in the isoprene chains. Chemically "transforming" double bonds from *cis* to *trans* causes a much greater change in the dimensions of a diene chain already having a considerable fraction of 0° states about other bonds, since for the all-*trans* diene chain $\lim_{n \to \infty} \langle r^2 \rangle_0 / nl^2 = \infty$. Similarly, the presence of *trans* double bonds enhances the effect of the suppression of ±120° states about CH_2–CH_2 bonds. Thus, whereas the characteristic ratio of *cis*-PIP should be only 7% greater than that of *cis*-PBD, the value for *trans*-PIP exceeds that for *trans*-PBD by 27%.

The moderately large negative temperature coefficient $d \ln \langle r^2 \rangle_0 / dT = -0.65 \times 10^{-3}$ deg^{-1} for *trans*-PBD results primarily from transitions 180°, 0° → 180°, ±120° about CH–CH_2–CH_2 bond pairs as the temperature is increased. In *trans*-PIP, an additional contribution arises from the increase in the number of ±120° states about CH_2–CH_2 bonds with increasing temperature, and a large, negative coefficient $d \ln \langle r^2 \rangle_0 / dT = -1.4 \times 10^{-3}$ deg^{-1} is predicted. In contrast, the positive temperature coefficients of $\langle r^2 \rangle_0$ found for *cis*-PBD and *cis*-PIP are due primarily to transitions to more extended conformations about CH_2–C_2H_2–CH_2 bond pairs (*e.g.*, ±60°, ±60° → 0°, ±60°) as the temperature is increased.

Thus, on the basis of presently available experimental results, a reasonably consistent interpretation of the configurational properties of both the *trans* and *cis* forms of PBD and PIP has been obtained.

Both *trans*-PBD and *trans*-PIP crystallize in a number of forms, but the configurations of the chains in these modifications have not yet been unambiguously determined.[37–41] Comparison of the configurations of these chains in a crystalline lattice with the configuration of lowest intramolecular energy, in an attempt to determine the effect of intermolecular interactions,[2] is therefore not possible at the present time.

Acknowledgment. The author wishes to thank the Computer Center of the Polytechnic Institute of Brooklyn for the use of their facilities.

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Valence-Shell Calculations on Polyatomic Molecules. I. CNDO SCF Calculations on Nitrogen and Oxygen Heterocyclics[1a]

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Received July 10, 1967

Abstract: Valence-shell SCF MO calculations using the complete neglect of differential overlap approximation (CNDO/2) are described for the azines and for five-membered heterocyclic compounds containing carbon, nitrogen, and oxygen. The calculated dipole moments, unlike those of similar calculations using the extended Hückel method, are in excellent agreement with experiment. The total charge densities, but not the π-electron densities, are shown to be linearly related to C^{13} chemical shifts. The Karplus–Pople theory of chemical shifts is found to lead to a less satisfactory agreement between theory and experiment, especially for carbon atoms without hydrogens.

There have been many attempts to explain the ground-state properties of molecules, *e.g.*, dipole moments and nmr chemical shifts, in terms of molecular orbital (MO) theory. Most of these attempts have been confined to the behavior of the π electrons, or the effect of the σ electrons has been included in a very empirical fashion.[2,3] Originally this was necessary because of limitations of computation, but this is no longer so, and *ab initio* calculations of the Roothaan LCAO type, including all the electrons and using large basis sets of atomic orbitals, are now possible for quite large molecules.[4] However, such calculations are very expensive to carry out and, although pilot calculations on a few reference molecules of different types will be of enormous value,[5] it seems likely that there will always be room for semiempirical methods in the interpretation and prediction of relationships between structure and properties.

There are at present two categories of methods: (1) numerous versions of the extended Hückel theory (EHT), which includes overlap but neglects electron-repulsion integrals completely have been suggested,[6–9]

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(1) (a) Presented in part at the 153rd National Meeting of the American Chemical Society, Miami Beach, Fla., April 1967; (b) abstracted in part from the M.S. Thesis of D. L. Breen, June 1967, University of Virginia.

(2) (a) A. Streitwieser, Jr., "Molecular Orbital Theory for Organic Chemists," John Wiley and Sons, Inc., New York, N. Y., 1961; (b) G. Del Re, *J. Chem. Soc.*, 4031 (1957); (c) H. Berthod and A. Pullman, *J. Chim. Phys.*, **62**, 942 (1965); (d) L. E. Orgel, T. L. Cottrell, W. Dick, and L. E. Sutton, *Trans. Faraday Soc.*, 47, 113 (1951).

(3) H. F. Hameka and A. M. Liquori, *Mol. Phys.*, **1**, 9 (1958); *Tetrahedron*, **18**, 985 (1962).

(4) I. G. Csizmadia, M. C. Harrison, J. W. Moskowitz, and B. T. Sutcliffe, *Theoret. Chim. Acta*, **6**, 191, 217 (1966).

(5) Calculations of this type have been reported on molecules such as benzene, fluorobenzene, pyridine, and pyrazine (ref 27 of E. Clementi and D. R. Davis, *J. Computational Phys.*, **1**, 223 (1966)), but no details were given.

(6) R. Hoffmann, *J. Chem. Phys.*, **39**, 1397 (1963).

(7) M. D. Newton, F. P. Boer, and W. N. Lipscomb, *J. Am. Chem. Soc.*, **88**, 2353 (1966).

and (2) SCF MO methods which, by introducing the approximation of zero differential overlap into the SCF LCAO formalism of Roothaan, simplify the computational problems enormously by eliminating all but one- and two-center repulsion integrals.[10-12] The final choice of a semiempirical method is, at present, largely a matter of personal preference, a preference which must then be justified by the value of the results. We have chosen to investigate methods of the second category, particularly of the type described by Pople, *et al.*[11] (the CNDO/2 method).

In this first paper we present the results obtained for the calculations of dipole moments, chemical shifts, and ionization potentials of a number of heterocyclic molecules containing carbon, hydrogen, oxygen, and nitrogen. Prior to this study, reports in the literature of applications of the CNDO/2 method have been restricted to small molecules of the AB, AB_2, and AB_3 type and to ethane and formaldehyde.[11,12] The calculations presented here are therefore the first to be reported which take explicitly into account, within a general SCF method, the effect of σ electrons on the dipole moments and chemical shifts of polyatomic organic molecules containing heteroatoms. However, while this work was in progress, three studies of heterocyclic molecules using methods of category (1), *i.e.*, EHT methods, have been published.[7,8,13] In the first and most extensive of these studies,[7] calculations on the dipole moments and ionization potentials of a large number of molecules, including some of the heterocyclics included in this study, were reported. In the second of these studies,[8] the charge densities of a number of heterocyclic compounds were calculated and used to interpret proton and C^{13} chemical shifts.

In the third study,[13] the dipole moments of a large number of heterocyclics were reported. We believe that the gross exaggeration of charge transfer, generally found in EHT methods, which results in the extremely high dipole moments reported in this latter study, compared to the excellent agreement with experiment we have found (Tables I and II), is a very strong vindication for our preference for the CNDO/2 method over the EHT method.

Table I. Dipole Moments of Azines

Molecule	Axis[a]	μ^{σ} [a]	μ^{π} [b]	μ_{SP}[c]	μ_{LP}[d]	μ_{calcd}[e]	μ_{obsd}[f]
Pyridine		0.321	−0.329	1.54	1.80	2.19	2.21
Pyridazine		0.198	0.644	2.92	1.83	3.76	3.97
Pyrimidine		0.229	0.308	1.72	1.81	2.26	2.40
Quinoline	*X*	−0.137	0.127	0.11	1.82	2.34	2.31
	Y	0.171	0.56	1.56			
Isoquinoline	*X*	−0.29	−0.25	0.74	1.80	2.20	2.75
	Y	0.06	0.06	1.38			

[a] For quinoline and isoquinoline the *Y* axis passes through the two central C atoms (no. 9 and 10) and the *X* axis through the midpoint of the 9–10 bond. [b] Calculated from appropriate atomic densities and eq 4. [c] Calculated from eq 5. [d] Calculated using only P_{2s2p} values in eq 5 for the nonbonded pair on one of the N atoms. [e] Calculated from eq 6. [f] Taken from ref 21.

Method

The details of the method and the actual parameters have been fully discussed elsewhere.[11] One of the most desirable features of the CNDO/2 method is that, once the parameters have been chosen to reproduce the one-electron properties of simple diatomic molecules, there is no way of empirically adjusting the parameters, even though in polyatomic molecules the environment of the atom may be quite different from the original molecule chosen for calibration. For example, the only empirical parameter for molecular nitrogen is the bonding parameter β^0_{A-A}, which was chosen by comparison of CNDO calculations with minimal-basis-set LCAO SCF calculations on the N–H radical.[14] The core integrals are obtained using valence-state ionization potentials and are the values used in ref 11. The repulsion integrals were calculated using Slater 2s atomic orbitals. When the SCF MO's had been obtained, the dipole moments were calculated in terms of the bond order matrix $P_{\mu\nu}$ using eq 1–6,[11] where μ_{at} is the atomic

$$P_{\mu\nu} = 2\sum_{i}^{\text{occ}} C_{i\mu}C_{i\nu} \tag{1}$$

$$P_A = \sum_{\substack{\text{orbitals}\\ \text{on A}}} P_{\mu\mu} \tag{2}$$

$$Q_A = Z_A - P_A \tag{3}$$

$$\mu_{at}(X) = 2.5416 \sum_{A}^{\text{atoms}} Q_A(X_A) \tag{4}$$

$$\mu_{sp}(X) = -7.337 \sum_{A}^{\text{atoms}} \frac{P_{2s(A)2p_x(A)}}{Z_A} \tag{5}$$

$$\mu_{total} = \mu_{at} + \mu_{sp} \tag{6}$$

charge contribution to the total moment, and μ_{sp} contains the contributions from the atomic dipoles including the very important lone-pair contributions.

The C^{13} chemical shifts were calculated using the theory of Karplus and Pople.[15] Thus, the chemical shift at a carbon atom A is considered to be made up of a diamagnetic contribution $\sigma_d{}^A$ and a paramagnetic contribution $\sigma_p{}^A$. These contributions are defined in terms of the excess charge density Δq^A and the bond orders between 2p atomic orbitals $P_{\mu\nu}$ (eq 7–12).

$$\Delta q^A = 4 - \sum_{\mu}^{\substack{\text{orbitals}\\ \text{on atom A}}} P_{\mu\mu} \tag{7}$$

$$\sigma_d{}^A = 57.85 - 8.23\Delta q^A \tag{8}$$

$$\sigma_p{}^A = \frac{-(103.57 + 33.46\Delta q^A)}{\Delta E} \sum_{B}^{\substack{\text{all}\\ \text{atoms}}} Q_{BA} \tag{9}$$

$$Q_{AA} = \frac{4}{3}(P_{X_AX_A} + P_{Y_AY_A} + P_{Z_AZ_A}) - \frac{2}{3}P_{Y_AY_A}P_{Z_AZ_A} + P_{Z_AZ_A}P_{X_AX_A} + P_{X_AX_A}P_{Y_AY_A} + \frac{2}{3}(P_{X_AX_A}P_{Y_AY_A}) \tag{10}$$

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Table II. Dipole Moments of Five-Membered Heterocyclic Compounds

Molecule	μ^{σ} [a]	μ^{π} [a]	μ_{SP} [b]	μ_{LP} [c]	μ_{calcd} [d]	μ_{obsd} [e]
Pyrrole	1.42	−2.62	−0.80	. . .	−2.00	2.20
Furan	2.09	−1.83	0.53	1.35	0.8	0.72
Indole					1.86	2.0
Imidazole					−4.09	4.02
Pyrazole					2.71	2.21[f]
Isoxazole					3.17	3.01
Oxazole					1.34	1.40
1,2,5-Oxadiazole	−1.96	0.66	−2.23	1.87 (N) 1.27 (O)	−3.52	3.36[g]
1,3,4-Oxadiazole	−1.74	2.40	2.24	1.88 (N) 1.27 (O)	2.89	3.0[h]
1,2,3-Triazole					4.30	1.79
1,2,5-Triazole	−1.25	1.82	−0.36	1.89	0.20	
1,2,3,5-Tetrazole					2.35	5.15
1,2,3,4-Tetrazole					5.23	

[a] Calculated using eq 4. [b] Calculated using eq 5. [c] Calculated using P_{2s2p} on heteroatom indicated. [d] Calculated using eq 6. [e] Taken from ref 21. [f] Taken from ref 22. [g] Taken from ref 23. [h] Taken from ref 24.

$$\sum_{B \neq A} Q_{AB} = -\frac{2}{3}(P_{Y_A Y_B} P_{Z_A Z_B} + P_{Z_A Z_B} P_{X_A X_B} + P_{X_A X_B} P_{Y_A Y_B}) - \frac{4}{3}(P_{X_A Y_B} P_{Y_A X_B}) \quad (11)$$

$$\delta(C^{13})_{calcd} = \sigma_p{}^A + \sigma_d{}^A \quad (12)$$

In evaluating the Q_{AB} terms it was found that Q_{AB} terms for nonneighbor atoms made a negligible contribution and could be safely ignored.

The actual calculations were carried out on a Burroughs 5500 computer using an ALGOL program similar to the FORTRAN program of Pople and Segal.[16] The only input data required were the atomic numbers and atomic coordinates which, wherever possible, were calculated from actual geometry. When this was unavailable, the necessary bond lengths and angles were estimated from related known structures.[17]

Dipole Moments

The dipole moment of a molecule is a very sensitive test of the reliability of molecular wave functions, and their calculation have been the subject of very many studies. Much of the earlier work has recently been reviewed[18] and will not be discussed in detail, except to point out that, in the great majority of previous studies on organic molecules, the atomic dipole contribution μ_{sp} defined by eq 5 was completely neglected, even though it has been known for many years[19] that such terms are extremely important in determining the dipole moments of small molecules such as water and ammonia. An important exception is the work of Hameka, *et al.*,[3] who included lone-pair moments explicitly, but the method was calibrated by allowing only the π-electron dipole moments to change from molecule to molecule. The π-electron densities were calculated by simple Hückel calculations by adjusting the heteroatom parameters for certain reference molecules until their experimental dipole moments could be reproduced exactly, and then using these parameters for other molecules. This method of adjusting parameters to fit the dipole moments of reference molecules appears to be a universal feature of the great majority of successful dipole moment calculations prior to the present study.

We have used the CNDO/2 method to calculate the dipole moments of six-membered heterocyclic rings containing pyridine nitrogen atoms, *i.e.*, the azines (Table I) and for five-membered heterocyclics containing both nitrogen and oxygen (Table II).

Azines. In Table II the data on six-membered heterocyclics, containing aza nitrogens (the azines) are summarized. It is seen that agreement with experimental values is generally good. A breakdown of the various contributions to the dipole moment shows that, in comparison with previous calculations, the CNDO/2 method predicts a much smaller π contribution (μ^{π}) to the total moment for the azines. Thus we predict for pyridine a μ^{π} of 0.33 D., whereas Brown,[20] using a VESCF π-electron method, predicted a μ^{π} of 1.21 D., and Hameka[3] predicted a value of 0.83 D. For quinoline and isoquinoline the CNDO/2 calculations predict values very close to those of pyridine. Whereas this is reasonable for quinoline, it is not so for isoquinoline, for which a vapor-phase value of 2.75 D. has been reported.[21] Isoquinoline, therefore, would seem to be one of those molecules for which the CNDO/2 method is inadequate. It should, however, be pointed out that previous calculations, even when specially calculated to fit the dipole moments of azines, also failed to predict a high value for this molecule compared to quinoline and pyridine. Perhaps then, before taking this discrepancy too seriously, confirmation of the high experimental value should be obtained.

The dipole moments of the three polar aza heterocyclics, pyridine, pyridazine, and pyrimidine, may be calculated purely on the basis of vector addition of lone-pair moments, assuming the dipole moment of pyridine is entirely due to the replacement of the C–H bond by the pyridine nitrogen with complete neglect of polarization of the σ and π electrons. This series, therefore, is

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(16) G. A. Segal, Program No. 91 of the Quantum Chemistry Program Exchange, Indiana University, Bloomington, Ind.

(17) Copies of the atomic coordinates used and the eigenvectors and eigenvalues obtained from these calculations are available on request.

(18) R. D. Brown and B. A. W. Coller, *Theoret. Chim. Acta*, 7, 259 (1967).

(19) J. A. Pople, *Proc. Roy. Soc.* (London), **A202**, 323 (1950).

(20) R. D. Brown and M. L. Heffernan, *Australian J. Chem.*, **12**, 554 (1959).

(21) A. L. McClellan, "Tables of Experimental Dipole Moments," W. H. Freeman and Co., San Francisco, Calif., 1963.

not a good test for calculations of dipole moments, except that fairly good agreement with experiment means that, in our calculations, the magnitude for the lone-pair moment is reasonable and also that we are not either grossly under- or overestimating the charge redistribution in the rest of the molecule on introducing a pyridine-type nitrogen atom.

Derivatives of Pyrrole and Furan. In contrast to the six-membered ring aza heterocyclics, the dipole moments (Table II) of five-membered ring compounds containing nitrogen and oxygen heteroatoms are very good examples for testing MO theories of charge densities, since no one factor dominates the dipole moments in all cases and the simple vector addition model soon breaks down. An additional feature which makes calculations on small heterocyclic molecules worthwhile is the availability of an increasing number of accurately known dipole moment measurements using microwave techniques on the vapor.[21–24] The theoretical calculations would otherwise be of marginal value, because the interpretation of solution dielectric constant data is notoriously difficult for heterocyclic molecules, as the result of complexities due to molecular association and tautomerism.

In Table II we summarize the calculations on pyrrole and furan and on some derivatives containing one and two pyridine-type nitrogen atoms. Apart from the case of pyrazole, agreement between experiment and theory is astonishingly good, especially for those molecules for which accurate dipole moment measurements are available.[21–24]

For furan and pyrrole, there are large π-electron contributions to the total dipole moment. In the former case this is offset by both a large σ-bond moment and a lone-pair moment so that the final dipole moment direction is with the negative end on the oxygen lone pair. For pyrrole, however, the absence of a lone-pair moment and the smaller σ-bond moment result in a dipole with its positive end on the nitrogen.

Comparison between the corresponding pyrrole and furan derivatives shows up very clearly the different factors governing the dipole moments in each type of derivative. The introduction of the electronegative pyridine nitrogen into the ring enhances the migration of π electrons from the pyrrole nitrogen and the furan oxygen, giving rise to an increase in the π-electron dipole moment. In imidazole, this moment is reinforced by the pyridine nitrogen lone-pair moment, resulting in a fairly large dipole moment of 4.09 D. In oxazole the situation is quite complex. While in furan the π moment, μ^π is somewhat smaller than μ^σ and is also in the opposite direction, in oxazole μ^π is larger than μ^σ because of the extra π-electron negative charge on the 3 position caused by electronegative pyridine nitrogen. However, the pyridine nitrogen lone-pair moment opposes the oxygen lone-pair moment, so that the resultant dipole moment is somewhat larger than in furan and in quite a different direction. The magnitude of the predicted moment is in good agreement with experiment. Unfortunately, the experimental direction of the moment is not known.

A similar complex interaction of the different contributions to the dipole moment of isoxazole results in a high moment of 3.17 D. in agreement with the experimental value of 3.01 D. For the corresponding pyrrole derivative, pyrazole, however, the agreement between experiment and theory is not particularly good. A very recent microwave determination[22] for this molecule gives a dipole moment of 2.2 D. compared to our predicted value of 2.71 D. Prior to this new result from vapor-phase measurements, our predicted value was in good agreement with the solution value of 2.52 D.[21] However, pyrazole is very subject to molecular association, and solution data are consequently very sensitive to concentration and solvent effects, and most unreliable compared to microwave vapor-phase measurements. A further experimental study which would determine the components of the dipole moment about the principal axes of inertia would help to clarify the discrepancy between theory and experiment.

For molecules containing three heteroatoms the interpretation of the experimental results for pyrrole derivatives is complicated by the possible existence of tautomers. For furan derivatives, however, this is not so, and we are fortunate in having experimental microwave dipole moments for 1,2,5-oxadiazole (furazan) and 1,3,4-oxadiazole.[23,24] The agreement with experiment is excellent in both cases (Table II).

The 1,2,3- and 1,2,5-triazoles can exist in tautomeric equilibrium with each other.

Dipole moment measurements of the corresponding phenyl and methyl derivatives have been made,[21] and these are in fair agreement with our calculations taking into consideration the difficulties associated with interpreting solution dipole moment measurements for this type of compound. The report of a moment of 1.77 D. for 1,2,3-triazole in benzene solution[21] could be interpreted, by comparison with our calculated values, as indicating the presence of both tautomers, but, in view of the discrepancy found for pyrazole between experiment and theory, such a conclusion is somewhat specious.

A similar situation hold for the tetrazoles. Here our dipole moment prediction for 1,2,3,4-tetrazole is very close to that of the reported value, supporting the general conclusion that the 1,2,3,4-tetrazole is the predominant isomer in solution. However, the calculated total energy of 1,2,3,4-tetrazole was found to be 0.3 ev higher than the 1,2,3,5 isomer. This indicates either that the theoretical method is unreliable for predicting conformers energetically or that the 1,2,3,4-tetrazole is stabilized preferentially by the solvent, a conclusion which would be reasonable in view of its much higher dipole moment.

A previous SCF π-electron calculation[25] which neglects σ-bond polarization has been used to predict dipole moments for the two isomers. The value given for the total moment of 1,2,3,4-tetrazole of 5.22 D. is very similar to ours, but the breakdown into contributing components is different, having a π moment of only 1.17 D. compared to our value of 3.19 D. The μ^π value

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(23) E. Saegebarth and A. P. Cox, *J. Chem. Phys.*, **43**, 166 (1965).

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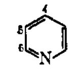
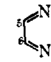
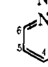
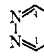
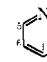
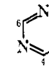
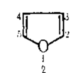
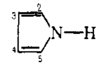
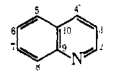
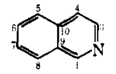
(25) J. B. Lounsbury, *J. Phys. Chem.*, **67**, 721 (1963).

Table III. C^{13} Chemical Shifts for Heterocyclic Compounds with One Ring[a]

Molecule	Atom no.	Δq^{π}	Δq^{total}	Q_{AB}[b]	σ_p, ppm	σ_d, ppm	$\sigma_p + \sigma_d$,[c] ppm	$\delta(C^{13})_{\text{exptl}}$[a] ppm
	2	+0.0346	+0.100	2.548	−272.55	57.01	−215.54	−21.85
	3	−0.0280	−0.029	2.514	−258.04	58.09	−199.953	+4.29
	4	+0.0489	+0.045	2.528	−267.52	57.60	−209.91	−7.63
	2	+0.0042	+0.066	2.540	+268.62	57.30	−211.32	−17.4
	3	+0.0116	+0.066	2.549	−269.57	57.30	−212.27	−23.9
	4	+0.0221	+0.002	2.525	−261.87	57.83	−204.04	+1.08
	1	+0.0276	+0.131	2.539	−273.71	55.94	−217.77	−33.4
	3	+0.0702	+0.20	2.563	−281.84	56.14	−225.70	−30.48
	1	+0.0773	+0.130	2.555	−275.55	56.75	−218.80	−28.87
	6	−0.0506	−0.062	2.508	−254.59	58.35	−196.24	+6.1
	2	+0.1184	+0.223	2.579	−285.62	55.93	+229.69	−39.0
	2	−0.0568	+0.107	2.479	−265.87	56.95	−208.92	−14.4
	3	−0.0671	−0.046	2.471	−252.24	58.24	−194.01	+1.6
	2	−0.0914	+0.042	2.357	−263.15	57.49	−205.66	+9.4
	3	−0.0941	−0.052	2.527	−254.75	58.27	−196.50	+19.6

[a] Relative to benzene. Taken from ref 28 and 8. [b] Calculated using eq 10 and 11. [c] Calculated using eq 8, 9, and 12.

Table IV. C^{13} Chemical Shifts for Quinoline and Isoquinoline

Molecule	Atom no.	Δq^{π}	Δq^{total}	Q_{AB}[a]	$\sigma_p^A + \sigma_d^A$ [b]	$\delta(C^{13})_{\text{exptl}}$, ppm[c]
	2	+0.0475	+0.102	2.55	−215.66	−23.0
	3	−0.0267	−0.033	2.51	−199.42	+8.0
	4	+0.0462	+0.034	2.52	−206.27	−7.5
	5	+0.0100	+0.005	2.52	−203.35	+2.0
	6	−0.0091	−0.004	2.45	−196.00	+2.5
	7	+0.0136	+0.013	2.42	−193.80	−1.0
	8	−0.0221	−0.032	2.54	−202.28	−1.5
	9	+0.033	+0.118	2.71	−234.72	−21.5
	10	−0.003	+0.010	2.69	−221.88	+1.0
	1	+0.0397	+0.096	2.55	−215.12	−24.5
	3	+0.0258	+0.091	2.54	−214.34	−15.0
	4	−0.0394	−0.045	2.52	−198.55	−7.0
	5	−0.0220	−0.017	2.48	−197.88	+2.2
	6	+0.0199	+0.014	2.52	−204.48	−2.0
	7	−0.0111	−0.006	2.52	−203.05	+1.2
	8	+0.0096	+0.008	2.34	−185.55	+1.0
	9	−0.0058	+0.007	2.69	−221.58	−0.03
	10	+0.0474	+0.066	2.68	−226.56	−7.4

[a] Calculated using eq 10 and 11. [b] Calculated using eq 8, 9, and 12. [c] Taken from the graph in ref 8.

for the 1,2,3,5 isomer in our calculation is only slightly smaller than the other isomer, but in the π-electron-only calculation, it was predicted to be only 0.49 D. This is an example of a case, unfortunately, very common in dipole moment calculations, where quite different calculations can give very similar total dipole moments although the individual components of the moment are quite dissimilar. It is for this reason that we stress the importance of trying to use the same calculation to predict more than one type of property.

C^{13} Chemical Shifts. Three different approaches have been suggested for the relationship between electronic distributions and C^{13} chemical shifts. It has been suggested that the chemical shift is related (1) to the π-electron density,[26,27] (2) to the total electron density,[27,28] and (3) to a more complex expression involving the electron density and the bond-order terms (Q_{AB} terms) as described by eq 7–12.[15] A recent very systematic experimental study of C^{13} chemical shifts of azines by Lauterbur[28] and by Grant[29] has made it possible to investigate the relative success of these fairly thoroughly.

The three relationships are expressed graphically in Figures 1–3, using data summarized in Tables III and IV. Figure 1 shows that the relationship between the π-electron densities and C^{13} chemical shifts is very poor, at least for the compounds of the present study. As might be expected if σ-electron densities are mainly short-range effects, there does appear to be a relationship between the Δq^{π} values for carbon atoms not flanked by nitrogen atoms (open circles) while the other carbon atoms are scattered in an irregular fashion below the line which could be drawn through the open circles. When we consider the total electron density as in Figure 2, however, the relationship is much better. The positions adjacent to the nitrogens in quinoline and isoquinoline are the ones with the largest observed shifts

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(26) H. Spiespecke and W. G. Schneider, *Tetrahedron Letters*, **14**, 468 (1961).

(27) J. B. Stothers, *Quart. Rev.* (London), **19**, 144 (1965).

(28) P. C. Lauterbur, *J. Chem. Phys.*, **43**, 360 (1965).

(29) D. M. Grant and R. J. Pugmire, unpublished work, reported in ref 8 and 13.

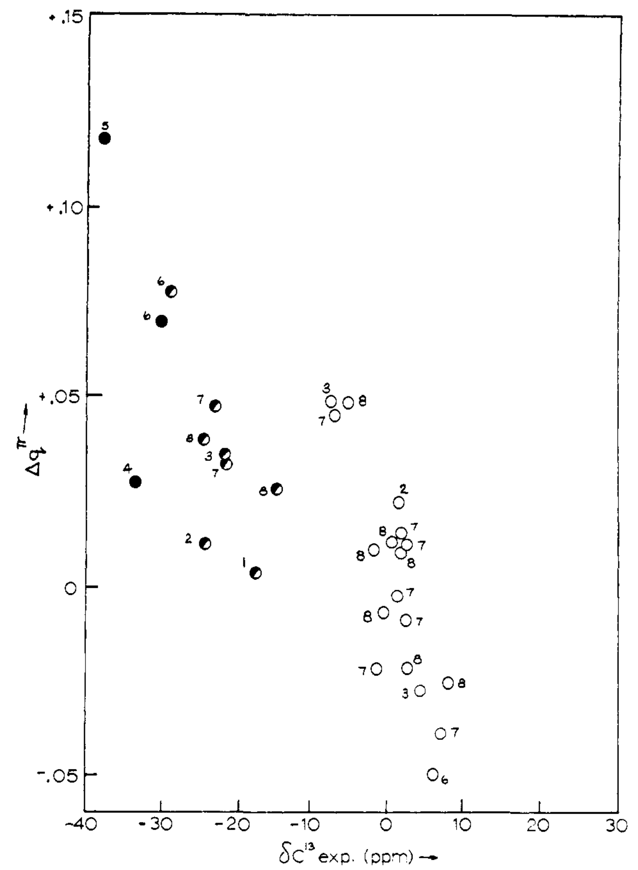

Figure 1. Δq^{π} values and C^{13} chemical shifts ($\delta(C^{13})_{exptl}$): (1) pyrazine, (2) pyridazine, (3) pyridine, (4) tetrazine, (5) triazine, (6) pyrimidine, (7) quinoline, (8) isoquinoline. ○, carbon atoms flanked by two nitrogen atoms; ◑, C atoms adjacent to one N atom; ●, C atoms with no adjacent heteroatoms.

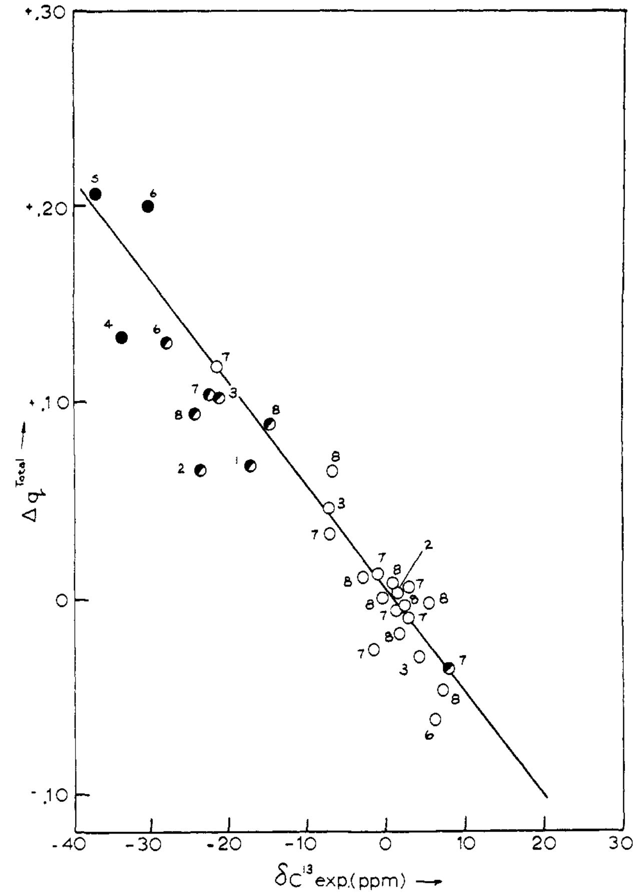

Figure 2. Δq^{total} values and C^{13} chemical shifts ($\delta(C^{13})_{exptl}$); numbering and circle designation as in Figure 1.

(Table IV). Whereas these points would be off the regression line for open circles in the relationship between Δq^{π} and the chemical shift (Figure 1), when the total electron density is used (Figure 2), these points fall into line with the data for the other positions.

When, however, we calculate the chemical shifts by the Karplus–Pople theory (eq 12) using a constant ΔE value of 10 ev (Figure 3), we find that there is quite definitely no improvement over the simple relationship of Figure 2. In fact, the inclusion of the Q terms causes considerably greater deviations for quinoline and isoquinoline than if we had ignored them. The central carbon atoms of these two molecules have considerably higher Q values than the other atoms, resulting in a predicted chemical shift considerably greater in magnitude than that actually found.

In addition to the azines, C^{13}-shift data are available for pyrrole and furan,[30] and we have included these in Table III. Again the major changes relative to benzene are in the electron densities. The chemical shift for the 3 positions of pyrrole and furan is predicted correctly to be more shielded than benzene, but the magnitude of the observed shift is grossly underestimated. For the 2 position of pyrrole, even the sign of the shift is not correct. After finding the best straight line in Figure 3 for the azines, we then included these two compounds on the graph (points 7 and 8). It appears that the calculated furan shifts give a much closer fit than the pyrrole shifts. It is evident that more experimental data on five-membered heterocyclics are needed before further calculations may reveal the reason for these discrepancies.

Ionization Potentials. By neglecting any energy changes due to the reorganization of the remaining electrons on the ionization of one electron, it is possible to equate the negative of the energy of the highest occupied MO [E(HOMO)] to the ionization potential (Koopmans' theorem).[31,32] Usually a correction factor has then to be added,[32] which is constant within a series of similar molecules. The E(HOMO) values for some of our heterocyclic molecules are compared with experimental values in Table V.

Table V. Ionization Potentials

Compound	IP_{calcd}, ev	IP_{exptl}, ev[a]
Pyridine	-13.023 (σ)	9.28
Pyrazine	-12.495 (σ)	9.27
Pyridazine	-12.709 (σ)	8.91
Pyrimidine	-12.816 (σ)	9.47
Pyrrole	-11.859 (π)	8.22
Furan	-12.321 (π)	8.77

[a] Experimental values taken from Turner.[32]

The theory predicts the HOMO for pyridine to be σ in character, and not confined to the nitrogen lone pair,

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(30) T. F. Page, Jr., T. Alger, and D. M. Grant, *J. Am. Chem. Soc.*, 87, 5333 (1965).

(31) T. Koopmans, *Physica*, 1, 104 (1933).

(32) D. W. Turner, *Advan. Phys. Org. Chem.*, 4, 31 (1966).

but rather considerably delocalized over the other σ orbitals of the ring.

For azines containing two ring nitrogens, the HOMO is found to be σ in type in all cases, but the predicted order of the ionization potentials does not agree with the experimental order,[32] although the individual differences are small. A still different order for these ionization potentials is found using an EHT method.[7]

In the cases of pyrrole and furan, the method predicts a difference between the ionization potentials of 0.46 ev in good agreement with the observed value of 0.50 ev. One interesting feature of the calculations on these two molecules is that the HOMO is of π character and antisymmetric with respect to the plane of symmetry through the heteroatom perpendicular to the molecular plane. In contrast, MO calculations using only π electrons[32] have found the HOMO's of these two molecules to be symmetrical π-type MO's, possessing higher energies because of the mixing in of heteroatom $2p\pi$ wave functions.

Conclusions

1. The CNDO/2 method is shown to be excellent for calculating dipole moment of heterocyclic molecules. Analysis of the results shows the presence of considerable σ-core polarization and confirms the importance of lone-pair and hybridization moments, factors which have often been neglected in calculations of dipole moments.

2. The calculated total electron densities are linearly related to C^{13} chemical shifts. The results of our calculations on heterocyclic compounds and of similar calculations on monosubstituted benzenes[33] show that suggestions that C^{13} chemical shifts are related to π-electron densities[26,27] are not very satisfactory compared to the use of the total electron density.[27,28] On the other hand, the inclusion of bond-order terms (Q terms), as suggested by Karplus and Pople,[15] leads to a deterioration in the relationship between theory and experiment.

3. The relationship between the energy of the highest occupied MO and observed ionization potentials is not very satisfactory and at best is only qualitative in nature. The HOMO of pyridine is predicted to be σ in character and delocalized over the whole molecule rather than an orbital involving only the nitrogen atomic orbitals.

In other work[33] we have found that the CNDO/2 method in general is much more successful in calculating properties depending on the wave functions (eigenvectors) than properties dependent on energies (eigenvalues). For example, the dipole moment of nitroethylene is predicted correctly, but the predicted barrier to internal rotation is calculated to be much smaller (by a factor of 10) than the recently measured value.[34] Attempts to correct this deficiency are in progress.

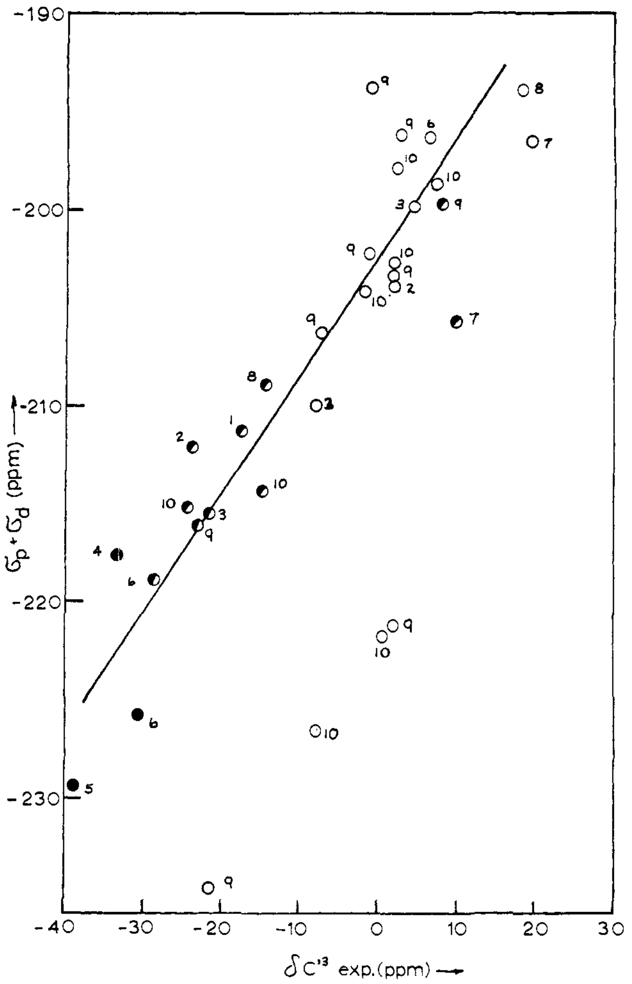

Figure 3. Calculated (eq 12) and experimental C^{13} chemical shifts: (1) pyrazine, (2) pyridazine, (3) pyridine, (4) tetrazine, (5) triazine, (6) pyrimidine, (7) pyrrole, (8) furan, (9) quinoline, (10) isoquinoline. Circles as in Figures 1 and 2.

Acknowledgments. This research was partially supported by the National Institutes of Health (Grant MH 12951). We are also grateful for a grant of computer time from the University of Virginia Computer-Science Center (supported partly by the National Science Foundation).

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(33) J. E. Bloor, D. L. Breen, and F. P. Billingsley, unpublished work.

(34) K. R. Loos and H. S. H. Günthard, *J. Chem. Phys.*, **46**, 1200 (1967).