JOURNAL OF PHYSICAL ORGANIC CHEMISTRY
J. Phys. Org. Chem. 2001; **14:** 884–889
DOI:10.1002/poc.438

Dual-parameter correlations on rate of dehydration step of a condensation reaction in aqueous solutions of ethanol, propan-2-ol and 2-methylpropan-2-ol

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Received 26 March 2001; revised 12 June 2001; accepted 2 July 2001

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ABSTRACT: The reaction kinetics of the dehydration step of the condensation reaction between phenylhydrazine and benzaldehyde was studied spectrophotometrically in aqueous solutions of ethanol, propan-2-ol and 2-methylpropan-2-ol at pH 11.5 at 25 °C. The apparent second-order rate constants $k_{2\mathrm{app}}$ of the reaction increase with mole fraction of water in all the aqueous solutions. Single-parameter correlations of log $k_{2\mathrm{app}}$ versus π^* (dipolarity/polarizability), α (hydrogen-bond donor acidity) and E_T^N (normalized polarity parameter) were obtained in all the aqueous solutions. In all cases, the correlations versus E_T^N and π^* are acceptable, but correlations versus α are poor (e.g. in aqueous solutions of ethanol the correlation coefficients are 0.964, 0.967, and 0.751 respectively). The results of dual-parameter correlations of log $k_{2\mathrm{app}}$ versus π^* and α in all cases represent improvements with regard to the single-parameter models (in aqueous solutions of ethanol $r = 0.995$; s.e. = 0.038; $n = 13$). The apparent second-order rate constants of the reaction increase with α, π^* and E_T^N. Increasing the hydrogen-bond donor acidity of the solvent stabilizes the activated complex of the reaction *via* hydrogen-bond formation with the intermediate. A dual-parameter equation of log $k_{2\mathrm{app}}$ versus π^* and α was obtained for the combined data set of aqueous solutions ($n = 35$, $r = 0.989$, s.e. = 0.062, $F_{2,32} = 719.21$), in which π^* has the major effect on the reaction rate relative to α. Copyright © 2001 John Wiley & Sons, Ltd.

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KEYWORDS: dual-parameter correlation; solvent effect; condensation reaction; polarity parameter; dipolarity/polarizability; hydrogen-bond donor acidity; kinetics

INTRODUCTION

The considerable influence of solvents on chemical and physical processes (reaction rates, selectivity, chemical equilibria, position and intensity of spectral absorption bands) has been well established.[1] Modeling of solvent effects is one of the most useful methods to obtain information about the mechanism of organic reactions.[1] The study of physicochemical properties that depend on solute–solvent interactions is much more complex in mixed than in pure solvents. On the one hand, the solute can be preferentially solvated by any of the solvents present in the mixture. On the other hand, solvent–solvent interactions can strongly affect solute–solvent interactions.[2] The complexity of even the simplest reactions in aqueous solutions is significantly greater than those of reactions in organic solvents.[3] The structure of liquid water is dominated by strong intermolecular hydrogen bonding. The complexity of the chemistry of life *in vivo*, which occurs primarily in an aqueous environment, should encourage chemists to focus more attention on the reactions in aqueous solutions. Therefore, there is a continuing challenge to understand the role of water in these processes. In many cases, these reactions involve complicated and large organic molecules. However, we know that our understanding of these phenomena can be improved by studying in depth the rates of relatively simple organic reactions in aqueous solutions.[4–8]

In addition to the nonspecific Coulombic, inductive and dispersion interactions, other interactions, such as specific hydrogen bonding, electron pair donor–acceptor and solvophobic interactions, may play a part in the solvent effects.[9–18] The problem is to identify and to assess the relative importance of these various factors on the reaction rate and other chemical and physical processes.

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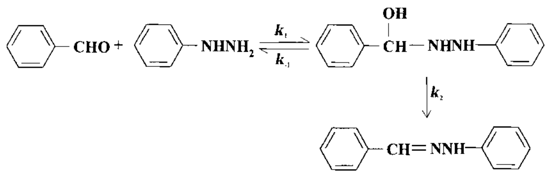

Scheme 1

In continuing our study into solvent effects on organic reactions,[12–15] we were interested in solvent effects on reaction kinetics of condensation reaction of amines with aldehydes. Surprisingly, for these reactions little attention has been paid to the study of solvent effects.[19–21] Phenylhydrazone formation from benzaldehyde, like many condensation reactions, occurs with rate-determining attack of the nucleophile reagent under acidic conditions and rate-determining dehydration of the carbinolamine intermediate under neutral and basic conditions (Scheme 1).[22–26] Each of these steps may be catalyzed by acids and (or) bases (general and specific).[25,26] At pH >10, dehydration of the intermediate of the reaction is the rate-determining step. In the reaction there is no base catalysis.[27]

The purpose of our work is to identify the medium effects on the dehydration step of the condensation reaction, so our study was carried out at pH 11.5. In these conditions we have:

$$\text{rate} = k_2[\text{Interm}] \tag{1}$$

$$[\text{Interm}] = k_1[\text{Ald}][\text{Hyd}]/(k_{-1} + k_2) \tag{2}$$

Because $k_2 \ll k_{-1}$, then:

$$\text{rate} = k_2 k_1/k_{-1}[\text{Ald}][\text{Hyd}] = k_2 K_1[\text{Ald}][\text{Hyd}] \tag{3}$$

$$= k_{2\text{app}}[\text{Ald}][\text{Hyd}] \tag{4}$$

in which K_1 and $k_{2\text{app}}$ are equilibrium constant of the first step and the apparent second-order rate constant of the reaction, respectively. Solvent effects on the kinetics of the dehydration step of the condensation reaction of phenylhydrazine with benzaldehyde were studied in aqueous solutions of ethanol, propan-2-ol and 2-methylpropan-2-ol at pH 11.5 at 25 °C.

RESULTS AND DISCUSSION

The apparent second-order rate constants for the reaction were obtained in aqueous solutions of ethanol, propan-2-ol and 2-methylpropan-2-ol in pH 11.5 at 25 °C (Tables 1–3).

As can be seen, the rate constants of the reaction increase rapidly with the mole fraction of water. The results show that the increase of the reaction rate in

Table 1. Apparent second-order rate constants of the reaction in aqueous solutions of ethanol at 25 °C[a]

X_w	$10^4\ k_{2\text{app}}$ (dm^3 mol^{-1} s^{-1})	E_T^N	α	π^*
0.266	5.01 ± 0.17	0.673	0.93	0.66
0.449	6.91 ± 0.21	0.702	0.89	0.75
0.583	7.80 ± 0.18	0.717	0.89	0.81
0.685	8.49 ± 0.27	0.731	0.88	0.88
0.765	12.8 ± 0.24	0.761	0.86	0.96
0.830	18.2 ± 0.34	0.784	0.87	1.03
0.850	20.0 ± 0.37	0.800	0.90	1.06
0.884	27.8 ± 0.41	0.842	0.92	1.10
0.900	31.6 ± 0.45	0.861	0.97	1.11
0.929	33.0 ± 0.52	0.907	1.05	1.12
0.967	39.5 ± 0.48	0.959	1.17	1.13
0.984	46.4 ± 0.74	0.980	1.21	1.13
1.000	49.4 ± 0.72	1.000	1.26	1.14

[a] At least three runs were averaged. Errors are standard deviations.

Table 2. Apparent second-order rate constants of the reaction in aqueous solutions of propan-2-ol at 25 °C[a]

X_w	$10^4\ k_{2\text{app}}$ (dm^3 mol^{-1} s^{-1})	E_T^N	α	π^*
0.321	2.89 ± 0.08	0.583	0.78	0.65
0.516	3.72 ± 0.10	0.639	0.81	0.72
0.646	3.84 ± 0.12	0.667	0.81	0.77
0.740	5.06 ± 0.16	0.685	0.82	0.81
0.810	7.14 ± 0.19	0.714	0.82	0.87
0.825	8.91 ± 0.18	0.719	0.82	0.89
0.865	11.3 ± 0.24	0.740	0.82	0.95
0.909	18.1 ± 0.25	0.790	0.85	1.06
0.925	22.4 ± 0.35	0.833	0.88	1.12
0.945	31.5 ± 0.32	0.901	0.96	1.14
0.975	37.1 ± 0.48	0.957	1.12	1.15
1.000	49.4 ± 0.72	1.000	1.26	1.14

[a] At least three runs were averaged. Errors are standard deviations.

Table 3. Apparent second-order rate constants of the reaction in aqueous solutions of 2-methylpropan-2-ol at 25 °C[a]

X_w	$10^4\ k_{2\text{app}}$ (dm^3 mol^{-1} s^{-1})	E_T^N	α	π^*
0.369	1.51 ± 0.05	0.546	0.69	0.63
0.568	3.09 ± 0.10	0.600	0.74	0.69
0.692	3.63 ± 0.09	0.631	0.78	0.73
0.778	4.27 ± 0.08	0.653	0.80	0.77
0.800	5.01 ± 0.11	0.660	0.81	0.78
0.840	5.80 ± 0.10	0.672	0.81	0.81
0.888	8.34 ± 0.12	0.701	0.82	0.86
0.900	10.0 ± 0.15	0.710	0.83	0.88
0.925	13.4 ± 0.17	0.741	0.85	0.96
0.955	27.8 ± 0.24	0.846	0.90	1.10
0.979	36.6 ± 0.27	0.861	1.09	1.16
1.000	49.4 ± 0.72	1.000	1.26	1.14

[a] At least three runs were averaged. Errors are standard deviations.

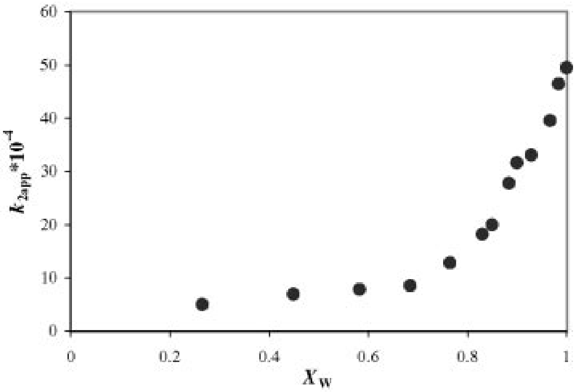

Figure 1. Plot of apparent second-order rate constants of the reaction versus mole fraction of water in aqueous solutions of ethanol at 25 °C

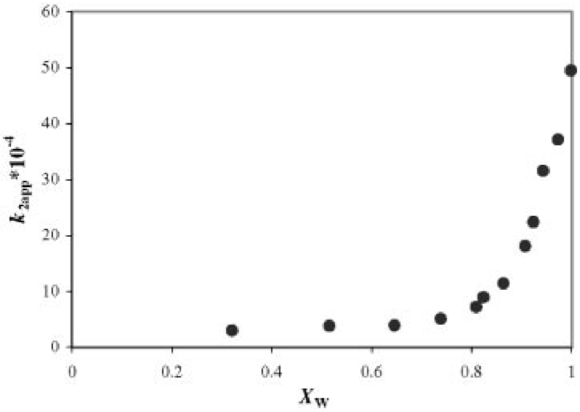

Figure 2. Plot of apparent second-order rate constants of the reaction versus mole fraction of water in aqueous solutions of propan-2-ol at 25 °C

aqueous solutions of ethanol is higher than those of propan-2-ol and 2-methylpropan-2-ol (Figs 1–3).

The normalized polarity parameter E_T^N of the aqueous solutions increases with mole fraction of water in aqueous solutions of the alcohols. At the same mole fractions of water, E_T^N of aqueous solutions follows the sequence ethanol > propan-2-ol > 2-methylpropan-2-ol.[28]

Single-parameter correlations of log k_{2app} versus E_T^N in aqueous solutions of ethanol, propan-2-ol and 2-methylpropan-2-ol give acceptable results:

EtOH–H_2O

$$\log k_{2app} = -5.182(\pm0.206) + 2.968(\pm0.248)E_T^N \quad (5)$$

$$(n = 13, r = 0.964; \text{s.e.} = 0.095; F_{1,11} = 144.79)$$

2-PrOH–H_2O

$$\log k_{2app} = -5.396(\pm0.162) + 3.169(\pm0.208)E_T^N \quad (6)$$

$$(n = 12, r = 0.979; \text{s.e.} = 0.090; F_{1,10} = 231.61)$$

t-BuOH–H_2O

$$\log k_{2app} = -5.619(\pm0.197) + 3.537(\pm0.270)E_T^N \quad (7)$$

$$(n = 12, r = 0.972; \text{s.e.} = 0.114; F_{1,10} = 171.72)$$

By combining all the data for the aqueous solutions the following correlation of log k_{2app} with E_T^N was obtained:

$$\log k_{2app} = -5.518(\pm0.109) + 3.381(\pm0.142)E_T^N \quad (8)$$

$$(n = 35, r = 0.972, \text{s.e.} = 0.097; F_{1,33} = 568.08)$$

All of the above correlations show that the rate of the reaction increases with increasing normalized polarity parameter of the media. We concluded that the polarity of the activated complex of the reaction for the dehydration step is higher than that of the intermediate of the reaction.

The normalized polarity parameter is a blend of dipolarity/polarizability and hydrogen-bond donor acidity.[29–31] Because in the intermediate of the reaction the OH group will be eliminated, it might be expected that formation of a hydrogen bond to the leaving group would increase the propensity for elimination. However, single-parameter correlations of log k_{2app} with α do not give satisfactory results in all aqueous solutions (Table 4).

Linear correlations of log k_{2app} with π^* (dipolarity/polarizability) give good results for all three types of aqueous solution:

EtOH–H_2O

$$\log k_{2app} = -4.718(\pm0.159) + 2.002(\pm0.158)\pi^* \quad (9)$$

$$(n = 13, r = 0.967; \text{s.e.} = 0.090; F_{1,11} = 160.37)$$

2-PrOH–H_2O

$$\log k_{2app} = -5.108(\pm0.144) + 2.288(\pm0.150)\pi^* \quad (10)$$

$$(n = 12, r = 0.979; \text{s.e.} = 0.090; F_{1,10} = 231.45)$$

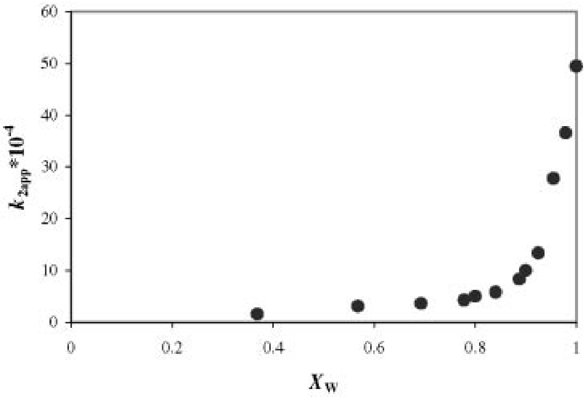

Figure 3. Plot of apparent second-order rate constants of the reaction versus mole fraction of water in aqueous solutions of 2-methylpropan-2-ol at 25 °C

Table 4. Regression coefficients and slopes of solvent parameters in aqueous solutions of ethanol, propan-2-ol, 2-methylpropan-2-ol and for all the aqueous solutions at 25 °C

Aqueous solution	E_T^N	π^*	α	Intercept	r
Ethanol	2.968	–	–	−5.182	0.964
	–	2.002	–	−4.718	0.967
	–	–	1.813	−4.520	0.751
	–	1.659	0.688	−5.056	0.995
Propan-2-ol	3.169	–	–	−5.396	0.979
	–	2.288	–	−5.108	0.979
	–	–	2.340	−5.055	0.822
	–	1.872	0.708	−5.351	0.994
2-methylpropan-2-ol	3.537	–	–	−5.619	0.972
	–	2.576	–	−5.333	0.991
	–	–	2.604	−5.330	0.890
	–	2.327	0.321	−5.393	0.993
All aqueous solutions	3.381	–	–	−5.518	0.972
	–	2.298	–	−5.078	0.975
	–	–	2.555	−5.245	0.831
	–	1.870	0.758	−5.362	0.989

t-BuOH–H_2O

$$\log k_{2\mathrm{app}} = -5.333(\pm 0.097) + 2.576(\pm 0.109)\pi^* \quad (11)$$

$$(n = 12, r = 0.991; \text{s.e.} = 0.064; F_{1,10} = 560.18)$$

As anticipated, a good correlation between log $k_{2\mathrm{app}}$ and π^* exists for the aqueous solutions' combined data set:

$$\log k_{2\mathrm{app}} = -5.078(\pm 0.086) + 2.298(\pm 0.091)\pi^* \quad (12)$$

$$(n = 35, r = 0.975, \text{s.e.} = 0.092; F_{1,33} = 632.78)$$

These equations show that with increasing dipolarity/polarizability of the medium the reaction rate increases, because the activated complex of the reaction is stabilized with increasing dipolarity/polarizability of media. In order to show the relative effects of dipolarity/polarizability π^* and hydrogen-bond donor acidity α on the reaction rate, dual-parameter correlations of log $k_{2\mathrm{app}}$ versus α and π^* were obtained in aqueous solutions of ethanol, propan-2-ol and 2-methylpropan-2-ol. These correlations gave interesting results:

EtOH–H_2O

$$\log k_{2\mathrm{app}} = -5.056(\pm 0.082) + 1.659(\pm 0.082)\pi^* + 0.688(\pm 0.096)\alpha \quad (13)$$

$$(n = 13; r = 0.995; \text{s.e.} = 0.038; F_{2,10} = 472.87)$$

2-PrOH–H_2O

$$\log k_{2\mathrm{app}} = -5.351(\pm 0.093) + 1.872(\pm 0.119)\pi^* + 0.708(\pm 0.144)\alpha \quad (14)$$

$$(n = 12, r = 0.994; \text{s.e.} = 0.049; F_{2,9} = 393.88)$$

t-BuOH–H_2O

$$\log k_{2\mathrm{app}} = -5.393(\pm 0.104) + 2.327(\pm 0.213)\pi^* + 0.321(\pm 0.240)\alpha \quad (15)$$

$$(n = 12, r = 0.993; \text{s.e.} = 0.062; F_{2,9} = 303.07)$$

A dual-parameter correlation of log $k_{2\mathrm{app}}$ versus π^* and α was obtained for the aqueous solutions' combined data set:

$$\log \ k_{2\mathrm{app}} = -5.362(\pm 0.073) + 1.870(\pm 0.091)\pi^* + 0.758(\pm 0.118)\alpha \quad (16)$$

$$(n = 35; r = 0.989; \text{s.e.} = 0.062; F_{2,32} = 719.22)$$

The regression coefficients and standard deviations of the equations are satisfactory. As can be seen, dipolarity/polarizability has the major effect on the reaction rate. In fact, in the dehydration step of the reaction, the polarity of the intermediate increases with progressing of the reaction; as a result, the activated complex leading to the product of the reaction is expected to be favored by an increasing dipolarity/polarizability of the media. The solvent also stabilizes the activated complex of the dehydration step of the reaction by hydrogen-bond interactions.

In order to show the efficiency of the suggested dual-parameter equation, the experimental values of log $k_{2\mathrm{app}}$ versus the calculated values from Eqn. (16) are plotted for all the aqueous solutions of alcohols (Fig. 4).

As can be seen, the experimental and calculated values of log $k_{2\mathrm{app}}$ are in good agreement for all aqueous solutions. A plot of %residual versus experimental values of log $k_{2\mathrm{app}}$ shows that the suggested dual-parameter equation is statistically acceptable, because the distribu-

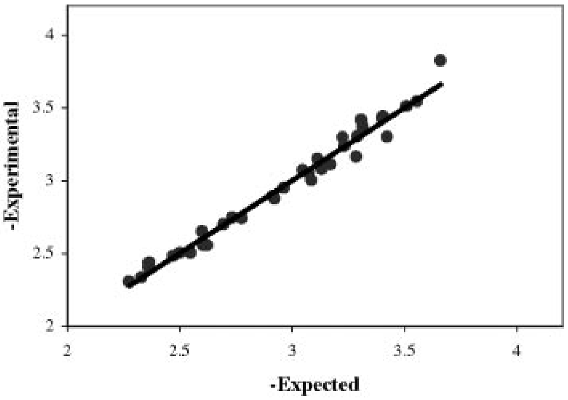

Figure 4. Plot of the experimental values of log k_{2app} versus the expected values obtained from Eqn. (16) for all the aqueous solutions at 25 °C

tion of points is random (Fig. 5):

$$\%\text{residual} = 100 \times (\log\ k_{2\,\text{exp}} - \log\ k_{2\text{expec}})/\log\ k_{2\,\text{exp}} \qquad (17)$$

in which log $k_{2\,\text{exp}}$ and log $k_{2\text{expec}}$ are the experimental and expected values of log $k_{2\text{app}}$ respectively.

The rates of some organic reactions in water and several aqueous solutions are higher than those in organic solvents. Hydrophobic interactions often play an important role in these cases.[32–36] In this study, the effect of hydrophobic interactions is negligible, because the rate-determining step of the reaction is dehydration step. As a result, there is no reduction in the surface accessible in the activation process and hydrophobic interactions do not play a role. This is fully consistent with the dual-parameter correlations of log $k_{2\text{app}}$ with π^* and α in all the aqueous solutions discussed here.

EXPERIMENTAL

Reagents and solvents

Phenylhydrazine and benzaldehyde were distilled under reduced pressure. All solvents were of the highest quality available and were purified as usual. Water was redistilled in a quartz distillation unit.

Kinetic procedures

The kinetics of the reaction was studied spectrophotometrically. The spectrophotometer was coupled to a PC with an interface that allowed recording of absorbance measurements versus time (four readings per second). Absorbances were recorded at wavelengths from 342 nm (in water) to 348 nm (in 2-methylpropan-2-ol) at pH 11.5 at 25 °C. Pseudo-first-order (k_{obs}) conditions were used in all cases. The rate constants were obtained by the initial rate method. All the kinetic runs were carried out at least in triplicate. Concentrations of phenylhydrazine were 4×10^{-3} to 20×10^{-3} M and those of benzaldehyde were 5×10^{-5} to 4×10^{-4} M. Water was circulated around the cell of the spectrophotometer and the water temperature was thermostatically maintained at 25 ± 0.1 °C.

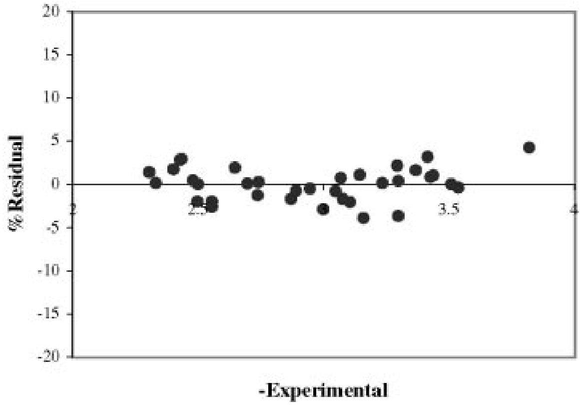

Figure 5. Plot of the %residual versus the experimental values of log $k_{2\text{app}}$ for all the aqueous solutions at 25 °C

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