and N-Ethyl-β-mercaptocinnamamide and Thiodibenzoylmethane by Solvent Extraction Method

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Stability constants and extraction constants of some metal complexes of N-phenyl-β-mercaptocinnamamide (I), N-ethyl-β-mercaptocinnamamide (II) and thiodibenzoylmethane (III) were determined by solvent extraction method in chloroform and isoamyl acetate systems. The stability constants in I-metal and II-metal systems are smaller than those in III-metal system whereas the extraction constants in the former two systems are larger than those in the latter. In these ligands, the difference in the partition coefficients is regarded as a main factor to cause the remarkable difference in the extractability of metal ion. It was shown that the introduction of amide group makes the reagent hydrophilic to react readily with the metal ions in the extraction system.

In the preceding paper,[3] the determination of the acid dissociation constants by the solvent extraction method was reported on N-phenyl-β-mercaptocinnamamide (I), N-ethyl-β-mercaptocinnamamide (II) and thiodibenzoylmethane (III) which have been applied to the determination of various metal ions.[4] The presence of amide group did not show any effect on the selectivity of the ligand towards the metal ions.[5] The influence of the amide group on the acid dissociation constant was observed scarcely in I and II in comparison with III. On the contrary, the partition coefficient was remarkably lowered by the presence of the amide group.[3] This result may offer an expectation that the presence of the amide group will give remarkable effect on the extractability of the complexes into the organic solvent. This paper deals with the determination of the extraction constants and the stability constants of some metal complexes of I, II and III.

Experimental

Materials——N-Phenyl-β-mercaptocinnamamide (I), N-ethyl-β-mercaptocinnamamide (II) and thiodibenzoylmethane (III) were synthesized and purified by the method reported previously.[6] The solvent, $CHCl_3$ and isomyl acetate were shaken with water and distilled over Linde 4A molecular sieves for the immediate use. Deionized and redistilled water was used. All other chemicals were of analytical grade. The pH of the aqueous phase was adjusted to a desired value with 0.005—0.1 M solution of the following reagents: $HClO_4$ (pH 1.2—3.5), HOAc–NaOAc (pH 3.5—7.0) and KH_2PO_4–$Na_2B_4O_7$ (pH 7.0—9.0). The ionic strength was adjusted to 0.1 by $NaClO_4$. The radioisotopes, ^{65}Zn, ^{63}Ni, ^{203}Hg, ^{59}Fe and ^{60}Co were obtained from New England Nuclear Corp., U.S.A. as HCl solution of $ZnCl_2$, from the Radiochemical Center, Amersham, England,

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[1] Part XXXVIII: S. Kawanishi, A. Yokoyama and H. Tanaka, *Chem. Pharm. Bull.* (Tokyo), 21, 2653 (1973).

[2] Location: *Yoshida, Shimoadachi-cho, Sakyo-ku, Kyoto.*

[3] A. Yokoyama, M. Chikuma and H. Tanaka, *Chem. Pharm. Bull.* (Tokyo), **20**, 2000 (1972).

[4] *a*) A. Yokoyama, M. Chikuma, H. Hayashi and H. Tanaka, *Bunseki Kagaku*, **18**, 24 (1969); *b*) H. Tanaka, N. Nakanishi, Y. Sugiura and A. Yokoyama, *ibid.*, **17**, 1428 (1969); *c*) E. Uhlemann and B. Schuknecht, *Anal. Chim. Acta*, **63**, 236 (1973); *d*) E. Uhlemann and H. Müller, *ibid.*, **41**, 311 (1968); *e*) *Idem*, *ibid.*, **48**, 115 (1969).

[5] A. Yokoyama, S. Kawanishi and H. Tanaka, *Chem. Pharm. Bull.* (Tokyo), **18**, 356 (1970).

[6] A. Yokoyama, S. Kawanishi, M. Chikuma and H. Tanaka, *Chem. Pharm. Bull.* (Tokyo), **15**, 540 (1967); A. Yokoyama, K. Ashida and H. Tanaka, *ibid.*, **12**, 690 (1964).

as HCl solution of $NiCl_2$ and as crystal of $Hg(OAc)_2$, from Union Carbide Corp., U.S.A., as HCl solution of $FeCl_3$ and from Oak Ridge Laboratory, U.S.A., as HCl solution of $CoCl_2$, respectively. The stock solution of 1×10^{-6} M of each metal ion was prepared by the dilution of the original radioisotope solution with 0.1 M $NaClO_4$ solution which contains 0.01 M $HClO_4$ and by the addition of the carrier. The solution of the ligand was prepared freshly for each experiment. Scintillator solution was prepared by mixing 1000 ml of toluene, 500 ml of Triton X-100, 5 g of 2,5-diphenyloxazole (Dojindo Co., Ltd.) and 300 mg of 1,4-bis[2-(4-methyl-5-phenyloxazoly)] benzene (Beckman Instruments Inc.).

Measurement of Radioactivity——Measurements of γ-radioactivity of ^{59}Fe, ^{60}Co, ^{65}Zn and ^{203}Hg were made with a Fujitsu well-type scintillation counter Model ATS-121 (NaI) on 2 ml aliquots taken from the aqueous and the organic phases. Radioactivity of ^{63}Ni was measured with a Beckman liquid scintillation system LS-233. It is necessary to measure the activities in the organic and the aqueous phases with the same efficiency for the determination of the distribution ratio between two phases. The two channel ratio method was chosen for the purpose and the following treatments were confirmed to give the satisfactory results. In the case of the aqueous phase, 0.5 ml aliquot was taken and mixed with the scintillator solution in a vial. In the case of the organic phase, the concentration of the ligand was made less than 2×10^{-3} M by the dilution, and 0.5 ml aliquot was taken into a vial and the scintillator solution was added after the solvent was distilled off under reduced pressure.

Procedure——Five ml of aqueous sample solution, which contains 1×10^{-7} M metal ion was adjusted to a desired pH, and was shaken with 5 ml of chloroform or isoamyl acetate solution of the ligand at 22±0.5° until the extraction equilibrium was reached. A solution of KCN as a masking agent was added to the aqueous phase in the extraction system containing Hg(II) to reduce the apparent extractability of Hg(II). After two phases were separated, the activities of both phases were measured. The remaining aqueous phase was applied to pH measurement.

The distribution ratio of metal was defined as follows.

$$D = \frac{\text{total radioactivity of metal in organic phase}}{\text{total radioactivity of metal in aqueous phase}}$$

Result

When the metal in the organic phase is present as MR_n and the metal in the aqueous phase as ion (M^{n+}) and the concentrations of other species are negligibly small, the distribution ratio is expressed as equations (1) and (2).

$$D = \frac{\text{total concentration of metal in organic phase}}{\text{total concentration of metal in aqueous phase}} \qquad (1)$$

$$= \frac{[MR_n]_0}{[M^{n+}]} \qquad (2)$$

In these equations, the subscript "0" refers to the organic phase and R^- represents the anion of the ligand (HR). The equation (2) can be expressed as equation (3) by the introduction of the extraction constant (K_{ex}) which is defined in equation (4).

$$\log D = \log K_{ex} + n \log [HR]_0 + n\text{pH} \qquad (3)$$

$$K_{ex} = \frac{[MR_n]_0[H^+]^n}{[M^{n+}][HR]_0{}^n} \qquad (4)$$

$$(M^{n+}) + n(HR)_0 \overset{K_{ex}}{\rightleftarrows} (MR_n)_0 + n(H^+)$$

The extraction constant can be also expressed as equation (5),

$$K_{ex} = P_c \cdot \beta_n \cdot K_a{}^n \cdot P_r{}^{-n} \qquad (5)$$

P_c, β_n, K_a and P_r are expressed as follows.

$$P_c = \frac{[MR_n]_0}{[MR_n]} \qquad \beta_n = \frac{[MR_n]}{[M^{n+}][R^-]^n}$$

$$K_a = \frac{[H^+][R^-]}{[HR]} \qquad P_r = \frac{[HR]_0}{[HR]}$$

The values of P_r/K_a are shown in Table I.

Table I. Values of log (P_r/K_a)

Solvent	I	II	III
Chloroform[a]	9.48	8.98	12.52
Isoamyl acetate	9.86	8.85	11.93
Benzene[a]	9.15	8.65	12.31

a) lit.[3]

(1) Nickel [II] and Zinc [II] Complex

Fig. 1, 2 and 3 show the influence of the ligand concentration on the distribution of the metal. Under the experimental condition shown in Fig. 1, 2 and 3, that the large excess of ligand is present, equation (3) can be applied to the distribution of the metal. Thus, log D is a linear function of log $[HR]_0$ and the slope represents n. The value of n was obtained to be 2 in the nickel [II] and zinc [II] systems.

When the successive formation of the complex cannot be neglected in the extraction system of nickel [II] and zinc [II], the distribution ratio, D can be expressed as equation (6).

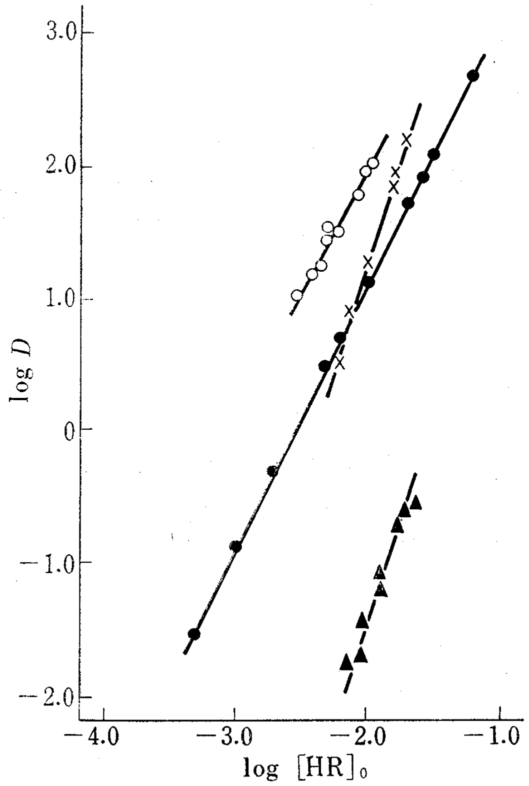

Fig. 1. Distribution of Metal Complexes of N-Phenyl-β-mercaptocinnamamide as a Function of Concentration of Ligand in $CHCl_3$–Water System[a]

ionic strength: 0.1 ($NaClO_4$),
○: Ni (pH 5.30) ●: Zn (pH 4.00)
×: Co (pH 4.64) ▲: Fe (pH 1.22)
a) data in benzene-water system in the case of Fe [III]

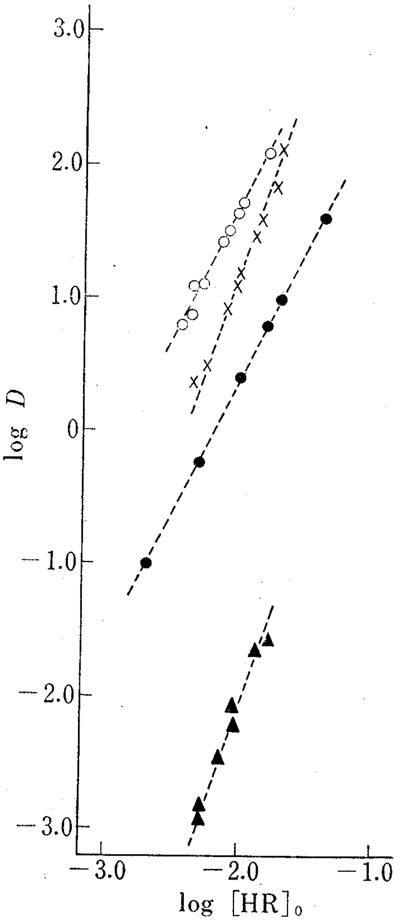

Fig. 2. Distribution of Metal Complexes of N-Ethyl-β-mercaptocinnamamide as a Function of Concentration of Ligand in $CHCl_3$–Water System[a]

ionic strength: 0.1 ($NaClO_4$),
○: Ni (pH 5.52) ●: Zn (pH 3.47)
×: Co (pH 4.73) ▲: Fe (pH 1.22)
a) data in benzene-water system in the case of Fe [III]

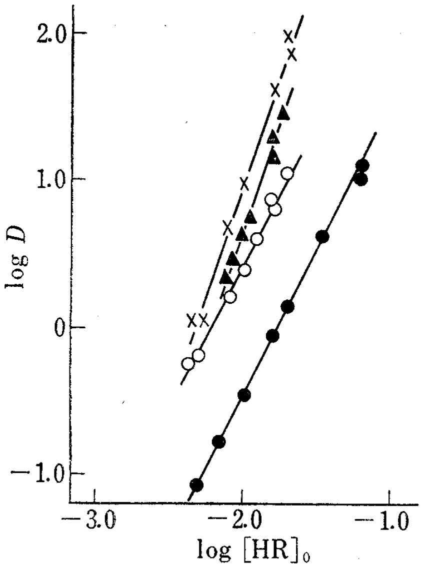

Fig. 3. Distribution of Metal Complexes of Thiodibenzoylmethane as a Function of Concentration of Ligand in $CHCl_3$–Water System

ionic strength: 0.1 ($NaClO_4$)
○: Ni (pH 5.56) ●: Zn (pH 4.03)
×: Co (pH 4.74) ▲: Fe (pH 5.30)

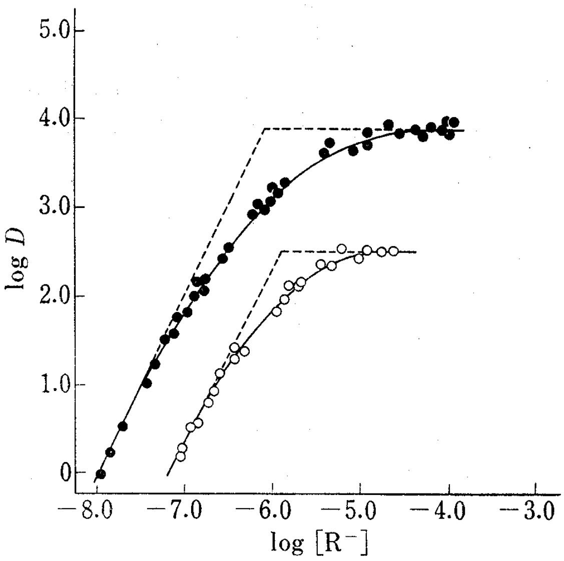

Fig. 4. Distribution of Nickel and Zinc Complexes of N-Phenyl-β-mercaptocinnamamide as a Function of Concentration of Ligand Anion in $CHCl_3$–Water System

○: Ni, ●: Zn, ---: asymptote
The full-drawn curves represent
$\log D=\log 10^{14.33}[R^-]^2-\log(1+10^{6.00}[R^-]+10^{11.84}[R^-]^2)$ for nickel [II] system
and $\log D=\log 10^{16.06}[R^-]^2-\log(1+10^{6.86}[R^-]+10^{12.16}[R^-]^2)$ for zinc [II] system.

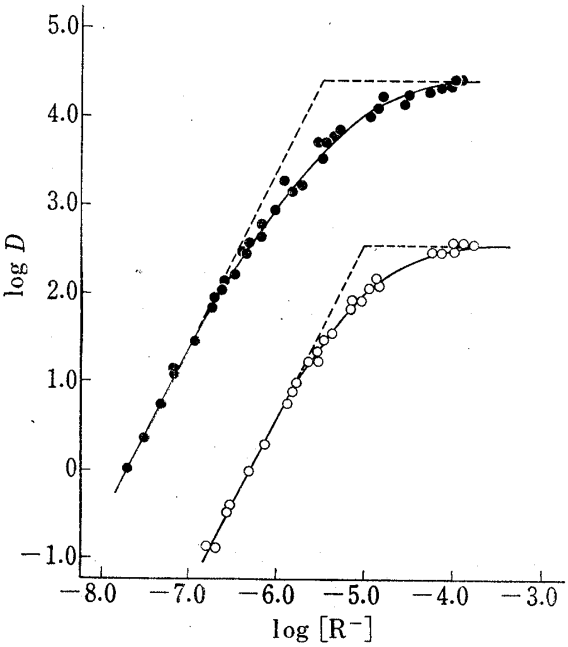

Fig. 5. Distribution of Nickel and Zinc Complexes of N-Ethyl-β-mercaptocinnamamide as a Function of Concentration of Ligand Anion in $CHCl_3$–Water System

○: Ni, ●: Zn, ---: asymptote
The full-drawn curves represent
$\log D=\log 10^{12.62}[R^-]^2-\log(1+10^{5.34}[R^-]+10^{10.07}[R^-]^2)$
for nickel [II] system
and $\log D=\log 10^{15.40}[R^-]^2-\log(1+10^{6.22}[R^-]+10^{10.99}[R^-]^2)$
for zinc [II] system.

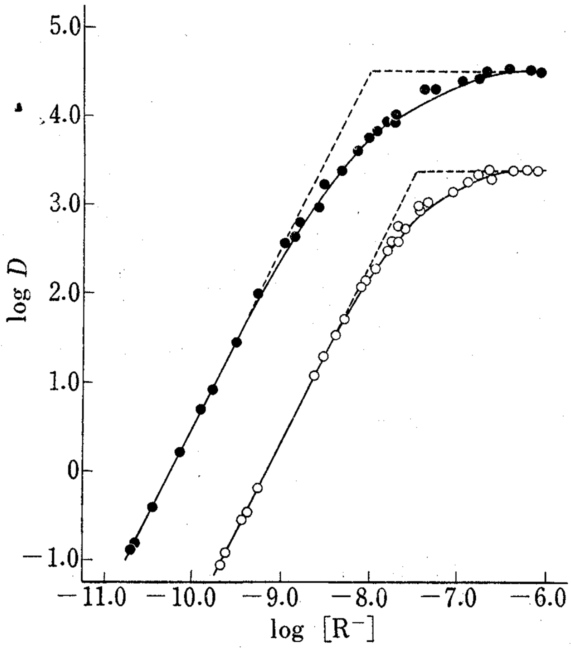

Fig. 6. Distribution of Nickel and Zinc Complexes of Thiodibenzoylmethane as a Function of Concentration of Ligand Anion in $CHCl_3$–Water System

○: Ni, ●: Zn, ---: asymptote
The full-drawn curves represent
$\log D=\log 10^{18.32}[R^-]^2-\log(1+10^{7.51}[R^-]+10^{14.94}[R^-]^2)$
for nickel [II] system and
$\log D=\log 10^{20.50}[R^-]^2-\log(1+10^{8.60}[R^-]+10^{15.99}[R^-]^2)$
for zinc [II] system.

$$D = \frac{P_c\beta_2[R^-]^2}{1+\beta_1[R^-]+\beta_2[R^-]^2} \qquad (6)$$

By the use of curve-fitting method[7] for obtaining the successive formation constants, equation (6) is rewritten as equation (7) by rearranging it, where v and p represent $[R^-]\sqrt{\beta_2}$ and $\beta_1/\sqrt{\beta_2}$ respectively.

$$\log D = \log P_c\cdot v^2 - \log(1+pv+v^2) \qquad (7)$$

The experimental data plotted as log D against log $[R^-]$ can be compared with the normalized curves. For the normalization of equation (7), the curves which are expressed by X and Y indicated as follows are used.

$$X = \log v$$
$$Y = \log(1+pv+v^2)$$

The value of $[R^-]$ was calculated from equation (8).

$$[R^-] = \frac{C_{HR}}{1+(P_r+1)K_a^{-1}[H^+]} \qquad (8)$$

In equation (8), C_{HR} is the total concentration of the ligand, namely $([HR]_0+[HR]+[R^-])$. The plots of log D against log $[R^-]$ are shown in Fig. 4, 5 and 6. The values of β_1 obtained are collected in Table II.

TABLE II. Successive Stability Constants obtained by Curve-fitting Method

Metal	I	II	III
	log β_1	log β_1	log β_1
Ni	6.00	5.34	7.51
Zn	6.86	6.22	8.60

(2) Mercury [II] Complex

The stability constants of mercury [II] complexes of I and II were determined by the method reported previously,[8] which was applied to the determination of the stability constant of mercury [II] complex of III. When potassium cyanide is used as an auxiliary complexing agent for the reduction of the apparent extractability of mercury [II], the extraction is expressed as equation (9). In equation (9), β_x is expressed as follows.[8]

$$\beta_x = \frac{[Hg(CN)_x{}^{(2-x)+}]}{[Hg^{2+}][CN^-]^x}$$

$$K_{ex} = \frac{D[H^+]^2\{1+\sum\beta_x[CN^-]^x\}}{[HR]_0{}^2} \qquad (9)$$

The distribution behavior of mercury [II] in the presence of cyanide ion is shown in Fig. 8. The partition coefficients (P_c) of mercury [II] complexes of I and II were measured in the absence of cyanide ion (Fig. 7).

(3) Cobalt [II] Complex

As shown in Fig. 1, 2, 3 and 7, the equation (3) can be applied apparently under these conditions and the slope (n) is 3. The formations of cobalt [III] complexes by the reactions of cobalt [II] ion with these ligands are suggested from these results.

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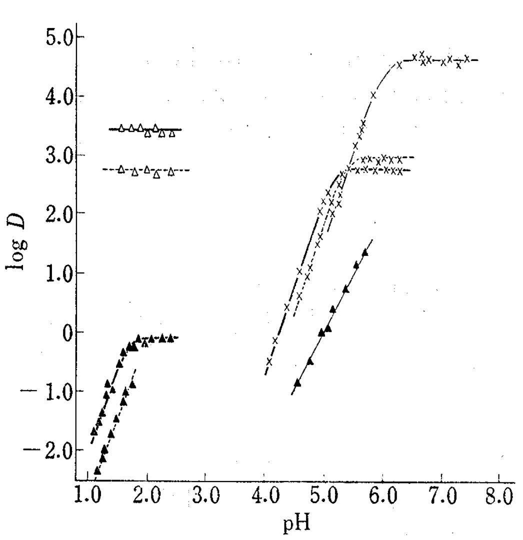

Fig. 7. Distribution of Metal Complexes as a Function of pH in $CHCl_3$–Water System[a]

× : Co ▲ : Fe △ : Hg

——: N-phenyl-β-mercaptocinnamamide

------: N-ethyl-β-mercaptocinnamamide

——: thiodibenzoylmethane

ionic strength: 0.1 ($NaClO_4$)

initial concentration of each ligand: 1×10^{-2}M

a) data in benzene-water system in the cases of Fe-N-phenyl- and N-ethyl-β-mercaptocinnamamide

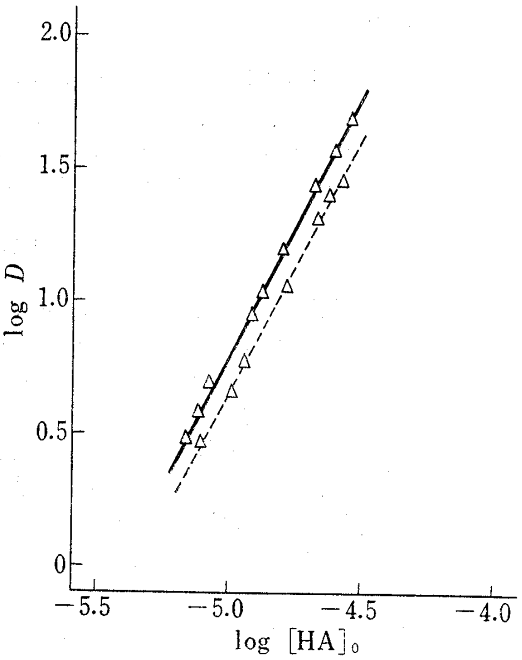

Fig. 8. Distribution of Mercury [II] Complexes as a Function of Concentration of Ligand in the Presence of KCN in $CHCl_3$–Water System

——: N-phenyl-β-mercaptocinnamamide, pH 4.66

------: N-ethyl-β-mercaptocinnamamide, pH 4.70

concentration of KCN: 5×10^{-2}M

(4) Iron [III] Complex

In the cases of I and II, the extraction constants were able to be measured in the pH region where the formation of hydroxocomplex of iron [III] was negligible as shown in Fig. 1, 2 and 7. On the other hand, in the case of III, the distribution curves shown in Fig. 2 and 7 were obtained. These curves may be explained by the application of the following equation (10) in which the term of the hydrolysis of iron [III] ion is introduced.

$$K_{ex}=\frac{D[H^+]^3\{1+\sum\beta_h[OH^-]^h\}}{[HR]_0{}^3} \qquad (10)$$

In equation (10), β_h is expressed as follows.

TABLE III. Extraction Constants and Stability Constants

Metal	Solvent	n	I			II			III		
			log K_{ex}	log P_c	log β_n	log K_{ex}	log P_c	log β_n	log K_{ex}	log P_c	log β_n
Ni	$CHCl_3$	2	−4.62	2.49	11.84	−5.34	2.55	10.07	−6.72	3.38	14.94
	$C_5H_{11}OAc$		−5.81	2.40	11.51	−5.16	2.40	10.14			
Zn	$CHCl_3$	2	−2.90	3.90	12.16	−2.56	4.41	10.99	−4.54	4.51	15.99
	$C_5H_{11}OAc$		−3.00	4.13	12.59	−1.92	4.78	11.00			
Hg	$CHCl_3$	2	21.8	3.47	37.3	21.7	2.77	36.9	19.3[a]	3.4[a]	41.8[a]
Co	$CHCl_3$	3	(−6.69)[b]	2.80	(18.95)[b]	(−7.02)[b]	2.96	(16.96)[b]	(−7.24)[b]	4.70	(25.6)[b]
	$C_5H_{11}OAc$		(−7.95)[b]	2.67	(18.96)[b]	(−6.32)[b]	2.81	(17.42)[b]			
Fe	$CHCl_3$	3	0.9[c]	−2.0[c]	28.6[c]	0.2[c]	—[d]	—[d]	−4.8	—[d]	—[d]

a) lit.[8]

b) The values in parentheses were calculated as the apparent values.

c) values in benzene and water system

d) Determination of these values was impossible for the hydrolysis of Fe [III] ion.

$$\beta_h = \frac{[Fe(OH)_h{}^{(3-h)+}]}{[Fe^{3+}][OH^-]^h}$$

When $10^{11.0}$ and $10^{21.7}$, which are the first and second stability constants of the hydroxo complex of iron [III],[9] were used as the value of β_h, the calculation of the extraction constant was achieved. The extraction constants and the overall stability constants of nickel, zinc, mercury, cobalt and iron complexes are listed in Table III.

Discussion

As shown in Table III, the extraction constants of the metal complexes with I and II which contain amide group are higher than those of the metal complexes with III. This fact may be regarded as a result of the decrease of the partition coefficient in I and II[3] caused by the hydrophilic character of the amide group. On the other hand, the order of the stability constants of the metal complexes is I, II$<$III.

Considerable differences were observed in the solubility into the organic solvent among the metal complexes of I. Cobalt [III] complex of I was not completely soluble in chloroform when the concentration of cobalt ion was more than 10^{-4}M but could be extracted in isoamyl acetate, and I was found to be available for the extraction-spectrophotometric determination of cobalt by the use of isoamyl acetate.[4a] The solubility of the isolated nickel [II] complex of I in chloroform and isoamyl acetate are about 1.0×10^{-2}M and 1.3×10^{-1}M at 25° respectively. The solubility of iron [III] complex of I is abnormally low in both solvents. The ratio of metal to ligand in the extraction system was found to be 1:3 from the value of n. The elemental analysis of the iron [III] complex of I verified this ratio. The solubility of the metal complex of I generally is low into the noncoordinating solvents than into the coordinating solvents. The behavior of the metal complexes of II is almost similar to that of the metal complexes of I. On the other hand, the metal complexes of III are highly soluble in most of the organic solvent. This remarkable difference in the solubility in organic solvents between the metal complexes of III and those of I and II was observed in their cobalt [III], nickel [II] and iron [III] complexes, but not in their zinc [II] and mercury [II] complexes which are readily soluble in various organic solvents. It has been found that, generally, an undissociated amide group coordinates with metal through its carbonyl oxygen atom. X-ray crystallographic analysis of nickel [II] complex of I[10] verified a structure of S-O coordinating type and cis-configuration in nickel [II] thiodibenzoylmethanate.[11] The values of P_c, are not so different each other among the complexes of I, II and III, in spite of the remarkable difference in the solubility of these complexes into chloroform. As mentioned in the experimental part, the determination of P_c value was made on the extraction systems which contain trace amounts of metal ion, hence it is expected that the difference of the solubility among the metal complexes of I, II and III may be interpreted by the difference of the effect between secondary amide and carbonyl groups on the complex formation in solution. The values of P_c in Table III suggest that the nickel [II] and zinc [II] complexes of I and II are not solvated with isoamyl acetate molecules, because the values obtained in isoamyl acetate were not significantly different from those obtained in $CHCl_3$, which is a noncoordinating solvent.

The stability constants and the extraction constants of cobalt complexes can not be determined strictly because the valency of cobalt changes from II to III in the course of the chelate formation. However, provided that the equilibrium of the chelate formation (11) is valid the apparent stability constants and extraction constants may be calculated.

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9) A. Ringbom, "Complexation in Analytical Chemistry," (translated by N. Tanaka and H. Sugi) Sangyo Tosho, Tokyo, 1965, p. 272.

10) M. Shiro, private communication.

11) L.P. Eddy, J.W. Hayes, S.E. Livingstone, H.L. Nigam and D.V. Radford, *Australian J. Chem.*, **24**, 1077 (1971).

$$\mathrm{Co[III]^{3+} + 3HR \rightleftharpoons Co[III]R_3 + 3H^+} \quad (11)$$

The values obtained are designated in parentheses in Table III. Similar oxidation reaction of cobalt [II] ion in the course of the complex formation has been reported in other extraction systems containing sulfur coordinating ligands.[4c–e,12)]

The stability constants of zinc [II] complexes of these three ligands were larger than those of nickel [II] complexes. This order has been found generally in sulfur-oxygen coordinating ligands such as N-acetyl-penicillamine,[13)] mercaptosuccinic acid,[14)] thioglycollic acid.[15)] The extent of the back donation from metal to sulfur is probably an important factor for the order of stability in nickel [II] and zinc [II] complexes, as proposed in various types of sulfur-coordinating complexes.[16)] The stability constants of nickel [II] and zinc [II] complexes of III measured in dioxane–water system by the potentiometric method[17)] are inconsistent with our values and the value for nickel [II] complex was found to be greater than that of zinc [II] complex. We consider that the stability constants obtained by the solvent extraction method are more reliable than those obtained by the potentiometric method, based on our opinion proposed in the previous paper,[3)] which concerns with the advantageousness of the solvent extraction method for the determination of the stability constants of the metal complexes of the ligands such as I, II and III, which are hydrolysed and oxidized readily in the course of the complex formation.

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