# Stereochemistry of Allylic Rearrangements. 17. Acid-Catalyzed Rearrangement of *exo*- and *endo*-Bicyclo[3.2.1]oct-3-en-2-ol

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**Abstract:** The acid-catalyzed rearrangement of *exo*- and *endo*-bicyclo[3.2.1]oct-3-en-2-ol (**2**-OH) in 33.3% aqueous acetone results in interconversion of enantiomers (racemization) and exo $\rightleftharpoons$ endo isomerization. Rate constants have been determined for (1) loss of optical activity ($k_\alpha$), (2) exo $\rightleftharpoons$ endo equilibration ($k_{eq}$) and $^{18}O$ exchange ($k_{exc}$) for both isomers. Reactivity ratios ($k_{\alpha\text{-exo}}/k_{\alpha\text{-endo}}$) range from 247 at 20 °C to 136 at 49 °C. The partition ratio for solvent capture of the symmetrical bicyclo[3.2.1]oct-3-en-2-yl cation (**3**) ranges from 878 at 20 °C to 454 at 49 °C (exo capture favored). The activation energy difference for exo and endo capture ($\Delta E_a$) is 4.4 ± 0.5 kcal mol$^{-1}$ and $\Delta H$ for the exo $\rightleftharpoons$ endo equilibration is 0.6 ± 0.2 kcal mol$^{-1}$ (exo isomer more stable). The $k_{exc}/k_\alpha$ ratio is 0.67 for *exo*-**2**-OH and 1 for *endo*-**2**-OH.

In earlier work we investigated the stereochemistry of the acid-catalyzed ($HClO_4$) rearrangement of *cis*- and *trans*-5-methyl-2-cyclohexenol (**1**-OH) in aqueous acetone and determined the pseudo-first-order rate constants for (1) loss of optical activity ($k_\alpha$, eq 1), (2) geometric isomerization ($k_i$, eq 2), and (3) $^{18}O$ exchange ($k_{exc}$, eq 3).[1a] It was found that for both isomers $k_\alpha$ is several times larger than $k_i$. Thus there is excess rearrangement with preservation of geometric configuration (racemization) which means that the isomeric alcohols give different symmetrical 5-methyl-2-cyclohexenyl intermediates. This memory effect was interpreted in terms of the quasi-axial conformations being stereoelectronically favored for bond cleavage and formation.[1a] Thus, the isomers are related to symmetrical intermediates that differ conformationally and the initially formed intermediate is predisposed to re-form racemic alcohol with preservation of geometric configuration as illustrated in Scheme I. Presumably isomerization involves conformational change of the cation.

$$\text{active } cis\text{- or } trans\text{-}\mathbf{1}\text{-OH} \xrightarrow{k_\alpha} \text{inactive product} \qquad (1)$$

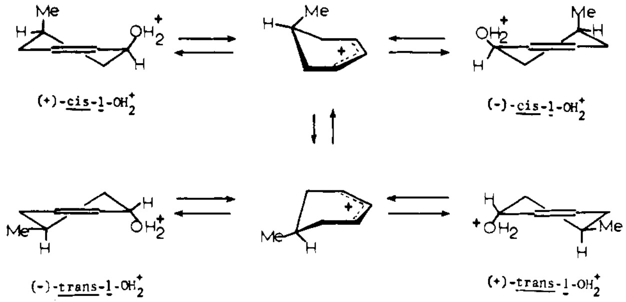
$$cis\text{-}\mathbf{1}\text{-OH} \underset{k_{\text{i-trans}}}{\overset{k_{\text{i-cis}}}{\rightleftharpoons}} trans\text{-}\mathbf{1}\text{-OH} \qquad (2)$$

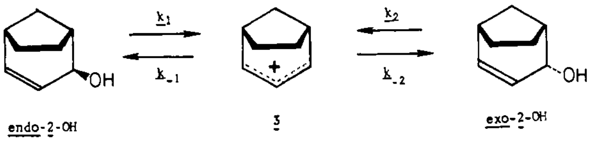
$$R^{18}OH \xrightarrow{k_{exc}} ROH \qquad (3)$$

It was also observed that the rate of $^{18}O$ exchange ($k_{exc}$), relative to the rate of formation of a symmetrical intermediate ($k_\alpha$), differs for the isomers. For *cis*-**1**-OH, racemization is essentially intramolecular (little exchange associated with interconversion of enantiomers). With *trans*-**1**-OH, interconversion of enantiomers results in almost complete exchange. This was attributed to differences in the way that the conformationally different cations are solvated.[1a]

We now report the results of a similar investigation of *exo*- and *endo*-bicyclo[3.2.1]oct-3-en-2-ol (**2**-OH). In this rigid bicyclic system the hydroxyl group is quasi-axial in the exo isomer and quasi-equatorial in the endo isomer and the symmetrical allylic cation (**3**) is structurally independent of the epimer from which it is derived. In this work we have determined (a) the difference in exo and endo transition state energies, (b) the reactivity ratios ($k_{\alpha\text{-exo}}/k_{\alpha\text{-endo}}$), and (c) the partition ratios ($k_{-1}/k_{-2}$) for the acid-catalyzed ($HClO_4$) rearrangement of **2**-OH in 33.3% (v/v) aqueous acetone (reaction 4).

Scheme I





(4)

The method is similar to that used in earlier investigations[1b] and involves comparison of rates and activation parameters of two reactions for each isomer. These are loss of optical activity (reaction 5) and exo $\rightleftharpoons$ endo isomerization (reaction 6). The pseudo-first-order rate of reaction 5 corresponds to the rate of formation of the carbonium ion—in other work it was shown that irreversible solvent capture gives completely racemic products.[3] The difference between the pseudo-first-order rate constants for loss of optical activity ($k_\alpha$) and isomerization ($k_i$) is the constant for racemization with preservation of geometric configuration, i.e., $k_{rac} = k_\alpha - k_i$. The rate of $^{18}O$ exchange (reaction 3) was also determined for each isomer for comparison with the rate of generation of the symmetrical intermediate.

$$exo\text{- or } endo\text{-}\mathbf{2}\text{-OH} \xrightarrow{k_\alpha} \text{inactive products} \qquad (5)$$

$$exo\text{-}\mathbf{2}\text{-OH} \underset{k_{\text{i-endo}}}{\overset{k_{\text{i-exo}}}{\rightleftharpoons}} endo\text{-}\mathbf{2}\text{-OH} \qquad (6)$$

The measurable pseudo-first-order constants for reactions 5 and 6 and the equilibrium constant for reaction 6 ($K_{eq}$) are related to the constants in eq 4 as follows:

$$k_{\alpha\text{-exo}} = k_1 \text{ and } k_{\alpha\text{-endo}} = k_2 \qquad (7)$$

$$k_{\text{i-exo}} = k_1k_{-2}/(k_{-1} + k_{-2})$$

$$k_{\text{i-endo}} = k_2k_{-1}/(k_{-1} + k_{-2})$$

$$K_{eq} = (k_1/k_2)(k_{-2}/k_{-1}) \qquad (8)$$

Thus

$$k_{\alpha\text{-exo}}/k_{\text{i-exo}} = (k_{-1}/k_{-2}) + 1 \qquad (9)$$

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**Table I.** Catalytic Second-Order Constants for Acid-Catalyzed Loss of Optical Activity ($k_\alpha$) for *exo*- and *endo*-Bicyclo[3.2.1]oct-3-en-2-ol in 33.33% Acetone[a]

| temp, °C | $k_{\alpha\text{-exo}}$,[b] $M^{-1}\ h^{-1}$ | $k_{\alpha\text{-endo}}$,[c] $M^{-1}\ h^{-1}$ |
|---|---|---|
| 12.67 | 5.59 ± 0.13[d] | |
| 19.90 | 15.0 ± 0.1[e] | |
| 29.87 | 51.0 ± 0.2[f] | 0.258 ± 0.002[g] |
| 40.14 | 166 ± 0.1[h] | 1.04 ± 0.04[i] |

[a] Ionic strength constant; $[HClO_4] + [LiClO_4] = 0.182$ M. Concentrations (25 °C) not corrected for thermal expansion. [b] $E_a = 21.8 \pm 0.2$ kcal mol$^{-1}$. [c] $E_a = 25.6 \pm 0.8$ kcal mol$^{-1}$. [d] Average of four determinations; $[HClO_4] = 0.01–0.09$ M; [*exo*-**2**-OH] = 0.02–0.025 M. [e] Average of three determinations; $[HClO_4] = 0.007–0.036$ M; [*exo*-**2**-OH] = 0.02 M. [f] Average of seven determinations; $[HClO_4] = 0.007$ M; [*exo*-**2**-OH] = 0.015–0.059 M. [g] Average of two determinations; $[HClO_4] = 0.182$ M; [*endo*-**2**-OH] = 0.13 M. [h] Average of two determinations; $[HClO_4] = 0.003$ M; [*exo*-**2**-OH] = 0.014 and 0.027 M. [i] Average of three determinations; $[HClO_4] = 0.09–0.182$ M; [*endo*-**2**-OH] = 0.12 M.

From this it can be seen that $k_\alpha$ corresponds to generation of the symmetrical cation (**3**). Thus the $k_{\alpha\text{-exo}}/k_{\alpha\text{-endo}}$ ratio corresponds to the reactivity ratio ($k_1/k_2$) and the partition ratio for capture of the cation ($k_{-1}/k_{-2}$) can be determined from the $k_{\alpha\text{-exo}}/k_{\text{i-exo}}$ ratio.

The rates of reactions 3, 5, and 6 are proportional to acid concentration, which shows that each involves reversible protonation by lyonium ion ($SH^+$) followed by first-order reaction of the conjugate acid of the alcohol as illustrated by reactions 10–13.

$$\text{2-OH} + \text{SH}^+ \overset{K_{10}}{\rightleftharpoons} \text{2-OH}_2^+ + \text{S} \tag{10}$$

$$\text{2-OH}_2^+ \xrightarrow{k_\alpha'} \text{inactive products} \tag{11}$$

$$\text{2-OH}_2^+ \xrightarrow{k_i'} \text{geometric isomer} \tag{12}$$

$$\text{2-}^{18}\text{OH}_2^+ \xrightarrow{k'_{\text{exc}}} \text{2-OH}_2^+ \tag{13}$$

From this it can be seen that the pseudo-first-order constants ($k_\alpha$, eq 5; $k_i$, eq 6; and $k_{\text{exc}}$, eq 3) are directly proportional to the corresponding first-order constants for the protonated alcohols ($k'$, eq 11–13) as shown by the equation

$$k = K_{10}[SH^+]k' \tag{14}$$

## Results

Racemic and optically active *exo*-bicyclo[3.2.1]oct-3-en-2-ol (*exo*-**2**-OH) and racemic *endo*-**2**-OH were prepared as reported earlier.[3] Optically active *endo*-**2**-OH was obtained by resolution of the acid phthalate derivative.

Oxygen-18 labeled *exo*-**2**-OH was prepared by solvolysis of bicyclo[3.2.1]oct-3-en-2-yl bromide (**2**-Br)[3] in $^{18}$O-enriched water (1.5% excess $^{18}$O). Recrystallization of the *p*-nitrobenzoate derivative followed by saponification gave isomerically homogeneous $^{18}$O-labeled *exo*-**2**-OH containing 1.3% excess $^{18}$O. Oxygen-18 labeled endo alcohol was prepared from $^{18}$O-enriched bicyclo[3.2.1]oct-3-en-2-one.[3] The labeled ketone was obtained by acid-catalyzed exchange of the carbonyl oxygen with $^{18}$O-enriched water[1a] and converted to $^{18}$O-labeled *endo*-**2**-OH by reduction with $LiAlH_4$.[3] After purification[3] the product consisted of 93% of the endo isomer and contained 1.8% excess $^{18}$O. Configurational compositions were determined by capillary GC.

Rate constants for loss of optical activity ($k_\alpha$) in 33.33% aqueous acetone (v/v, 25 °C) are presented in Table I and data for the acid-catalyzed *exo*-**2**-OH ⇄ *endo*-**2**-OH equilibration are summarized in Table II. All rates are cleanly pseudo first order and at constant ionic strength, directly proportional to acid concentration over the range 0.007–0.182 M. Solvents were prepared by mixing stock solutions of 0.182 M $HClO_4$ and 0.182 M $LiClO_4$ in 33.33% acetone. Thus in all cases the ionic strength was 0.182 M. Control experiments showed that $k_{\alpha\text{-exo}}$, the constant that can be determined with the greatest accuracy, increases ~30% with an increase in ionic strength from 0.007 to 0.182 M and is unchanged by varying solvent composition from 33.3 to 32.3% acetone.

The rate constants in the tables are catalytic second-order constants obtained by dividing observed pseudo-first-order constants by acid concentration. Or to put it another way, these are pseudo-first-order constants normalized for $[HClO_4] = 1$ M.

Rate constants for loss of optical activity ($k_\alpha$) were determined as described earlier. In most cases reactions were carried out in a thermostated polarimeter tube and followed by observing the decreasing rotation at appropriate intervals. In some experiments aliquots were quenched with base at intervals during the reaction and $k_\alpha$ was determined from the observed static rotations. Both methods gave the same results. Reactions were followed to >75% completion and in all cases decay of optical activity was first order. Rotations could be determined to ±0.002° and the total change in rotation was >1° for *exo*-**2**-OH and >0.5° for *endo*-**2**-OH. Thus $k_\alpha$ can be determined with good precision. As indicated by the footnotes, the values in Table I are average (and average deviations) of from two to seven independent kinetic experiments.

The rate constants for isomerization ($k_{\text{i-exo}}$ and $k_{\text{i-endo}}$) in Table II were determined from the pseudo-first-order constant for irreversible equilibration of the isomers ($k_{\text{eq}}$, reaction 15) and the equilibrium constant ($K_{\text{eq}}$, reaction 6). As shown by

**Table II.** Equilibrium ($K_{\text{eq}}$) and Catalytic Second-Order Constants for the Acid-Catalyzed Equilibration of *exo*- and *endo*-Bicyclo[3.2.1]-oct-3-en-2-ol in 33.3% Acetone

| temp, °C | $10K_{\text{eq}}$[a,b] | $10k_{\text{eq}}$, $M^{-1}\ h^{-1}$ | $10k_{\text{i-exo}}$,[c,d] $M^{-1}\ h^{-1}$ | $10k_{\text{i-endo}}$,[d,e] $M^{-1}\ h^{-1}$ |
|---|---|---|---|---|
| 29.87 | 2.99 ± 0.08 | 3.24 ± 0.07[f] | 0.74 ± 0.02 | 2.49 ± 0.07 |
| 40.14 | 3.05 ± 0.07 | 13.0 ± 0.02[g] | 3.0 ± 0.1 | 9.9 ± 0.01 |
| 49.03 | 3.12 ± 0.03 | 40.4 ± 0.5[h] | 9.6 ± 0.1 | 30.8 ± 0.3 |
| 78.47 | 3.42 ± 0.05 | | | |
| 99.36 | 3.57 ± 0.05 | | | |

[a] $K_{\text{eq}}$ = *endo*-**2**-OH/*exo*-**2**-OH; $\Delta H = 0.57 \pm 0.13$ kcal mol$^{-1}$. $\Delta G = 0.74$ kcal mol$^{-1}$ $\Delta S$ 0.53 eu at 50 °C. [b] Average and average deviation of five (29.87 °C), three (40.14 and 99.36 °C), and two (49.03 and 78.47 °C) determinations. Equilibrium approached from both directions at each temperature. [c] $E_a = 25.8 \pm 0.3$ kcal mol$^{-1}$. [d] Uncertainties derived from uncertainties in $K_{\text{eq}}$ and $k_{\text{eq}}$. [e] $E_a = 25.5 \pm 0.3$ kcal mol$^{-1}$. [f] Average of five determinations starting with 0.09–0.133 M *endo*-**2**-OH, $[HClO_4] = 0.182$ M. [g] Average of two determinations starting with 0.182 M *endo*-**2**-OH, $[HClO_4] = 0.182$ M. [h] Average of determinations with $[HClO_4] = 0.012$ and 0.036 M; [*endo*-**2**-OH] = 0.12 M.

**Table III.** Comparison of Rate Constants for Reactions Associated with the Acid-Catalyzed Rearrangement of *exo*- and *endo*-Bicyclo[3.2.1]oct-3-en-2-ol in 33.33% Acetone at 29.87 °C

| | *exo*-**2**-OH [a] | *endo*-**2**-OH [a] |
|---|---|---|
| $10k_{\alpha}$ $M^{-1}$ $h^{-1}$ | $510 \pm 4$ | $2.58 \pm 0.02$ |
| $10k_{eq}$, $M^{-1}$ $h^{-1}$ | $3.24 \pm 0.07$ | $3.24 \pm 0.07$ |
| $10k_{i}$, $M^{-1}$ $h^{-1}$ | $0.74 \pm 0.02$ | $2.49 \pm 0.07$ |
| $10k_{rac}$, $M^{-1}$ $h^{-1}$ [b] | $510 \pm 4$ | $0.09 \pm 0.09$ |
| $10k_{exc}$, $M^{-1}$ $h^{-1}$ | $348 \pm 15$[c] | $2.47 \pm 0.04$[d] |

[a] Data taken from Tables I and II. [b] Determined from rate constants for total loss of optical activity and geometric isomerization, i.e., $k_{rac} = k_{\alpha} - k_{i}$. [c] Average of three determinations, $[HClO_4] = 0.007$; [*exo*-**2**-OH] = 0.11–0.25 M. [d] Average of two determinations, $[HClO_4] = 0.1$ M; [*endo*-**2**-OH] = 0.11 M.

**Table IV.** Partition ($k_{\alpha\text{-exo}}/k_{i\text{-exo}}$) and Reactivity ($k_{\alpha\text{-exo}}/k_{\alpha\text{-endo}}$) Ratios for Acid-Catalyzed Isomerization of *exo*- and *endo*-Bicyclo[3.2.1]oct-3-en-2-ol in 33.33% Aqueous Acetone

| temp, °C | $k_{\alpha\text{-exo}}/k_{i\text{-exo}}$ [a,b] | $k_{\alpha\text{-exo}}/k_{\alpha\text{-endo}}$ [b] |
|---|---|---|
| 19.90 | $878 \pm 39$[c] | 247[d] |
| 29.87 | $689 \pm 24$ | $198 \pm 2$ |
| 40.14 | $553 \pm 19$ | $159 \pm 7$ |
| 49.03 | $454 \pm 15$[e] | 136[f] |

[a] $\Delta E_a = 4.4 \pm 0.3$ kcal mol$^{-1}$; $\Delta\Delta S^{\ddagger} = 1.9 \pm 1.4$ eu. [b] Uncertainties estimated from limiting values of constants. [c] Determined by extrapolation of $k_{i\text{-exo}}$. [d] Determined by extrapolation of $k_{\alpha\text{-endo}}$. [e] Determined by extrapolation of $k_{\alpha\text{-exo}}$. [f] Determined by extrapolation of both constants.

eq 16 and 17, these measurable constants correspond to the sum and ratio of the individual constants.

$$exo\text{- or } endo\text{-}\mathbf{2}\text{-OH} \xrightarrow{k_{eq}} \text{equilibrium mixture} \quad (15)$$

$$k_{eq} = k_{i\text{-exo}} + k_{i\text{-endo}} \quad (16)$$

$$K_{eq} = endo\text{-}\mathbf{2}\text{-OH}/exo\text{-}\mathbf{2}\text{-OH} = k_{i\text{-exo}}/k_{i\text{-endo}} \quad (17)$$

The equilibrium constant for equilibration ($K_{eq}$) was determined from the isomeric composition of **2**-OH (approached from both sides) after >8 half-periods for equilibration. Equilibration (reaction 15) was cleanly first order. The reaction was followed by isolation of **2**-OH from aliquots quenched with base and determining configurational compositions by capillary GC. The exo isomer predominates at equilibrium; thus, starting with *endo*-**2**-OH results in the largest measurable change in composition (~75 percentage units) and the more precise determination of $k_{eq}$. The values in Table II are average of two to five independent experiments starting with *endo*-**2**-OH. Optically active *endo*-**2**-OH was used in one experiment and the same aliquots were used to determine both $k_{eq}$ and $k_{\alpha\text{-endo}}$. A control experiment showed that, under conditions for exo ⇄ endo equilibration, **2**-OH is not consumed by irreversible parasitic reactions—there was no detectable change in the **2**-OH concentration after 5 half-lives for equilibration.

The *endo*-**2**-OH → *exo*-**2**-OH isomerization constants ($k_{i\text{-endo}}$) in Table II were determined from $K_{eq}$ and $k_{eq}$ with eq 18, which is derived from eq 16 and 17. The isomerization constant for *exo*-**2**-OH ($k_{i\text{-exo}}$) was determined from $k_{i\text{-endo}}$ and $K_{eq}$ (eq 17).

$$k_{i\text{-endo}} = k_{eq}/(K_{eq} + 1) \quad (18)$$

Pseudo-first-order rate constants ($[HClO_4] = 1$ M) for the various acid-catalyzed transformations of *exo*- and *endo*-**2**-OH at 29.87 °C are compared in Table III. All constants were measured at this temperature and thus extrapolations are not involved in these comparisons.

Rate constants for $^{18}O$ exchange (reaction 3) are included in Table III. Exchange is cleanly pseudo first order and reactions were followed by isolation of the alcohol fraction from quenched aliquots by gas chromatography and determination of the $^{18}O$ contents.[1a,4] Since the ratio of exchangeable oxygen in the solvent to alcohol is ~140, reincorporation of $^{18}O$ is negligible and for all practical purposes exchange is irreversible as indicated by reaction 3. Control experiments showed that isolation did not alter the $^{18}O$ content of the alcohol and that samples of isolated **2**-OH were representative of the alcohol in the reaction mixture at the time of quenching. For $^{18}O$ labeled *endo*-**2**-OH, both $k_{eq}$ and $k_{exc}$ were determined from the same isolated samples of **2**-OH.

The data in Table III show the relative rates at which the conjugate acid of each isomer undergoes the various transformations. For *exo*-**2**-OH, $k_{\alpha}$ is much larger than $k_{i}$, which means that the exo/endo capture ratio of 3 ($k_{-1}/k_{-2}$) is high and ionization results primarily in racemization ($k_{rac}$). Comparison of $k_{exc}$ and $k_{\alpha\text{-exo}}$ shows that ionization does not result in complete exchange. About 30% of the racemization is intramolecular (i.e., there is excess recapture of the water molecule formed in the ionization step) and this fraction of the ionization is undetected by $k_{exc}$. The possibility of an isotope effect is important in comparing $k_{exc}$ with the other constants because only the fraction of alcohol with $^{18}O$ is observed in the exchange experiments. In an earlier paper[1a] we gave reasons for concluding that the $^{18}O$ isotope effect for acid-catalyzed reactions of alcohols is negligible and we now present evidence that this is the case. Because of the high exo/endo capture ratio, ionization of the conjugate acid of *endo*-**2**-OH results in isomerization and exchange as well as loss of optical activity. Thus exchange can be compared with two independent measures of total ionization, and, as shown in Table III, for *endo*-**2**-OH, $k_{exc}$ is within the combined uncertainties of $k_{i}$ and $k_{\alpha}$.

The $k_{\alpha\text{-exo}}/k_{i\text{-exo}}$ and $k_{\alpha\text{-exo}}/k_{\alpha\text{-endo}}$ ratios for various temperatures are presented in Table IV. As shown by eq 7 and 9 the first is the partition ratio ($k_{-1}/k_{-2}$) and the latter the reactivity ratio ($k_{1}/k_{2}$). The reactivity ratio can also be determined from the partition ratio and the equilibrium constant (eq 8). This method uses only data for *exo*-**2**-OH and gives values within the combined uncertainties of the $k_{\alpha\text{-exo}}/k_{\alpha\text{-endo}}$ ratios in Table IV. This agreement establishes that *exo*- and *endo*-**2**-OH give a common symmetrical intermediate as illustrated by eq 4.

The $k_{\alpha\text{-exo}}/k_{i\text{-exo}}$ ratios show the magnitude and temperature dependence of the exo/endo capture ratio of the bicyclic allylic cation (**3**). These data are in agreement with earlier solvolytic investigations which showed that the exo/endo ratio for irreversible capture of **3** is >200.[2,3]

The constants in the reactivity ratio ($k_{1}/k_{2}$), unlike those in the partition ratio ($k_{-1}/k_{-2}$), are composites and include a proportionality constant ($K_{10}$) connecting the observed pseudo-first-order constants and the first-order constants for ionization of the conjugate acid of the alcohol. Since $K_{10}$ is unknown and presumably differs for the two isomers, $k_{\alpha\text{-exo}}/k_{\alpha\text{-endo}}$ does not necessarily correspond to the ratio for ionization of the two conjugate acids. Thus, the reactivity ratios do not provide interpretable mechanistic information regarding structure and reactivity.

Apparent activation energies ($E_a$) for generation of the common intermediate (**3**) and $\Delta H$ for equilibration are shown in Figure 1. These parameters were determined from data in Tables I and II and in all cases logarithmic plots were linear. The pseudo-first-order constants have not been corrected for thermal expansion of the solvent. However, from data for 30%

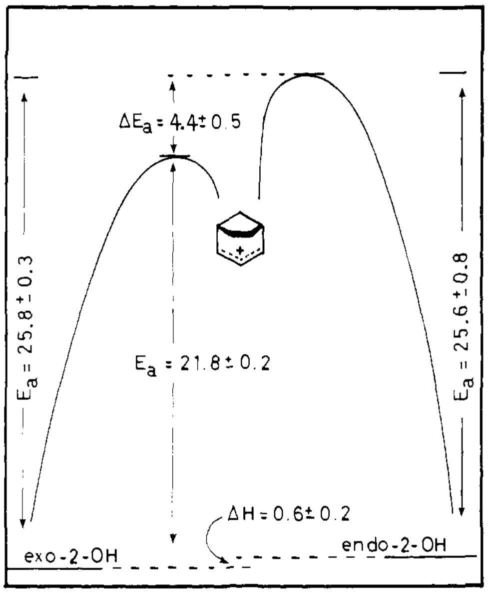

**Figure 1.** Apparent activation energies ($E_a$) for acid-catalyzed reactions of *exo*- and *endo*-**2**-OH and $\Delta H$ for equilibration.

aqueous acetone,[5] it is evident that corrections would be minor. The density of the solvent changes about 2.3% over the temperature range of 13-40 °C. This would increase the range of the rate constants by the same amount and increase $E_a$ about 0.15 kcal mol$^{-1}$, less than experimental error.

The energy gap for the exo and endo transition states ($\Delta E_a$) in Figure 1 was determined directly from a logarithmic plot of the $k_{\alpha\text{-exo}}/k_{i\text{-exo}}$ ratio. This value is in good agreement with that obtained indirectly from $\Delta H$ and $E_a$ for $k_{\alpha\text{-endo}}$ and $k_{\alpha\text{-exo}}$. The magnitude of $\Delta E_a$ for this system is about the same as that observed earlier for the acid-catalyzed equilibration of *exo*- and *endo*-1,2-dimethylnorbornanol ($\Delta E_a$ = 4.3 kcal mol$^{-1}$) and substantially less than that for equilibration of *exo*- and *endo*-norbornanol ($\Delta E_a$ = 7.8 kcal mol$^{-1}$).[1c] Presumably, in the present system the greater stability of the exo transition state results primarily from steric factors (cf. the 1,2-dimethylnorbornanol system) with a possible contribution from stereoelectronic factors (better overlap for formation of the exo quasi-axial bond[1a]).[6]

## Experimental Section

**Materials.** Dry acetone was prepared by fractionation of reagent grade acetone from Drierite. A stock solution of 0.182 M $HClO_4$ in 33.33% acetone was prepared by mixing two volumes of conductivity water containing the required amount of $HClO_4$ with one volume of dry acetone. Volumes were measured at 25 °C and the concentration was determined by titration of the solution. A stock solution of 0.182 M $LiClO_4$ in 33.33% acetone was prepared by mixing two volumes of conductivity water containing the required amount of $LiClO_4$ and one volume of dry acetone. The calculated amount of $LiClO_4$ was corrected for volume shrinkage (~2%) on mixing of two volumes of water and one of dry acetone.

Solvents of various acid concentrations were prepared by mixing the proper amounts of the two stock solutions. All concentrations were determined by titration with standard sodium hydroxide and are for 25 °C.

Isomeric compositions of **2**-OH were determined with a 100-ft UCON Polar polyglycol (LB-550-X) capillary column at 100 °C. Base line resolution of *exo*- and *endo*-**2**-OH was observed and control experiments with standard mixtures gave accurate and reproducible results and showed that decomposition does not occur. Complete resolution of *exo*- and *endo*-**2**-OH was also observed with a 100-ft TCEP column. With these columns base line resolution was observed for any combination of the following: *exo*- and *endo*-**2**-OH and their saturated analogues, and the corresponding saturated and $\alpha,\beta$-unsaturated bicyclic ketones.

**Racemic and Optically Active *exo*-Bicyclo[3.2.1]oct-3-en-2-ol (*exo*-2-OH).** Racemic *exo*-**2**-OH, mp 81.5-83.5 °C (lit.[3] mp 82.7-83.7 °C), >99% exo isomer, was prepared and purified (recrystallization of *p*-nitrobenzoate derivative) as reported earlier.[3] Optically active *exo*-**2**-OH, $[\alpha]^{25}{}_D$ 204° (*c* 0.5, $CHCl_3$) (93% optically pure),[3] was also prepared as reported earlier. Capillary GC showed this material to be homogeneous *exo*-**2**-OH. Anal. ($C_8H_{12}O$) C, H.

**Racemic and Optically Active *endo*-Bicyclo[3.2.1]oct-3-en-2-ol (*endo*-2-OH).** *dl*-Bicyclo[3.2.1]oct-3-en-2-one was prepared by chromic acid oxidation[7] of *dl*-*exo*-**2**-OH and converted to *endo*-**2**-OH by reduction with $LiAlH_4$ as reported earlier.[3] After purification by recrystallization of the *p*-nitrobenzoate derivative followed by silver nitrate extraction to separate saturated contaminants,[3] the product was >97% *endo*-**2**-OH, mp 67.4-68.5 °C, and the only impurity was the exo isomer.

Previous attempts[3] to resolve the endo acid phthalate were unsuccessful. In this work *endo*-**2**-OH was resolved as follows. The acid phthalate derivative of *endo*-**2**-OH, mp 99-104 °C, and an equivalent amount of cinchonidine were dissolved in acetone. After five recrystallizations the head crop of the cinchonidine salt was converted[2] to *endo*-**2**-OH. After purification by chromatography on silica gel as described earlier,[3] *endo*-**2**-OH (>99% endo isomer) was obtained which had mp 67.6-68.2 °C, $[\alpha]^{24}{}_D$ 7.28° (*c* 1.47, $CHCl_3$). From the absolute rotation[3] it can be seen that this material was 50% optically pure. Anal. ($C_8H_{12}O$) C, H.

***exo*-Bicyclo[3.2.1]oct-3-en-2-ol-$^{18}O$.** Solvolysis of the crude allylic bromide derived from 0.69 mol of bicyclo[3.2.1]oct-2-ene[3] in 140 mL of $H_2{}^{18}O$ (~1.5% excess $^{18}O$) over 110 g of calcium carbonate gave a product which after purification[3] was >99.5% *exo*-**2**-OH, mp 79.8-82 °C, 1.3% excess $^{18}O$.

***endo*-Bicyclo[3.2.1]oct-3-en-2-ol-$^{18}O$.** A solution of 45 g of bicyclo[3.2.1]oct-3-en-2-one[3] and 12 mL of $^{18}O$-enriched water (~3.3% excess $^{18}O$) containing a trace of *p*-toluenesulfonic acid was stirred at room temperature for 50 h. The solution was extracted with ether and dried ($MgSO_4$). The ether was removed under reduced pressure and the residual ketone fraction distilled at 20 mm (93 °C) and reduced with $LiAlH_4$.[3] This product consisted of 81% *endo*-**2**-OH, 10% *endo*-bicyclo[3.2.1]octan-2-ol,[8] 2% *exo*-**2**-OH, and 3% *exo*-bicyclo[3.2.1]octan-2-ol.[8] The saturated components were removed by extraction with silver nitrate solution.[3] This extraction results in partial racemization of active *endo*-**2**-OH[3] and presumably in slow oxygen exchange. Extraction was carried out rapidly to minimize loss of $^{18}O$. The purified product used to determine $k_{exc}$ consisted of 91% *endo*-**2**-OH (1.83% excess $^{18}O$), 2% *exo*-**2**-OH, and 7% of the saturated analogues. Since *exo*-**2**-OH is much more reactive than *endo*-**2**-OH and the saturated analogues are unreactive, these impurities do not interfere with determination of $k_{exc}$.

**Kinetic Experiments. A. Polarimetric Rates.** Rate constants for loss of optical activity ($k_\alpha$) were determined as described earlier[1a] except that a 1-dm all-glass jacketed polarimeter tube was used and rotations were determined with a Perkin-Elmer Model 141 polarimeter. Reactions were followed to ~80% completion and in all cases good first-order behavior was observed. Rate constants for loss of optical activity were also determined for both *exo*- and *endo*-**2**-OH by quenching aliquots with aqueous sodium hydroxide at appropriate times and determining the static rotations of the quenched solutions. Values of $k_\alpha$ determined by the two methods were indistinguishable.

**B. Isomerization Experiments.** Two methods were used to determine the rate constant for irreversible exo $\rightleftharpoons$ endo equilibration ($k_{eq}$). In the first, a solution of *endo*-**2**-OH was placed in a thermostated 10-mL volumetric flask. At appropriate intervals aliquots were withdrawn with a syringe and added to aqueous sodium hydroxide to quench the reaction. The basic solutions were extracted with pentane and the dried ($MgSO_4$) pentane extract was concentrated under a stream of dry nitrogen. In the second method the reaction mixture was distributed into glass ampules. At appropriate times ampules were taken from the constant-temperature bath, chilled, and added to base to quench the reaction and the aliquots were extracted with pentane. Configurational compositions of the alcohol fractions were determined by capillary GC and the first-order constant for equilibration ($k_{eq}$) was determined as described earlier.[1a] Reactions were followed to ~75% completion and good first-order behavior was observed in all cases. Control experiments with standard solutions showed that isolation of the alcohol fraction does not result in fractionation and that the analytical method gives the correct configurational compositions.

It was also shown that under conditions of equilibration (the slowest reaction investigated) there is no destruction of **2**-OH. In this experiment a solution of an equilibrium mixture of *exo*- and *endo*-**2**-OH in 33.3% acetone containing 0.182 M $HClO_4$ and 1,1,2,2-tetrachloroethane as an internal standard was heated for >5 half-lives for equilibration. During this period relative peak areas of the alcohol and internal standard remained constant.

**C. Oxygen-Exchange Experiments.** Rate constants for oxygen exchange of $^{18}O$-labeled *exo*- and *endo*-**2**-OH ($k_{exc}$) were determined as described earlier[1a] except that platinized carbon[4] was substituted for carbon black in the Unterzaucher apparatus. Samples were isolated from quenched reaction mixtures by extraction with pentane. After drying ($MgSO_4$), the pentane was removed with a spinning band column and the residual **2**-OH purified by sublimation (90 °C, 6 mm). Control experiments showed that the isolation procedure did not alter the $^{18}O$ content or the configurational composition of the alcohol fraction. Reactions were followed to >75% completion and good first-order behavior was observed.

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## References and Notes

(1) (a) Goering, H. L.; Josephson, R. R. *J. Am. Chem. Soc.* **1962,** *84,* 2779. (b) Goering, H. L.; Schewene, C. B. *ibid.* **1965,** *87,* 3516. (c) Goering, H. L.; Brown, C.; Schewene, C. B. *ibid.* **1968,** *90,* 6214.

(2) Goering, H. L.; Towns, D. L. *J. Am. Chem. Soc.* **1963,** *85,* 2295.

(3) Goering, H. L.; Anderson, R. P. *J. Am. Chem. Soc.,* in press.

(4) Taylor, J. W.; Chen, I.-J. *Anal. Chem.* **1970,** *42,* 224.

(5) Schwers, A. *J. Chim. Phys. Phys.-Chim. Biol.* **1911,** *9,* 15.

(6) Jefford, C. W.; Yne, E. H. *Tetrahedron* **1967,** *23,* 4549.

(7) Brown, H. C.; Garg, C. P. *J. Am. Chem. Soc.* **1961,** *83,* 2952.

(8) Georing, H. L.; Fickes, G. N. *J. Am. Chem. Soc.* **1968,** *90,* 2862.

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# Superacid-Catalyzed Transfer Nitration of Benzene, Toluene, and Mesitylene with 9-Nitroanthracene and Pentamethylnitrobenzene. Evidence for the Reversibility of Electrophilic Aromatic Nitration in Specific Cases[1]

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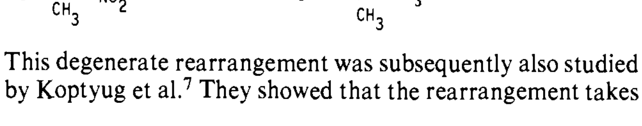
**Abstract:** The reversibility of electrophilic aromatic nitration was proven in certain, specific cases of superacid-catalyzed transfer nitration reactions. In the reaction of benzene, toluene, and mesitylene with 9-nitroanthracene and pentamethylnitrobenzene, nitrobenzene, nitrotoluenes, and nitromesitylene were obtained in the presence of superacids. The primary kinetic hydrogen isotope effect, $k_H/k_D$, of 2.25 ± 0.05 in the nitration of anthracene-$d_{10}$ with nitronium hexafluorophosphate in nitromethane solution is also in accord with the reversibility of the nitration of anthracene.

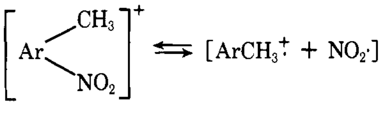
## Introduction

Acid-catalyzed nitration is a typical electrophilic aromatic substitution reaction.[2,3] Nitration is generally considered to be an irreversible reaction, and nitroaromatics, in general, do not undergo rearrangement or isomerization under the reaction conditions.[4] Relative reactivities and substituent effects have been extensively studied in nitration of aromatics, assuming that the nitro compounds formed do not revert to starting materials. Otherwise, relative reactivities calculated on the basis of product ratios would become affected and require a more rigorous treatment, taking into account the reverse reaction.

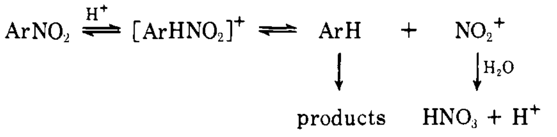
Examples of acid-catalyzed migration of nitro groups have been reported.[4] The mechanisms of these migrations, however, are not always fully established and may involve, at least in some instances, radical cations (as in the case of nitramine rearrangements[5]). In the case of hexamethylnitrobenzenium ion, the intramolecular nitro group migration was directly studied by NMR spectroscopy, and the complex was also found capable of transnitrating aromatics:[6]



This degenerate rearrangement was subsequently also studied by Koptyug et al.[7] They showed that the rearrangement takes place through successive intramolecular 1,2 migrations of the nitro group and ruled out the possibility of an alternative pathway proceeding through the formation of a radical cation.



In the course of acid-catalyzed nitration of aromatics, the reversibility of nitration was, however, until now not directly established. Gore[8] reported in 1957 that heating of 9-nitroanthracene with 12 N sulfuric acid in trichloroacetic acid solution for 15 min at 95 °C gave a dark solution in which, after dilution with water, an 81% yield of free nitric acid was estimated. Anthracene itself could not be isolated, but a 20% yield of anthraquinone and appreciable amounts of polymer and soluble sulfonic acids were obtained. Experiments conducted in the presence of added nitrobenzene to detect cross-nitration products failed. The system was represented as a protolytic denitration reaction:



$$\mathrm{ArNO_2} \overset{\mathrm{H^+}}{\rightleftharpoons} [\mathrm{ArHNO_2}]^+ \rightleftharpoons \underset{\substack{\downarrow \\ \text{products}}}{\mathrm{ArH}} + \underset{\substack{\downarrow \mathrm{H_2O} \\ \mathrm{HNO_3} + \mathrm{H^+}}}{\mathrm{NO_2^+}}$$

Oxidative degradation, however, could also account for the observed products.

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