Formation Constants and Metal-to-Ligand Ratios for Tris(hydroxymethyl)aminomethane–Metal Complexes

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The ligand, Tris(hydroxymethyl)aminomethane, has various other names including Trizma Base, Tris, or THAM (*1*). It is used as a primary standard for HCl even though it has a fairly low equivalent weight of 121.1 and it is a weaker base than might be desired, $K_B = 1.26 \times 10^{-6}$. THAM is pure and nonhydroscopic (*2*). It is used primarily as a buffering medium in many biochemical applications. The buffering pH range is 7 to 9.

Other uses of Tris have also been noted in the literature. Manfredi et al. published the dramatic results they obtained when Tris was injected intravenously into patients suffering from CO_2 retention (*3*). Nahas reported the use of Tris in the treatment of artifically induced apnea in dogs (*4*). In 1958, McFarland and Norris published their results using Tris to reduce the mortality of fish in transport (*5*). The inhibition of the enzyme horseradish peroxidase by various metal ions in THAM buffer to give a 50% inhibition of the activity is noted in Guilbault's paper (*6*).

Since THAM is quite an important product, it is felt that there may be some significance in studying the complexing ability of the ligand.

There are several methods for the determination of formation constants and metal-to-ligand ratios of metal complexes. Let us consider the methods of polarography and potentiometry.

The polarographic method for investigating metal complexes (*7*) is based on the fact that the half-wave potential of metal ion is shifted cathodically when that ion enters into complex formation. Measuring the shift of the half-wave potential as a function of the complexing agent's concentration produces information concerning both the formula and stability constant of the metal complex. If we represent the reduction of the metal ion at a mercury electrode as $M^{n+} + Hg + ne^- \rightarrow M(Hg)$, the corresponding Nernst equation at 25 °C is:

$$E = E^\circ - \frac{0.0591}{n} \log \frac{M(Hg)}{[M^{n+}]} \quad (1)$$

If there are ligands that will complex with the metal ion, the complexation equation can be written as follows:

$$M^{n+} + pX \rightleftharpoons MX_p \quad (2)$$

$$K_f = \frac{(MX_p)}{(M^{n+})(X)^p}$$

therefore:

$$(M^{n+}) = \frac{(MX_p)}{K_f(X)^p} \quad (3)$$

Substituting Equation 3 in Equation 1 gives:

$$E = E^\circ - \frac{0.0591}{n} \log \frac{K_f(X)^p M(Hg)}{(MX_p)} \quad (4)$$

The half-wave potential for each wave is the potential when the concentrations of both the oxidized and reduced forms at the electrode surface are equal. Thus, for the uncomplexed system,

$$E_{1/2} = E^\circ - \frac{0.0591}{n} \log 1 = E^\circ \quad (5)$$

and for the metal complex,

$$E_{1/2} = E^\circ - \frac{0.0591}{n} \log K_f(X)^p \quad (6)$$

From the equations stated above, one can see that the difference in half potentials is given by

$$\Delta E_{1/2} = E_{1/2} \text{ (simple ion)} - E_{1/2} \text{ (complex ion)}$$

$$= \frac{0.0591}{n} \log K_f[X]^p$$

or

$$\Delta E_{1/2} = \frac{0.0591}{n} \log K_f + \frac{0.0591\, p}{n} \log [X] \quad (7)$$

Since the log K_f is constant, one may plot $\Delta E_{1/2}$ vs. log [X] to evaluate p. Once p has been evaluated, the above equation can be solved for log K_f. Thus, this method can be used for determining formation constants and ligand-to-metal ratios for metal–ligand complexes.

The potentiometric determination of formation constants depends on the free metal ions, which contribute to the voltage. If a large excess of ligand is present, essentially all the metal ion is in the form of the complex, and the free ligand concentration equals initial ligand concentration. The following equations will be used for determination of formation constants.

$$E = E^\circ + \frac{0.0591}{n} \log [M] \quad (8)$$

$$K_f = \frac{MX_p}{[M][X]^p} \quad (9)$$

$$\log K_f = \log (MX_p) - \log (M^{n+}) - p \log (X) \quad (10)$$

$$\log (M^{n+}) = \log (MX_p) - \log K_f - p \log (X) \quad (11)$$

A plot of log metal ion concentration vs. log of ligand concentration is a straight line. From the intercept, the slope of this straight line, and from Equation 11, we can determine ligand-to-metal ratios and formation constants.

EXPERIMENTAL

All solutions were made with distilled and deionized water. All chemicals were of reagent grade purity.

Polarographic Determination. *Stock Solutions.* These included Tris(hydroxymethyl)aminomethane (THAM) 2.00M, cadmium nitrate 0.0200M, lead nitrate 0.0200M, and sodium perchlorate 2.00M.

Apparatus. A Sargent model XV recording polarograph was used for all polarographic measurements with a DME and SCE electrode system.

Procedure. A solution consisting of 0.5 ml of 0.0200M metal ion and 50 ml of 2.00M NaClO$_4$ were placed in a 100-ml volumetric flask and diluted to the mark with distilled water. The polarograph was set on a sensitivity of 0.01 μA per millimeter and the voltage was scanned from 0 to -2 volts vs. a saturated calomel electrode. On succeeding experiments, increasing amounts of 2.00M THAM in 1.00M NaClO$_4$ were added and their polarograms established.

Potentiometric Determinations. *Stock Solutions.* THAM 1.00M, cupric nitrate 0.0100M, silver nitrate 0.0100M, and sodium perchlorate 1.00M were used.

Apparatus. A Corning model 7 pH meter was used for all millivolt measurements.

Table I. Polarographic Determination of K_f for the THAM–Lead Complex

Lead	
Log (X)	$\Delta E_{1/2}$(volts)
−2.00	0.043
−1.52	0.070
−1.00	0.097
−0.52	0.126
$p = 1.86$	$\log K_f = 5.22$

Table II. Polarographic Determination of K_f for the THAM–Cadmium Complex

Cadmium	
Log (X)	$\Delta E_{1/2}$ (volts)
−2.00	0.012
−1.52	0.042
−1.00	0.080
−0.52	0.112
$p = 2.20$	$\log K_f = 5.46$

Table III. Potentiometric Determination of K_f for the THAM–Silver Complex

Log (X)	E_{measd} (volts)	E_{Ag} (volts)	Log (Ag$^+$)
−2.00	0.168	0.414	−6.53
−1.70	0.133	0.379	−7.12
−1.53	0.113	0.359	−7.46
−1.40	0.097	0.343	−7.73
−1.31	0.086	0.332	−7.92
−1.23	0.078	0.324	−8.05
−1.17	0.071	0.317	−8.17
−1.11	0.063	0.309	−8.31
−1.06	0.058	0.304	−8.39
−1.02	0.053	0.299	−8.47
$p = 1.96$	$\log K_f = 6.45$		

Procedure. A platinum wire electrode was plated from an acidic medium with the metal whose THAM complex was to be studied. Then 1 ml of the $0.0100M$ metal solution was diluted to 100 ml and transferred to a beaker. The plated electrode and a saturated calomel electrode were placed in the beaker, then THAM was put into the beaker from a buret and the potential was recorded.

DISCUSSION

Polarographic and potentiometric techniques were used to determine the formation constants and ligand-to-metal ratios of THAM complexes in this study. Data are presented in Tables I–V.

Lead(II) and cadmium(II) both have a polarographic wave between 0 and −2 volts vs. SCE, and were studied by this technique. Silver(I) and copper(II) were studied by a potentiometric technique.

In the polarographic study, sodium perchlorate was used as the inert electrolyte. It was chosen because the perchlorate ion has been shown to be a very weak complexing ligand, if it complexes at all. This would allow the ligand THAM to not have to compete with ClO_4^- for the metal's coordinating sites.

In both of these studies, large excesses of ligand were employed for two reasons: to ensure that only one complex was formed (highest complex), and to be able to equate the analytical initial concentrations of the ligand to the equilibrium concentrations.

Table IV. Potentiometric Determination of K_f for the THAM–Copper Complex

Copper			
Log (X)	E_{measd} (volts)	E_{Cu} (volts)	Log (Cu^{2+})
−2.00	−0.137	0.109	−7.73
−1.70	−0.162	0.084	−8.58
−1.53	−0,173	0.073	−8.95
−1.40	−0.186	0.060	−9.39
−1.31	−0.197	0.049	−9.76
$p = 2.88$	$\log K_f = 9.55$		

Table V. Results

Complex	$\log K_f$	p
THAM–Pb^{2+}	5.22	1.86
THAM–Cd^{2+}	5.46	2.20
THAM–Ag$^+$	6.45	1.96
THAM–Cu^{2+}	9.55	2.88

The structure of the ligand, THAM, is $H_2NC(CH_2OH)_3$. Since the ligand is neutral, the activity of the ligand was assumed to be equal to its concentration. The activities of the various metals were obtained directly from the various potentiometric determinations, since electrode potentials measure directly the activity of a metal ion.

Tris(hydroxymethyl)aminomethane can have four coordination sites, one lone pair on nitrogen and three lone pairs on the individual oxygens. Silver(I)–THAM complex has a ligand-to-metal ratio of essentially 2 (1.86 experimentally), $\log K_f = 5.22$. This would probably indicate a linear complex of SP hybridization; coordination most likely, from the lone pair on the nitrogen in each ligand (8).

Cadmium(II)–THAM complex has a ligand-to-metal ratio of 2 (2.20 experimentally), $\log K_f = 5.46$. The cadmium complex is also a linear complex. Copper(II)–THAM complex has a ligand-to-metal ratio of 3 (2.88 experimentally), $\log K_f = 9.55$. The copper complex is probably octahedral, forming chelates with five-membered rings, utilizing the lone pair on nitrogen and one lone pair on one of the oxygens.

Lead(II)–THAM complex has a ligand-to-metal ratio of 2 (1.86 experimentally), $\log K_f = 5.22$. This may imply a linear SP hybrid structure (coordination through N lone pairs alone).

Since the K_f for Cu^{2+} is greater than 10^8, it may be a suitable ligand for the complexometric titration of Cu^{2+} (9).

LITERATURE CITED

(1) Sigma Technical Bulletin, No. 106B, Sigma Chemical Co., St. Louis, Mo.
(2) J. S. Fritz and G. H. Schenk, "Quantitative Analytical Chemistry", third edition, Allyn and Bacon, Inc. Boston, 1974, pp 187, 656.
(3) F. Manfredi, H. Sieker, A. Spoto, and H. Saltzman, *J. Am. Med. Assoc.*, **173**, 999 (1960).
(4) G. Nahas, *Science*, **129**, 782 (1959).
(5) W. McFarland and K. Norris, *Calif. Fish Game*, **44** (4), 291 (1958).
(6) G. Guilbault, P. Brignac, Jr., and M. Zimmer, *Anal. Chem.*, **40**, 190 (1968).
(7) C. N. Reilley and D. T. Sawyer, "Experiments for Instrumental Analysis", McGraw-Hill, New York, 1961, pp 70, 176.
(8) F. Cotton and G. Wilkinson, "Advanced Inorganic Chemistry", Interscience, New York, 1966, pp 482, 610, 898.
(9) H. Laitinen, "Chemical Analysis", McGraw-Hill, New York, 1960, p 231.