

## CHAPTER VIII

### **Electrical Conductance of Some Tetraalkylammonium and Alkali Metal Bromides and Sodium Tetraphenylborate in Ethane-1, 2-diol and Formamide at 298.15 K**

---

The electrical conductivities of solutions of salts, ( $\text{Me}_4\text{NBr}$ ,  $\text{Et}_4\text{NBr}$ ,  $\text{Pr}_4\text{NBr}$ ,  $\text{Bu}_4\text{NBr}$ ,  $\text{NaBPh}_4$ ,  $\text{LiBr}$ ,  $\text{NaBr}$ , and  $\text{KBr}$ ) in pure molecular liquids (ethane-1, 2-diol and formamide) have been measured at 298.15 K. The conductance data have been analyzed by the 1978 Fuoss-concentration equation in terms of limiting molar conductance ( $\Lambda^0$ ), the association constant ( $K_A$ ), and the association diameter ( $R$ ). The limiting molar conductances of studied electrolytes in molecular liquid (ethane-1, 2-diol) have been compared well with the same electrolytes in another molecular liquid (formamide). The limiting ionic conductance have been estimated from the appropriate division of the limiting molar conductivity value of the "reference electrolyte"  $\text{Bu}_4\text{NBPh}_4$ . The results have been analyzed in terms of ion-ion and ion-solvent interactions.

---

**Keywords:** Limiting molar conductance; association constant; association diameter; solvation; Walden product; ethane-1, 2-diol (EG); formamide (FM).

#### **8.1. Introduction**

The studies of transport properties of electrolytes in different molecular liquids are of considerable importance for the information they provide on the behavior of ions in solutions. The Conductometric method is well suited to investigate the ion-solvent and ion-ion interactions in electrolyte solutions<sup>(1-4)</sup>.

Recently, we have initiated a comprehensive program to study solvation and association behavior of several 1:1 electrolytes in different non-aqueous

and mixed molecular liquids from the measurements of various transport and thermodynamic properties<sup>(5-8)</sup>.

In this paper, an attempt is made to unravel the nature of various types of interactions prevailing in solutions of some selected tetraalkylammonium bromides and alkali metal salts, e.g., tetramethylammonium bromide ( $\text{Me}_4\text{NBr}$ ), tetraethyl ammonium bromide ( $\text{Et}_4\text{NBr}$ ), tetrapropylammonium bromide ( $\text{Pr}_4\text{NBr}$ ), tetrabutylammonium bromide ( $\text{Bu}_4\text{NBr}$ ), sodium tetraphenylborate ( $\text{NaBPh}_4$ ), lithium bromide ( $\text{LiBr}$ ), sodium bromide ( $\text{NaBr}$ ), and potassium bromide ( $\text{KBr}$ ) in ethane-1, 2-diol (EG) and in formamide (FM) - from precise conductivity measurements.

## 8.2. Experimental

### Chemicals

The tetraalkylammonium bromides were purum or puriss grade (Fluka) and were purified as described in the literatures<sup>(9-10)</sup>. These salts were purified by recrystallization and the higher homologs recrystallized twice to ensure maximum purity. The recrystallized salts were dried in *vacuo* at elevated temperatures for 12 h. Sodium tetraphenylborate was precipitated from acetone and dried in *vacuo* at 80°C for 72 h. Lithium bromide, sodium bromide and potassium bromide were dried under vacuum at high temperature for 48 h and were used without further purifications.

Formamide (FM) and ethane-1, 2 diol (EG) were purified as described in the literatures<sup>(11)</sup>. The purified molecular liquids properties (FM: density 1.1292 g-cm<sup>-3</sup>, and coefficient of viscosity 3.302 m Pa-s; EG: density 1.1098 g-cm<sup>-3</sup>, and coefficient of viscosity 16.9 m Pa-s) are in good agreement with the literature values<sup>(12)</sup>.

### **Apparatus and Procedure**

The densities ( $\rho$ ) were measured with an Ostwald-Sprengel type pycnometer having a bulb volume of 25 cm<sup>3</sup> and an internal diameter of the capillary of about 0.1 cm. The pycnometer was calibrated at 298.15 K with doubly distilled water and benzene. Details have been described earlier<sup>(13-14)</sup>.

A stock solution for each salt was prepared by mass, and the working solutions were obtained by mass dilution. The conversion of molality to molarity was done using the density values.

In order to avoid moisture contamination, the solutions were prepared in a dehumidified room with utmost care.

The viscosities were measured by means of a suspended level Ubbelohde viscometer at the desired temperature (accuracy  $\pm 0.01^\circ\text{C}$ ). The precision of the viscosity measurement was  $\pm 0.05\%$ . Details have been described earlier<sup>(15)</sup>.

Conductance measurements were carried out on a Systonic-306 conductivity bridge using a dip-type cell (cell constant = 1.01 cm<sup>-1</sup>) with an accuracy of  $\pm 0.01\%$ . The cell was calibrated by the method of Lind and co-workers<sup>(16)</sup> using aqueous potassium chloride solutions. Measurements were made as described earlier<sup>(17)</sup>.

Several independent solutions were prepared and conductance measurements were performed with each of these to ensure the reproducibility of the results. Corrections were made for the specific conductance of the solvent.

The relative permittivity of formamide (FM) ( $\epsilon = 109.5$ ) and ethane-1, 2-diol (EG) [ $\epsilon = 40.7$ ] at 298.15 K was taken from the literature<sup>(12)</sup>.

### 8.3. Results and Discussion

Molar conductances ( $\Lambda$ ), of electrolyte solutions as a function of molar concentrations ( $c$ ) are given in Table I in different molecular liquids at 298.15 K. The plots of molar conductances  $\Lambda$  versus square root of molar concentrations  $\sqrt{c}$  of all the electrolytes in both the molecular liquids are presented in fig.1-2.

The conductance data have been analyzed by the 1978 Fuoss conductance-concentration equation<sup>(18-19)</sup>. For a given set of conductivity values ( $c_j, \Lambda_j, j = 1 \dots, n$ ), three adjustable parameters, the limiting molar conductivity ( $\Lambda^0$ ), the association constant ( $K_A$ ) and the distance of closest approach of ions ( $R$ ) are derived from the following set of equations:

$$\Lambda = p [\Lambda^0 (1 + R_X) + E_L] \quad (1)$$

$$p = 1 - \alpha (1 - \gamma) \quad (2)$$

$$\gamma = 1 - K_A c \gamma^2 f^2 \quad (3)$$

$$-\ln f = \beta \kappa / 2(1 + K_R) \quad (4)$$

$$\beta = e^2 / \epsilon k_B T \quad (5)$$

$$K_A = K_R / (1 - \alpha) = K_R (1 + K_S) \quad (6)$$

where  $R_X$  is the relaxation field effect,  $E_L$  is the electrophoretic countercurrent,  $\epsilon$  is the relative permittivity of the solvent,  $e$  is the electron charge,  $k_B$  is the Boltzmann constant,  $\gamma$  is the fraction of solute present as an unpaired ion,  $c$  is the molarity of the solution,  $f$  is the activity coefficient,  $T$  temperature in absolute scale, and  $\beta$  is the twice the Bjerrum distance.

The computations were performed on a compute using the program suggested by Fuoss. The initial  $\Lambda^0$  values for the iteration procedure were obtained from Shedlovsky extrapolation of the data<sup>(20)</sup>. Input for the program is

the set  $(c_j, A_j; j = 1, \dots, n)$ ,  $n$ ,  $\epsilon$ ,  $\eta$ ,  $T$ , initial values of  $\Lambda^0$ , and an instruction to cover a preselected range of  $R$  values.

In practice, calculations are performed by finding the values of  $\Lambda^0$  and  $\alpha$ , which minimize the standard deviation,  $\sigma$ ,

$$\sigma^2 = \sum_{j=1}^n [A_j(\text{calcd}) - A_j(\text{obsd})]^2 / (n - 2) \quad (7)$$

for a sequence of  $R$  values and then plotting  $\sigma$  against  $R$ ; the best-fit  $R$  corresponds to the minimum in  $\sigma$  versus  $R$  curve. However, since a rough scan using unit increment of  $R$ -values from 4 to 20 gave no significant minima in the  $\sigma$  (%) vs.  $R$ -curves, the  $R$  values was assumed to be  $R = a + d$ , where  $a$  is the sum of the crystallographic radii of the ions and  $d$  the average distance corresponding to the side of a cell occupied by a solvent molecule. The distance  $d$  is given by

$$d = 1.183 (M/\rho)^{1/3} \quad (8)$$

where  $M$  is the molecular weight of molecular liquid and  $\rho$  is its density.

The values of  $\Lambda^0$ ,  $K_A$  and  $R$  obtained by this procedure are reported in Table II.

In order to investigate the specific behavior of the individual ions comprising these electrolytes, it is necessary to split the limiting molar electrolyte conductances into their ionic components. In the absence of accurate transport number data of the system, we have used the "reference electrolyte" method for the division of  $\Lambda^0$  into their ionic components.  $\text{Bu}_4\text{NBPh}_4$  has been used as the "reference electrolyte" <sup>(21)</sup>.  $\text{Bu}_4\text{NBPh}_4$  was also used as the reference electrolyte by Fuoss and Hirsch <sup>(22)</sup> to evaluate the limiting ionic conductances in several organic solvents. We have divided the  $\Lambda^0$  values of  $\text{Bu}_4\text{NBPh}_4$  into ionic components using a method similar to that proposed by Krumgalz <sup>(23)</sup> for division of viscosity  $B$  - coefficients:

$$\lambda^{\circ}(\text{Bu}_4\text{N}^+)/\lambda^{\circ}(\text{Ph}_4\text{B}^-) = r(\text{Ph}_4\text{B}^-) / r(\text{Bu}_4\text{N}^+) = 5.35 / 5.00 = 1.07 \quad (9)$$

The  $r$ -values have taken from the work of Gill et al <sup>(24-25)</sup>.

The limiting molar conductivity  $\lambda^{\circ}$  of the "reference electrolyte"  $\text{Bu}_4\text{NBPh}_4$  was obtained by considering the Kohlrausch rule that allows the calculation of the  $\lambda^{\circ}$  value for a given electrolyte by the appropriate combination of others. The  $\lambda^{\circ}$  values of  $\text{Bu}_4\text{NBr}$ ,  $\text{NaBPh}_4$  and  $\text{NaBr}$  obtained in the present solvent media have been used to obtain the  $\lambda^{\circ}$  value of  $\text{Bu}_4\text{NBPh}_4$  through the following equation:

$$\lambda^{\circ}(\text{Bu}_4\text{NBPh}_4) = \lambda^{\circ}(\text{Bu}_4\text{NBr}) + \lambda^{\circ}(\text{NaBPh}_4) - \lambda^{\circ}(\text{NaBr}) \quad (10)$$

The limiting ionic conductances calculated from the above equation are recorded in table III.

Table I show that the limiting equivalent conductivity for tetraalkylammonium bromides decreases with increasing length of the alkyl chain. This is in agreement with earlier findings in several pure and mixed solvents <sup>(5, 26-29)</sup>. We also see that  $\lambda^{\circ}$  values of alkali metal salts of common anion follow the sequence:  $\text{Li}^+ < \text{Na}^+ < \text{K}^+ <$ . This trend is in agreement with earlier finding in different pure and mixed molecular liquids <sup>(26, 30)</sup>. The above-mentioned trend is same in both the pure molecular liquids. But the values of  $\lambda^{\circ}$  are less in EG as compared with FM.

The association constants  $K_A$  of these electrolytes (cf., Table II) indicate that these salts are slightly associated in EG. This implies that a preponderant portion of each salt remains dissociated in this liquid. This quite expected because of the rather low permittivity of the liquid. For all these tetraalkylammonium bromide, the associations constants are found to be very close to each other, thus indicating these salts are almost similarly associated at any given concentration. Furthermore, the process of ionic association of these

electrolytes does not exhibit the simple dependence upon ionic size predicted by electrostatic theory. This behavior of tetraalkylammonium salts has also been reported in other molecular liquids <sup>(28, 29)</sup> and may be attributed to the specificity of the nature of interactions in these electrolyte solutions.

The association constants ( $K_A$ ) listed in Table II for all electrolytes in FM is practically negligible (i.e.,  $K_A < 10$ ). Therefore, the numerical values of  $K_A$  should not be taken seriously <sup>(31)</sup>. One can only conclude that these electrolytes exist as free ions in FM. This is expected because the relative permittivity of the pure FM is very high ( $\epsilon = 109.5$ ).

In both the liquids medium, the limiting ionic equivalent conductances of tetraalkylammonium ions decreases in the order  $\lambda_{Me_4N^+}^0 > \lambda_{Et_4N^+}^0 > \lambda_{Pr_4N^+}^0 > \lambda_{Bu_4N^+}^0$ , i.e., the limiting ionic conductivity values decreases with increasing size of these ions. On the other hand, the limiting ionic conductivity of alkali metal, bromide and tetraphenyl ions follows the sequence:  $\lambda_{Br^-}^0 > \lambda_{K^+}^0 > \lambda_{Na^+}^0 > \lambda_{Li^+}^0 > \lambda_{BPh_4^-}^0$ . This indicates that besides the relative permittivity and the viscosity of the media, the specific interaction of the ions with liquid media has a profound influence on their mobilities.

The Walden products  $\lambda_{\pm}^0 \eta^0$  of the ions are also included in Table III, which are usually employed to discuss the interactions of the ions with the liquid medium. From this table, we see that for the tetraalkylammonium ions, the Walden products decreases from tetramethylammonium to tetrabutylammonium ion and, for alkali metal ions, it follows the reverse trend. This points out that the electrostatic ion-solvent interaction is very weak in tetraalkylammonium ions, apparently due to the very low surface charge density on these tetraalkylammonium ions. On the other hand, the alkali – metal ions are small enough to possess high charge density, resulting in strong ion-solvent interactions <sup>(27)</sup>.

## References

1. B.E. Conway, J. O' M. Bockris, and E. Yeager, *Comprehensive Treatise of Electrochemistry*, Plenum Press, New York, vol. 5 (1983).
2. N. Islam and A.A. Ansari, *Bull. Chem. Soc. Jpn.*, **62**, 309-316 (1989).
3. R. Fernandez-Prini, *Physical Chemistry of Organic Solvents*, A.K. Covington and T. Dickinson, Wiley, New York, (1973).
4. N. Papadopoulos, *Can. J. Chem.*, **67**, 1624-1627 (1989).
5. M.N. Roy and A. Jha, *J. T. R. Chem.*, **11(2)**, 61-73 (2004).
6. M.N. Roy and A. Choudhury, *Pak. J. Sci. Ind. Res.*, **43(8)**, 162-166(2005).
7. D. Ghosh and B. Das, *J. Chem. Eng. Data*, **49**, 1771-1774 (2004).
8. M.C. Blanco, D.C. Champeney and M. Kameche, *Phys. Chem. Liq.*, **19**, 163-169 (1989).
9. D.F. Evans, C. Zawoyski, and R.L. Kay, *J. Phys. Chem.*, **69**, 3878-3885 (1965).
10. R.L. Kay, C.Z. Zawoyski, and D.F. Evans, *J. Phys. Chem.*, **69**, 4208-4215 (1965).
11. *Purification of laboratory chemicals*, 3<sup>rd</sup> edition, D.D. Perrin and W.L.F. Armarego, 1988, Pergamon Press Plc, Headington Hill Hall, Oxford, OX30BW, England.
12. A.K. Covington and T. Dickinson, *Physical Chemistry of Organic Solvent Systems*, Plenum Publishing Company Ltd, London.
13. M.N. Roy, B. Sinha, R. Dey, and A. Sinha, *International J. Therm. Phys.*, **26(5)**, 1549-1563 (2005).
14. B.B. Gurung, A. Choudhury and M.N. Roy, *J. Indian Chem. Soc.*, **81**, 330-334 (2004).
15. M.N. Roy, A. Jha and A. Choudhury, *J. Chem. Eng. Data*, **49**, 291-296 (2004).



16. J.E. Lind, J.J. Zwolenik, and R.M. Fuoss, *J. Am. Chem. Soc.*, **81**, 1557 (1959)
17. B. Das and D. K. Hazra, *J. Solution Chem.*, **27**, 1021-1030 (1998).
17. R.M. Fuoss, *Proc. Natl. Acad. Sci., U.S.A.* **75**, 16-20 (1978a).
18. R.M. Fuoss, *J. Phys. Chem.*, **82**, 2427-2440 (1978b).
19. R.M. Fuoss and T. Shedlovsky, *J. Amer. Chem. Soc.*, **71**, 1496-1498 (1949).
20. B.S. Krumgalz, *J. Chem. Soc. Faraday Trans.1*, **79**, 571 (1983).
21. R.M. Fuoss and E. Hirsch, *J. Amer. Chem. Soc.*, **82**, 1013 (1960).
22. B.S. Krumgalz, *J. Chem. Soc. Faraday Trans.1*, **76**, 1275 (1980).
23. D.S. Gill and A.N. Sharma, *J. Chem. Soc. Faraday Trans.1*, **78**, 475-484 (1982).
24. D.S. Gill and M.B. Singh, *J. Chem. Soc. Faraday Trans. I*, **78**, 119-125 (1982).
25. M.N. Roy, D. Nandi, and D.K. Hazra, *J. Indian Chem. Soc.*, **70**, 121-124 (1993).
26. D. Das, B. Das and D.K. Hazra, *J. Solution Chem.*, **32(1)**, 77-83 (2003).
27. D. Das, B. Das and D.K. Hazra, *J. Solution Chem.*, **31(5)**, 425-431 (2002).
28. B. Das and N. Saha, *J. Chem. Eng. Data*, **45**, 2-5 (2000).
29. A. Hammadi, D.C. Chameney, *J. Chem. Eng. Data.*, **45**, 1116-1120 (2000).
30. L. Bahadur, M.V. Ramanamurti, *Can. J. Chem.*, **62**, 1051-1055 (1984).

**Table I. Equivalent Conductances and Corresponding Molarities of Electrolytes in EG and FM at 298.15 K**

$10^4 c$ (mol-dm <sup>-3</sup> )	$\Lambda \times 10^4$ (S.m <sup>2</sup> -mol <sup>-1</sup> )	$10^4 c$ (mol-dm <sup>-3</sup> )	$\Lambda \times 10^4$ (S.m <sup>2</sup> -mol <sup>-1</sup> )
<b>Electrolyte + EG</b>			
<b>Me<sub>4</sub>NBr</b>		<b>Et<sub>4</sub>NBr</b>	
18.36	7.84	18.13	7.15
33.67	7.37	33.23	6.79
46.62	7.15	46.02	6.61
57.71	7.04	56.97	6.43
67.33	6.91	66.47	6.36
83.18	6.73	82.11	6.24
95.68	6.64	94.45	6.20
105.80	6.57	104.40	6.01
114.20	6.54	112.70	6.05
121.20	6.48	119.60	5.88
<b>Pr<sub>4</sub>NBr</b>		<b>Bu<sub>4</sub>NBr</b>	
19.20	6.66	19.07	6.05
35.20	6.38	34.97	5.79
48.74	6.18	48.42	5.67
60.34	6.05	59.94	5.51
70.40	5.93	69.93	5.40
86.96	5.78	86.39	5.30
100.00	5.71	99.38	5.22
110.60	5.63	109.90	5.17
119.40	5.59	118.60	5.13
126.70	5.55	125.90	5.01
<b>LiBr</b>		<b>NaBr</b>	
22.65	6.71	26.08	7.52
41.52	6.58	47.82	7.31
57.49	6.50	66.21	7.17
71.17	6.43	81.97	7.06
83.03	6.36	95.63	7.96
93.41	6.33	107.60	6.90
102.60	6.29	118.10	6.84
118.00	6.23	127.50	6.81
130.50	6.19	143.45	6.74
140.80	6.15	156.50	6.68

<b>KBr</b>		<b>NaBPh<sub>4</sub></b>	
21.48	8.79	18.56	4.15
39.38	8.48	34.02	4.03
54.53	8.25	47.10	3.93
67.51	8.08	58.31	3.86
78.77	7.95	68.03	3.80
88.61	7.84	84.04	3.71
97.30	7.75	96.68	3.65
111.90	7.60	106.9	3.61
123.80	7.50	115.40	3.58
133.56	7.42	122.46	3.56
<b>Electrolyte + FM</b>			
<b>Me<sub>4</sub>NBr</b>		<b>Et<sub>4</sub>NBr</b>	
387.70	33.48	377.10	31.97
480.00	32.28	466.90	30.88
560.00	31.48	544.70	30.18
630.00	31.03	612.80	29.46
691.70	30.23	672.80	29.01
746.60	29.40	726.20	28.64
795.70	29.22	817.00	27.78
840.00	28.64	891.30	27.15
916.30	28.54	923.60	26.95
979.90	27.75	953.20	26.87
<b>Pr<sub>4</sub>NBr</b>		<b>Bu<sub>4</sub>NBr</b>	
369.60	30.32	372.40	27.05
457.50	28.04	461.10	26.34
533.80	27.36	537.90	25.78
600.50	26.89	605.10	25.36
659.40	26.23	664.50	25.02
758.60	26.10	764.40	24.34
838.80	25.50	845.30	23.80
905.10	25.48	912.10	23.29
934.20	25.15	941.30	23.20
960.80	25.05	968.20	22.95
<b>LiBr</b>		<b>NaBr</b>	
373.90	29.14	382.00	34.80
462.90	28.63	472.90	34.00
540.10	28.17	551.70	33.32
607.60	27.74	620.70	32.77
667.20	27.40	681.60	32.33
720.10	27.16	735.60	31.86
767.50	26.80	784.00	31.44
810.20	26.55	827.60	31.16

883.80	26.20	867.00	30.98
945.20	25.80	935.50	30.25
	<b>KBr</b>		<b>NaBPh<sub>4</sub></b>
378.00	36.29	369.60	20.15
468.10	35.05	457.60	19.58
546.10	34.45	533.80	19.11
614.30	33.88	600.60	18.78
674.60	33.25	659.40	18.52
728.10	32.84	711.80	18.30
776.00	32.36	758.60	18.14
819.10	32.12	800.80	17.95
893.60	31.72	838.90	17.80
955.60	31.34	873.50	17.66

**Table II. Derived Conductivity Parameters of Electrolytes in EG and FM at 298.15 K**

Salt	$\Lambda^0 \times 10^4$ (S.m <sup>2</sup> .mol <sup>-1</sup> )	$K_A$ (dm <sup>3</sup> .mol <sup>-1</sup> )	$R$ (A <sup>0</sup> )	$\sigma$
<b>Electrolyte + EG</b>				
Me <sub>4</sub> NBr	8.34	39.55	9.94	0.06
Et <sub>4</sub> NBr	7.62	38.26	10.47	0.06
Pr <sub>4</sub> NBr	7.15	37.79	10.99	0.03
Bu <sub>4</sub> NBr	6.48	36.34	11.41	0.03
LiBr	7.04	11.41	7.08	0.01
NaBr	8.00	17.46	7.43	0.01
KBr	9.52	33.23	7.80	0.01
NaBPh <sub>4</sub>	4.44	30.87	9.68	0.01
<b>Electrolyte + FM</b>				
Me <sub>4</sub> NBr	42.56	9.62	9.46	0.19
Et <sub>4</sub> NBr	39.80	8.72	9.99	0.09
Pr <sub>4</sub> NBr	35.75	7.68	10.51	0.49
Bu <sub>4</sub> NBr	32.60	6.92	10.93	0.11
LiBr	33.54	4.62	6.60	0.09
NaBr	40.98	5.69	6.95	0.06
KBr	43.02	6.33	7.32	0.09
NaBPh <sub>4</sub>	23.69	6.04	9.20	0.03

**Table III. Limiting Ionic Conductances and Ionic Walden Products in EG and FM at 298.15 K**

Ions	$\lambda_{\pm} \times 10^4$ (S.m <sup>2</sup> .mol <sup>-1</sup> )	$\lambda_{\pm} \eta_0 \times 10^4$ (S.m <sup>2</sup> .mol <sup>-1</sup> .Pa.s)
<b>Electrolyte + EG</b>		
Me <sub>4</sub> N <sup>+</sup>	3.37	0.057
Et <sub>4</sub> N <sup>+</sup>	2.65	0.045
Pr <sub>4</sub> N <sup>+</sup>	2.18	0.037
Bu <sub>4</sub> N <sup>+</sup>	1.51	0.026
Li <sup>+</sup>	2.07	0.035
Na <sup>+</sup>	3.03	0.051
K <sup>+</sup>	4.55	0.077
Br <sup>-</sup>	4.97	0.084
BPh <sub>4</sub> <sup>-</sup>	1.41	0.024
<b>Electrolyte + FM</b>		
Me <sub>4</sub> N <sup>+</sup>	17.87	0.059
Et <sub>4</sub> N <sup>+</sup>	15.11	0.050
Pr <sub>4</sub> N <sup>+</sup>	11.06	0.037
Bu <sub>4</sub> N <sup>+</sup>	7.91	0.026
Li <sup>+</sup>	8.85	0.029
Na <sup>+</sup>	16.29	0.054
K <sup>+</sup>	18.33	0.061
Br <sup>-</sup>	24.69	0.082
BPh <sub>4</sub> <sup>-</sup>	7.40	0.024

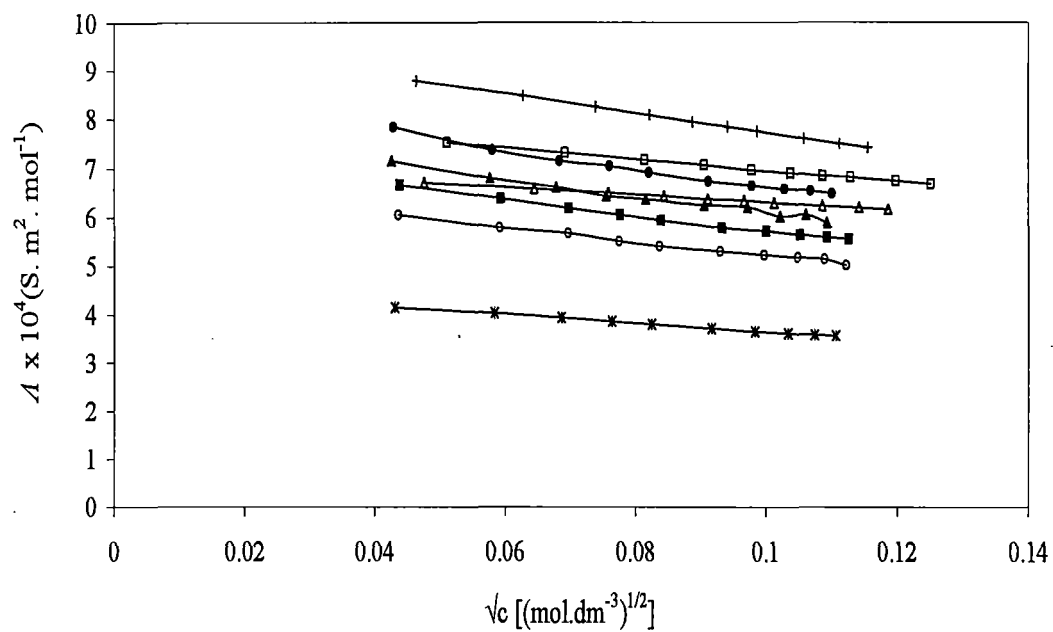


Fig.1. The plots of equivalent conductances  $\Lambda$  versus  $\sqrt{c}$  for different electrolytes in ethane-1,2-diol at 298.15 K. Experimental points: Me<sub>4</sub>NBr (●), Et<sub>4</sub>NBr (▲), Pr<sub>4</sub>NBr (■), Bu<sub>4</sub>NBr (○), LiBr (Δ), NaBr (□), KBr (+), NaBPh<sub>4</sub> (Ж).

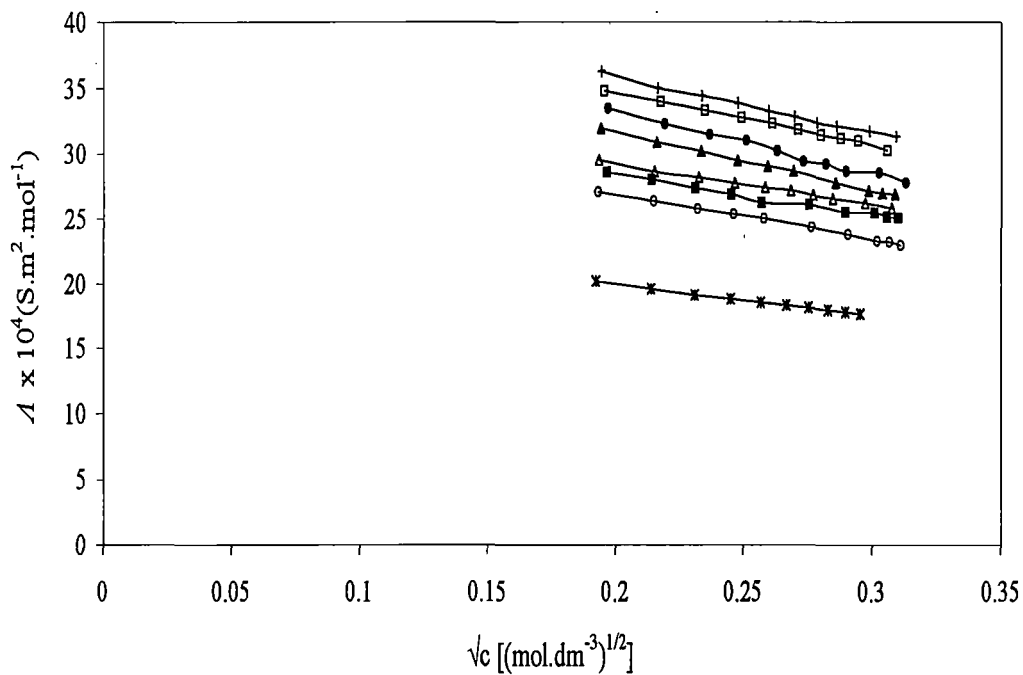


Fig.2. The plots of equivalent conductances  $\Lambda$  versus  $\sqrt{c}$  for different electrolytes in formamide at 298.15 K. Experimental points:  $\text{Me}_4\text{NBr}$  (●),  $\text{Et}_4\text{NBr}$  (▲),  $\text{Pr}_4\text{NBr}$  (■),  $\text{Bu}_4\text{NBr}$  (○),  $\text{LiBr}$  (Δ),  $\text{NaBr}$  (□),  $\text{KBr}$  (+),  $\text{NaBPh}_4$  (X)