Liquid Structure of Aqueous Solutions of Sodium Chloride: Solubility, Vapor Pressure, and Conductivity

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Solubility of NaCl in H₂O and vapor pressure and conductivity of such solutions are related to the structure of the solution. It is concluded from those properties that the NaCl dissolved is completely hydrolyzed into NaOH and HCl. The two molecules are arranged in pairs such that the Cl atom in the HCl molecule is separated by one H₂O molecule from the Na atom and by one H₂O molecule from the O atom in the NaOH molecule. Solubility is determined by the possibility of formation of NaCl · 2 H₂O which is precipitated.

1. INTRODUCTION

In a preceding paper¹ on the solubilities of silver and lead halides in water, the solubility was calculated from the formula (1) valid for ideal solutions and by applying the statistical theory of ideal mixtures to the hydrolysis of the salt.

The fundamental formula is

$$X = e^{-\frac{\Delta E_s}{R} \left(\frac{1}{T} - \frac{1}{T_f} \right)} \quad (1)$$

when X is the mole fraction of the salt dissolved, T_f its melting point, $-\Delta E_s$ the heat of solution. When the dissolved salt is hydrolyzed we have to add a factor W related to the configurational entropy of hydrolysis, and (1) takes the form

$$X = W e^{-\frac{\Delta E_s}{R} \left(\frac{1}{T} - \frac{1}{T_f} \right)} \quad (2)$$

The solubilities of those salts were taken from conductivity measurements. As has been pointed out elsewhere^{2,3} only dissociated solute molecules contribute to the conductivity. Hence, conductivity measurements do not give the total amount of salt dissolved but only the fraction of the salt which is hydro-

lyzed, *e.g.* in the case of AgCl, only the Ag present in the form of AgOH but not that in the form of AgCl.

It follows from these considerations that the slope of the plot of log solubility versus $\frac{1}{T}$ gives a heat of dissolution which is equal to ΔE_s divided by the number of molecules formed by the hydrolysis. Hence, the state of hydrolysis may be determined from that plot and the calorimetric value of ΔE_s . It was found that the states of hydrolysis are independent of the temperature. Thus, PbCl_2 and PbBr_2 give $\text{PbOHX} + \text{HX}$ but no $\text{Pb(OH)}_2 + 2 \text{HX}$, whereas PbI_2 gives $\text{Pb(OH)}_2 + 2 \text{HX}$ but no $\text{PbOHX} + \text{HX}$.

When the two atoms M and X occupy adjacent lattice sites, they form a molecule of MX which does not contribute to the conductivity. When they occupy non-adjacent sites, they are present in the form of MOH and HX, respectively. The number of MX molecules present is determined by (1) or (2) while the numbers of MOH and HX molecules are determined by the statistics of the liquid structure, *viz.* by the numbers of ways in which bonds may be broken and reformed and in which the M and X atoms may be distributed on the lattice sites.

In this communication we will apply similar arguments to aqueous solutions of sodium chloride. It turns out, however, that the structure of the solution is not random but that there is a preferred distribution of Na and Cl atoms relative to each other. In order to discuss this structure we also have to consider other properties of the solution, *viz.* conductivity and vapor pressure. The fundamentals of those properties have been treated elsewhere (conductivity², vapor pressure⁴) and reference will therefore be made to those papers in order to avoid repetition.

Unfortunately, the data available are not sufficient for a complete treatment of the subject. The data missing include the temperature-dependence of the conductivity. Further, the geometry of the structure is not known. According to X-ray determinations by Finbak⁵ the structure of NaOH is similar to that of HCl from which result we may conclude that the coordination numbers are the same in both structures.

2. SOLUBILITY

The solubilities of NaCl in H_2O and in $\text{H}_2\text{O-HCl}$ mixtures depend very little on the temperature as is shown by the data given by various investigators⁶. Hence, we have approximately $\Delta E_s = 0$ in both cases. The solubility is then determined almost entirely by the configurational entropy. Since the observed solubility includes Na both as NaOH and as NaCl and since $\Delta E_s = 0$, we can not conclude upon the degree of hydrolysis from the same arguments as in the case of the Ag and Pb halides.

Table 1 reproduces the data of Fricke and Brümmer^{6,7} on the solubility of NaCl in aqueous HCl solutions at 18°.

Table 1.

HCl mol/100 g solution	NaCl	HCl	NaCl	H ₂ O	$\frac{[\text{H}_2\text{O}] + [\text{NaCl}]}{[\text{HCl}] + [\text{NaCl}]}$
			mole fractions		
0.0	4.513	0.0	9.950	90.050	10.055
0.0438	4.474	0.0960	9.838	90.065	10.057
0.0870	4.433	0.1913	9.720	90.059	10.068
0.1756	4.346	0.3853	9.5365	90.078	10.040
0.3610	4.151	0.7879	9.0593	90.153	10.075
0.9263	3.605	1.9935	7.7583	90.248	10.049
1.6677	2.889	3.5247	6.1060	90.369	10.017
2.6740	2.055	5.5529	4.2675	90.180	(9.618)
3.5670	1.408	7.3209	2.8898	89.789	(9.070)
Average					10.053 $\pm$ 0.005

In the last column we have computed the ratio $\frac{[\text{H}_2\text{O}] + [\text{NaCl}]}{[\text{HCl}] + [\text{NaCl}]}$, the brackets denoting the mole fraction of the substance put into the solution. This ratio is a constant, 10.053 ± 0.005 , say 10, below 3.5 mol. % of HCl.

On the assumption that there are no NaCl molecules present in the solution, all NaCl dissolved being hydrolyzed into NaOH and HCl, we have the following actual concentrations in the solution:

$$\left. \begin{array}{l} \text{H}_2\text{O}: [\text{H}_2\text{O}] - [\text{NaCl}] \\ \text{HCl}: [\text{HCl}] + [\text{NaCl}] \\ \text{NaOH}: [\text{NaCl}] \end{array} \right\} (3)$$

corresponding to a reaction which we may write symbolically

indicating that the Na and Cl atoms are separated from each other by H₂O molecules in the structure of the solution.

From vapor pressure⁴ and conductivity² data we have concluded that when there are more than 10 H₂O molecules per HCl molecule on the average, each HCl molecule is surrounded by 10 neighbors, all of which are H₂O molecules and none of which is a HCl molecule. The NaOH molecule, like the AgOH molecule¹, may be assumed to occupy two lattice sites. We will also assume that a NaOH molecule may replace two H₂O molecules among the neighbors of the HCl molecule¹.

The number of sites occupied by H₂O and NaOH molecules is then according to (3)

$$[\text{H}_2\text{O}] - [\text{NaCl}] + 2 [\text{NaCl}] = [\text{H}_2\text{O}] + [\text{NaCl}]$$

and the number of sites occupied by HCl molecules is

$$[\text{HCl}] + [\text{NaCl}].$$

Hence, when

$$\frac{[\text{H}_2\text{O}] + [\text{NaCl}]}{[\text{HCl}] + [\text{NaCl}]} \geq 10 \quad (5)$$

each HCl molecule is surrounded by 10 H₂O or NaOH molecules and there is no H₂O or NaOH molecule belonging simultaneously to the two groups of 10 neighbors of two different HCl molecules.

The data in Table 1 show that (5) is the condition for solubility. Putting the ratio (5) equal to 10 we obtain for the solubility of NaCl in aqueous solutions of HCl

$$[\text{NaCl}] = \frac{1}{9} \{[\text{H}_2\text{O}] - 10 [\text{HCl}]\} \quad (6)$$

The form of (6) suggests the simple interpretation that the HCl molecules added to H₂O associate with themselves 10 H₂O molecules each and that the NaCl dissolves according to (4) in whatever water is left. This changes nothing in the previous argument, if the HCl molecules directly added are undistinguishable from those created by (4).

According to (6), no NaCl at all should dissolve in aqueous solutions of less than 10 H₂O per HCl. However, there is some slight solubility even at very high concentrations as shown by the data of Seidel and Fischer ^{6,8} at 0°. Those data are reproduced in Table 2.

Table 2.

HCl percent	NaCl by weight	HCl	NaCl mole fractions	H ₂ O	$\frac{[\text{H}_2\text{O}] + [\text{NaCl}]}{[\text{HCl}] + [\text{NaCl}]}$
0	26.2	0	9.86	90.14	10.14
22.6	1.15	12.72	0.40	86.88	6.65
26.7	0.41	15.30	0.15	84.55	5.48
39.0	0.047	24.01	0.018	75.97	3.16
43.9	0.034	27.89	0.014	72.10	2.58
44.5	0.032	28.38	0.012	71.61	2.52

Data solubilities of NaCl in H₂O-NaOH mixtures given by von Antropoff ^{6,9} show a very slight dependence upon temperature, i.e. approximately $\Delta E_s = 0$.

Table 3 gives the solubility at 25° according to Åkerlöf and Short ^{6,10}.

Table 3.

NaOH mol/1 000 g	NaCl H ₂ O	H ₂ O	$\frac{[\text{H}_2\text{O}] - 5 [\text{NaOH}] + [\text{NaCl}]}{[\text{NaCl}]}$	$\frac{[\text{H}_2\text{O}] - 9 [\text{NaCl}]}{[\text{NaOH}]}$
0.920	5.613	55.506	10.07	5.42
1.858	5.048		10.21	5.43
2.714	4.672		9.97	4.96
3.070	44.86		9.95	4.93
3.812	4.048		10.00	5.01
4.798	3.610		9.73	4.80
5.721	3.206		9.39	4.66
6.688	2.635		9.37	4.75
7.141	2.442		8.86	4.70
8.031	2.173		8.06	4.48
8.876	1.956		6.69	4.27
9.824	1.558		5.10	4.22
11.437	1.280		-0.28	3.85
11.582	1.253		-0.9	3.82

In Table 3 we have also given the ratios

$$\frac{[\text{H}_2\text{O}] - 5 [\text{NaOH}] + [\text{NaCl}]}{[\text{NaCl}]} \quad (7)$$

and

$$\frac{[\text{H}_2\text{O}] - 9 [\text{NaCl}]}{[\text{NaOH}]} \quad (8)$$

The ratio (7) should be 10, just as in the case of H_2O -HCl solutions, the ratio (8) should be 5, *viz.* the number of H_2O molecules bound to one NaOH molecule. At low concentrations, below 4 *N* NaOH, those values are actually obtained. At higher concentrations, they are lower, just as was found in H_2O -HCl mixtures. This phenomenon will be discussed presently.

We find, thus, that the same condition of solubility is valid whether we add HCl or NaOH to the solution.

For the hydrolysis of AgX we have found the relation ¹

$$c_{\text{AgOH}} \cdot c_{\text{HX}} = 10 c_{\text{AgX}} \quad (9)$$

the factor 10 being derived from the probability that an Ag atom and a X atom occupy adjacent lattice sites. This is the equilibrium equation of hydrolysis, the factor 10 being the equilibrium constant. However, we may also look upon (9) as an expression for the solubility product, if

$$c_{\text{AgX}} = \frac{55.55}{1215} e^{-\frac{\Delta E_s}{R} \left(\frac{1}{T} - \frac{1}{T_f} \right)} \quad (10)$$

is inserted for c_{AgX} in (9). We would therefore expect the solubility of AgX, *i.e.* c_{AgOH} , to decrease when HCl is added to the solution. Experimental data show, on the contrary, that the solubility of AgX increases steadily with the concentration of HCl. This deviation from the rule of the constant solubility product is accompanied by a marked reduction of ΔE_s with increasing HCl concentration as shown by the data of Dede and Walther ^{6,11}.

In the case of NaCl solutions we find that the solubility of NaCl decreases when HCl or NaOH is added to the solution. However, the rule of the constant solubility product is still not applicable, although $\Delta E_s = 0$ is unaffected by those additions, because the hydrolysis does not follow the same statistics, owing to a preferred distribution of NaOH and HCl molecules relative to each other (section 4).

In order that the solubility product be constant, two conditions have to be fulfilled:

1. A random distribution of all molecules on the lattice sites.
2. A constant value of ΔE_s , independent of the concentrations of the various species in the solution.

Both conditions exclude the formation of soluble complex salts, since this corresponds to a preferred distribution and because the formation of the complex is connected with a heat of formation.

Similarly, deviations from any law of solubility which does not take into account changes in distribution and in ΔE_s , may be due to those causes. Thus, according to von Antropoff^{6,9}, the solubility of NaCl in H_2O —NaOH mixtures increases by 0.5 % between 20° and 30° at 5 % NaOH and by 25 % at 50 % NaOH by weight. The increase of the solubility of AgCN with the concentration of added HCN (Randall and Halford^{6,12}) is attributed to formation of the acid $HAg(CN)_2$ with the complex ion $Ag(CN)_2^-$. The same applies to AgCl.

Corresponding complexes of NaCl with HCl and NaOH do not seem to be known. It is therefore difficult to explain why more NaCl than given by the rule just discussed is dissolved at high concentrations of NaOH or HCl. However, the essence of the structure of NaCl in H_2O and its relation to solubility is covered by the agreement at low concentrations of NaOH and HCl.

3. VAPOR PRESSURE

The partial vapor pressure of water above aqueous solutions of NaCl has been measured by a number of investigators¹³. The data show that the relative partial vapor pressure p , i. e. the actual partial vapor pressure above the solution divided by that above pure water of the same temperature, is independent of the temperature. Hence, the differential heat of vaporization is $\Delta E_{H_2O} = 0$.

For aqueous solutions of HCl we have found⁴ that $\Delta E_{H_2O} = 0$ for $m > 10$. This condition is fulfilled for all NaCl solutions, if we in the number of H_2O molecules m include twice the number of NaOH molecules as discussed in section 2, because this total number of H_2O and NaOH neighbors of the HCl molecule is always > 10 . For NaOH solutions⁴ $\Delta E_{H_2O} = 0$ for $m > 5$.

Fig. 1 shows $p = f\left(\frac{1}{m+1}\right)$ for NaCl solution at 25° on the data of Olvyk and Gordon¹⁴ and Robinson¹⁵. At low concentrations we find

$$p = 1 - \frac{2}{m+1} = \frac{m-1}{m+1}, \quad m > 24 \quad (11)$$

Fig. 2 shows the difference

$$\Delta p = \frac{m-1}{m+1} - p \quad (12)$$

For $m > 24$, $\Delta p = 0$.

For HCl solutions we have found⁴

$$p = 1 - \frac{3}{m+1} = \frac{m-2}{m+1}, \quad m > 26 \quad (13)$$

When mixing one NaCl with m H_2O we obtain in the solution as discussed in section 2, one NaOH, one HCl, and $(m-1)$ H_2O . Thus, the concentration of H_2O is $\frac{m-1}{m+1}$, and (10) is therefore identical with Raoult's law.

An analysis of Δp for NaCl solutions is difficult owing to the small range of $\Delta p \neq 0$. We will observe, however, that the sudden increase of Δp at $m = 24$ for NaCl solutions corresponds to that at $m = 26$ for HCl solutions.

Fig. 1. $p = f\left(\frac{1}{m+1}\right)$ for NaCl solutions.
+ Olnyk & Gordon, O Robinson.

Fig. 2. $\Delta p = f\left(\frac{1}{m+1}\right)$ for NaCl solutions
on the data of Olnyk & Gordon (+) and
Robinson (O).

The interpretation given for the fact that $\Delta E_{H_2O} = 0$ for $m > 10$ in the case of HCl solutions is that 10 H_2O molecules are bound to each HCl molecule and that, thus, when $m > 10$ there are H_2O molecules which are bound to other H_2O molecules only⁴. If $m = 28$ is substituted for $m = 10$ the same applies to H_2SO_4 solutions⁴. Applying this interpretation to NaCl solutions we must conclude that there are H_2O molecules not bound to a HCl molecule as long as the sum of the number of H_2O molecules and twice that of NaOH molecules is greater than 10. Hence, two of the 10 sites around the HCl molecule are preferentially occupied by NaOH molecules, *i. e.*, since there is one NaOH molecule per HCl molecule, two of the 10 sites are always occupied by an NaOH molecule independently of the number of H_2O molecules per HCl molecule.

4. CONDUCTIVITY

The data employed are listed in Table 4.

Table 4.

Solute	Experimental data by	Temp.	Fig.	α	ΔS_0
NaCl	Kohlrausch & Grotrian ^{13,16}	18°	3	7.64	1.52
KOH	Kohlrausch ^{13,17}	15°	4	7.53	1.49
NaOH	Kohlrausch ^{13,17}	18°	5	11.59	2.30
	Bousfield & Lowry ^{13,16}	18°	5	11.59	2.30

Fig. 3. $\ln \Lambda = f\left(\frac{1}{m+1}\right)$ for NaCl solutions at 18°.

Fig. 4. $\ln \Lambda = f\left(\frac{1}{m+1}\right)$ for KOH solutions at 15°.

Figs. 3—5 show the plots of $\ln \Lambda = f\left(\frac{1}{m+1}\right)$ when Λ is the equivalent conductance and m is the number of water molecules per molecule of solute. At low concentrations those plots are of the form ²

$$\ln \Lambda = A - \frac{x}{m+1}. \quad (14)$$

The values of x obtained from the curves are given in Table 3. From x we have computed ΔS_0 in the formula for the entropy of activation ²

$$\Delta S = \frac{10}{m+1} \Delta S_0. \quad (15)$$

The data given by Darken and Meier ¹⁹ can not be represented by (14), because they fall much quicker at low concentrations, giving a non-linear curve for $\ln \Lambda = f\left(\frac{1}{m+1}\right)$ at low concentrations. The same effect was found in some data on acid solutions ².

The coordination numbers may be assumed to be the same in the aqueous solutions of NaOH, KOH, and HCl ⁵.

For NaCl the formula (14) is valid in the entire range of solubility. For NaOH and KOH the validity is less extended. For HCl it is valid in the entire range of complete dissociation, $m > 5$. If we may call ΔS_0 for NaCl and KOH normal, ΔS_0 for NaOH is abnormal. Without further data, particularly on the temperature-dependence of Λ , it is difficult to account for this anomaly. However, we may conclude that the presence of HCl restores normality to NaOH.

Fig. 5. $\ln \Lambda = f\left(\frac{1}{m+1}\right)$ for NaOH solutions at 18°. ○ Kohlrausch, + Bousfield & Lowry.

The HCl molecule is surrounded by 5 nearest neighbors. For HCl we have derived ²

$$\Delta S_0 = R \ln \left(\frac{4}{5} \cdot \frac{1}{5} \right)^{1/2} = -1.90 \text{ cal./mole}^0. \quad (16)$$

The KOH molecule conducts with its OH radical only just as an acid molecule conducts with its proton only ². The O atom has 4 nearest neighbors one of which is the K atom. Hence,

$$\Delta S_0 = R \ln \left(\frac{2}{3} \cdot \frac{1}{3} \right)^{1/2} = -1.49 \text{ cal./mole}^0 \quad (17)$$

in agreement with the empirical values for KOH and NaCl.

From these considerations we may conclude that the OH group is bound in the same manner in the two cases to which (17) applies. Hence, we must conclude that the NaOH molecule is not bound to a HCl molecule with its OH group adjacent to the Cl atom in the range where (17) is valid, *i. e.* below and at saturation.

In the structure of HCl in H₂O there are two coordination numbers, 5 and 10. The HCl molecule can form hydrogen bonds with 5 H₂O molecules directly and bind another 5 H₂O molecules indirectly. The number 5 has been found in the following properties: conductivity ², vapor pressure ⁴, dissociation ²⁰, all indicating pairs of HCl molecules when $m < 5$ and no such pairs when $m \geq 5$. The number 10 has been found in the heats of vaporization of H₂O and HCl ⁴ and in the solubilities of halides ¹ (and section 2). We must conclude that the HCl molecule is surrounded by two shells of H₂O molecules, each consisting of 5 H₂O molecules. For H₂O molecules within the outer shell, the binding energy is greater than that of H₂O molecules outside the outer shell, $\Delta E_{\text{H}_2\text{O}}$ being zero for $m > 10$ and increasing rapidly with decreasing m in the range $m < 10$ ⁴. Since $\Delta E_{\text{H}_2\text{O}} = 0$ for NaCl solutions too, we must conclude

that there are H_2O molecules outside the outer shell in the entire range of solubility. This implies that, at least when $m < 10$, the NaOH molecules preferentially populate the outer shell around the HCl molecule. Assuming that this preference remains at all concentrations, we are led to assume that the NaOH molecule occupies two sites in the outer shell, *i. e.* the Na atom and the O atom are both located in the outer shell, and occupies thus a fixed position and a fixed orientation relative to the HCl molecule.

This discussion on the structure of NaCl solutions may explain why the presence of HCl molecules affects the value of ΔS_0 of the NaOH molecule, although it does not explain why it is rendered normal. Nor, of course, does it explain the anomaly of ΔS_0 in absence of HCl molecules.

5. CONCLUSIONS

The solubility data show that pairs of Na and Cl atoms, *i. e.* NaCl molecules, are not present in the solution, the NaCl dissolved being completely hydrolyzed into NaOH and HCl . This result is supported by vapor pressure data. The heat of vaporization of H_2O shows a preferential arrangement of the NaOH and HCl molecules, probably due to attraction between dipoles. From the activation entropy of conductivity it follows further, that not only the Na but also the O atom in the NaOH molecule is separated from the Cl atom by a H_2O molecule. The structure contains, thus, pairs of one NaOH molecule and one HCl molecule in a fixed position and a fixed orientation relative to one another, the two molecules in the pair being separated from each other by two H_2O molecules, one between Cl and Na , the other between Cl and O .

When two such pairs are separated by water they can not combine. But if the water molecules are not many enough, such pairs may combine to form a complex which we may write in the form $2(\text{HCl} \cdot 2 \text{H}_2\text{O} \cdot \text{NaOH})$. This complex may form when the ratio (5) is equal to 10. When it is less than 10, one H_2O molecule may be missing in the complex which may then precipitate as two molecules of the di-hydrate, $\text{NaCl} \cdot 2 \text{H}_2\text{O}$. Thus, the condition of solubility (4) implies that the precipitate is $\text{NaCl} \cdot 2 \text{H}_2\text{O}$ which is also actually found at low temperatures. Above 0.1° the precipitate contains $\text{NaCl} \cdot 2 \text{H}_2\text{O} + \text{NaCl}$, above 20° it contains NaCl only²¹. The two-phase precipitate $\text{NaCl} \cdot 2 \text{H}_2\text{O} + \text{NaCl}$ seems to indicate that the process

is sluggish in the temperature interval 0.1 to 20° , and that equilibrium was not attained in the experiments quoted.

At equilibrium the free energy of solution is

$$\Delta F_s = \Delta E_s - T \Delta S_s = 0.$$

In this case ΔE_s and ΔS_s are both zero, and the entropy, *i. e.* the configuration, and the energy are both the same in the solution and in the precipitate.

Both structures consist, thus, at low temperatures, of aggregates of 2 Na atoms, 2 Cl atoms, and 4 H₂O molecules held together by hydrogen bonds. At higher temperatures, the precipitate gives off its crystal water.

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REFERENCES

1. Owe Berg, T. G. *Acta Chem. Scand.* **7** (1953) 1313.
2. Owe Berg, T. G. *J. chim. phys.* **50** (1953) 247.
3. Owe Berg, T. G. *J. chim. phys.* *In press.*
4. Owe Berg, T. G. *Acta Chem. Scand.* **7** (1953) 1045.
5. Finbak, Chr. *Tidskr. Kjemi, Bergvesen, Met.* **3** (1943) 40.
Finbak, Chr., and Viervoll, H. *Fra Fysikens Verden I* (1945) 153.
6. Seidell, A. *Solubilities, I and III*, New York 1953.
7. Fricke, R., and Brümmer, F. *Z. anorg. Chem.* **223** (1935) 397.
8. Seidel, W., and Fischer, W. *Z. anorg. Chem.* **247** (1941) 367.
9. von Antropoff, A. *Z. Elektrochem.* **30** (1924) 457.
10. Åkerlöf, G., and Short, O. *J. Am. Chem. Soc.* **59** (1937) 1912.
11. Dede, L. and Walther, Th. *Z. anorg. Chem.* **163** (1927) 185.
12. Randall, M., and Halford, J. O. *J. Am. Chem. Soc.* **52** (1930) 178.
13. Landolt-Börnstein, *Physik.-Chem. Tabellen*, Berlin 1926–1936.
14. Olnyk, P., and Gordon, A. R. *J. Am. Chem. Soc.* **65** (1943) 224.
15. Robinson, R. A. *Trans. Roy. Soc. New Zealand* **75** (1945/46) 203.
16. Kohlrausch, F., and Grotrian, *Poggendorfs Ann. Physik* **154** (1875) 1.
17. Kohlrausch, F. *Wiedemanns Ann. Physik* **6** (1879) 1.
18. Bousefield and Lowry, *Phil. Trans.* **204** (1904) 253.
19. Darken, L. S., and Meier, H. F. *J. Am. Chem. Soc.* **64** (1942) 621.
20. Ochs, L., Guéron, J., and Magat, M. *J. phys. radium* **1** (1940) 85.
21. Seidell, A. *Solubilities I*, New York 1953, p. 1218.

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