

An AM1 Study of an α -Nucleophile:

Geometries and Interconversion Modes of Oximate Anion Stereoisomers

S. Hoz^a and E. Bunce^b

^a Department of Chemistry, Bar-Ilan University, Ramat Gan, Israel 52100

^b Department of Chemistry, Queen's University, Kingston, Canada K7L 3N6

(Received in USA 18 August 1988)

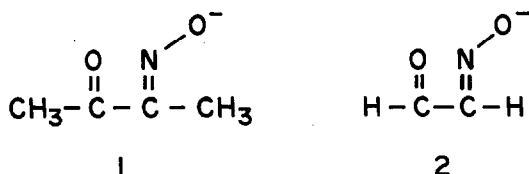
ABSTRACT

The geometries, the relative stabilities and the interconversions of the four isomers of glyoxal monoximate anion (2) were studied at the AM1 level. In the ground state all four geometries were found to be planar, with the stability order S1 > W > S > U. There are three factors that tend to stabilize U: (a) non-bonded interactions; b) $n_N-\sigma_{C-C}^*$ interaction; and (c) electrostatic repulsion between n_N and the partial negative charge on the carbonyl oxygen. In contrast, in the all-carbon system 3, W is more stable than S. The reversal in stability in the case of 2 probably results from an $n_N-\sigma_{C-C}^*$ interaction which favours S1 over W. The barrier to rotation of the carbonyl about the C-C bond is ca. 12 kcal/mol, reflecting the partial double bond character in the latter. Rotation and in-plane-inversion are the two possible paths for tilting the oximate oxygen from one side to the other. In the case of formaldoximate anion, the activation energies for these two processes are nearly the same. In 2, however, rotation is strongly assisted by the presence of the carbonyl group, which lowers its TS energy to ca. 23 kcal/mol as compared to 51 kcal/mol for in-plane-inversion. The AM1 results were corroborated, with one minor exception, by MNDO calculations.

INTRODUCTION

α -Nucleophiles are distinguished from other (normal) nucleophiles in having a lone-pair of electrons in a position vicinal to the nucleophilic site.¹ This unique electronic feature, as well as the enhanced nucleophilicity (α -effect) of these nucleophiles, has stimulated intensive studies of their physical properties.² One of these properties is the geometry, which for most α -nucleophiles is rather straightforward. It is very simple for nucleophiles such as ClO^- , and somewhat more complex for nucleophiles such as hydrazine and its derivatives, which have been extensively studied.³ The geometry of oximate anions has received much less attention, although oximate anions have frequently been used in α -effect studies over nearly three decades.⁴⁻⁶

One of the more popular members of this group of oximate anions is 2,3-butanedione monoximate, 1. This nucleophile is potentially ambidentate in nature, since in principle



both oxygen atoms can partake in nucleophilic reactions. However experimental evidence points to the oximate-oxygen being the only nucleophilic site.⁷

The butanedione monoximate anion may exist as an equilibrium mixture of four major isomeric structures. The relationship among these is essential to understanding of the chemical reactivity of this nucleophile. In the present paper, we report a study, using the semi-empirical method AM1, on the dynamics of the geometric interconversions of glyoxal monoximate 2, which is derived from 1 by replacing the two methyl groups by hydrogen atoms. All calculations pertaining to 2 were also performed using the MNDO method.

RESULTS AND DISCUSSION

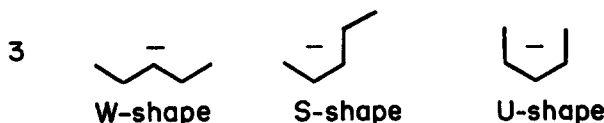
Computational Methods

The calculations were carried out on an IBM-3081 computer using the standard AM1 procedure as implemented in the AMPAC program version 2.10.⁸ The geometries of all four isomers were completely optimized. Transition states were located by using either the NLLSQ or the SIGMA routines (implemented in AM1) and were characterized as such by having a single negative force constant. Most of the calculations were repeated with the MNDO program.^{8,9}

Ground States

The geometries and heats of formation of the four isomers W, S1, S2 and U are given in Fig. 1a. All isomers were found to be planar. Their stability decreases along the series S1 > W > S2 > U.

Pertinent to the present study is a comparison with the analogous all-carbon system 3 (pentadienyl anion). Contrary to earlier suggestions that for 3 the U geometry is the most stable,¹⁰ it was later shown (nmr studies and theoretical calculations) that the W-shape is more stable than the S-shape, which in turn is more stable than the U-shape.¹¹



We have found that in the case of 2, U is indeed the least stable isomer. However, one of the S isomers (S1) is now more stable than the W.

In the following we will discuss the relative stability of U and the reversal in the stability order between S1 and W in the present system. The U isomer is the most suitable

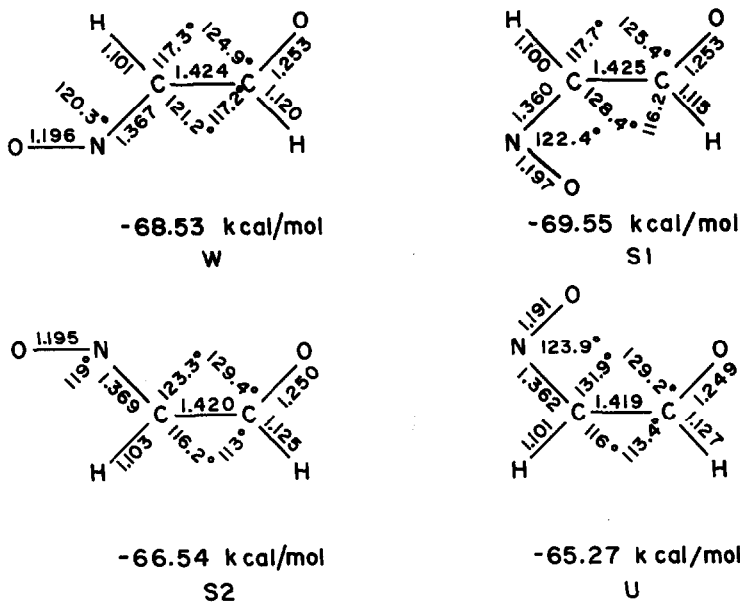


Figure 1a. Heats of formation and structural parameters (AM1, energies in kcal/mol, angles in degrees and bond lengths in Å) for the pair isomers of **2**.

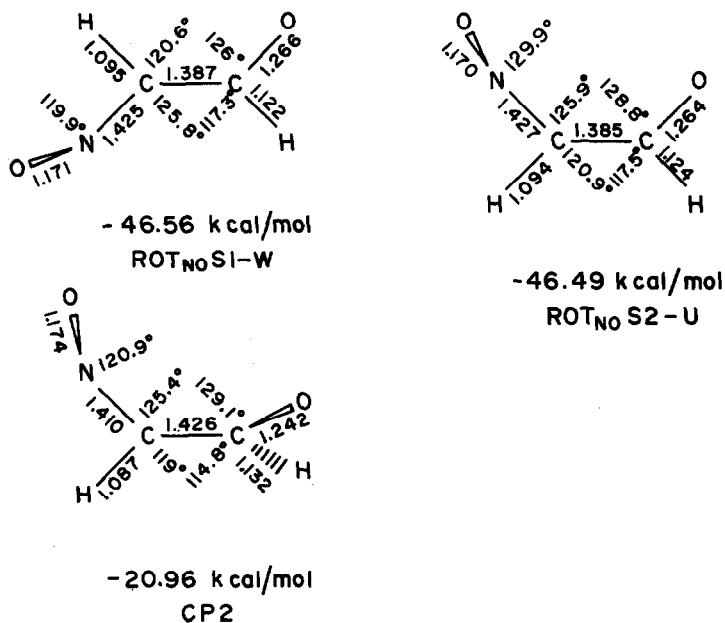


Figure 1b. Heats of formation and structural parameters (AM1, energies in kcal/mol, angles in degrees and bond lengths in Å) for oximate rotation and the center point in the rotation-rotation potential surface of **2**.

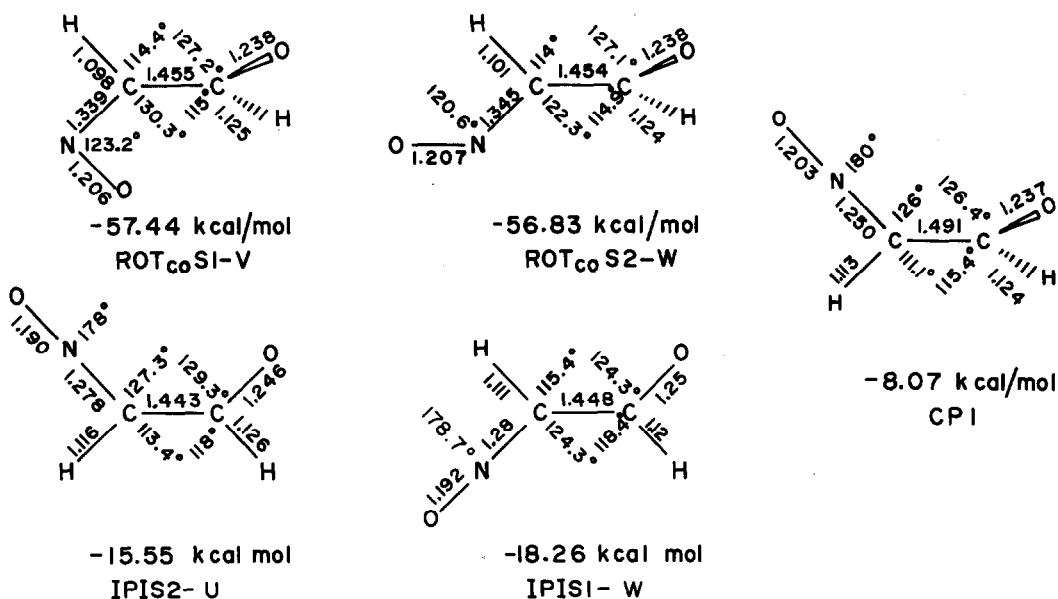


Figure 1c. Heats of formation and structural parameters (AM1, energies in kcal/mol, angles in degrees and bond lengths in Å) for transition structures on the rotation-inversion potential surface of 2.

candidate for attractive non-bonded interactions. Inspection would suggest this configuration of the anion to be capable of exhibiting homoaromaticity.^{10b} On this account one would expect it to be more stable than the other isomers. An additional effect which should further stabilize it, at least with respect to S2, is the $n_{\text{N}}-\sigma_{\text{C-C}}^*$ interaction. It was shown by Epiotis and coworkers that the favored geometry for this interaction is the one in which the lone-pair is located anti to the σ bond,¹² and that this interaction is stronger with σ^* of C-C bonds than with that of C-H bonds.¹³ In the present case, this is probably much more so since the carbonyl-carbon bears a partial positive charge (0.24 according to Mulliken population analysis), even though the molecule is negatively charged overall. Therefore, the $n-\sigma^*$ interaction is likely to be stronger in U, where the lone-pair on the nitrogen is located anti to the C-C bond, than in S2, where it is suitably located to interact with the σ^* orbital of the C-H bond. Furthermore, the interaction of a lone-pair with a negative charge, as suggested by Houk et al.,¹⁴ may also play an important role in determining the stabilities of the isomers. This factor should also destabilize S2 relative to U, since in the former the nitrogen lone-pair is directed toward the negatively charged carbonyl oxygen. Nevertheless, U was found to be the least stable isomer.

We believe that the main reason for the relative instability of U is the repulsive electrostatic interaction between the two terminal oxygen atoms, each of which bears ca. $0.5e^-$. This becomes evident from examination of the internal bond angles in U which, as can be seen from Fig. 1a, are significantly larger than the corresponding angles in the most relaxed structure W. In fact, with the exception of a single angle in one isomer (OCC in S2), all angles in U are larger than the corresponding angles in all the other isomers.

An additional indication of a strong electrostatic repulsion in U was obtained from calculation of the energies of the corresponding radicals in the W and U geometries. In the fully optimized geometries of the radical species, the net negative charge on each of the oxygen atoms is found to be less than half that in the corresponding anions. The electrostatic repulsion is, therefore, drastically reduced and, as expected, in the radical, the U isomer was indeed found to be more stable than the W, by ca. 8 kcal/mol.

Finally, an oversimplified calculation of the sum of the electrostatic interactions over the pairs O,O, O,N and O,C, assuming point charges and a dielectric constant of unity, shows that U is less stable than W by 3.2 kcal/mol.

It has thus been shown that in 2 as well as in the pentadienyl anion 3, the U structure is the least stable one. However, in 2 there are additional effects which are absent in 3. The lone-pair on the nitrogen induces a relative stabilization of U, as shown above, but this is countered by the larger electrostatic repulsion which stems from the substantial negative charges on the terminal oxygen atoms, thus leaving U as the least stable isomer.

The reversal in stability of S1 and W, in 2 versus 3, may again be due to a better $n-\sigma^*$ interaction in S1 as compared to W. The magnitude of this effect in systems such as vinyl ether or formate esters ranges from 1.5 to 6 kcal/mol.¹³ In the absence of this stabilizing effect, the energy of S1 would have probably neared that of S2 and may even exceed that of U, as found in 3.

This reversal in stability disappeared when calculations were performed at the MNDO level, which led to the stability order W > S1 > S2 > U. In the absence of suitable experimental data, it remains unclear whether this reversal is real or an artifact of the method(s) used.

Introduction of the methyl groups into 2, to produce 1, will probably not alter greatly the stability order. However, the two less stable isomers, S2 and U, in which the Me groups are syn to each other, will be somewhat more destabilized and may also to some extent deviate from planarity.

Isomerization

Transformation of one isomer of 2 to the other, may involve one or both of the following processes. The first is rotation of the carbonyl relative to the oxime function, about the C-C bond. The second process is a tilt of the oximate oxygen, from one side

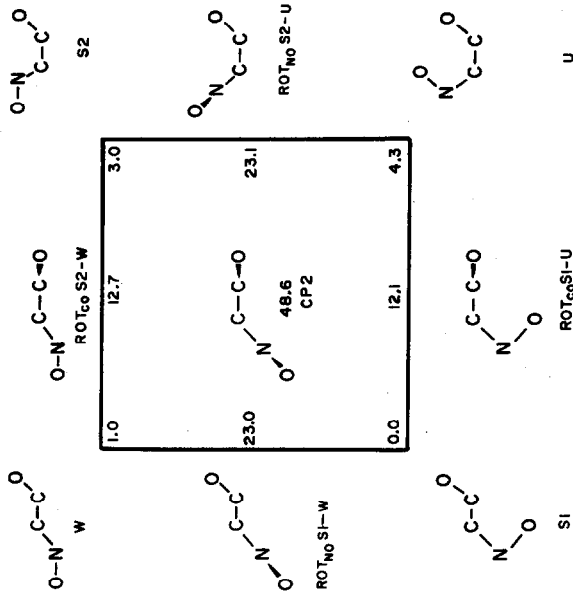


Figure 3. Rotation-rotation potential surface for 2. Energies (kcal/mol) are relative to S1.

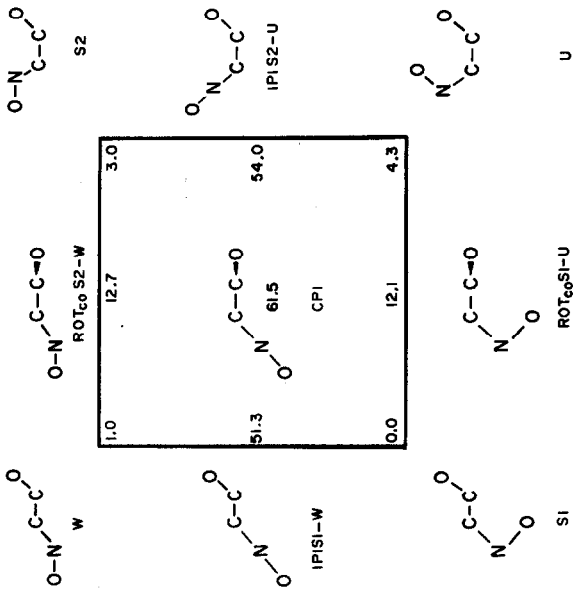


Figure 2. Rotation-inversion potential surface for 2. Energies (kcal/mol) are relative to S1.

of the C=N bond to the other. The pairs S1 and U, and S2 and W, are interconvertible by rotation of the carbonyl function (transition states ROT_{CO}S1-U and ROT_{CO}S2-W, respectively, in Fig. 2). The activation barrier in both cases is of similar height (Fig. 2). The relatively high value of the barrier for rotation about this formal single bond clearly results from partial double bond character between the two carbon atoms.

For the tilt of the oxygen, from one side of the C=N bond to the other, two mechanistic paths are available: in-plane-inversion (IPI, transition states: IPIS1-W and IPIS2-U, Fig. 2), and rotation (ROT) about the C=N bond (transition states: ROT_{NO}S1-W and ROT_{NO}S2-U, Fig. 3). It is found that, for the analogous formaldoximate system 4, the transition states of the two processes are of nearly the same energy (Fig. 4).

We first analyze the IPI transition state. The activation energy found for IPI in 2 is of the same magnitude as that calculated for 4. A rotation-inversion potential surface composed of carbonyl rotation and IPI offers, for interconversion of two isomers located at opposite corners of the potential surface (Fig. 2), two alternative extreme paths. One involves motion along the edges of the potential surface with consecutive rotation of the C=O function and in-plane-inversion of the oximate-oxygen, the latter process being associated with the higher energy barrier. The other is the path crossing diagonally through the center of the diagram.

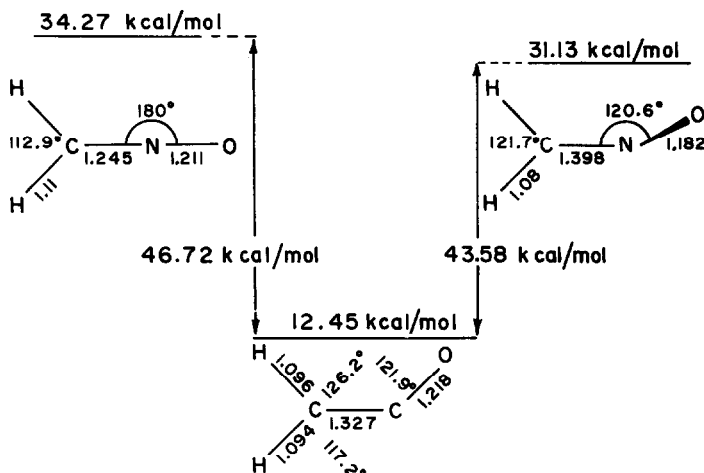
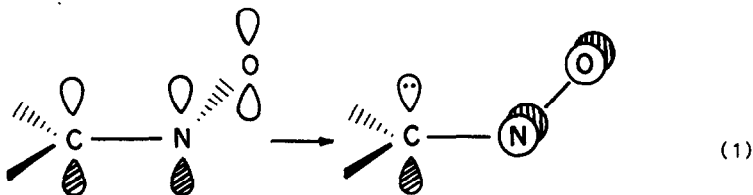


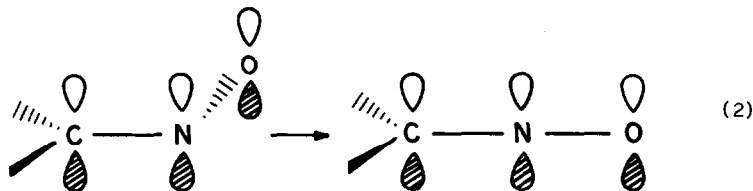
Figure 4. Heats of formation and geometrical parameters for ground state, and rotation and inversion transition states, of formaldoxime anion (AM1, energies are in kcal/mol, angles in degrees and bond lengths in Å).

In general, in cases where there is no differential coupling between ground states and transition states, the central point in the surface will reflect the instabilities of the transition states by which it is flanked. Thus the energy of the central point will be nearly the sum of the energies of the individual transition states. As seen from Fig. 2, the relative energy of the central point (CP1, 61.5 kcal/mol) is, indeed, nearly the sum of the relative energies of the transition states for carbonyl rotation (12 kcal/mol) and oximate-oxygen IPI (51 kcal/mol). However, making use of the other process (ROT) for the tilt of the oximate oxygen, leads to an entirely different result. This process in 2 is strongly coupled and assisted by the carbonyl function. As a result, the activation energy for ROT becomes drastically reduced, from ca. 44 kcal/mol in 4 to 23 kcal/mol in 2. The differential coupling is also reflected in the energy of the central point (CP2) in the rotation-rotation potential surface (Fig. 3), which is higher by 13 kcal/mol than the sum of two individual edgewise transition state energies.

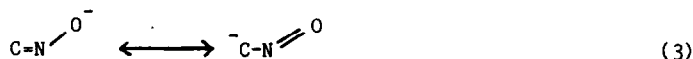
The fact that while the activation energy for IPI is nearly unaffected by the presence of the carbonyl group, the ROT process strongly benefits from it, is worth commenting upon. Focusing on the C-N-O portion of the π -system (allylic-like), it is seen (eq. 1) that in the IPI transition state, this π -system is not significantly affected:



Therefore, disrupting the conjugation by rotation of the carbonyl group (horizontal motion in Fig. 2) results in practically the same energy change, whether one starts from one of the isomers or from the IPI transition state. On the other hand, for the ROT process, the extent of delocalization is drastically affected upon going from the ground state to the transition state. In the ground state the negative charge resides mainly on the oxygen, due to its higher electronegativity. Rotation about the C-N bond will orthogonalize the two p orbitals on N and O that form the N=O double bond with respect to the p orbital on the carbon, which will then be doubly occupied (eq. 2):



In valence bond description, this structure has a single bond between the two carbons, thus enabling "free" rotation about this bond (eq. 3):



The negative charge on the carbon can now be delocalized onto the carbonyl function, causing significant stabilization of the transition state. Carbonyl rotation out-of-plane will leave the negative charge localized on the carbon atom, thus increasing the energy of the whole system (central point in Fig. 3).

Heats of formation and major structural parameters of all stationary points are given in Figs. 1a-1c. Optimized geometries of acetaldehyde imine anion are given in Fig. 5 together with Houk's values using 4-31G.¹⁴

MNDO calculations of the heats of formation of the transition states for isomerization of 2, relative to AM1, gave the same results with an average deviation of 1.3 kcal/mol.

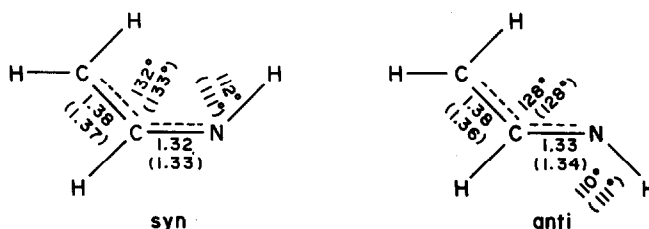


Figure 5. AM1 optimized geometries of acetaldehyde imine anion. Values in parentheses were taken from Houk's study (4-31G, ref. 14).

SUMMARY AND CONCLUSIONS

Semiempirical calculations using AM1 show that in the glyoxal monoximate anion 2, the order of isomer stability is S₁ > W > S₂ > U. This order resembles the one found for the all-carbon analog-pentadienyl anion 3, with the variation that S₁ is slightly more stable than W. This can probably be traced to a $n_{\text{N}}-\sigma^*_{\text{C-C}}$ interaction which stabilizes the former. The instability of U in the case of 2 most probably results from electrostatic repulsion between the two terminal oxygen atoms. This is inferred from: (a) analysis of the bond angles which in general are larger in U; (b) comparison with the uncharged analogous radical system; and (c) calculations of electrostatic interactions.

The isomerizations W \rightleftharpoons S₂ and U \rightleftharpoons S₁ are achieved by rotation about the C-C bond which is associated with an activation energy of ca. 12 kcal/mol. For the isomerizations W \rightleftharpoons S₁ and U \rightleftharpoons S₂, the transition state involves rotation about the C-N bond. The latter process ($E_a \approx 20$ kcal/mol) was found to be lower by ca. 30 kcal/mol than the transition state for an in-plane-inversion of the oximate-oxygen. The presence of the conjugated carbonyl function reduces the activation energy for rotation of this oxygen by ca. 20 kcal/mol, as compared to the unactivated system 4.

REFERENCES AND FOOTNOTES

1. Reviews:
 - (a) Fina, N.J.; Edwards, J.O. Int. J. Chem. Kinetic. 1973, 5, 1.
 - (b) Grekov, A.P.; Veselov, V.Y. Usp. Khim. 1978, 47, 1200.
 - (c) Hoz, S.; Buncel, E. Isr. J. Chem. 1985, 26, 313.
2.
 - (a) Jencks, W.P.; Gilchrist, M. J. Am. Chem. Soc. 1968, 90, 2622.
 - (b) Dixon, J.E.; Bruce, T.C. J. Am. Chem. Soc. 1972, 94, 2052.
 - (c) Dessolin, M.; Laloi-Diard, M.; Vilkas, M. Tetrahedron Lett. 1974, 2405.
 - (d) Nelsen, S.F.; Buschek, J.M. J. Am. Chem. Soc. 1974, 96, 2392.
 - (e) Foster M.S.; Williamson, A.D.; Beachamp, J.L. Int. J. Mass. Spectrom. Ion. Phys. 1974, 15, 429.
 - (f) Bierbaum, V.M.; Schmitt, R.J.; DePuy, C.H. J. Am. Chem. Soc. 1981, 103, 6262.
 - (g) Buncel, E.; Chuaqui, C.; Wilson, H. Int. J. Chem. Kinetic. 1982, 14, 823.
 - (h) Evanseck, J.D.; Blake, J.F.; Jorgensen, W.L. J. Am. Chem. Soc. 1987, 109, 2349.
3.
 - (a) Nelsen, S.F.; Frigo, T.B.; Kim, Y.; Thompson-Colon, J.A. J. Am. Chem. Soc. 1986, 108, 7926.
 - (b) Nelsen, S.F. Acc. Chem. Res. 1981, 14, 131.
 - (c) Bock, H.; Solouki, B. Angew. Chem. Int. Ed. Engl. 1981, 20, 427.
4.
 - (a) Jencks, W.P.; Carriuolo, J. J. Am. Chem. Soc. 1960, 82, 1778.
 - (b) Edwards, J.O.; Pearson, R.G. J. Am. Chem. Soc. 1962, 84, 16.
 - (c) Aubort, J.D.; Hudson, R.F. Chem. Commun. 1969, 1342; ibid., 1970, 937.
 - (d) Aubort, J.D.; Hudson, R.F.; Woodcock, R.C. Tetrahedron. Lett. 1973, 2229.
 - (e) Laloi-Diard, M.; Verchere, J.F.; Gosselin, P.; Terrier, F. Tetrahedron Lett. 1984, 25, 1267.
 - (f) Buncel, E.; Um, I.H. J. Chem. Soc. Chem. Commun. 1986, 595.
5. There are several reports regarding the structure of oxime derivatives: Huige, C.J.M.; Hezemans, A.M.F.; Rasmussen, K. J. Comput. Chem. 1987, 8, 204; Nguyen, M.T.; Ha, T.K. J. Mol. Struct. Theochem. 1982, 88, 127; Leroy, G.; Nguyen, M.T.; Sana, M.; Villaveces, J.L. Bull. Soc. Chim. Belg. 1980, 89, 1023; Lhamyani-Chraibi, M.; Gelize-Duvigneau, M.; Arriau, J.; Elguero, ibid., 1980, 89, 71. The latter reports also CNDO/2 calculations on the geometry of the formaldoximate ion.
6. Buncel, E.; Um, I.H.; Hoz, S. J. Am. Chem. Soc. in press.
7. Buncel, E.; Um, I.H., to be published.
8. Dewar, M.J.S.; Zoebisch, E.G.; Healy, E.F.; Stewart, J.J.P. J. Am. Chem. Soc., 1985, 107, 3902; Dewar, M.J.S.; Rzepa, H.S. J. Am. Chem. Soc. 1978, 100, 58.
9. In a preliminary study, we have calculated using AM1 the fully optimized geometry of the two isomers of acetaldehyde imine anion, a system previously studied at the 4-31G level by Houk and coworkers.¹⁴ The syn isomer was found to be more stable than the anti by 5.2 kcal/mol, as compared to 4.7 kcal/mol reported by Houk.¹⁴ In addition to this good agreement between AM1 and the ab initio split basis set energy calculations, the geometrical parameters (Fig. 5) were also found to be in a very good agreement.
10.
 - (a) Bates, R.B.; Carnighan, R.H.; Staples, C.E. J. Am. Chem. Soc. 1963, 85, 3031;
 - (b) Hoffmann, R.; Olofson, R.A. J. Am. Chem. Soc. 1966, 88, 943.
11. Dewar, M.J.S.; Fox, M.A.; Nelson, D.J. J. Organometal. Chem. 1980, 185, 157; Yasuda, H.; Yamauchi, M.; Ohnuma, Y.; Nakamura, A. Bull. Chem. Soc. Jpn. 1981, 54, 1481.
12. Epiotis, N.D.; Yates, R.L.; Larson, J.R.; Kirmaier, K.; Bernardi, F. J. Am. Chem. Soc. 1977, 99, 8379; Epiotis, N.D.; Cherry, W.R.; Shaik, S.; Yates, R.L.; Bernardi, F. Top. Curr. Chem. 1977, 70.
13. Larson, J.R.; Epiotis, N.D.; Bernardi, F. J. Am. Chem. Soc. 1978, 100, 5713.
14. Houk, K.N.; Strozler, R.W.; Rondan, N.G.; Fraser, R.R.; Chuaqui-Offermans, N. J. Am. Chem. Soc. 1980, 102, 1426.