Kinetics and Mechanism of the Acid-Catalyzed Bromination of Ring-Substituted Acetophenones in Methanol. Thermodynamics of the Ketone-Acetal-Enol Ether System in Methanol and Water[1]

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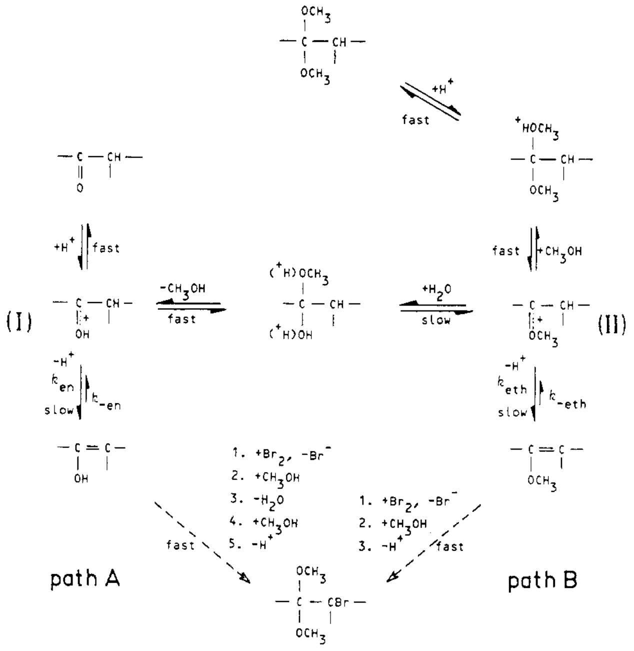
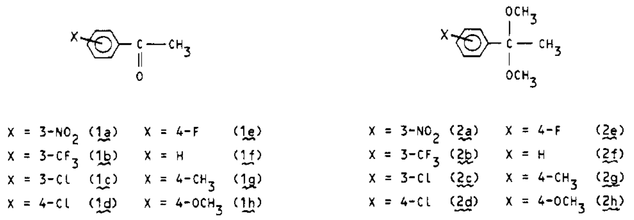
The kinetics of HBr-catalyzed bromination of ring-substituted acetophenones in equilibrium with acetals have been studied in methanol containing small amounts of water at very low halogen concentration (ca. 10^{-6} M). Under these conditions, the rate is zeroth order with respect to the halogen and first order with respect to the $CH_3OH_2^+$ ion and the overall carbonyl compound concentrations. The relative concentration of acetal at equilibrium ($x_{eq} = [\text{acetal}]_{eq}/([\text{acetal}]_{eq} + [\text{ketone}]_{eq})$) has been varied by increasing the water concentration; the lyonium ion catalytic constant depends linearly on x_{eq} with an $x_{eq} = 0$ intercept corresponding to bromination via intermediate enol formation and with an $x_{eq} = 1$ intercept corresponding to bromination via enol ether. In the case of 3-nitroacetophenone (**1a**), these extrapolated rate constants are identical with those directly measured from the rates for bromination of **1a** and of **1a** dimethyl acetal, before ketone-acetal equilibration. In most cases, the enol ether path is highly predominant. Substituent effects on the constants associated with the two paths are examined. The rate constants for enol ether formation are combined with rate constants for the reverse process and with acetal formation equilibrium constants to provide acetal-to-enol ether and ketone-to-enol ether equilibrium constants in methanol. Taking into account data on Gibbs free energies of transfer from methanol to water for the different species involved, equilibrium data in water are also deduced.

Some years ago we showed that the acid-catalyzed bromination of acetone in methanol occurs mainly via the formation of the corresponding methyl enol ether,[2] instead of via the enol tautomer as for acid-catalyzed bromination in other solvents.[3] This mechanism accounts for significant pecularities in regioselectivity[4] and stereoselectivity[5] for ketone halogenation in alcoholic media. The evidence was based on two kinds of kinetic experiments: (i) a study of the kinetic behavior for the bromination of dynamic ketone-acetal systems, i.e., when ketone was added to bromine solutions under conditions such that ketone was simultaneously transformed to acetal, and (ii) a study of the rate of zeroth-order bromination of equilibrated ketone-acetal mixtures, the relative proportions of the two species being varied by the addition of small amounts of water. From this work it was concluded that the enol path (path A) and the enol ether path (path B) are usually in competition, with rate ratios depending on the water content, but that the enol ether path is exclusive in the pure anhydrous solvent. Data were interpreted in terms of the mechanism described in Scheme I, where the enol and enol ether intermediates are assumed to be formed from the hydoxycarbenium (I) and alkoxycarbenium (II) ions (oxocarbenium ions), respectively.[6] The fact that the rate of bromine consumption was observed to be the sum of partial rates, associated with competitive enol and enol ether formation, and that these partial rates could be linearly related to ketone and acetal concentrations was explained by considering that the concentrations of I and II should be proportional to those of ketone and acetal, respectively, as the consequence of fast equlibria between the ions and the starting materials.

Scheme I



Scheme II



To examine further the factors that make one of the two reaction paths predominant, we have extended the study of the bromination of equilibrated ketone-acetal mixtures in methanol to ring-substituted acetophenones **1a-h** in equilibrium with the corresponding acetals **2a-h** (Scheme II). Besides its significance for a better understanding of the reaction mechanism, a major interest of this work lies in the fact that by combining the rate constants for $CH_3OH_2^+$-catalyzed enol ether formation with those cor-

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(1) Some of this material appeared in a preliminary form: Toullec, J.; El-Alaoui, M. *Tetrahedron Lett.* **1979**, 3085-3088.

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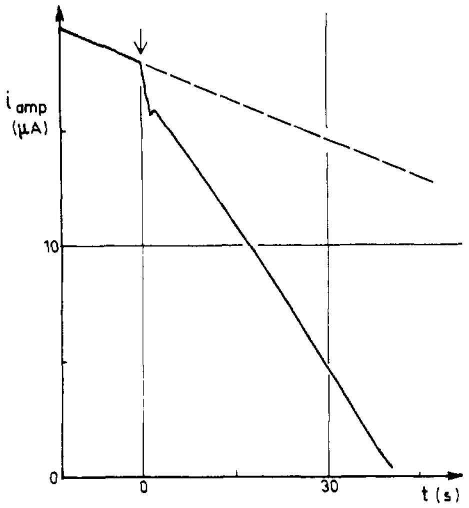

Figure 1. Kinetics of the initial bromintion of **1a** in acidic methanol ([**1a**] = 1.8×10^{-2} M; $[Br_2]_{st} = 5.3 \times 10^{-6}$ M; [HBr] = 1.97×10^{-2} M; $[H_2O] = 0.0091$ M).

responding to lyonium ion catalyzed methanol addition to α-methoxystyrenes (the methyl enol ethers derived from acetophenones)[1,7] and equilibrium constants for dimethyl acetal formation,[8] it is possible to derive equilibrium constants for the ternary ketone–acetal–enol ether system and to develop a detailed examination of substituent effects on rates and equilibria.

Results

The rates of bromination of equilibrated ketone–acetal mixtures and, in the special case of **1a**, those of bromination of unequlibrated ketone and acetal were measured in acidic (HBr) methanol by an amperometric method,[2,9] which depends on the measurement of a diffusion current (i_{amp}) between two platinum electrodes with a constant potential difference (150 mV).

Initial Rates for the Bromination of 1a in Methanol. Figure 1 shows the i_{amp}–time curve observed when a solution of **1a** in chlorobenzene was added to a solution of bromine at low concentration (ca. 5×10^{-6} M) in acidic methanol. Since i_{amp} is linearly related to bromine concentration (see Experimental Section), it can be seen that, after an initial bromine uptake (which can be due to the reaction of the enol tautomer or, more likely, to that of small amounts of impurities), bromine consumption was linear. When the mixture was allowed to stand for a few minutes and when bromine was rapidly regenerated by electrolysis, it was observed that the slope of the bromine concentration–time straight line was increased. This can easily be related to the occurence of acid-catalyzed formation of acetal which, under the experimental conditions, attains equilibrium concentration after 15–30 min (as calculated from data on the kinetics of ketone–acetal equilibration).[10] It follows that the initial rate (but just after the fast bromine uptake) can be attributed to the bromination of pure ketone and the time-independent rate after 15–30 min to that of the bromination of the equilibrated ketone–acetal mixture.

Table S1 (supplementary material) lists rate constants k_ψ, defined as the ratios between the initial pseudo-zeroth-order rate constants and **1a** concentrations (eq 1), where $[Br_2]_{st}$ represents the sum of Br_2 and Br_3^- concentrations and where α is the experimental ratio between $[Br_2]_{st}$ and i_{amp}.

$$-d[Br_2]_{st}/dt = -\alpha\, di_{amp}/dt = k_\psi[\text{ketone}] \qquad (1)$$

According to the mechanism in Scheme I, bromination of ketones in the absence of acetal should occur exclusively along path A. Since, as assumed previously[2] and confirmed below, the concentration of II is always proportional to acetal concentration, the reaction cannot involve enol ether formation. It follows that bromination rates should depend only on the concentration of I, formed by fast protonation. Accordingly, k_ψ can be expressed by eq 2, where $(K_a)_I$ is

$$k_\psi = k_{en}[H^+]/(K_a)_I \qquad (2)$$

the *apparent* acidity constant in methanol containing small amounts of water and where $[H^+]$ is the stoichiometric hydrogen ion concentration. It was previously shown[2] that because of the partition of protons between water and methanol molecules (eq 3),[11] the concentration of the $CH_3OH_2^+$ ion can be easily calculated (eq 4) and that

$$CH_3OH_2^+ + H_2O \overset{K_r}{\rightleftharpoons} CH_3OH + H_2O \qquad (3)$$

$$[CH_3OH_2^+] = a_m[H^+]/(a_m + K_r[H_2O]) \qquad (4)$$

the $[H^+]/(K_a)_I$ ratios are in fact equal to $[CH_3OH_2^+]/[(K_a)_I]_m$, where $[(K_a)_I]_m$ is the acidity constant of ion I in *pure anhydrous methanol.* [in eq 4, a_m is the activity of methanol ($a_m = 1$ in the pure solvent, but corrected for water concentration when methanol contains significant amounts of water); $K_r = 4.94$ M^{-1}.[11]] Table S1 (supplementary material) also gives the values of $k_\psi/[CH_3OH_2^+]$, the lyonium ion catalytic constant [$(1.82 \pm 0.04) \times 10^{-4}$ M^{-1} s^{-1})].

Initial Rates for the Bromination of 2a in Methanol. Similar experiments were performed with solutions of **2a**. As for **1a**, an initial very fast reaction was observed, but bromine uptake was far larger [this was probably due to fast reaction of bromine with the corresponding enol ether, as expected from the relatively large amounts of enol ether at equilibrium (vide infra)]. After a fast regeneration of bromine, a linear decrease of concentration was observed. Since it can be calculated that the amount of acetal hydrolyzed during the time-lag between acetal addition and

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(6) The assumption that lyonium ion catalyzed enolization in methanol occurs by a stepwise mechanism, i.e., via the formation of a discrete hydroxycarbenium ion, is in agreement with the classical mechanism in water (see ref 3), which was confirmed recently (Albery, W. J.; Gelles, J. S. *J. Chem. Soc., Faraday Trans. 1* **1982**, *78*, 1569–1578. Albery, W. J. *Ibid.* **1982**, *78*, 1579–1590), but not with the concerted mechanism suggested for the hydronium ion catalyzed ketonization of vinyl alcohols derived from aldehydes (Capon, B.; Zucco, C. *J. Am. Chem. Soc.* **1982**, *104*, 7567–7572. Capon, B.; Siddhanta, A. K.; Zucco, C. *J. Org. Chem.* **1985**, *50*, 3580–3584). However, since protonated ketones are certainly far more stable than protonated aldehydes, essentially because of the hyperconjugative effect of the additional alkyl group, it is very likely that ketone enolization occurs by the stepwise mechanism even in methanol. This is also in keeping with the similar solvent isotope effects observed for acetone enol ketonization in water and for methyl isopropenyl ether hydrolysis[3] and with the similar structure isotope effects observed for acetone bromination in water and in methanol containing significant amounts of water.[2]

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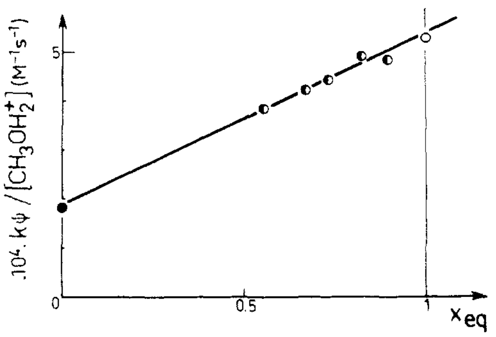


Figure 2. Plot of the catalytic constant for **1a**/**2a** bromination vs. the relative amount of acetal: (●) initial bromination of **1a**; (○) initial bromination of **2a**; (◐) bromination of equilibrated ketone–acetal mixtures at different water concentrations.

curve recording was negligible, these kinetics can be associated with acetal bromination only. Table S2 (supplementary material) lists the pseudo-first-order rate constants, calculated as the slopes of the $[Br_2]_{st}$ vs. time plots divided by the acetal concentration. The dependence of k_ψ on the concentration of the small amounts of water can be ascribed to a decrease in lyonium ion concentration when $[H_2O]$ increases. Since the concentration of ion II can be written as $K^{II}_{ac.H^+}[CH_3OH_2^+][acetal]/[(K_a)_{ac.H^+}]_m$, where $[(K_a)_{ac.H^+}]_m$ and $K^{II}_{ac.H^+}$ are the acidity constant of acetal conjugate acid in pure methanol and the standard state equilibrium constant for the formation of ion II, respectively, the fact that $k_\psi/[CH_3OH_2^+]$ is constant $[(5.3 \pm 0.2) \times 10^{-4}\ M^{-1}\ s^{-1}]$ is in agreement with the mechanism described by Scheme I. Bromination of unequilibrated acetal occurs exclusively via enol ether.

Bromination of Equilibrated Ketone–Acetal Systems. The rates of bromination of ring-substituted acetophenones **1a–h** in equilibrium with the corresponding dimethyl acetals **2a–h** were also measured by the amperometric method. For $[Br_2]_{st}$ = ca. $(1–5) \times 10^{-6}$ M and [HBr] = $(5–50) \times 10^{-3}$ M (depending on the substrate; see the Experimental Section for the factors which limit the choice of acid concentrations), bromine consumption was linear, corresponding to a zeroth order with respect to bromine (eq 5). (Under these conditions, bromine con-

$$-d[Br_2]st = k_\psi([ketone]_{eq} + [acetal]_{eq}) = k_\psi[ketone]_T \qquad (5)$$

centration was so low that the ketone and acetal concentrations could be considered as constant throughout the kinetic runs.) Table S3 (supplementary material) lists the k_ψ values observed for **1a–h** bromination at different HBr and water concentrations, as well as the calculated molar fractions of acetal at equilibrium ($x_{eq} = [acetal]_{eq}/[ketone]_T$) and lyonium ion concentrations. Variations of $k_\psi/[CH_3OH_2^+]$ vs. x_{eq} are shown in Figures 2 and 3 for **1a** and **1f**, respectively. These are typical of two different behaviors: (i) for acetophenones with electron-withdrawing substituents (**1a–c**), straight lines with nonzero intercepts are observed, in agreement with eq 6, where $(k_A)_m$ and

$$k_\psi/[CH_3OH_2^+] = (k_A)_m + [(k_B)_m - (k_A)_m]x_{eq} \qquad (6)$$

$(k_B)_m$ are the values of $k_\psi/[CH_3OH_2^+]$ for $x_{eq} = 0$ and $x_{eq} = 1$, respectively; (ii) for **1d–h**, the intercept is very close to the origin, and it is not possible to derive $(k_A)_m$ values.

It is noteworthy that, in the special case of **1a**, the extrapolated $(k_A)_m$ and $(k_B)_m$ values correspond to those measured for direct bromination of **1a** and **2a**, respectively. This means that $(k_A)_m$ and $(k_B)_m$ are lyonium ion catalyzed rate constants for enol and enol ether formation and confirms the mechanistic assumption, namely, that the partial rates for the two competing paths are linearly related to the concentrations of ions I and II. According to this assumption, k_ψ can be written as in eq 7. Moreover,

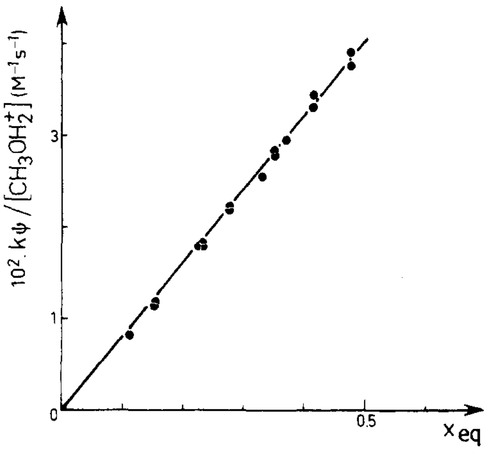


Figure 3. Plot of the catalytic constant for the bromination of **1f**/**2f** equilibrated mixtures vs. the relative amount of acetal.

Table I. Lyonium Ion Catalyzed Rate Constants for Enol Formation, $(k_A)_m$, and Enol Ether Formation, $(k_B)_m$, in Anhydrous Methanol[a]

X in 1	$10^2[H^+]$,[b] M	$10^3(k_B)_m$,[c,d] M^{-1} s^{-1}	$10^3(k_A)_m$,[c,d] M^{-1} s^{-1}
4-OCH_3	0.47–0.49	1250 ± 70	(0.74)
4-CH_3	1.04–1.07	288 ± 15	(0.80)
H	1.01–1.02	77.5 ± 2.2	(0.63)
4-F	1.02	42.0 ± 1.0	(0.46)
4-Cl	1.04	15.3 ± 0.7	(0.41)
	1.99–2.01	17.6 ± 1.8	(0.43)
3-Cl	1.92–1.95	4.84 ± 0.07	0.334 ± 0.023
3-CF_3	1.94–2.04	1.79 ± 0.09	0.287 ± 0.033
3-NO_2	2.00–2.07	0.53 ± 0.02	0.182 ± 0.004

[a] At 25 ± 0.1 °C. [b] Stoichiometric acid concentration (the sum of H_3O^+ and $CH_3OH_2^+$ concentrations). [c] Calculated by plotting the first-order rate constants for bromination of equilibrated ketone–acetal mixtures vs. the molar fraction of acetal (see eq 6), except for values in parentheses, which were estimated from the log $(k_A)_m$–σ^n plot for **1a–c** and corrected for the differences in electrolyte effects (see the text). [d] Errors are standard deviations.

$$k_\psi = \left[\frac{k_{en}}{(K_a)_I}(1 - x_{eq}) + \frac{k_{eth}K^{II}_{ac.H^+}}{(K_a)_{ac.H^+}}x_{eq}\right][H^+] \qquad (7)$$

when considering that the ratios $[H^+]/(K_a)_I$ and $[H^+]/(K_a)_{ac.H^+}$ are equal to $[CH_3OH_2^+]/[(K_a)_I]_m$ and $[CH_3OH_2^+]/[(K_a)_{ac.H^+}]_m$, respectively, the two extrapolated rate constants $(k_A)_m$ and $(k_B)_m$ can be related by eq 8 and 9 to the rate and equilibrium constants of the individual steps.

$$(k_A)_m = k_{en}/[(K_a)_I]_m \qquad (8)$$

$$(k_B)_m = k_{eth}K^{II}_{ac.H^+}/[(K_a)_{ac.H^+}]_m \qquad (9)$$

Table I lists $(k_A)_m$ and $(k_B)_m$ values for the acetophenone series. When $(k_A)_m$ values could not be determined as the intercepts of the $k_\psi/[CH_3OH_2^+]$ graphs, they were calculated by assuming that the linear Hammett plot observed for **1a–c** was valid for all substituents (vide infra).

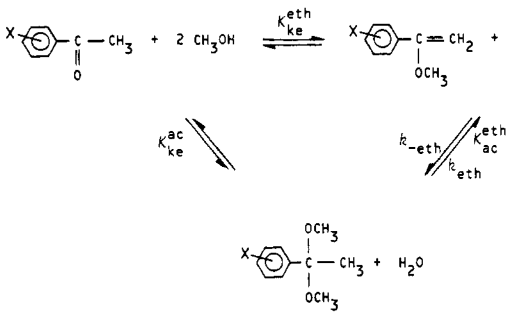
Equilibrium Constants for the Ternary Ketone–Dimethyl Acetal–Enol Ether System. The $(k_B)_m$ values listed in Table I, corresponding to the catalytic constants for α-methoxystyrene formation from acetals in pure methanol, can be combined with those corresponding to $CH_3OH_2^+$-catalyzed acetal formation from α-methoxystyrenes[1,7] to provide equilibrium constants for the ace-

Table II. Calculated Equilibrium Constants for Enol Ether Formation from 1a–h and 2a–h in Methanol (25 °C)

X in 1 or 2	$10^3(k_B)_m{}^0$,[a] $M^{-1}\,s^{-1}$	$(k_{-eth})_m{}^0$,[b] $M^{-1}\,s^{-1}$	$10^4K_{ac}^{eth}$,[c]	$10^3K_{ke}^{ac}$,[d] M	$10^6K_{ke}^{eth}$,[e] M
4-OCH_3	1130	1180	9.58	2.16	2.07
4-CH_3	250	273	9.16	5.93	5.43
H	67.5	81.1	8.32	15.5	12.9
4-F	36.6	45.2	8.10	16.9	13.7
4-Cl	13.3[f]	18.1	7.35	32.8	24.1
3-Cl	4.00	6.95	5.75	69.3	39.8
3-CF_3	1.48	2.73	5.42	107	58.0
3-NO_2	0.436	0.81	5.38	244	131

[a] Ionic strength corrected rate constants for $CH_3OH_2^+$-catalyzed enol ether formation from acetal in pure anhydrous methanol. [b] Ionic strength corrected rate constants for $CH_3OH_2^+$-catalyzed acetal formation from α-methoxystyrenes in pure anhydrous methanol (from ref 1). [c] $K_{ac}^{eth} = (k_B)_m{}^0/(k_{-eth})_m{}^0$ = [enol ether]$_{eq}a_m$/[acetal]$_{eq}$ with $a_m = 1$. [d] From ref 8. [e] $K_{ke}^{eth} = K_{ke}^{ac}K_{ac}^{eth}$ = [enol ether]$_{eq}[H_2O]$/[ketone]$_{eq}a_m$. [f] From data at [HBr] = 1.04 × 10^{-2} M.

Scheme III



tal–enol ether interconversion.[3,12] Nevertheless, since kinetic measurements were performed at different HBr concentrations, i.e. at different ionic strengths, $(k_B)_m$ data in Table I should be corrected for electrolyte effects. These were deduced from ionic strength effects observed for the reverse reaction, i.e. for $CH_3OH_2^+$-catalyzed methanol addition to α-methoxystyrenes.[7] For this latter reaction, in the special case of the 3-nitro derivative, it was observed that the catalytic constant depends on HBr concentration according to eq 10, where $(k_{-eth})_m{}^0$ is the rate constant at $\mu = 0$. (Similar effects have also been reported for

$$\log (k_{-eth})_m = \log (k_{-eth})_m{}^0 + 0.58\,[\mathrm{HBr}]^{1/2} \quad (10)$$

methanol addition to cyanoketene dimethyl acetal.[13]) Since no electrolyte effects are expected on equilibrium constants, it follows that $(k_B)_m$ values should depend on [HBr] according to the same equation. The corrected rate constants, $(k_B)_m{}^0$, are listed in Table II; it was assumed that the sensitivity coefficient to electrolyte effects does not depend on substituent. Table II also lists the calculated constants in methanol for the acetal–enol ether and ketone–enol ether equilibria (Scheme III). These latter constants were calculated by taking into account the ketone-to-acetal equilibrium constants reported previously.[8]

Gibbs free energies of transfer (Table III) from methanol to water for α-methoxystyrenes $(\Delta G°_{m\to w})_{eth}$ were determined from partition coefficients of these species between methanol and dodecane, and between dodecane and water.[14] These data and similar ones for **1a–h**, $(\Delta G°_{m\to w})_{ke}$, and **2a–h** $(\Delta G°_{m\to w})_{ac}$,[8] make it possible to calculate the ketone-to-enol ether and acetal-to-enol ether equilibrium constants in *water*. Indeed, the standard Gibbs free energy increments for these equilibria can be expressed by eq 11 and 12, where $(\Delta G°_{m\to w})_w$ (−0.23 kcal

$$(\Delta G°_w)_{ke}^{eth} = (\Delta G°_m)_{ke}^{eth} + (\Delta G°_{m\to w})_{eth} - (\Delta G°_{m\to w})_{ke} + (\Delta G°\mathrm{m}_{\to w})_w - (\Delta G°\mathrm{m}_{\to w})_m \quad (11)$$

$$(\Delta G°_w)_{ke}^{eth} = (\Delta G°_m)_{ac}^{eth} + (\Delta G°_{m\to w})_{eth} - (\Delta G°_{m} \rightarrow {}_{w})_{ac} + (\Delta G_{m} \rightarrow {}_{w})_m \quad (12)$$

mol^{-1}) and $(\Delta G°_{m\to w})_m$ (0.19 kcal mol^{-1}) are the Gibbs free energies of transfer to water and methanol from methanol to water, calculated from the vapor pressures of methanol and water, and from the Henry coefficients of the two components in methanol–water mixtures.[8]

Discussion

Fast Enol and Enol Ether Bromination in Methanol. Despite the low bromine concentrations in methanol, a zeroth-order rate with respect to bromine was observed. This means that, under our conditions, enol ether bromination and enol bromination are faster than the protonation processes.

Data on the bromination of α-methoxystyrenes in methanol[15] make it possible to calculate the second-order rate constants for enol ether bromination (by Br_2 and Br_3^-) under our conditions and the bromination/protonation rate ratios. These are always higher than 30 at $[Br_2]_{st} = 10^{-7}$ M, i.e., for more than 90% reaction extent, and are in agreement with the observed pseudo-zeroth-order when the reaction occurs mainly via enol ether.

From comparative studies on the bromination and the iodination of acetophenones,[16] it was concluded that halogen addition to enol in water is diffusion controlled. This was recently confirmed.[17] Moreover, Br_3^- reacts with enol with a rate constant that is only half of that for bromination by free bromine. If one assumes a similar behavior in methanol (in agreement with structure–reactivity realtionships),[15] at least for the most reactive enols, the overall rate constant for enol bromination by Br_2 and Br_3^- in methanol can be calculated as high as 10^9 $M^{-1}\,s^{-1}$. Hence, by taking into account that lyonium ion catalyzed enol ketonization in methanol should only be a little faster than hydronium ion catalyzed ketonized in water (the ratio between the rate constants for α-methoxystyrene protonation by $CH_3OH_2^+$ and H_3O^+ in methanol and water, respectively, lies in the 1–3 range, depending on the substituent),[1] one can expect a ratio higher than 10 at $[Br_2]_{st} = 10^{-6}$ M between enol bromination and ketonization rates.[18] Since a zeroth-order was always observed, even for bromination via enols, it should even be larger in the case of **1a–c**.

Relative Importance of Enol and Enol Ether Paths vs. Water Concentration. Since the molar fraction of acetal at equilibrium depends on the water content, the percentages of the two paths can be calculated from the $(k_A)_m$ and $(k_B)_m$ values (measured or estimated) as a

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[18] The rate constants for the H_3O^+-catalyzed ketonization of acetophenone enols, calculated by combining data from ref 16 and 17c, are: 82 (**1a**), 307 (**1b**), 339 (**1c**), 632 (**1d**), 1020 (**1f**), 2430 (**1g**), and 6220 (**1h**) ($M^{-1}\,s^{-1}$; 25 °C).

Table III. Calculated Equilibrium Constants for the Acetal-to-Enol Ether and Ketone-to-Enol Ether Equilibria in Water (25 °C)

X[a]	$\log P_{m\to d}$[b]	$\log P_{d\to w}$[c]	$(\Delta G°_{m\to w})_{eth}$,[d] cal mol^{-1}	$(\Delta G°_{m\to w})_{ac}$,[e] cal mol^{-1}	$(\Delta G°_{m\to w})_{ke}$,[e] cal mol^{-1}	$10^2K_{ac}^{eth}$,[f] M	$10^9K_{ke}^{eth}$,[g] M^{-1}
4-OCH_3	0.215 ± 0.007	−1.918 ± 0.004	2808	3215	2390	7.7	1.5
4-CH_3	0.108 ± 0.001	−2.777 ± 0.019	4126	4428	3521	6.1	2.9
H	−0.070 ± 0.010	−2.196 ± 0.002	3575	3926	2954	6.0	6.8
4-F	−0.107 ± 0.001	−2.571 ± 0.017	4137	4523	3203	6.2	4.2
4-Cl	0.082 ± 0.002	−3.110 ± 0.106	4615	5050	3788	6.1	8.9
3-Cl	0.072 ± 0.002	−3.374 ± 0.063	4988	5646	4128	7.0	14
3-NO_2	−0.526 ± 0.011	−2.220 ± 0.026	4230	4597	2865	4.0	19.7

[a] X in $XC_6H_4COCH_3$, $XC_6H_4C(OCH_3)_2CH_3$, or $XC_6H_4C(OCH_3)CH_2$. [b] Partition coefficients of α-methoxystyrenes between methanol and dodecane; errors are standard deviations. [c] Partition coefficients of α-methoxystyrenes between dodecane and water; errors are standard deviations. [d] Gibbs free energies of transfer of α-methoxystyrenes from methanol to water, calculated as $-RT\ln(P_{m\to d}P_{d\to w}V_w^*/V_m^*)$, where V_m^* and V_w^* are the mole volumes of methanol and water, respectively. [e] From ref 8. [f] K_{ac}^{eth} = [enol ether]$_{eq}$[CH_3OH]$_{eq}$/[acetal]$_{eq}$; equilibrium constants on the basis of mole fractions were calculated from eq 12 and converted to equilibrium constants on the basis of molarities by multiplying by $10^3/V_w^*$. [g] K_{ke}^{eth} = [enol ether]$_{eq}a_w$/[ketone]$_{eq}$[CH_3OH]$_{eq}$ with $a_w = 1$ in pure water; equilibrium constants on the basis of mole fractions were calculated from eq 11 and converted to equilibrium constants on the basis of molarities by dividing by $10^3/V_w^*$.

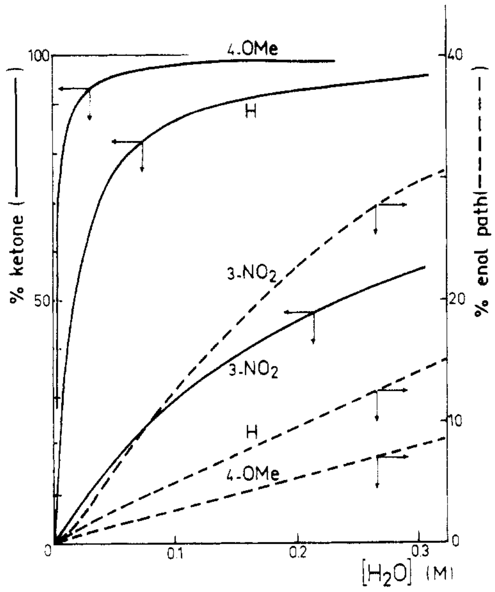

Figure 4. Plot of the percentage of free ketone at equilibrium in methanol (—) and of the percentage of enol path (- - -) vs. water concentration, in the cases of **1a**/**2a**, **1f**/**2f**, and **1h**/**2h**.

function of water concentration. Figure 4 shows the % path A–[H_2O] curves for **1a**, **1f**, and **1h**, compared with the [ketone]$_{eq}$–[H_2O] curves.[19] In the same way, the following percentages of enol path can be calculated for [H_2O] = 0.02 M and compared with the percentages of ketone at equlibrium (enclosed in parentheses): **1a**, 2.74 (7.6); **1b**, 2.90 (15.7); **1c**, 1.95 (22.4); **1d**, 1.59 (37.8); **1e**, 1.29 (54.2); **1f**, 1.04 (56.4); **1g**, 0.93 (77.1); **1h**, 0.55 (90.3). It is noteworthy that contrary to what is intuitively expected, the importance of the enol path is not larger when a substituent decreases the relative stability of the acetal. *The less abundant the ketone for the same water concentration, the more favored is the enol path.*

Figures 2 and 3 show that eq 6 accounts for the variations of k_ψ; they are only related to acetal and lyonium ion concentrations, which in turn depend on the water content. As emphasized previously,[2] catalysis by H_3O^+, which would correspond to deprotonation of I and II by water molecules, is negligible. This is in keeping with the absence of hydronium ion catalysis terms in the rate law for the reaction of acidic methanol with enol ethers.[1,7,13] The linearity of the $k_\psi/[CH_3OH_2^+]$ graphs stems from the fact that ions I and II are in fast equilibria with ketone and acetal, respectively. The fast ketone ⇌ I equilibrium is due to fast proton transfer to and from the carbonyl group, whereas, as shown below, the fast acetal ⇌ II equilibrium is a consequence of the larger rate for methanol addition to II than for proton abstraction by methanol.

A recent study from our laboratory showed that the rate-limiting step for acetal hydrolysis in methanol containing small amounts of water corresponds to water addition to II.[10] The combination of the lyonium ion catalyytic constants for hydrolysis with the rate constants $(k_B)_m$ for enol ether formation makes it possible to calculate the ratios between the rate constants for water addition to II (k_{H_2O}) and proton abstraction by methanol molecules ($k_{eth}' = k_{eth}/[CH_3OH]$); they are in the (2.5–3.6) × 10^4 range and depend only slightly on the substituent. On the other hand, it was observed with Cordes and Wenthe in a study of 2,2-dimethoxypropane hydrolysis in an equimolecular mixture of D_2O and CD_3OD that the OCH_3/OCD_3 group exchange occurs at roughly the same rate as acetone formation by hydrolysis.[20] These data mean that methanol attachment to II is about as fast as water attachment. Since the same behavior can be assumed for the methoxycarbenium ions arising from acetophenones, the water attachment/deprotonation ratio can be taken as a measure of that between the second-order rate constants for methanol attachment and deprotonation. Therefore, *methanol addition to II is about 10^4 faster than proton abstraction.*

The ratios of ca. 10^4 between water attachment rate constants and those for proton abstraction by methanol molecules mean that these two processes are in competition when the water concentration is low (e.g., for [H_2O] = 10^{-2} M, $k_{H_2O}[H_2O]/k_{eth}'[CH_3OH] \approx 10$) and that the ketone–acetal equilibratrion process usually interferes with enol ether formation. This complicating factor, which was exploited in order to extract important information from data on the bromination of dynamic ketone–acetal systems,[2] can be avoided, as in this work, by using a very small bromine concentration. Consequently, bromine uptake is only accompanied by small variations in the relative amounts of acetal and ketone.

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[19] As in the case of $(k_B)_m$, $(k_A)_m$ values were corrected for the ionic strength effect to obtain $(k_A)_m^0$, the enol formation rate constants in pure anhydrous methanol. Because of the close analogy between the two processes, it was assumed that the ionic strength effect is the same on enol formation as on enol ether formation. The calculated $(k_A)_m^0$ values are 0.150 × 10^{-3}, 0.237 × 10^{-3}, and 0.276 × 10^{-3} M^{-1} s^{-1} for **1a**, **1b**, and **1c**, respectively. For the other substituents, $(k_A)_m^0$ was estimated by assuming the validity of the $\log (k_A)_m^0 = 0.79\sigma^n - 3.260$ relationship observed for **1a**–**1c**.

[20] Wenthe, A. M.; Cordes, E. H. *J. Am. Chem. Soc.* **1965**, *87*, 3173–3180. Cordes, E. H. *Prog. Phys. Org. Chem.* **1967**, *4*, 1–44. Cordes, E. H.; Bull, H. G. *Chem. Rev.* **1974**, *74*, 581–603.

Table IV. Regression Parameters[a–c]

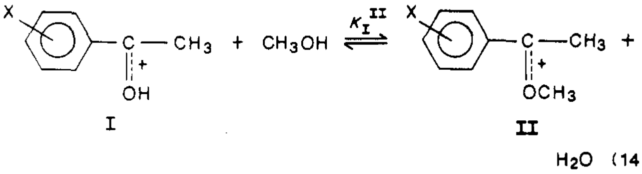
rate or equilibrium constants	solv	ρ^n	ρ^r	$\mathcal{R}^2$ [d]	s [e]	F [f]
$(k_B)_m{}^0$	CH_3OH	−3.12 ± 0.09	−1.39 ± 0.12	0.998	0.06	1313
		(−2.38 ± 0.18)[g]		(0.966)[g]	(0.23)[g]	(172)[g]
		(−3.78 ± 0.37)[h]		(0.946)[h]	(0.29)[h]	(105)[h]
K_{ac}^{eth}	CH_3OH	−0.29 ± 0.05	−0.09 ± 0.06	0.936	0.03	37
		(−0.33 ± 0.04)[h]		(0.906)[h]	(0.03)[h]	(58)[h]
	H_2O	−0.18 ± 0.12	−0.06 ± 0.15	0.53	0.07	2
K_{ke}^{eth}	CH_3OH	1.45 ± 0.07	0.88 ± 0.09	0.995	0.05	542
		(1.21 ± 0.07)[g]		(0.982)[g]	(0.08)[g]	(328)[g]
		(1.87 ± 0.24)[h]		(0.912)[h]	(0.18)[h]	(62)[h]
	H_2O	0.79 ± 0.13	0.86 ± 0.15	0.972	0.08	70
		(0.82 ± 0.06)[g]		(0.971)[g]	(0.07)[g]	(172)[g]

[a] From eq 15, except when otherwise stated. [b] For the meaning of the different statistical parameters, see: Draper, N. R.; Smith, H. *Applied Regression Analysis*; Wiley: New York, 1966. [c] Errors are standard deviations. [d] Squares of the correlation coefficients for simple and multiple regressions. [e] Standard deviations for regression. [f] Fischer parameters ($F = [\sum(y_i - \bar{y})^2 - \sum(y_i - \hat{y}_i)^2/(p-1)]/s^2$, where y_i values are the experimental log k or log K values, $\bar{y}$ is the mean value of y_i, $\hat{y}_i$ values are the calculated y_i, and p is the number of independent variables. [g] Regression parameters for simple regression with σ^+. [h] Regression parameters for simple regression with σ^n.

Since the two oxocarbenium ions are always in fast equilibria with ketone and acetal, the ratio between enol formation and enol ether formation rates can be expressed as in eq 13, where K_I^{II} is the equilibrium constant for the

$$\% \text{ path A}/\% \text{ path B} = k_{en}[H_2O]/K_I^{II}k_{eth}[CH_3OH] \quad (13)$$

I-to-II isodesmic process (eq 14) ($K_I{}^{II} = [II][H_2O]/[I][CH_3OH]$). In other words, the importance of the two reaction paths depends (i) on the ratio of the elementary rate constants for proton abstraction and (ii) on the relative amounts of the two precursor ions.



(14)

A further step in understanding the selectivity factors between the two competing pathways results from an evaluation of the k_{en}/k_{eth} ratios. In ref 2 we assumed that k_{en} and k_{eth} are almost identical (because the two ions are similar) and deduced a K_I^{II} value for acetone (0.07). Similar results can be obtained for acetophenones with $K_I{}^{II}$ ranging from 0.027 (**1a**) to 0.147 (**1h**). However, recent data on the rate constants for the two reverse processes [namely, enol ketonization (k_{-en}) and methyl enol ether hydrolysis (k_{-eth})], which have shown that the k_{-en}/k_{-eth} ratios in the water range from 67 (**1a**) to 16 (**1h**),[16–18] suggest that this assumption is too crude and that variations in k_{en}/k_{eth} should be taken into account. These ratios can be evaluated from k_{-en}/k_{-eth} values since the following pieces of evidence suggest the existence of a general Marcus–Brønsted relationshp accounting for variations in the catalyst and in the substrate for both enol and enol ether formations: (i) k_{eth} depends on protonation equilibrium constants according to a Brønsted relation with a β exponent (0.63) very close to the α values observed when the catalyst is varied,[21] (ii) α values for enol ketonization and enol ether hydrolysis are closely similar,[3] and (iii) there is a linear relationship in water between the log k_{en} and log K_I^{en}, with a slope (0.4) in agreement with that expected.[3,16] By assuming a mean value of 0.6 for the Brønsted exponent and a similar behavior in methanol and water, it follows that the k_{-en}/k_{-eth} data result in k_{en}/k_{eth} ratios in the 0.06 (**1a**) to 0.16 (**1h**) range and that K_I^{II} varies from 0.0016 to 0.023. The proportions of path A and B depend on rate and equilibrium factors which partly compensate each other with, however, a predominance of the effects of K_I^{II}. The fact that the enol path is more favored for **1a** than for **1h**, for example, is due to a larger difference in the stability of the two ions. An electron-withdrawing substituent makes the two ions less stable but also increases the difference between the stabilizing effects of the OH and OCH_3 groups.

Bromination in Methanol vs. Bromination in Water. Ketone bromination in methanol is always very much faster than ketone bromination in water (e.g., for **1f** in methanol containing 0.02 M water, $k_\psi/[CH_3OH_2{}^+] = 3.38 \times 10^{-2}$ M^{-1} s^{-1}, whereas $k_\psi/[H_3O^+] = 1.22 \times 10^{-5}$ M^{-1} s^{-1} in water).[16,17b] This is mainly due to the differences in reaction mechanisms and probably to the fact that enol ether is much more stable in methanol than enol in water (e.g., [enol ether]/[ketone] = 6.4 × 10^{-4} for **1f** in methanol containing 0.02 M water, whereas [enol]/[ketone] = (1.1–1.2) × 10^{-8} for **1f** in water).[17] It is more surprising that the rate constants for enol formation in methanol are rather larger than in water. For example, the ratio of the rate constants for H_3O^+- and $CH_3OH_2{}^+$-catalyzed enolization of acetone in H_2O and CH_3OH, respectively, is 60; its value is 24, 27, and 30 for **1a**, 1b, and **1c**, respectively.

Substituent Effects on Rates and Equilibria. Table IV lists the ρ^n and ρ^r parameters calculated from the Young–Jencks modified Yukawa–Tsuno–Sawada equation (eq 15).[22–25] According to Young and Jencks, this equation

$$\log k \text{ or } \log K = \rho^n\sigma^n + \rho^r(\sigma^+ - \sigma^n) + i \quad (15)$$

makes it possible to separate polar effects, due to the substituted ring, from resonance effects stemming from conjugation between the reaction center and the substituted ring. In most cases a one-parameter treatment using σ^n (or σ^o)[1] or σ^+ is insufficient to deal with the reported data.

The low ρ^n value observed for the acetal-to-enol ether equilibrium shows that polar effects on acetal and enol

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(21) Toullec, J. *Tetrahedron Lett.* **1979**, 3089–3092. Toullec, J.; Kleffert, P., to be published.

(22) (a) Young, P. R.; Jencks, W. P. *J. Am. Chem. Soc.* **1977**, *99*, 8238–8247; (b) **1979**, *101*, 3288–3294. See ref 8, 10, and 16 and: Young, P. R.; McMahon, P. E. *J. Am. Chem. Soc.* **1979**, *101*, 4678–4681.

(23) Yukawa, Y.; Tsuno, Y.; Sawada, M. *Bull. Chem. Soc. Jpn.* **1966**, *39*, 2274–2286.

(24) The σ^n values are those of van Bekkum et al. (van Bekkum, H.; Verkade, P. E.; Wepster, B. M. *Recl. Trav. Chim. Pays-Bas* **1959**, *78*, 815–850): −0.111 (4-OCH_3), −0.129 (4-CH_3), 0.168 (4-F), 0.238 (4-Cl), 0.373 (3-Cl), 0.467 (3-CF_3), 0.710 (3-NO_2).

(25) The σ^+ parameters are those of Brown and Okamoto (Brown, H. C.; Okamoto, Y. *J. Am. Chem. Soc.* **1958**, *80*, 4979–4987): −0.778 (4-OCH_3), −0.311 (4-CH_3), −0.073 (4-F), 0.114 (4-Cl), 0.399 (3-Cl), 0.520 (3-CF_3), 0.674 (3-NO_2).

ether energies are small or similar, and the low ρ^r value shows that the conjugation involving the enol ether double bond and the substituted ring is of minor importance. This latter result was expected because of the predominant conjugation between the double bond and the alkoxy group.

The fairly large ρ^n values observed (1.45 in methanol and 0.79 in water) for ketone-to-enol ether equilibria can easily be ascribed to the large polar effects on carbonyl compound stability;[26] the ρ^r value, close to unity (0.88 in methanol, 0.86 in water), is typical of what is usually observed for reactions of aromatic carbonyl compounds yielding products without significant through-resonance effects.[8] The $RT\rho^r(\sigma^+ - \sigma^n)$ term represents the loss in resonance energy on going from acetophenones to α-methoxystyrenes. As observed for acetal formation,[8] ρ^r does not depend significantly on solvent.

The rate constant for enol ether formation is strongly depressed by electron-withdrawing substituents. These large effects suggest that the transition state to some extent resembles the alkoxycarbenium ion intermediate. Bromination of CH_3COCH_3 and CD_3COCD_3 shows a large primary kinetic isotope effect (6.9),[2] closely similar to that observed for bromination in water (6.7).[27] This indicates that the change in solvent does not significantly alter the degree of proton transfer in the transition state, even though bromination in methanol essentially involves enol ether formation from II instead of enol formation from I. It follows that the conclusion that the proton is about half-transferred in the transition state for enolization in water should also be valid for enol ether formation in methanol. Since it can be assumed that the difference in solvents should not modify conjugative effects significantly,[8] it is of interest to compare the ρ^r value calculated for the rate of enol ether formation (−1.39) with that estimated by Young and Jencks for the acetal-to-ion II equilibria in water (−2.2);[22a] the ratio 1.39/2.2 (=0.63) is close to that expected when one assumes a transition state with a little less than half-transferred proton. In contrast, the ρ^n value calculated from $(k_B)_m{}^\circ$ data (−3.12) is only a little smaller than that estimated for ion II formation (−3.6); this should not invalidate the conclusion that the proton is about half-transferred at the transition state, but should instead be related to the fact that the differences in dielectric properties of methanol and water probably make the ρ^n value for the acetal–ion II equilibrium lower in water than in methanol. It is also noteworthy that a close analogy in rate constants was observed between proton addition to α-methoxystyrenes in methanol and that in water[1] and that α Brønsted exponents ($\approx$0.6), solvent isotope effects, and rate–equilibria relationships observed for enol ether protonation in water indicate that the proton is only a little more than half-transferred in the transition state.[3,21,28]

The ρ^n and ρ^r values observed when eq 15 is applied to the rate constants for enol ether formation are very close to those observed for acetal hydrolysis in methanol containing small amounts of water ($\rho^n = -3.3$; $\rho^r = -1.1$).[10] This suggests that the transition states for these two reactions, which correspond to water addition to and proton abstraction from the same ions, are closely analogous as far as charge distribution is concerned.

The small contribution of the enol formation path to the rate of acetophenone bromination in methanol precludes a detailed examination of substituent effects on $(k_A)_m$. For the three values derived, ρ^n can be calculated as −0.8. This slope is significantly larger than that derived from enolization rate constants in water which fit a linear log k–σ^n relationship with $\rho^n = -0.37$;[16] it probably means that the cancellation of polar effects on the transition state by those on the starting materials is less complete than in water. This could be a consequence of a larger substituent dependence of the stability of the hydroxycarbenium ion in methanol than in water or of a little closer resemblance between the transition state and the ionic intermediate.

Summary

We have shown that acid-catalyzed bromination of acetophenones in slightly aqueous methanol occurs predominantly via enol ethers instead of via enols but there is a reverse relationship between the amount of acetal at equilibrium and the proportion of enol ether path. When acetal is the more abundant species, as in the case for acetophenones bearing withdrawing substituents, the enol path is followed to a significant extent, but when acetal is very unstable the enol ether path is highly favored. This unexpected behavior is explained by considering substituent effects on the stability ratio between hydroxy and methoxycarbenium ion intermediates. Substituent effects on enol and enol ether formation rates are accounted for by transition states with some resemblance to the intermediate ions. The derivation of equilibrium constants for the ternary ketone–acetal–enol ether system, both in methanol and in water, has shown that substituent effects on equilibria stem mainly from those on carbonyl compound stability. This result is consistent with that found for enol formation in water.

Experimental Section

Reagents. All the ketones **1a–h** are commercially available. Compounds **1b–g** were purified by VPC (Varian Aerograph Series 700). The purity was always better than 99.5% and was checked by VPC (Varian Aerograph Series 1200, $^3/_8$ in. × 10 ft column packed with 30% Carbowax on Chromosorb W). Compounds **1a** and **1h** were purified by recrystallization from hot methanol, mp 39–40 and 80 °C, respectively. **2a** was prepared by the reaction of trimethyl orthoformate in methanol with **1a**[29] and recrystallized from hot methanol (mp 44–45.5 °C). Methanol (Baker Analyzed Reagent) was distilled with magnesium under nitrogen. Methanolic solutions of hydrobromic acid were prepared by bubbling gaseous HBr (99.8%, Baker Chemicals) under nitrogen and subsequent dilution in a drybox. Bulk solutions were stored at −18 °C. Volumetric acid titrations were performed with aqueous solutions of sodium hydroxide (indicator: methyl red). The water content of methanolic HBr solutions without deliberately added water or when the added amounts of water were small was determined, as previously described,[30] by measuring the relative concentrations of **1g** and **2g** at equilibrium by UV spectroscopy. Water concentrations were calculated from the absorbance at 252 nm by eq 16, with $\epsilon_{ke} = 14\,600\ M^{-1}\ cm^{-1}$, $\epsilon_{ac} = 190\ M^{-1}\ cm^{-1}$, and

$$[H_2O] = k_{ke}^{ac}a_m{}^2(A_{252} - \epsilon_{ac}[\text{ketone}]_T)/(\epsilon_{ke}[\text{ketone}]_T - A_{252}) \quad (16)$$

$K_{ke}^{ac} = 5.94 \times 10^{-3}$ M.[8] It can be noted that the thus-determined water concentrations did not include water associated with the proton in H_3O^+. Standard deviations on water determinations were 2–3 mg L^{-1}. When water was deliberately added to HBr solutions, the water concentration was calculated from the concentration in the absence of added water and from the amount added. It was taken into account that the water addition provoked a displacement of the $CH_3OH_2{}^+/H_3O^+$ equilibrium (eq 3) and that the water concentration should accordingly be corrected. Water

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[27] Toullec, J.; Dubois, J.-E. *J. Am. Chem. Soc.* **1974**, *96*, 3524–3532.

[28] Kresge, A. J.; Sagatys, D. S.; Chen, H. L. *J. Am. Chem. Soc.* **1977**, *99*, 7228–7233.

[29] Loudon, G. M.; Smith, C. K.; Zimmerman, J. E. *J. Am. Chem. Soc.* **1974**, *96*, 465–479.

[30] Toullec, J.; El-Alaoui, M. *Anal. Chim. Acta* **1979**, *109*, 187–190.

concentrations were also corrected for water formed by acetal formation, both in water determinations and in kinetic measurements.

Rate Measurements. The two-electrode electrochemical system used for the kinetic measurements was described previously.[2,9,31] Argon was passed over the solution and then interrupted to avoid evaporation of bromine during the reaction. Bromine was produced by bromide ion electrolysis (constant current and measurement of the electrolysis time). It was observed that in the absence of added ketone, bromine concentration (ca. 10^{-6} M) decreased fairly rapidly (see Figure 1) with a roughly zeroth-order rate which depended on acid and water concentrations. The lower the water concentration and the higher the HBr concentration, the faster the blank reaction was; consequently, kinetic studies were restricted to narrow ranges of acid concentrations such that the largest water concentration effects could be examined. Another restriction on the choice of acid concentrations also came from the necessity to make the ketone-to-acetal equilibrium sufficiently fast. The ratios between $[Br_2]_{st}$ and i_{amp} (ca. 10^{-7} M/μA) were determined by jumps in bromine concentration due to successive electrolyses and i_{amp} measurements. In most cases, since i_{amp} was not strictly linear with concentration, corrections were introduced for rate constant determinations [this was due to the fact that the fixed electrode (anode) did not stay at a rigorously constant potential, contrary to what is required by the amperometric method]. Aliquots of the methanolic HBr solutions were transferred to the UV–vis spectrophotometry cell for water determination just before the kinetic measurements or just before water addition. Solutions of **1a**–**h** (15–50 μL) in chlorobenzene were introduced into the cell containing 20–40 mL of the acidic solution of bromine by using a weighed microsyringe. For measurements on equilibrated ketone–acetal systems, the mixture was allowed to equilibrate for some minutes (this time lag was calculated from kinetic data on ketone–acetal equilibration),[10] bromine was produced by electrolysis, and the i_{amp}–time curve was recorded (Esterline Angus, 1101S model). The zeroth-order rate constant was calculated from the slope and corrected for the blank reaction. The first-order rate constant was obtained by dividing the corrected slope by the overall ketone concentration. For most experiments, the fraction of overall ketone which reacted before the kinetic run itself was small, but in some cases it was necessary to correct the concentration. It was assumed that the α-bromo ketone dimethyl acetal formed was much less reactive than the starting materials, as expected from the polar effects of the bromine atom. Usually several successive runs could be performed in the same reaction mixture, after successive electrolyses, without significant changes in the first-order rate constants being observed.

Registry No. **1a**, 121-89-1; **1b**, 349-76-8; **1c**, 99-02-5; **1d**, 99-91-2; **1e**, 403-42-9; **1f**, 98-86-2; **1g**, 122-00-9; **1h**, 100-06-1; **2a**, 73585-54-3; **2b**, 73589-85-2; **2c**, 73585-53-2; **2d**, 72360-69-1; **2e**, 73585-52-1; **2f**, 4316-35-2; **2g**, 53578-01-1; **2h**, 27150-99-8; CH_3OH, 67-56-1; H_2O, 7732-18-5; $4\text{-}CH_3OC_6H_4C(OCH_3){=}CH_2$, 51440-56-3; $4\text{-}CH_3C_6H_4C(OCH_3){=}CH_2$, 51440-57-4; $C_6H_5C(OCH_3){=}CH_2$, 4747-13-1; $4\text{-}FC_6H_4C(OCH_3){=}CH_2$, 67471-38-9; $4\text{-}ClC_6H_4C(OCH_3){=}CH_2$, 67471-39-0; $3\text{-}ClC_6H_4C(OCH_3){=}CH_2$, 89726-07-8; $3\text{-}O_2NC_6H_4C(OCH_3){=}CH_2$, 89726-05-6.

Supplementary Material Available: Tables of primary kinetic results (Tables S1–S3) (8 pages). Ordering information is given on any current masthead page.

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(*E*)- and (*Z*)-1-Hydroxy-1,3-butadiene: New Kinetically Unstable C_4H_6O Isomers

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(*E*)- and (*Z*)-1-hydroxy-1,3-butadiene were prepared by the retro-Diels–Alder decomposition of the corresponding 3-*exo*-vinylbicyclo[2.2.1]hept-5-en-2-ols at 750 °C (10^{-6} torr). The ionization energies of the *E* and *Z* isomers were determined as 8.51 ± 0.03 and 8.47 ± 0.03 eV, respectively. The ionized (*E*)- and (*Z*)-dienols were generated from the same precursors by dissociative ionization, and their appearance energies were measured as 9.00 ± 0.05 and 8.95 ± 0.06 eV, respectively. The heats of formation of the ionic and neutral dienols were assessed as $\Delta H_f{}^\circ{}_{298}$ $[(E)\text{-}CH_2{=}CHCH{=}CHOH]^{\bullet +})$ = 733 ± 5 kJ·mol^{-1}, $\Delta H_f{}^\circ{}_{298}$ $((E)\text{-}CH_2{=}CHCH{=}CHOH)$ = −88 ± 5 kJ·mol^{-1}, $\Delta H_f{}^\circ{}_{298}$ $([(Z)\text{-}CH_2{=}CHCH{=}CHOH]^{\bullet +})$ = 728 ± 6 kJ·mol^{-1}, and $\Delta H_f{}^\circ{}_{298}$ $((Z)\text{-}CH_2{=}CHCH{=}CHOH)$ = −89 ± 9 kJ·mol^{-1}, in good agreement with semiempirical MNDO calculations. The neutral dienols are less stable than (*E*)-2-butenal ($\Delta\Delta H_f{}^\circ{}_{298}$ = 17–19 kJ·mol^{-1}), but *more stable* than 3-butenal ($\Delta\Delta H_f{}^\circ{}_{298}$ = −8 to −10 kJ·mol^{-1}). The transition state for the unimolecular 1,5-sigmatropic isomerization of (*Z*)-1-hydroxy-1,3-butadiene to (*Z*)-2-butenal was located by MNDO calculations at 236 kJ·mol^{-1} above the dienol. The corresponding barrier to isomerization in the cation radical was assessed as being lower than 187 kJ·mol^{-1}. The mass and collision-induced decomposition spectra are briefly discussed.

Prototropic isomerization of enols is thought to begin with proton attack at the electron-rich terminus of the enol system.[1] In dienols containing a terminal hydroxy group there are, in principle, two possibilities for the enol–keto conversion (Scheme I). Protonation of the dienol system at C-4, followed by hydroxyl hydrogen abstraction would yield an α,β-unsaturated aldehyde (e.g., 2-butenal, **2**, R = H), while protonation of the enol subsystem at C-2 would

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