

Physical Organic Chemistry of Transition Metal Carbene Complexes. 13.¹ Kinetics of Proton Transfer from (5-Methyl-2-oxacyclopentylidene)pentacarbonylchromium(0) and Hydrolysis of Its Conjugate Anion in Aqueous Acetonitrile

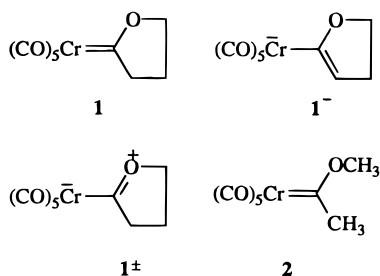
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The $pK_a = 13.41$ of the title compound (**5**), determined by a kinetic method, is about 1.3 units lower than the pK_a of (2-oxacyclopentylidene)pentacarbonylchromium(0) (**1**) in 50% MeCN–50% water at 25 °C. The acidifying effect of the methyl group is attributed to its stabilizing effect on the C=C double-bond resonance structure of the anion (**5**⁻). The rate constant for deprotonation of **5** by OH⁻ is about the same as for deprotonation of **1**, despite the higher acidity of **5**. This means that the *intrinsic* rate constant for proton transfer from **5** is significantly lower than that from **1**. This reduction in the intrinsic rate constant is mainly the result of the imbalanced nature of the transition state which deprives the transition state from the stabilizing effect of the methyl group. Based on precedents, the most likely mechanism for the hydrolysis of **5** involves rate-limiting protonation of **5**⁻ concerted with metal–carbon bond cleavage. Even though the relatively small kinetic solvent isotope effect cannot rigorously exclude an alternative nucleophilic substitution mechanism, a comparison of the reactivity of **5** and **1** allows one to rule out this latter mechanism.

A study of the thermodynamic and kinetic acidity of (2-oxacyclopentylidene)pentacarbonylchromium(0), **1**, in 50% acetonitrile–50% water has been reported recently.² The pK_a of **1** was found to be 14.47, which is

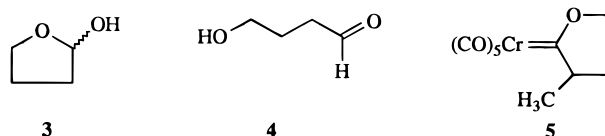


about two units higher than the pK_a of **2** (12.50) in the same solvent.³ This reduction in acidity results from the combined effect of a lower rate of deprotonation of **1** and an enhanced rate of reprotonation of the conjugate anion (**1**⁻), which was attributed to extra stabilization of the neutral carbene complex through π -donation from the ring oxygen (**1**[±]). This π -donor effect is more effective in stabilizing **1** than the π -donor effect of the methoxy group in **2**, mainly because the cyclic structure locks the oxygen into a position of better π -overlap with the (CO)₅Cr moiety. Such enhanced π -donation in **1**[±]

had been suggested based on ⁵³Cr NMR data reported prior to the pK_a determination.⁴

The study also provided kinetic data on the hydrolysis of **1**, which, presumably via **1**⁻, leads to the formation of 2-hydroxytetrahydrofuran (**3**) in equilibrium with small amounts of 4-hydroxybutanal (**4**).⁵

In this paper we report a similar investigation of (5-methyl-2-oxacyclopentylidene)pentacarbonylchromium(0), **5**. There are two main points of interest. The first



is to evaluate the effect of the methyl group on the thermodynamic and kinetic acidity of **5**. According to Casey and Brunsvold,⁶ the thermodynamic acidities of **5** and **1** in THF are approximately the same. On the other hand, in a recent study which compared **6** and **7** with **2**, it was found that in aqueous acetonitrile the α -methyl groups enhance the thermodynamic acidities of **6** and **7** significantly, while the rate constants for the deprotonation of **6** and **7** by various bases are *lower* than

(1) Part 12: Bernasconi, C. F.; Flores, F. X.; Kittredge, K. W. *J. Am. Chem. Soc.* **1998**, *120*, 7983.

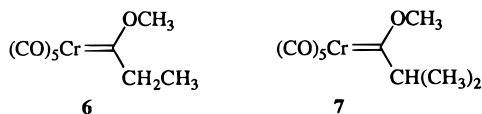
(2) Bernasconi, C. F.; Leyes, A. E. *J. Am. Chem. Soc.* **1997**, *117*, 5169.

(3) Bernasconi, C. F.; Sun, W. *J. Am. Chem. Soc.* **1993**, *115*, 12526.

(4) Hafner, A.; Hegedus, L. S.; de Weck, G.; Hawkins, B.; Dötz, K. H. *J. Am. Chem. Soc.* **1988**, *110*, 8413.

(5) Bernasconi, C. F.; Leyes, A. E. *J. Chem. Soc., Perkin Trans. 2* **1997**, 1641.

(6) Casey, C. P.; Brunsvold, W. R. *J. Organomet. Chem.* **1976**, *118*, 309.

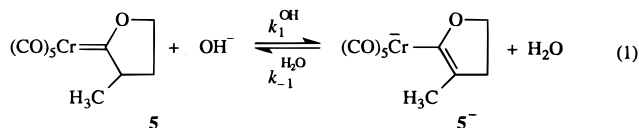


for the deprotonation of **2**.⁷ We wish to examine whether in aqueous acetonitrile the methyl group has a similar effect on the pK_a and proton-transfer rate constants of **5** relative to **1**, in which case the findings in THF⁶ would not be representative for the behavior in more polar solvents.

The second point of interest is the effect of the methyl group on the rate of hydrolysis of **5** and whether this can shed some light on a mechanistic ambiguity in the hydrolysis of **1**.⁵

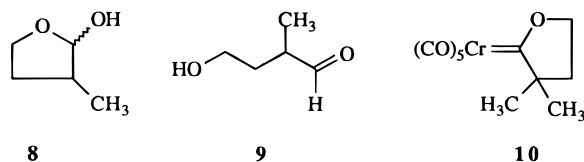
Results

General Features. The reaction of **5** with KOH in 50% MeCN–50% water (v/v) is characterized by two distinct kinetic processes. Both manifest themselves by a decrease in absorbance around 370 nm, which is quite similar to the behavior of **1**. The first reaction occurs on a time scale of a few seconds or less and represents reversible deprotonation of **5** by OH^- , eq 1. This process

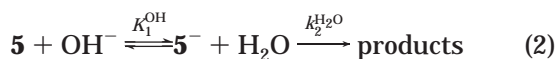


can be observed at KOH concentrations above 2×10^{-3} M; at lower concentrations the associated change in absorbance becomes too small to be measurable, indicating that the equilibrium disfavors 5^- .

The second process is substantially slower, with half-lives of several seconds to several minutes depending on [KOH]. No product analysis was performed for this process, but as will be shown below, its kinetic characteristics are consistent with hydrolysis of **5**, presumably via 5^- , to form **8** and/or **9** and $(\text{CO})_5\text{CrOH}^-$, the same



type of reaction shown to occur with **1**,⁵ **2**,⁸ and several other carbene complexes^{8,9} subjected to the same conditions. This process is described by eq 2, where the proton-transfer step now represents a fast preequilibrium.



Further changes in the absorbance occur on a still longer time scale and probably involve reactions of $(\text{CO})_5\text{CrOH}^-$. This process was not investigated further.

(7) Bernasconi, C. F.; Sun, W.; García-Río, L.; Yan, K.; Kittredge, K. W. *J. Am. Chem. Soc.* **1997**, *119*, 5583.

(8) Bernasconi, C. F.; Flores, F. X.; Sun, W. *J. Am. Chem. Soc.* **1995**, *119*, 4875.

(9) Bernasconi, C. F.; Sun, W. *Organometallics* **1995**, *14*, 5615.

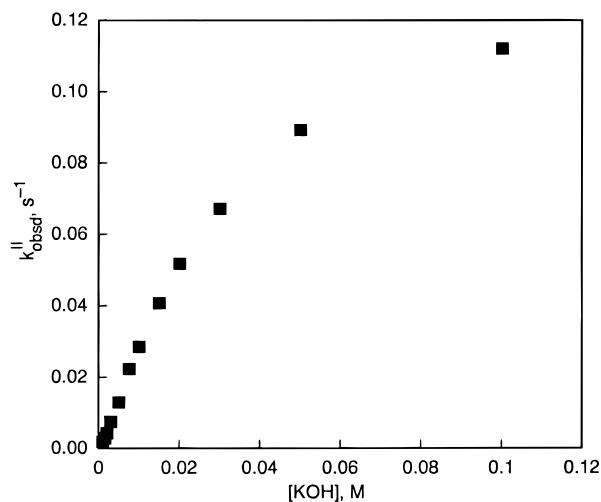


Figure 1. Plot of $k_{\text{obsd}}^{\text{II}}$ vs [KOH] according to eq 4.

One complication present in our study is that our sample of **5** contained approximately 12% of the dimethyl derivative, **10**; attempts to remove this contaminant were not successful.¹⁰ Since **10** cannot lose a proton, its presence does not affect the kinetics of proton transfer. On the other hand, **10** undergoes slow hydrolysis, which has the potential of affecting the kinetic data of **5**. However, independent measurements of the hydrolysis of **10** showed that its rate is much slower than that of the hydrolysis of 5^{11} and hence does not interfere with the kinetics of this latter reaction.

Kinetics in KOH Solutions. The pseudo-first-order rate constants for proton transfer ($k_{\text{obsd}}^{\text{I}}$) show a linear dependence on KOH concentration between 0.002 and 0.1 M (Figure 1S of the Supporting Information).¹² A fit to eq 3 yields $K_1^{\text{OH}} = 39.5 \pm 0.4 \text{ M}^{-1} \text{ s}^{-1}$ and $k_{-1}^{\text{H}_2\text{O}} = 0.66 \pm 0.04 \text{ s}^{-1}$.

$$k_{\text{obsd}}^{\text{I}} = K_1^{\text{OH}}[\text{OH}^-] + k_{-1}^{\text{H}_2\text{O}} \quad (3)$$

From $K_1^{\text{OH}} = k_1^{\text{OH}}/k_{-1}^{\text{H}_2\text{O}} = 59.8 \pm 4.5 \text{ M}^{-1}$ one obtains $K_a^{\text{CH}} = K_1^{\text{OH}} K_w = (3.86 \pm 0.29) \times 10^{-14}$ or $pK_a^{\text{CH}} = 13.41 \pm 0.03$.¹³ A plot of $k_{\text{obsd}}^{\text{II}}$ for the hydrolysis reaction is shown in Figure 1. A fit to eq 4 based on the scheme of

$$k_{\text{obsd}}^{\text{II}} = \frac{K_1^{\text{OH}} k_2^{\text{H}_2\text{O}} [\text{OH}^-]}{1 + K_1^{\text{OH}} [\text{OH}^-]} \quad (4)$$

eq 2 yields $K_1^{\text{OH}} = 21.3 \pm 1.5 \text{ M}^{-1}$ and $k_2^{\text{H}_2\text{O}} = 0.168 \pm 0.017 \text{ s}^{-1}$. A second set of data (not shown) yielded $K_1^{\text{OH}} = 26.0 \pm 3.0 \text{ M}^{-1}$ and $k_2 = 0.154 \pm 0.024 \text{ s}^{-1}$. The agreement between K_1^{OH} obtained from eq 4 and that from $k_1^{\text{OH}}/k_{-1}^{\text{H}_2\text{O}}$ is reasonably good. We shall use K_1^{OH} determined as $k_1^{\text{OH}}/k_{-1}^{\text{H}_2\text{O}}$ and the average k_2 value ($0.161 \pm 0.016 \text{ s}^{-1}$) in our further discussions.

Kinetics of Hydrolysis in Buffer and HCl Solutions. The hydrolysis was also investigated in triethyl-

(10) Wulff, W. D. Personal communication (**5** was a gift from Professor W. D. Wulff).

(11) For example, at 0.1 M KOH, $k_{\text{obsd}} = 0.112 \text{ s}^{-1}$ and $6.36 \times 10^{-3} \text{ s}^{-1}$ for hydrolysis of **5** and **10**, respectively; at 3×10^{-3} M KOH, $k_{\text{obsd}} = 7.40 \times 10^{-3} \text{ s}^{-1}$ (**5**) and $6.73 \times 10^{-5} \text{ s}^{-1}$ (**10**), respectively.

(12) See paragraph concerning Supporting Information at the end of this paper.

(13) $pK_w = 15.19$.³

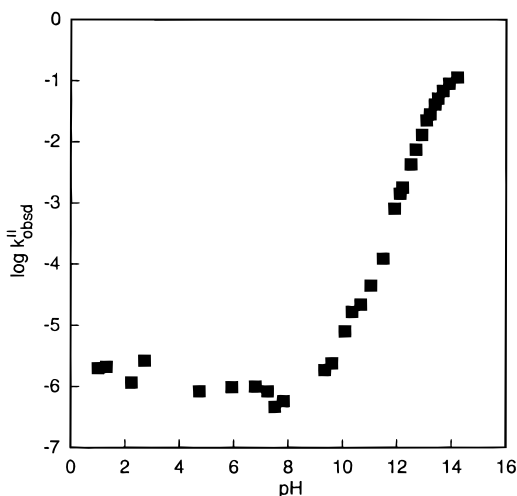
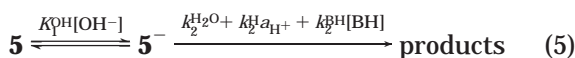


Figure 2. pH-profile for the hydrolysis of **5**. Points obtained in buffered solutions (pH ≤ 11.47) have been extrapolated to zero buffer concentration.

amine (pH 11.47, 11.27, 11.02, 10.66, 10.34, 10.08, 9.59, and 9.33), *N*-methylmorpholine (pH 8.38, 7.81, 7.49, 7.23, and 6.78), acetate (pH 5.92), and methoxyacetate buffers (pH 4.72) as well as in HCl solutions (pH 2.70, 2.22, 1.92, 1.62, 1.30, and 1.00). Figure 2 shows a pH-rate profile of all hydrolysis data, including those obtained in KOD solutions.

In buffer solutions the reaction is subject to general base catalysis; the data in Figure 2 represent $k_{\text{obsd}}^{\text{II}}$ values that have been extrapolated to zero buffer concentration. In some cases catalysis was substantial, which led to uncertainties in the extrapolated values and is probably the main reason for the significant scatter in the pH-rate profile. Another factor that contributed to the experimental error in the pH 1–8 region is the extreme slowness of the reaction, rendering a clear separation from the third kinetic process more difficult.

The results in buffer and in HCl solutions are consistent with two additional pathways for the conversion of **5**[−] to products, namely, general acid catalysis by hydronium ion ($k_2^{\text{H}} a_{\text{H}^+}$) and by buffer acid ($k_2^{\text{BH}} [\text{BH}]$), as shown in eq 5; $k_{\text{obsd}}^{\text{II}}$ is then given by eq 6, which for



$K_1^{\text{OH}}[\text{OH}] \ll 1$ (pH $\ll \text{p}K_a^{\text{CH}}$) simplifies to eq 7. The

$$k_{\text{obsd}}^{\text{II}} = \frac{K_1^{\text{OH}}[\text{OH}^-]}{1 + K_1^{\text{OH}}[\text{OH}^-]} (k_2^{\text{H}_2\text{O}} + k_2^{\text{H}} a_{\text{H}^+} + k_2^{\text{BH}} [\text{BH}]) \quad (6)$$

$$\begin{aligned} k_{\text{obsd}}^{\text{II}} &= K_1^{\text{OH}}[\text{OH}] (k_2^{\text{H}_2\text{O}} + k_2^{\text{H}} a_{\text{H}^+} + k_2^{\text{BH}} [\text{BH}]) \\ &= K_1^{\text{OH}} k_2^{\text{H}_2\text{O}} [\text{OH}^-] + K_a^{\text{CH}} k_2^{\text{BH}} + \frac{K_a^{\text{CH}}}{K_a^{\text{BH}}} k_2^{\text{BH}} [\text{B}] \quad (7)^{14} \end{aligned}$$

plateau from pH 1–8 thus corresponds to $K_a^{\text{CH}} k_2^{\text{H}}$, the sloping portion corresponds to $K_1^{\text{OH}} k_2^{\text{H}_2\text{O}} [\text{OH}^-]$, and the slopes of the plots of $k_{\text{obsd}}^{\text{II}}$ vs buffer base concentration

(14) Note: $K_1^{\text{OH}}[\text{OH}^-] k_2^{\text{H}} a_{\text{H}^+} = K_1^{\text{OH}} K_w k_2^{\text{H}} = K_a^{\text{CH}} k_2^{\text{H}}$; $K_1^{\text{OH}}[\text{OH}^-] k_2^{\text{BH}} [\text{BH}] = K_1^{\text{OH}}[\text{OH}^-] k_2^{\text{BH}} (a_{\text{H}^+}/K_a^{\text{BH}}) [\text{B}] = (K_a^{\text{CH}}/K_a^{\text{BH}}) k_2^{\text{BH}} [\text{B}]$.

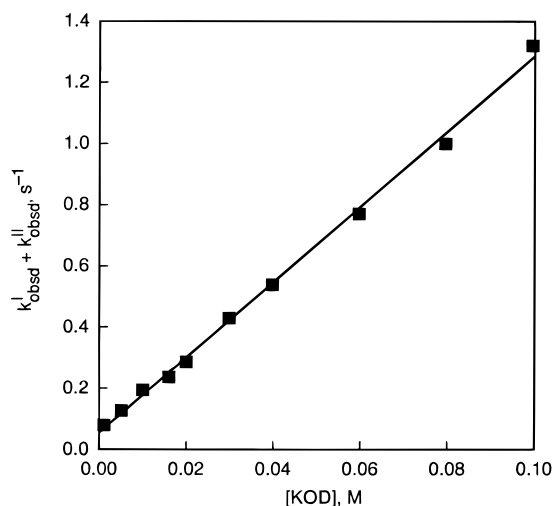
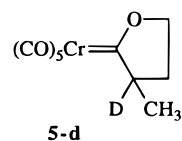


Figure 3. Plot of $k_{\text{obsd}}^{\text{I}} + k_{\text{obsd}}^{\text{II}}$ vs $[\text{KOD}]$ according to eq 8.

correspond to $K_a^{\text{CH}} k_2^{\text{BH}}/K_a^{\text{BH}}$. The k_2^{H} , $k_2^{\text{H}_2\text{O}}$, and k_2^{BH} values are summarized in Table 2.

Kinetics of the Reaction of **5-d** in KOD Solutions.

A number of experiments were carried out with the deuterated substrate, **5-d**, in KOD–MeCN–D₂O. The



$k_{\text{obsd}}^{\text{I}}$ and $k_{\text{obsd}}^{\text{II}}$ values are reported in Table 1S.¹² Compared to the reactions of **5** in KOH–MeCN–H₂O, $k_{\text{obsd}}^{\text{I}}$ is substantially reduced, while the reduction in $k_{\text{obsd}}^{\text{II}}$ is much smaller. As a result, the $k_{\text{obsd}}^{\text{I}}/k_{\text{obsd}}^{\text{II}}$ ratios at $[\text{KOH}]$ between 0.005 and 0.04 M become relatively small (< 10), implying significant coupling between the two kinetic processes. This makes eq 3 and 4 less than satisfactory approximations for the first-order rate constants. A better way to treat the data is by plotting the sum and product, respectively, of $k_{\text{obsd}}^{\text{I}}$ and $k_{\text{obsd}}^{\text{II}}$ vs $[\text{KOD}]$;¹⁵ as seen in Figures 3 and 4, such plots yield straight lines according to eqs 8 and 9. Note that in

$$k_{\text{obsd}}^{\text{I}} + k_{\text{obsd}}^{\text{II}} = k_{1,\text{D}}^{\text{OD}}[\text{KOD}] + k_{-1}^{\text{D}_2\text{O}} + k_2^{\text{D}_2\text{O}} \quad (8)$$

$$k_{\text{obsd}}^{\text{I}} k_{\text{obsd}}^{\text{II}} = k_{1,\text{D}}^{\text{OD}} k_2^{\text{D}_2\text{O}} [\text{KOD}] \quad (9)$$

eqs 8 and 9 the superscripts refer to the deuterated solvent and base, while the subscript in $k_{1,\text{D}}^{\text{OD}}$ refers to the deuterated carbene complex. The plots of Figures 3 and 4 yield $k_{1,\text{D}}^{\text{OD}} = 12.7 \pm 0.2 \text{ M}^{-1} \text{ s}^{-1}$, $k_2^{\text{D}_2\text{O}} = (5.5 \pm 0.3) \times 10^{-2} \text{ s}^{-1}$, and $k_{-1}^{\text{D}_2\text{O}} = (4.5 \pm 1.0) \times 10^{-2} \text{ s}^{-1}$.

Discussion

Thermodynamic Acidity of 5. The $\text{p}K_a^{\text{CH}}$ of **5** is approximately 1.3 $\text{p}K_a$ units lower than the statistically corrected $\text{p}K_a^{\text{CH}}$ of **1** (Table 1), indicating a substantial acidifying effect of the α -methyl group. This is reminiscent of the acidity order **7** (12.27) $>$ **6** (12.62) $>$ **2** (12.98) in 50% MeCN–50% water,⁷ with the numbers

(15) Bernasconi, C. F. *Relaxation Kinetics*; Academic Press: New York, 1976; p 28.

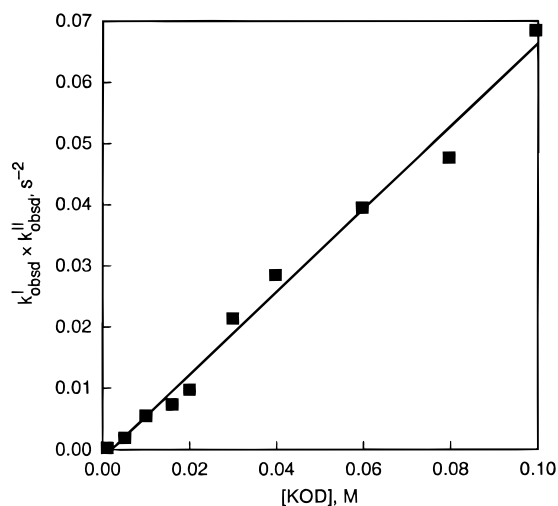


Figure 4. Plot of $k_{\text{obs}}^I / k_{\text{obs}}^{II}$ vs [KOD] according to eq 9.

in parentheses referring to the statistically corrected pK_a^{CH} values. As in the case of **7** and **6**, the most plausible explanation of the acidifying effect of the methyl group is that it reflects a stabilization of the anion whose dominant resonance form is the C=C double-bond structure **5**⁻. The increased stability of **5**⁻ relative to **1**⁻ is the result of the well-known stabilization of alkenes by methyl groups,¹⁶ as discussed in more detail elsewhere.⁷

The fact that the pK_a^{CH} difference between **5** and **1** is significantly larger than that between **6** and **2**, or **7** and **6**, is probably due to the cyclic structure of **5**⁻ and **1**⁻, which locks the sp²-hybridized α -carbon into a position of better π -overlap with the (CO)₅Cr moiety and enhances the dominance of the C=C double-bond resonance structure **5**⁻. Note this is the same reasoning given for the enhanced π -donation by the oxygen in **1**[±], which explains the high pK_a^{CH} of **1** relative to that of **2**.²

The strongly enhanced acidity of **5** compared to that of **1** in 50% MeCN–50% water contrasts with the observation that in THF the ionization constants of **1** and **5** are approximately the same.⁶ These latter ionization constants were determined by reacting the carbene complexes with the PPN⁺ salt¹⁷ of *p*-cyanophenol. In view of the fact that in THF ion pairing is very strong, the “leveling” of the acidities is probably the result of differences in the tightness of the **1**⁻/PPN⁺ and **5**⁻/PPN⁺ ion pairs. That ion pairing in THF can have a dramatic effect on acidities of carbene complexes is even more apparent when comparing the “ pK_a ”¹⁸ value of **2** determined by using Li⁺/MeO⁻ as the proton acceptor (“ pK_a ” = 12)¹⁹ with that determined using PPN⁺/4-CNC₆H₄O⁻ (“ pK_a ” = 8).²⁰

Rates of Proton Transfer. The rate constant for deprotonation of **5** by OH⁻ is about the same as the statistically corrected k_1^{OH} value for **1** (Table 1). Since

(16) (a) McMurry, J. *Organic Chemistry*, 4th ed.; Brooks/Cole: Pacific Grove, CA, 1996; p 197. (b) Fox, M. A.; Whitesell, J. K. *Organic Chemistry*; Jones and Bartlett: New York, 1997; p 40. (c) Jones, M. J., Jr. *Organic Chemistry*; Norton and Co.: New York, 1997; p 134.

(17) PPN⁺ = bis(triphenylphosphine)iminium ion.

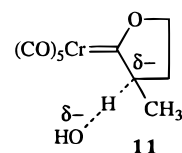
(18) We use quotation marks to indicate that these pK_a values are poorly defined in THF, although they can be used to give relative measures of the tendency of **2** to lose a proton in the presence of the respective bases.

(19) Anderson, R. L. Ph.D. Thesis, University of Wisconsin, 1974.

(20) Casey, C. P.; Anderson, R. L. *J. Am. Chem. Soc.* **1974**, *96*, 1230.

5 is more acidic than **1**, the fact that k_1^{OH} for **5** does not reflect this enhanced acidity implies that the *intrinsic* rate constant²¹ for the deprotonation of **5** is lower than for **1**. The same conclusion is reached based on the fact that the $k_{-1}^{\text{H}_2\text{O}}(\mathbf{5}^-)/k_{-1}^{\text{H}_2\text{O}}(\mathbf{1}^-)$ ratio of 0.047 is almost as small as the statistically corrected $K_1^{\text{OH}}(\mathbf{1})/K_1^{\text{OH}}(\mathbf{5})$ ratio of 0.044. One may estimate an approximate intrinsic rate constant as $\log k_0^{\text{OH}} \approx \log k_1^{\text{OH}} - 0.5 \log K_1^{\text{OH}}$ ²² = 0.70, which is about 0.66 log units lower than the statistically corrected $\log k_0^{\text{OH}}$ for **1** (Table 1).

A similar reduction in $\log k_0^{\text{OH}}$ by α -methyl groups was observed in the series **2** ($\log k_0^{\text{OH}} = 1.07$), **6** ($\log k_0^{\text{OH}} = 0.09$), and **7** ($\log k_0^{\text{OH}} = -0.99$).⁷ There are several factors that may contribute to these reductions. They are mainly related to the lag in charge delocalization behind proton transfer at the transition state (**11**), a lag that is characteristic for the transition states of all proton transfers from carbon acids leading to resonance-stabilized carbanions.^{24,25}



One consequence of this lag is that the feature that facilitates the strong stabilization of the anion (**5**⁻) by the methyl group, namely, the C=C double bond, is only weakly developed, and hence the transition state hardly benefits from this stabilization. The result is a reduction in k_0^{OH} . In addition to the absence of significant stabilization, there is *destabilization* of the transition state because of unfavorable interaction between the inductively electron donating methyl group and the negative charge that accumulates on the α -carbon. There is also a repulsive field effect between the methyl group and the partial negative charge on the approaching hydroxide ion that may further destabilize the transition state.^{24b} Finally, a steric effect may contribute to a lowering of k_0^{OH} ; with bulky amines this has been shown to be a significant factor,² but with OH⁻ as the base it is probably of minor importance.

Kinetic Solvent Isotope Effects. These measurements were mainly aimed at determining the kinetic isotope effect on the hydrolysis reaction because they are relevant to the mechanism of this reaction, as discussed below. These measurements were carried out with **5-d** rather than **5** in order to avoid complications that would arise from H/D exchange that gradually converts **5** into **5-d** during the course of the reaction. Note that the choice of **5-d** has no effect on the $k_2^{\text{D}_2\text{O}}$ step, which is the main focus of interest.

(21) The intrinsic rate constant for a reaction with the forward rate constant k_1 and the reverse rate constant k_{-1} is defined as $k_0 = k_1 = k_{-1}$ when $K_1 = k_1/k_{-1} = 1$. For proton transfers, statistical factors are usually taken into account.

(22) This is equivalent to applying the simplest version of the Marcus equation.²³

(23) Marcus, R. A. *J. Chem. Phys.* **1965**, *43*, 679.

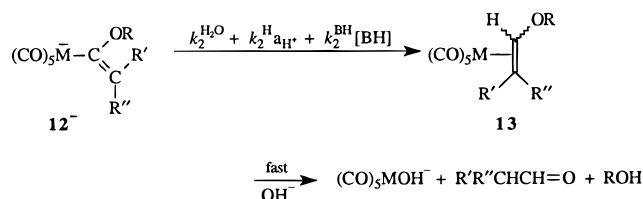
(24) (a) Bordwell, F. G.; Boyle, W. J., Jr. *J. Am. Chem. Soc.* **1972**, *94*, 3907. (b) Kresge, A. J. *Can. J. Chem.* **1974**, *52*, 1897. (c) Terrier, F.; Lelièvre, J.; Chatrousse, A.-P.; Farrell, P. *J. Chem. Soc., Perkin Trans. 2* **1985**, 1479.

(25) (a) Bernasconi, C. F. *Acc. Chem. Res.* **1987**, *20*, 301. (b) Bernasconi, C. F. *Acc. Chem. Res.* **1992**, *25*, 9. (c) Bernasconi, C. F. *Adv. Phys. Org. Chem.* **1992**, *27*, 119.

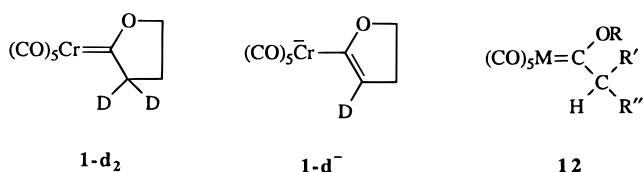
Table 1. Rate and Equilibrium Constants for the Reversible Deprotonation of **5 and **1** and of Conversion of **5**⁻ and **1**⁻ to Products in 50% MeCN–50% H₂O and 50% MeCN–50% D₂O**

	5		1 ^a	
	parameter	KIE ^b	parameter	KIE ^b
$k_1^{\text{OH}}, \text{M}^{-1} \text{s}^{-1}$	39.5 ± 0.04		74 (37) ^c	
$k_1^{\text{H}_2\text{O}}, \text{s}^{-1}$	0.66 ± 0.04		14	
$\log k_0^{\text{OH}}$	ca. 0.70		ca. 1.51 (1.36) ^c	
$K_1^{\text{OH}} = k_1^{\text{OH}}/k_{-1}^{\text{OH}}$	59.8 ± 4.5		5.29 (2.65) ^c	
$pK_{\text{a}}^{\text{CH}}$ ^d	13.41 ± 0.03		14.47 (14.77) ^c	
$k_2^{\text{H}_2\text{O}}$	0.161 ± 0.016		0.98	
$k_{1,\text{D}}^{\text{OD}}, \text{M}^{-1} \text{s}^{-1}$	12.4 ± 0.2	3.19 ^e	22.2 (11.1) ^c	3.36 ^e
$k_1^{\text{D}_2\text{O}}, \text{s}^{-1}$	(4.5 ± 1.0) × 10 ⁻²	14.7 ^f	1.33	10.5 ^f
$K_1^{\text{OD}} = k_{1,\text{D}}^{\text{OD}}/k_{-1}^{\text{OD}}$	(2.76 ± 0.65) × 10 ²		16.7 (8.35) ^c	
$k_2^{\text{D}_2\text{O}}$	(5.5 ± 0.3) × 10 ⁻²	2.93 ^g	0.32	3.06 ^g

^a References 2 and 5. ^b Kinetic isotope effect. ^c Statistically corrected for the two equivalent protons. ^d $pK_{\text{a}}^{\text{CH}} = pK_{\text{w}} - \log K_1^{\text{OH}}$ with $pK_{\text{w}} = 15.19$ (ref 3). ^e $k_{1,\text{H}}^{\text{OH}}/k_{1,\text{D}}^{\text{OD}}$ with $k_{1,\text{H}}^{\text{OH}} \equiv k_1^{\text{OH}}$. ^f $k_1^{\text{H}_2\text{O}}/k_1^{\text{D}_2\text{O}}$. ^g $k_2^{\text{H}_2\text{O}}/k_2^{\text{D}_2\text{O}}$.

Scheme 1

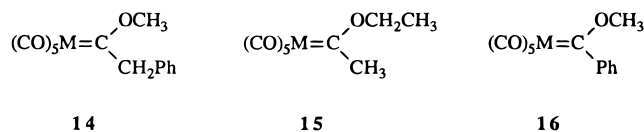
The experiments in the deuterated solvent also provided the rate constants $k_{1,\text{D}}^{\text{OD}}$ and $k_1^{\text{D}_2\text{O}}$ for deprotonation by OD⁻ and reprotonation by D₂O, respectively; they are reported in Table 1 along with the corresponding rate constants for **1**. The solvent isotope effect on the deprotonation of **5-d** (3.19) is quite comparable to that on the deprotonation of **1-d**₂ (3.36); the same is true



for protonation of the respective anions (14.7 for **5**⁻, 10.5 for **1-d**⁻). These ratios represent normal primary kinetic isotope effects, as discussed in more detail elsewhere.²

Hydrolysis of 5. The preferred hydrolysis mechanism for Fischer carbene complexes with hydrogen(s) on the α-carbon, **12**, appears to be the one shown in eq 5^{8,9} for **5**. The k_2 steps most likely involve protonation of the carbene carbon of the anion concerted with metal–carbon bond cleavage to yield a complex between the resulting vinyl ether and (CO)₅M, Scheme 1.²⁶ This complex, **13**, either dissociates into the free vinyl ether as the final organic product, which is the case for **14** (M = Cr or W),⁹ or rapidly hydrolyzes to the aldehyde (Scheme 1), as is the case for **2** and **15**.⁸ On the basis of the formation of vinyl ether products (from **14**)⁹ and kinetic solvent isotope effects^{8,9} it was concluded that for these compounds the mechanism of Scheme 1 is

(26) The possibility that conversion to **13** occurs in two steps, i.e., by protonation of **12**⁻ on the metal followed by a reductive elimination, cannot be rigorously excluded but is not likely.⁸



dominant. For example, the kinetic solvent isotope effect on the $k_2^{\text{H}_2\text{O}}$ step was found to be $k_2^{\text{H}_2\text{O}}/k_2^{\text{D}_2\text{O}} = 4.15$ for **2** and 7.0 for **15**; assuming a contribution by a secondary kinetic isotope effect of 1.41,⁵ these results imply a primary kinetic isotope effect of ca 2.9 for **2** and 5.0 for **15**.

The situation with **1** is more ambiguous because of a rather low value of the $k_2^{\text{H}_2\text{O}}/k_2^{\text{D}_2\text{O}}$ ratio (3.06), leading to an estimated primary isotope effect of only ca. 2.2.⁵ This raised the question as to whether the “classic” nucleophilic mechanism (Scheme 2) known from carboxylic acid chemistry and observed for carbene complexes lacking α-hydrogens^{27,28} (e.g., **16**) may be competitive in this case. According to this mechanism, which is kinetically indistinguishable from the mechanism of eq 5/Scheme 1, $k_{\text{obsd}}^{\text{II}}$ is given by eq 10 or, for $K_1^{\text{OH}}[\text{OH}^-] < < 1$, by eq 11.

$$k_{\text{obsd}}^{\text{II}} = \frac{1}{1 + K_1^{\text{OH}}[\text{OH}^-]} (k_3^{\text{OH}}[\text{OH}^-] + k_3^{\text{H}_2\text{O}} + k_3^{\text{B}}[\text{B}]) \quad (10)$$

$$k_{\text{obsd}}^{\text{II}} = k_3^{\text{OH}}[\text{OH}^-] + k_3^{\text{H}_2\text{O}} + k_3^{\text{B}}[\text{B}] \quad (11)$$

An analysis of the kinetic solvent isotope effect for **1** in terms of Scheme 2 yielded $k_3^{\text{OH}}/k_3^{\text{OD}} \approx 0.98$,⁵ which is within the normal range for such an isotope effect.^{28,29} This suggested that the mechanism of Scheme 2 might be operative, or at least be competitive, with Scheme 1.

For the hydrolysis of **5** the $k_2^{\text{H}_2\text{O}}/k_2^{\text{D}_2\text{O}}$ ratio of 2.93 is nearly identical to that for the hydrolysis of **1**, leaving the same mechanistic ambiguity. However, a comparison of the relative reactivities of **5** and **1** suggests that for these cyclic carbene complexes the mechanism of eq 5/Scheme 1 appears to be dominant after all, the low isotope effects notwithstanding. Table 2 summarizes the $k_2^{\text{H}_2\text{O}}$, k_2^{H} , and k_2^{BH} values calculated by assuming that hydrolysis occurs exclusively by the mechanism of eq 5/Scheme 1 and the k_3^{OH} , $k_3^{\text{H}_2\text{O}}$, and k_3^{B} values obtained by assuming that Scheme 2 is the actual mechanism.

If Scheme 2 were the prevailing mechanism, one would expect a reduction in the reactivity of **5** compared to **1** mainly due to the steric hindrance by the α-methyl group. This reduction should be substantial because nucleophilic reactions with Fischer carbene complexes are very sensitive to steric effects. For example, nucleophilic substitution of the methoxy group of **16** by secondary alicyclic amines is about 2 orders of magnitude slower than by primary aliphatic amines of the same pK_{a} .³⁰ This contrasts with reactions of sterically

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Scheme 2

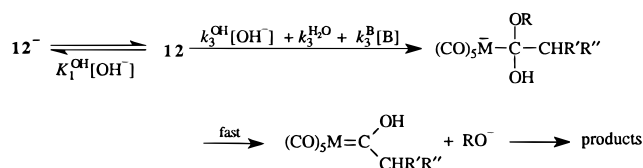


Table 2. Rate Constants for the Rate-Limiting Step in the Hydrolysis of 5 and 1: Interpretation in Terms of Eq 5 and Scheme 2^a

reactive agent (pK _a)	5 pK _a ^{CH} = 13.41 Eq 5 (<i>k</i> ₂ ^{H₂O} , <i>k</i> ₂ ^{BH} , <i>k</i> ₂ ^H)	1 ^b pK _a ^{CH} = 14.47	5/1
H ₂ O (16.63)	0.161 ^c	0.98 ^c	0.164
Et ₃ NH ⁺ (10.31)	0.142	2.80	0.051
<i>N</i> -MeMorH ⁺ (7.43)	15.5	2.20 × 10 ²	0.070
AcOH (5.93)	1.03 × 10 ³	2.19 × 10 ⁴	0.047
MeOCH ₂ COOH (4.73)	5.50 × 10 ³	3.30 × 10 ⁴	0.153
H ₃ O ⁺ (-1.44)	4.09 × 10 ⁷	2.90 × 10 ⁸	0.141
	Scheme 2 (<i>k</i> ₃ ^{OH} , <i>k</i> ₃ ^B , <i>k</i> ₃ ^{H₂O})		
OH ⁻ (16.63) ^c	9.63	5.20	1.85
Et ₃ N (10.31) ^c	1.13 × 10 ⁻⁴	1.94 × 10 ⁻⁴	0.58
<i>N</i> -MeMor (7.43) ^c	1.62 × 10 ⁻⁵	2.10 × 10 ⁻⁵	0.77
AcO ⁻ (5.93) ^c	3.32 × 10 ⁻⁵	6.32 × 10 ⁻⁵	0.53
MeOCH ₂ COO ⁻ (4.73) ^c	1.06 × 10 ⁻⁵	6.01 × 10 ⁻⁶	1.76
H ₂ O (-1.44) ^c	1.59 × 10 ⁻⁶ ^c	9.80 × 10 ⁻⁷ ^c	1.62

^a In units of M⁻¹ s⁻¹ unless otherwise stated. ^b Reference 5. ^c In units of s⁻¹. Estimated error in and *k*₂^{H₂O} and *k*₃^{OH}, ca. ±10%; in *k*₂^H and *k*₃^{H₂O} ± 50%; in *k*₂^{BH} and *k*₃^{BH} ca. ±20–30%.

unhindered or weakly hindered electrophiles where secondary alicyclic amines are typically *more* reactive than primary amines.³¹ The *k*₃^{OH}(5)/*k*₃^{OH}(1), *k*₃^{H₂O}(5)/*k*₃^{H₂O}(1), and *k*₃^B(5)/*k*₃^B(1) ratios are all close to unity, which is inconsistent with the expected steric effect. On the other hand, if eq 5/Scheme 1 is the dominant mechanism, one would expect 5⁻ to be significantly less reactive than 1⁻ because of the C=C double-bond stabilization of 5⁻ by the methyl group. This is borne out by the *k*₂^{H₂O}(5)/*k*₂^{H₂O}(1), *k*₂^H(5)/*k*₂^H(1), and *k*₂^{BH}(5)/*k*₂^{BH}(1) ratios that range between 0.047 and 0.164.

Experimental Section

Materials. (5-Methoxy-2-oxacyclopentylidene)pentacarbonylchromium(0), **5**, and (5,5-dimethyl-2-oxacyclopentylidene)pentacarbonylchromium(0), **10**, were gifts from Professor William D. Wulff. **5-d** was prepared by H/D exchange from **5** as follows. A 7 mg portion of **5** was dissolved in 0.7 mL of CD₃CN. After separation from some insoluble matter by centrifugation, a KOD solution in D₂O was added to give a 0.005 M concentration and solvent composition of 7:1 CD₃CN:

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D₂O. The water was removed after freezing the sample at -78 °C and thawing at room temperature and repeating this cycle several times. The ¹H NMR of the solution showed virtually complete elimination of the methine signal of **5** at 3.66 ppm. This solution was used directly as stock for the kinetic determinations. Triethylamine and *N*-methylmorpholine were refluxed over sodium under argon and then fractionally distilled. Acetic acid and methoxyacetic acid were used as received. HCl and KOH solutions were prepared by diluting prepackaged solutions (Baker Analytical). KOD was prepared by dissolving KOH in D₂O; the concentration of the resulting solution was determined by potentiometric titration. Acetonitrile (Fischer Scientific) was used as received. Water and acetonitrile used in the hydrolysis experiments were degassed by the freeze-pump-thaw method to minimize substrate oxidation.

Solutions and pH Measurements. All kinetic experiments were conducted in 50% MeCN–50% H₂O (v/v) or 50% MeCN–50% D₂O (v/v) solutions at 25 °C, μ = 0.1 M (KCl). All pH measurements were made with an Orion 611 pH meter equipped with a glass electrode and a "SureFlow" (Corning) reference electrode. Actual pH values were calculated by adding 0.18 to the measured pH, according to Allen and Tidwell.³² The pH of reaction solutions for stopped-flow experiments was adjusted in mock-mixing experiments which mimicked the stopped-flow runs. The pK_a values of the buffers were known from previous studies.^{3,9}

Kinetics. Typical carbene complex concentrations were (5–10) × 10⁻⁵ M. Rates were measured in an Applied Photophysics DX.17MV stopped-flow apparatus (fast rates) and in a Perkin-Elmer Lambda 2 or Hewlett-Packard 8452A diode-array UV-vis spectrophotometer (slow rates). Kinetics were followed by monitoring the disappearance of the carbene complex at 370 nm. Rate constants (*k*_{obsd}) were obtained by computer fit programs (Applied Photophysics and Enzfitter³³). Rates at pH < 9.3 were very low and determined by the initial rates methods: reactions were monitored for 3–12 h, after which enough 4 M KOH was added to neutralize the acidic component of the buffer and increase the pH of the solution to ≈12. Reactions were then further monitored until the infinity value (OD_∞) was reached. Pseudo-first-order rate constants were calculated according to eq 12, where *S* is the slope of the plot of OD vs time and ΔOD₀ = OD_∞ – OD₀.

$$k_{\text{obsd}} = S/\Delta OD_0 \quad (12)$$

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Supporting Information Available: Table S1 (kinetic data) and Figure S1 (plot of *k*_{obsd}¹ vs [KOH]) (2 pages). See any current masthead page for ordering information and Web access instructions.

OM980444S

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