222. *The Stability Constants of the Silver Complexes of Some Aliphatic Amines and Amino-acids.*

By S. P. DATTA and A. K. GRZYBOWSKI.

The stability constants of the silver complexes of 2-aminoethanol (ethanolamine), 2-aminoethyl phosphate, 2-aminoethyl sulphate, taurine, 2-amino-2-methylpropane-1 : 3-diol, sarcosine, and *NN*-dimethylglycine have been determined electrometrically at 25°, and for glycine and arginine over a wide temperature range. The thermodynamic quantities associated with the formation of the glycine and arginine silver complexes have been calculated, and their significance discussed.

DURING recent years a number of accurate determinations, over a wide temperature range, have been made in this laboratory of the acid dissociation constants of a number of aliphatic amines and amino-acids.[1] These determinations were made from e.m.f. measurements of cells without liquid junction containing hydrogen and silver–silver chloride electrodes and amine–chloride buffer solutions. In such cells there may be interactions between the silver chloride of the electrodes and the basic form of the amino-groups of the buffer acids. It was therefore necessary to know the stability constants of the silver complexes of these compounds in order to apply corrections to the e.m.f. results when necessary.

For most of the compounds determinations were made only at 25°, but with both arginine and glycine a wide temperature range was covered. The stability constants were determined from glass-electrode pH measurements during the titration with alkali of the acid form of the amine in the presence of silver nitrate. The values of the constants for glycine at 25° were determined by means of Irving and Rossotti's equation:[2]

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

All the others were determined from Rossotti and Rossotti's equation:[3]

$$[\mathrm{L}]^2\left(\frac{2}{\bar{n}}-1\right) = \frac{1}{K_2}[\mathrm{L}]\left(1-\frac{1}{\bar{n}}\right) + \frac{1}{K_1K_2} \quad . \quad . \quad . \quad . \quad (2)$$

where [L] is the ligand concentration, $\bar{n}$ is the average number of ligand molecules bound per silver atom, and the successive stability constants are: $K_1 = [\mathrm{LAg}]/[\mathrm{L}][\mathrm{Ag}^+]$ and $K_2 = [\mathrm{L_2Ag}]/[\mathrm{L}][\mathrm{LAg}]$. In all cases the term ligand refers to that form of the compound where the amino-group is unprotonated.

The hydrogen-ion concentration was calculated from the experimentally determined pH values by using the relation, $[\mathrm{H}^+] = \text{antilog}\,[-\mathrm{pH} - \log \gamma_{\mathrm{H}^+}]$, where γ_{H^+} was assumed to be equal to $\gamma_{\pm(\mathrm{HCl})}$ at the average ionic strength during the titration.

The constants of equations (1) or (2) were determined for glycine, arginine, and 2-amino-2-methylpropane-1 : 3-diol by the method of least squares after the linearity of the points used had been determined graphically. The graphical method only was used for the other ligands. The values of $\gamma_{\pm(\mathrm{HCl})}$ used were those of Bates and Bower;[4] the acid dissociation constants of the ligand amino-groups were determined from titration curves or calculated from the thermodynamic constants.

It was possible that silver ions might interact with the phosphate radical of 2-aminoethyl phosphate. That this is not so was established by a comparison of the titration

[1] Datta and Grzybowski, *Trans. Faraday Soc.*, 1958, **54**, 1179.
[2] Irving and Rossotti, *J.*, 1953, 3397.
[3] Rossotti and Rossotti, *Acta Chem. Scand.*, 1955, **9**, 1166.
[4] Bates and Bower, *J. Res. Nat. Bur. Stand.*, 1954, **53**, 283.

curves for this compound in the absence and in the presence of silver ions, the shapes of these curves indicating that the silver ions were only complexed by the basic form of the amino-group.

Results and Discussion.—The values of the stability constants obtained are shown in Tables 1 and 2. The values of log K_1 and log K_2 for the arginine and glycine complexes at different temperatures were fitted to Harned and Robinson's equation: [5]

$$\log K = A/T - D + CT \quad . \quad . \quad . \quad . \quad . \quad . \quad . \quad (3)$$

The values of the constants thus obtained are shown in Table 3.

TABLE 1. *Concentration stability constants at 25° of the silver complexes of some aliphatic amines and amino-acids.*

(Ligand refers to that ionic form of the compound in which the amino-group is not protonated.)

Ligand	Amino pK_a'	log K_1	log K_2	Mean I	Observed range of $\bar{n}$	Ligand : metal ratio	No. of points on titration curve
2-Aminoethanol	$9{\cdot}51_4$	$3{\cdot}29_3$	$3{\cdot}53_4$	$0{\cdot}015_2$	0—1·5	4	15
2-Aminoethyl phosphate...	$10{\cdot}31_6$	$3{\cdot}89_3$	$3{\cdot}22_8$	$0{\cdot}034_6$	0—1·4	4	16
2-Aminoethyl phosphate...	$10{\cdot}34_2$	$3{\cdot}87_9$	$3{\cdot}25_2$	$0{\cdot}026_6$	0—1·76	6	12
2-Aminoethyl sulphate ...	$9{\cdot}13_5$	$3{\cdot}42_8$	$3{\cdot}48_0$	$0{\cdot}013_6$	0·2—1·7	4	18
Taurine	$8{\cdot}95_0$	$3{\cdot}19_4$	$3{\cdot}19_9$	$0{\cdot}014_1$	0—1·5	4	13
2-Amino-2-methylpropane-1 : 3-diol.....	$8{\cdot}79_9$	$3{\cdot}20_0$	$3{\cdot}66_8$	$0{\cdot}027_5$	0·2—1·2	2	27
Sarcosine	$10{\cdot}11_0$	$3{\cdot}04_2$	$2{\cdot}84_9$	$0{\cdot}016_5$	0·2—1·0	4	11
NN-Dimethylglycine	$9{\cdot}86_8$	$2{\cdot}91_0$	$2{\cdot}81_8$	$0{\cdot}013_2$	0·4—0·9	4	7

TABLE 2. *Concentration stability constants of the silver complexes of glycine and arginine.*

[The values of log $K_{calc.}$ were derived from the parameters of equation (3) given in Table 3.]
Δ = log $K_{calc.}$ − log $K_{obs.}$ σ = Standard error.

Temp.	Amino-pK_a'	log $K_{1obs.}$	log $K_{1calc.}$	$10^3\Delta$	log $K_{2obs.}$	log $K_{2calc.}$	$10^3\Delta$	Observed range of $\bar{n}$	Ligand : metal ratio	No. of points on titration curve
				Glycine. Mean I = 0·01.						
5°	$10{\cdot}26_4$	$3{\cdot}47_0$	$3{\cdot}47_2$	+2	$3{\cdot}87_0$	$3{\cdot}85_3$	−17	0—1·8	3	20
15	$9{\cdot}96_8$	$3{\cdot}48_3$	$3{\cdot}48_4$	+1	$3{\cdot}63_4$	$3{\cdot}63_6$	+2	0—1·8	3	18
25	9·66	$3{\cdot}44_5$	$3{\cdot}43_1$	−14	$3{\cdot}37_2$	$3{\cdot}43_2$	+60	0—1·3	2	17
35	$9{\cdot}44_5$	$3{\cdot}29_9$	$3{\cdot}32_1$	+22	$3{\cdot}26_7$	$3{\cdot}23_9$	−28	0—1·5	3	16
45	$9{\cdot}21_4$	$3{\cdot}16_9$	$3{\cdot}15_7$	−12	$3{\cdot}10_1$	$3{\cdot}05_7$	−44	0—1·7	3	22
55	$8{\cdot}98_2$	$2{\cdot}94_4$	$2{\cdot}94_6$	+2	$2{\cdot}85_5$	$2{\cdot}88_3$	+28	0—1·3	3	16
σ5°, 55°			$0{\cdot}01_5$			$0{\cdot}04_5$				
σ25°			0·01			$0{\cdot}03_0$				
				Arginine. Mean I = $0{\cdot}023_7$.						
0	$9{\cdot}74_2$	$3{\cdot}65_6$	$3{\cdot}64_0$	−16	$4{\cdot}18_2$	$4{\cdot}22_2$	+40	0—1·1	4	6
10	$9{\cdot}41_7$	$3{\cdot}36_3$	$3{\cdot}43_4$	+71	$4{\cdot}03_1$	$3{\cdot}96_2$	−69	0—1·1	4	9
20	$9{\cdot}13_0$	$3{\cdot}37_6$	$3{\cdot}26_1$	−115	$3{\cdot}71_8$	$3{\cdot}74_5$	+27	0—1·3	4	10
25	—	—	$3{\cdot}18_7$	—	—	$3{\cdot}65_2$				
30	$8{\cdot}87_7$	$3{\cdot}04_2$	$3{\cdot}12_0$	+78	$3{\cdot}62_2$	$3{\cdot}56_7$	−55	0—1·0	4	9
40	$8{\cdot}65_6$	$3{\cdot}02_2$	$3{\cdot}00_6$	−16	$3{\cdot}32_1$	$3{\cdot}42_5$	+104	0—1·2	4	12
50	$8{\cdot}46_3$	$2{\cdot}91_9$	$2{\cdot}91_7$	−2	$3{\cdot}36_1$	$3{\cdot}31_4$	−47	0—1·0	4	8
σ0°, 50°			$0{\cdot}08_3$			$0{\cdot}08_0$				
σ25°			$0{\cdot}05_7$			$0{\cdot}05_5$				

The thermodynamic quantities associated with the formation of the 1 : 1 and 2 : 1 ligand : silver complexes of arginine at 25° and of glycine from 5° to 55° were calculated in the usual way from the constants of equation (3); these are shown in Tables 4 and 5, together with estimates of their errors.[6]

[5] Harned and Robinson, *Trans. Faraday Soc.*, 1940, **36**, 973.
[6] Please, *Biochem. J.*, 1954, **56**, 196.

The temperature coefficient of the constant for the formation of the 1 : 1 glycine : silver complex is remarkable insofar as there is apparently a temperature of maximum stability (11·7°). This is in contrast to the acid dissociation constant of the amino-group of glycine, which diminishes in magnitude with increasing temperature in a manner suggesting that

TABLE 3. *Parameters of equation (3) for the formation of silver complexes of arginine and glycine.*

(The charge on the ligand species is given by z.)

	Arginine: $I = 0{\cdot}023_7$; $z = 0$.			Glycine: $I = 0{\cdot}01$; $z = -1$.		
Reaction	*A*	*D*	*C*	*A*	*D*	*C*
$L^z + Ag^{1+} \rightleftharpoons LAg^{z+1}$ ...	3855·2	18·452	0·029208	−7686·3	−57·454	−0·094725
$LAg^{z+1} + L^z \rightleftharpoons L_2Ag^{2z+1}$ ...	4898·35	23·905	0·037321	1506·7	0·763	−0·002879
$LH^{z+1+} \rightleftharpoons L^z + H^{1+}$ ...	−4349·7	−12·654	−0·023694	—	—	—

there is a temperature of minimum stability of the protonated form, albeit far outside the experimental range.[1] This difference is interesting in view of the apparently similar electrostatic changes occurring on the attachment of a silver ion or a proton to the basic

TABLE 4. *Thermodynamic quantities associated with the formation of silver complexes and acid dissociation (amino-group) of arginine at 25°.*

Reaction	I	ΔG (kJ/mole)	ΔH (kJ/mole)	$-\Delta S$ (J/mole-deg.)
$L + Ag^+ \rightleftharpoons LAg^+$	0·0237	−18·19 ± 0·33	−24·1 ± 3·7	20 ± 12
$LAg^+ + L \rightleftharpoons L_2Ag^+$ ...	0·0237	−20·84 ± 0·32	−30·3 ± 3·6	32 ± 12
$L^+ \rightleftharpoons H^+ + L$	0·0237	51·37 ± 0·23	42·9 ± 2·6	28 ± 9
$L^+ \rightleftharpoons H^+ + L$	0	51·320 ± 0·007	44·89 ± 0·08	21·58 ± 0·29

TABLE 5. *Thermodynamic quantities associated with the formation of the silver complexes of glycine.*

Temp.	$-\Delta G$ (kJ/mole)	$-\Delta H$ (kJ/mole) 1 : 1 complex	$-\Delta S$ (J/mole-deg.)	$-\Delta C_p$ (J/mole-deg.)
5°	18·49 ± 0·09	−6·9 ± 2·2	−91 ± 8	1009 ± 86
15	19·22	3·4	−55	1045
25	19·59 ± 0·06	14·0 ± 0·8	−19 ± 3	1081 ± 92
35	19·59	25·0	18	1118
45	19·23	36·3	54	1154
55	18·51 ± 0·09	48·0 ± 2·6	90 ± 8	1190 ± 101
		2 : 1 complex		
25	19·59 ± 0·18	33·7 ± 2·5	47 ± 8	—

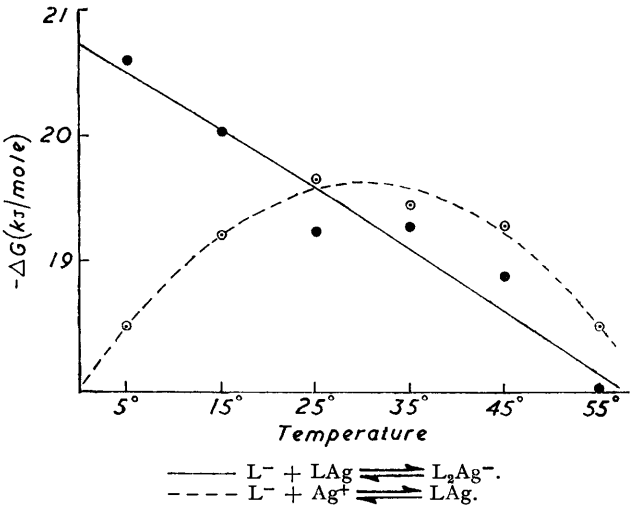
glycine amino-group. The value of ΔC_p for the formation of this complex is very large and negative, and the values given in Table 5 are significant despite the magnitude of their errors. Further, the value of ΔG for this process passes through a minimum at 30·1° (see Figure). Any explanation of these phenomena must await results of greater precision.

The values of the thermodynamic quantities associated with the second acid dissociation of arginine at $I = 0{\cdot}0237$, given in Table 4, were obtained from titration-curve data, their errors being similar to those for the values for the formation of the silver complexes. The fact that the former values are consistent, within the limits of error given, with those at $I = 0$ for the same acid dissociation obtained from measurements of high precision in cells without liquid junction [7] may be taken as evidence of the correctness of the estimates of the reliability of these results. The formation of the 1 : 1 arginine : silver complex does not show any of the phenomena described above for the corresponding glycine complex. The values of K_1 diminish continuously with increasing temperature and ΔG becomes

[7] Datta and Grzybowski, in preparation.

progressively more negative. The sign of ΔS is negative, being opposite to that for the corresponding reaction involving a proton.

The free energy of formation of Ag–*glycine complexes.*



The value of log K_2 for the arginine–silver complex is larger than log K_1 at all the temperatures investigated; with the glycine–silver complexes, on the other hand, log K_2 is larger than log K_1 only up to about 25°, above which temperature log K_1 is the larger.

TABLE 6. *Comparison of the stability constants of silver complexes of ethanolamine and glycine obtained in this work and by other investigators.*

Method *	I[20]	Temp.	pK_a'	log K_1	log K_2	Ref.
			Glycine			
gl.	0	25°	9·78	3·51	3·38	12
Ag	≈0·1	19	—	3·59	3·65	13
gl.	0·5M-KNO_3	20	9·76	3·7	3·3	14
sol.	0	25	9·78	4·28	—	15
gl.	0·01	25	9·66	3·43	3·43	This work
calc.	0·01	20	—	3·47	3·53	,,
			Ethanolamine			
gl.	0·5M-KNO_3	25	9·74	3·13	3·55	16
gl.	0·5M-KNO_3	25	9·60	3·11	3·57	17, 18
gl.	1M-$NaClO_4$	25	9·47	3·12	3·65	19
gl.	0·015	25	9·51	3·29	3·53	This work

* gl. = glass electrode; Ag = silver electrode; sol. = solubility; calc. = calculated.
[12] Monk, *Trans. Faraday Soc.*, 1951, **47**, 292, 297. [13] Dubois, *Compt. rend.*, 1957, **224**, 113. [14] Flood and Lovås, *Tidskr. Kjemi Bergvesen Met.*, 1945, **5**, 83. [15] Keefer and Reiber, *J. Amer. Chem. Soc.*, 1941, **63**, 689. [16] Bruehlman and Verhoek, *ibid.*, 1948, **70**, 1401. [17] Bjerrum, *Chem. Rev.*, 1950, **46**, 381. [18] Bjerrum and Refn, *Suomen Kem.*, 1956, **29**, *B*, 68. [19] Lotz, unpublished results from McIntyre, Doctoral Diss., Pennsylvania State Coll., 1953. Values at 25° have been obtained by interpolation from the data given in "Stability Constants, Part 1," The Chemical Society, 1958, p. 9. [20] The conventions used in indicating the ionic strength are those of *op. cit.*, p. xi.

The differences between the successive stability constants are probably due to an interplay of three factors: (*a*) the statistic effect, which would make log K_2 about 0·6 unit smaller than log K_1, (*b*) the electrostatic effect, which in most cases would tend to weaken the second complex (since the magnitude of this effect will increase with increasing tem-

perature it may account for the crossing over of K_1 and K_2 in the glycine complexes), and (*c*) changes in the electron orbitals involved in complex formation;[8] such changes (*c*) must be of considerable importance in the instances when K_2 is larger than K_1. The present results are compared with those of other workers in Table 6.

Experimental

Glass-electrode pH measurements were made with a Philips pH meter (type GM 4491) for 2-aminoethyl phosphate, 2-aminoethyl sulphate, *NN*-dimethylglycine, ethanolamine, and sarcosine; with a Tinsley vernier potentiometer and an impedance converter for glycine at 5°, 25°, and 45°, and arginine; and with a precision pH meter built in this laboratory by Mr. R. Leberman for glycine at 15°, 35°, and 55° and for 2-amino-2-methylpropane-1 : 3-diol. The cell used for the titrations consisted of a titration vessel and a saturated calomel reference electrode vessel, connected by a glass bridge which was filled with hot potassium nitrate solution which had been saturated at its b. p. On cooling, this bridge solution was practically solid. The ends of the bridge were closed by discs of sintered glass, fused in position.

The titration mixture contained, in most instances, the ligand and silver nitrate only, in the ratios shown in Tables 1 and 2. Arginine, ethanolamine, sarcosine, *NN*-dimethylglycine, and 2-amino-2-methylpropane-1 : 3-diol were only available as the hydrochlorides, and the titration mixtures were prepared by adding an appropriate excess of silver nitrate and filtering off the precipitated silver chloride. Titrations were effected with approx. 0·1M-potassium or sodium hydroxide (carbonate-free). During the titration the solution in the vessel was bubbled with nitrogen to effect mixing and to exclude carbon dioxide.

The cell was standardised in 0·05M-potassium hydrogen phthalate and either 0·01M- or 0·05M-borate.[9,10] All titrations were done in an oil-thermostat, the temperature being maintained within $\pm 0{\cdot}1°$ of the stated value.

All the ligands, except 2-aminoethyl phosphate which was prepared in this laboratory,[11] were purchased. They were purified by recrystallisation, some of them being previously converted into the hydrochlorides. Their purity was checked by titration and, where possible, by gravimetric chloride determinations.

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Department of Biochemistry, University College, London, W.C.1.
Department of Biochemistry and Chemistry,
Guy's Hospital Medical School, London, S.E.1. [*Received, October 15th*, 1958.]

[8] Burkin, *Quart. Rev.*, 1951, **5**, 1.
[9] Bower and Bates, *J. Res. Nat. Bur. Stand.*, 1957, **59**, 261.
[10] British Standards Institution, 1950, 1647.
[11] Clarke, Datta, and Rabin, *Biochem. J.*, 1955, **59**, 209.