

The Influence of the Solvent on Reaction Velocity

XXX. The Reaction between Benzyl Chlorides and Hydroxide Ion in Acetone-Water and Dimethyl Sulphoxide-Water Mixtures

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The rate constants for the reactions of hydroxide ion with benzyl chloride and *p*-methyl-, *p*-chloro-, and *m*-chlorobenzyl chlorides in acetone-water and dimethyl sulphoxide-water mixtures are reported as a function of both solvent composition and temperature. The reaction rate decreases with increasing acetone concentration, but increases with increasing dimethyl sulphoxide concentration. Solvent effects are discussed. Substituent effects are dependent both on the composition of the solvent and on the temperature.

Much information has recently been obtained on the solvolysis of benzyl chlorides in various solvents. In contrast, the data on the displacements



are scanty. Olivier¹ determined the bimolecular rate constants of the reaction for benzyl chloride and its methyl, bromo and cyano derivatives in 50 vol. % aqueous acetone at 30°. Beste and Hammett² measured the rate constant for benzyl chloride in 60.72 wt. % aqueous dioxan at 50°. Some values for the reaction in acetone-water mixtures are given in a previous paper of this series.³ The purpose of the present work was to investigate the effects of substituents and changes in solvent composition on the rate of the reaction in water-organic solvent mixtures. The organic solvents chosen were acetone and dimethyl sulphoxide (DMSO). A study of the reactions in these two solvent systems was attractive because the high rates of nucleophilic attack by certain anions in dipolar aprotic solvents have awakened considerable interest in recent years.⁴ For example, the reactions of methyl iodide with hydroxide, methoxide, and ethoxide ions,⁵ the reaction of fluoronitrobenzenes with hydroxide ion,⁶ and alkaline ester hydrolysis⁷⁻¹¹ are greatly accelerated by dimethyl sulphoxide. Acetone generally retards ester hydrolysis, but a rate increase, which, however, is only weak, is observed at high acetone concentrations.⁹ No corresponding effect has been observed in dioxan-water mixtures.⁹

The compounds investigated were benzyl chloride and its *p*-methyl, *p*-chloro, and *m*-chloro derivatives.

EXPERIMENTAL

The organic chlorides were the same substances as were used in the preceding work¹² of this series. Each of them was carefully redistilled at reduced pressure before use.

Baker's Analyzed acetone was used as received. Dimethyl sulphoxide (National Petrole France) was distilled at reduced pressure and crystallized repeatedly from the melt until its m.p. was 18.50–18.55°. The water used in preparing the solvent mixtures for the kinetic experiments was freshly redistilled, carbon dioxide-free water. The solvents for the organic chlorides were prepared by diluting a measured volume of water with DMSO or acetone to a definite volume in a volumetric flask. The sodium hydroxide solutions were prepared in a similar manner from a stock base solution and DMSO or acetone.

The kinetic experiments were carried out in sealed 8-ml or 12-ml ampoules of Pyrex glass into which the benzyl chloride and sodium hydroxide solutions were pipetted so that the total volume of the reaction mixture was 6 ml or 10 ml. The ampoules were immersed in electrically controlled thermostats, the temperatures of which remained constant within $\pm 0.02^\circ$. Ampoules withdrawn after suitable time intervals were chilled rapidly in ice-water and opened, and their contents were added to an excess of 0.02 N hydrochloric acid. The excess of hydrochloric acid was titrated with 0.02 N sodium hydroxide using cresol red as indicator. When precipitation occurred after mixing with the acid, some acetone was added to restore homogeneity. When the reaction was rapid, the experiments were carried out in two-compartment reaction vessels.¹³ Five-ml micro-burettes were used in the titrations. The endpoint was always sharp. The number of separate titrations in each run varied from six to eight.

Correction for solvolysis was made using the equation

$$dx/dt = k_2(a-x)^2 + k_1(a-x)$$

which gives^{14,15}

$$k_2 = \frac{k_1\alpha}{a(1-\alpha)}$$

where

$$\alpha = \frac{[\exp(-k_1t)] + [(x/a) - 1]}{(x/a) + \exp(-k_1t)}$$

The values of the first-order rate constants k_1 of the solvolysis were taken from previous papers.^{3,12} The equation always gave values of k_2 of good constancy. The values of k_1 for the chlorobenzyl chlorides are very small compared to the values of k_2 , and hence the latter could be computed from the simple second-order equation. The solvolysis was so slow in solvents containing 20 % or less of water that values of k_2 of good constancy were computed from the simple second-order equation, also for benzyl chloride.

Since the rate constants are greatly dependent on the initial concentrations,² the reactions were carried out as far as possible using equal initial concentrations. Thus the obtained rate constants are comparable with each other.

The second-order rate constants were corrected for the thermal expansion of the solvent, using for DMSO-water mixtures expansion coefficients given in the paper of Tommila and Murto.⁷ The expansion coefficients for acetone-water mixtures were derived from data given in Landolt-Börnstein's Handbook. The parameters of the Arrhenius equation, A and E , were calculated from the corrected k_2 values by the method of least squares. The Arrhenius equation was obeyed in all cases within the limits of experimental error. The activation entropies were calculated from the equation

$$A = e(kT/h)\exp(\Delta S^*/R)$$

and the quantities ΔG^* and ΔH^* from the equations $\Delta G^* = \Delta H^* - T\Delta S^*$ and $\Delta H^* = E - RT$.

RESULTS

The principal experimental data are summarized in Tables 1–4. The frequency factor A is expressed in litres mole⁻¹ sec⁻¹ and the activation

Table 1. The reaction $C_6H_5CH_2Cl + OH^- \rightarrow C_6H_5CH_2OH + Cl^-$ in DMSO-water mixtures. The initial concentrations, a , of benzyl chloride and NaOH were 0.02 mole/litre.

Water in the solvent			$10^4 k, \text{l mole}^{-1} \text{sec}^{-1}$					E cal	log A	ΔS^* (40°)	ΔG^* cal
ml/l	wt. %	x_w	20.00°	25.00°	40.00°	50.00°	60.00°				
600	56.6	0.849		1.87	10.0	26.8	70.8	20 440	11.264	- 9.1	22 670
500	46.4	0.790		1.94	10.4	28.4	72.8	20 440	11.280	- 9.0	22 650
400	36.7	0.715		3.05	14.0	40.6	96.4	19 610	10.860	-10.9	22 400
300	27.3	0.620	3.56	5.68	27.0	72.8		19 030	10.730	-11.5	22 010
200	18.1	0.490	12.7	22.6	86.7	210		17 390	10.080	-14.5	21 310
150	13.6	0.400	28.5	46.7	178	410		16 690	9.899	-15.3	20 870
100	9.1	0.302	91.1	141	492	1080		15 520	9.528	-17.0	20 230

^a $a = 0.011$.

Table 2. The reaction $XC_6H_4CH_2Cl + OH^- \rightarrow XC_6H_4CH_2OH + Cl^-$ in DMSO-water mixtures. $a = 0.02$ mole/litre.

X	H_2O ml/l	$10^4 k, \text{l mole}^{-1} \text{sec}^{-1}$				E cal	log A	ΔS^* (40°)	ΔG^* cal
		20.00°	25.00°	40.00°	50.00°				
$p\text{-CH}_3$	400				76.1				
»	300	6.67	11.1	44.2	111	17 540	9.900	-15.3	23 720
»	200				241				
$p\text{-Cl}$	400				76.1				
»	300	8.83	14.5	59.4	137	17 250	9.806	-15.8	21 560
»	200				373				
$m\text{-Cl}$	400				56.6				
»	300	6.67	11.4	46.7	113	17 620	9.997	-14.9	21 710
»	200				258				

Table 3. The reaction $C_6H_5CH_2Cl + OH^- \rightarrow C_6H_5CH_2OH + Cl^-$ in acetone-water mixtures. $a = 0.02$ mole/litre.

Water in the solvent			$10^5 k, \text{l mole}^{-1} \text{sec}^{-1}$				E cal	log A	ΔS^* (40°)	ΔG^* cal
ml/l	wt.-%	x_w	25.00°	40.00°	50.00°	60.00°				
600	63.6	0.849	3.54	16.5	43.6	108	19 300	9.694	-16.3	23 770
500	54.0	0.791	2.84	13.0	33.2	78.6	18 700	9.162	-18.7	23 930
400	44.2	0.719	2.21	10.0	25.3	62.7	18 830	9.146	-18.8	24 090

Table 4. The reaction $\text{XC}_6\text{H}_4\text{CH}_2\text{Cl} + \text{OH}^- \rightarrow \text{XC}_6\text{H}_4\text{CH}_2\text{OH} + \text{Cl}^-$ in acetone-water mixtures. $a = 0.02$ mole/litre. Temperature 50.00° .

Water in the solvent, ml/l	600	500	400
<i>p</i> -CH ₃ , $10^5 k$ l mole ⁻¹ sec ⁻¹	167 ^a	61.2	34.1
<i>p</i> -Cl, » »	44.1	38.2	30.0
<i>m</i> -Cl, » »	46.0	30.0	24.9

^a $a = 0.01$.

entropy ΔS^* relates to 40° , the approximate mean of the temperatures used.

The main results may be summarized as follows:

(1) Addition of DMSO to the solvent causes an increase in the reaction velocity. If there is any increase at low DMSO concentrations, the increase is very small. The rise is very steep when the mole fraction of DMSO exceeds 0.25 (Fig. 1).

(2) In the range of mixtures investigated, acetone retards the reaction.

(3) The increase in reaction velocity caused by DMSO is greater than in alkaline ester hydrolysis.^{7,9}

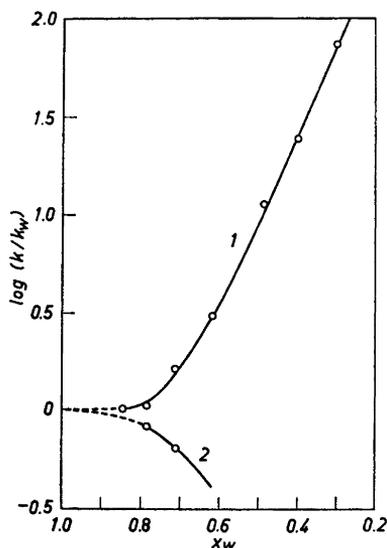


Fig. 1. Benzyl chloride, plot of $\log(k/k_w)$ against x_w , 25° . (1) DMSO-water, (2) acetone-water.

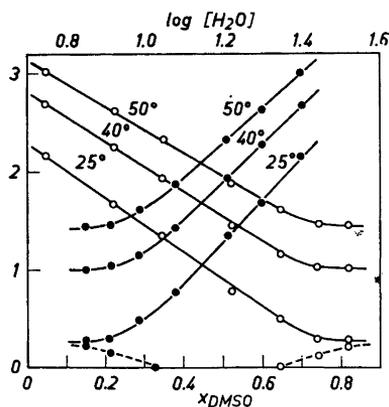


Fig. 2. Benzyl chloride, DMSO-water. Plots of $\log k$ against x_{DMSO} (●) and $\log [\text{H}_2\text{O}]$ (○). The slope of the linear parts decreases with increase in temperature. (---) plots in acetone-water.

Table 5. Benzyl chloride. The relative rates, k/k_w .

H ₂ O in the solvent ml/l	600	500	400	300	200	150	100
DMSO-water:							
25°	1	1.04	1.63	3.01	11.6	24.6	75.1
40°	1	1.04	1.51	2.79	8.78	18.1	50.5
50°	1	1.03	1.44	2.70	7.41	15.1	36.5
60°	1	1.03	1.38				
Acetone-water							
25°	1	0.818	0.627				
40°	1	0.778	0.605				
50°	1	0.753	0.588				
60°	1	0.731	0.573				

(4) When the mole fraction of DMSO is greater than 0.3, $\log k$ is a linear function of the mole fraction of DMSO, and the plot of $\log k$ against the logarithm of the molar water concentration is a straight line of negative slope (Fig. 2).

(5) In aqueous DMSO the solvent effect decreases with increase in temperature; in aqueous acetone the opposite is true (Table 5).

(6) An electron-repelling substituent (CH₃) as well as an electron-attracting substituent (Cl) in the *para*-position increases the reaction rate. Chlorine in the *meta*-position accelerates the reaction in aqueous DMSO, but retards it in aqueous acetone (Tables 6 and 7). These observations are in agreement with the results of Olivier, who, as already mentioned, measured the bimolecular rate constants of reaction (1) for benzyl chloride and its methyl, bromo, and cyano derivatives in 50 vol. % aqueous acetone at 30° and found that the substituents accelerate the reaction except methyl and bromine in *meta*-position that retard it. It may be noted that benzyl halides and nitrates usually show a rate minimum for the unsubstituted compound in second-order substitution reactions with anions.¹⁶

(7) The value of k_x/k_H , the ratio of the rate constant of a substituted compound to that of the unsubstituted benzyl chloride, increases with water content of the solvent (Tables 6 and 7). In acetone-water mixtures, however,

Table 6. The values of k_x/k_H . DMSO-water.

X	<i>p</i> -CH ₃			<i>p</i> -Cl			<i>m</i> -Cl			
	H ₂ O, ml/l	400	300	200	400	300	200	400	300	200
20°			2.06			2.74			2.08	
25°			1.95			2.57			1.97	
40°			1.66			2.13			1.69	
50°	1.93	1.49	1.15	1.93	1.90	1.84	1.43	1.53	1.27	

Table 7. The values k_X/k_H . Acetone-water, 50°.

H ₂ O, ml/l	600	500	400
<i>p</i> -CH ₃	3.83	1.84	1.34
<i>p</i> -Cl	1.01	1.15	1.19
<i>m</i> -Cl		0.91	0.93

the opposite is true for *m*- and *p*-chlorobenzyl chlorides. The substituent effects become weaker with increasing temperature.

(8) All the substituents studied lower the activation energy and frequency factor, at least in aqueous DMSO.

(9) Addition of DMSO or acetone to water lowers the activation energy and frequency factor. These changes in E and A have opposite influences on the reaction velocity. In acetone-water mixtures the decrease in E is more than offset by the decrease in A , and so the addition of acetone decelerates the reaction. In DMSO-water mixtures the decrease in E governs the rate, which thus increases.

(10) The activation energy and frequency factor for benzyl chlorides are lower in aqueous acetone than in aqueous DMSO of the same water content. (The mole fractions are almost identical for aqueous acetone and aqueous DMSO of equal water concentration, so that the data of Tables 1 and 3 refer also to identical mole fractions).

DISCUSSION

In aqueous acetone the activation energy is considerably smaller for reaction (1) than for solvolysis, as is seen by comparing the values of Table 3 with those found for the solvolysis, 19 390, 20 520, and 20 880 calories for the mixtures containing 600, 500, or 400 ml of water per litre, respectively.³ The greater velocity of the reactions with hydroxide ion follows from the lower activation energy, as is quite natural if the reaction rate varies as the repulsion between the seat of reaction and the attacking reagent. If the reactant is a water molecule instead of a negative ion, the repulsion should be greater, since the ion is more strongly attracted to the carbon atom at the seat of reaction than a neutral water molecule.

It was rather unexpected that, although the reaction is more than ten times faster in DMSO-water mixtures than in acetone-water mixtures, the activation energy is larger in the former. No less surprising is that activation energy of the hydroxide ion reaction (20 450 cal) is about the same or slightly higher than for the solvolysis (19 900 cal) in mixtures containing 500–600 ml of water per litre.¹² Only when the water concentration is less than 500 ml/l is the activation energy lower for the hydroxide ion reaction than for the solvolysis and decreases rapidly with the water concentration, whereas the changes in the activation energy of the solvolysis are small.

How can this variation of the activation energy be explained? The higher rates in dipolar aprotic solvents have in general been ascribed to an increased activity of negative ions resulting from their reduced solvation.^{4,5,7} Anions apparently are poorly solvated in pure dipolar aprotic solvents,^{4,17} and it is not impossible that the situation is similar also in the presence of water if its concentration is low. A reduction in the solvation of the initial state would thus result in a decrease in the energy of activation. Table 1 shows, in fact, that the activation energy decreases rapidly with the water concentration and soon becomes smaller than the activation energy of solvolysis in mixtures in which practically all the water molecules are bound to DMSO (to 2-1 or 1-1 complexes), *i.e.* when the mole fraction of DMSO is greater than 0.3. The reaction rate increases rapidly at higher mole fractions of DMSO. A reduction of the solvation of the initial state should also make the activation entropy more negative, as indeed is the case, as can be seen from Table 1.

As to the results obtained for the reactions in aqueous DMSO and aqueous acetone, the following may be noted. The solvation of hydroxide ions (due to hydrogen bonding with water molecules) is not greatly reduced before the DMSO concentration becomes high, as pointed out above, and the hydroxide ions are not much less solvated at high water concentrations in aqueous DMSO than in aqueous acetone. On the other hand, the solvation of benzyl chloride can not be much different in the two series of mixtures, and thus the total solvation of the initial state ($\text{OH}^- + \text{benzyl chloride}$) is on approximately the same level at water concentration from 400 to 600 ml/l in the two solvent series. Consequently the differences in the activation enthalpies and entropies must in the first instance be ascribed to differences in the solvation of the transition state in the two series of solvents. The negative charge originally on the hydroxide ion is shared by OH and Cl in this state and the charge density is accordingly reduced. Thus the interaction of the transition state with water is weaker than that between the hydroxide ion and water. In acetone-water mixtures this interaction is fairly constant down to low water contents. In DMSO-water, already at water concentrations from 600 to 400 ml/l, most of the water molecules are united with DMSO molecules by hydrogen bonds and are not available for solvation. Further, acetone molecules have a stronger specific attraction than DMSO molecules for the hydrophobic part of the transition state and also the planar acetone molecules fit better around the transition state than the pyramidal¹⁸ DMSO molecules. As a result, the transition state is more solvated in acetone-water than in DMSO-water mixtures at water concentrations from 300 to 600 ml/l, and this leads to a larger activation energy and a more positive activation entropy for the reaction in aqueous DMSO than in aqueous acetone.

It has been found that when there occurs dispersal of charge in the transition state, the activation energy is a maximum at a certain solvent composition. Examples are the reaction of hydroxide ion with ethyl bromide in acetone-water mixtures,¹⁵ the sulphonium salt solvolysis in ethanol-water mixtures,¹⁹ and the alkaline hydrolysis of 2,4-dinitrofluorobenzene in DMSO-water mixtures.⁶ Unfortunately, the present investigation was not extended to solvents of high water content. However, we can conclude on the basis of Table 1 that the activation energy of the reaction of benzyl chloride with

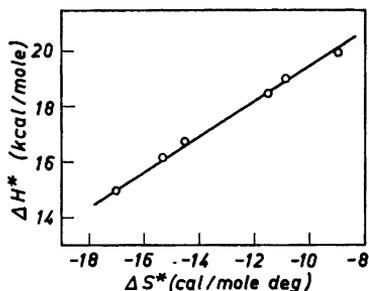


Fig. 3. Plot of ΔH^* against ΔS^* .

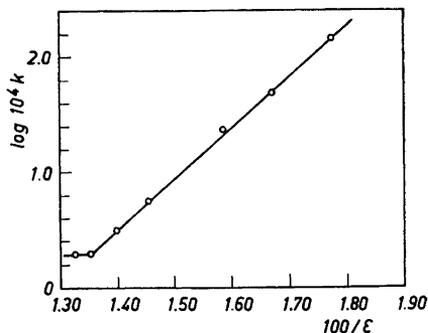


Fig. 4. Plot of $\log k$ against the reciprocal of the dielectric constant. The values of the dielectric constants were taken from Ref. 22.

hydroxide ion very likely shows a maximum when the water concentration is about 600 ml/l. (As is well known, E and A usually pass through a minimum as the composition of the solvent is continuously changed).

It has been often found that linear relationships exist between the enthalpy and entropy of activation. Fig. 3 shows that this is true also for the reaction of benzyl chloride with hydroxide ion.

The well-known theory of solvent polarity effects as formulated by Ingold²⁰ states that reactions in which the charge is dispersed undergo a small rate decrease with increase in solvent polarity. According to Laidler,²¹ a plot of $\log k$ against the reciprocal of the dielectric constant of a mixed solvent should be a straight line of positive slope for a reaction between an ion and a dipole. Contrary to these predictions, the rate decreases with decreasing dielectric constant in acetone-water mixtures. In aqueous DMSO, the plot of $\log k$ against $1/\epsilon$ is linear and its slope is positive when the water concentration is less than 500 ml/l (Fig. 4). The slope is much greater than would be expected on the basis of theory and cannot be caused by electrostatic effects alone. It can be taken for granted that the large value of the slope is chiefly due to the increase in the reactivity of the hydroxide ion with increasing DMSO concentration. Both this value and the similar values found for alkaline ester hydrolysis⁷ point to a reduced solvation of hydroxide ions at high DMSO concentrations in DMSO mixtures.

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