

## Charge calculations in molecular mechanics. IX.\* A general parameterisation of the scheme for saturated halogen, oxygen and nitrogen compounds

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### SUMMARY

The CHARGE2 program for the calculation of partial atomic charges has been amended to include bond parameters for a number of organic functional groups, including halogens, nitrogen and oxygen. These minor amendments to the original scheme produce dipole moments for the fluoro and chloro compounds which are in complete agreement with the observed values.

The less complete data sets for the bromo and iodo compounds are also well reproduced, and the dipole moments of a variety of mixed halo compounds are now in better agreement with experiment than previously.

The calculated dipole moments of the saturated nitrogen and oxygen compounds are now in much better agreement than in the original scheme, thus the revised parameterisation may be employed with confidence to predict the electrostatic energies of these compounds.

Furthermore, the revised scheme now gives a precise proportionality between the charge on the proton in a CH group and the  $^1\text{H}$  chemical shift of the corresponding proton, allowing the general prediction, in principle, of  $^1\text{H}$  chemical shifts. In addition, attempts to include variable electronegativity in the  $\alpha$  effect are described for fluoro compounds.

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### INTRODUCTION

In previous parts of this series [1–8], a computationally fast and conceptually simple approach to the problem of estimating partial atomic charges in molecules has been described, based on the inductive effect for sigma bonds and a Hückel treatment of  $\pi$  systems, with recourse at each stage to the experimental dipole moments of the molecules under consideration.

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The utility of this approach to the calculation of the electrostatic interactions both within and between molecules in a molecular mechanics format has been amply demonstrated by the adoption of this charge calculation in a number of available molecular modelling packages [9].

A complementary approach to this problem has been described by Gasteiger et al. [10,11], using the concept of variable electronegativity [12] and electronegativity equalisation on bonding. In this scheme a partial equalisation of the atomic electronegativity is allowed on bonding by the introduction of a damping factor to obtain convergence. This scheme was originally parameterised on molecular ionisation energies [10] but more recently they have adopted our practice and used molecular dipole moments [11].

Our attempts to include variable electronegativity within our charge scheme were not successful, though we did show that convergence of the charges could be obtained without the damping factor used by Gasteiger et al. [10,11]. It is probable that the two schemes are complementary in that the addition of separate  $\beta$  (over two bonds) and  $\gamma$  (over three bonds) effects in our charge scheme are practically equivalent to varying the electronegativity of the chain atoms. A more cogent defect of the variable electronegativity scheme is the questionable method of dealing with  $\pi$  systems, but a comprehensive treatment of  $\pi$  systems within the present charge scheme has already been given [5,7,8].

In the original paper in this series [2], which was concerned solely with haloalkanes, drastic simplifying assumptions concerning the various integrals in the parameterisation scheme were made. Despite this the scheme well reproduced the dipole moments of a wide variety of haloalkanes. However, when this was extended to saturated nitrogen, oxygen and sulphur compounds [3], not surprisingly the calculated dipole moments were in much poorer agreement with the observed values.

Furthermore, although the calculated charges on the hydrogen atoms of  $\text{CH}_3\text{X}$  compounds were found to give an excellent correlation with the corresponding  $^1\text{H}$  chemical shifts for  $\text{X} = \text{H}$ ,  $\text{CH}_3$ ,  $\text{Si}$ ,  $\text{N}$ ,  $\text{O}$  and  $\text{F}$  [1], the correlation failed when applied to  $\text{Cl}$ ,  $\text{Br}$  and  $\text{I}$ .

Here we address both these problems. We show that a simple relaxation of the too drastic previous assumption regarding integrals between bonds, so that now any  $\text{X}-\text{Y}$  bond has its own integral, produces dipole moments for all the saturated molecules considered in essentially complete agreement with the observed values. Also a reinvestigation of the electronegativities of the anomalous atoms noted above removes this anomaly and produces a good correlation of charge with proton chemical shift for all the molecules considered. These amendments are all within the basic framework given previously.

## THEORY

As the theory has been presented elsewhere [2], only a brief summary is given here. If we consider an atom  $\text{I}$  in a four-atom fragment  $\text{I}-\text{J}-\text{K}-\text{L}$ , the partial atomic charge on  $\text{I}$  is due to three effects: an  $\alpha$  effect from atom  $\text{J}$ ; a  $\beta$  effect from atom  $\text{K}$ , and a  $\gamma$  effect from  $\text{L}$ . The charge ( $q_i$ ) on atom  $\text{I}$  resulting from atom  $\text{J}$  (the  $\alpha$  effect) is given by Eq. 1:

$$q_i(\alpha) = (E_j - E_i)/a_{ij} \quad (1)$$

where  $E_i$  and  $E_j$  are the electronegativities of atoms  $\text{I}$  and  $\text{J}$  respectively and  $a_{ij}$  is a constant de-

pending on the exchange and overlap integrals for the bond I–J. Initially [2] only three values were used for  $a_{ij}$ : one if the bond was to hydrogen; one for bonds between first-row atoms, and one for bonds involving a second-row (or higher) atom. This method is now extended so that there exists a set of parameters for all the bonding pairs being considered. These are described as  $A(X,Y)$  parameters where  $X$  and  $Y$  are the atoms involved in the bond.

The  $\beta$  effect, from atoms two bonds removed (I–A–J), is considered to be proportional to both the electronegativity of atom  $J$  and the polarizability of atom  $I$ . Taking hydrogen as the base substituent, the  $\beta$  effect was defined as

$$q_i(\beta) = (E_J - E_H)P_i/b \quad (2)$$

In order to account for the variation of polarizability with charge, the  $\beta$  effect calculation is carried out iteratively according to Eq. 3:

$$P_i = P_i^0(1.0 + 3.0 \times (q_i^0 - q_i)) \quad (3)$$

where  $P_i$  is the polarizability of atom  $I$  with charge  $q_i$ ,  $P_i^0$  and  $q_i^0$  are the corresponding initial values.

The  $\gamma$  effect is considered to be directly proportional to the  $\beta$  effect and is given by Eq. 4:

$$q_i(\gamma) = B_{ij}/5.0 \quad (4)$$

The total charge is given by Eq. 5:

$$q_i = q_i(\alpha) + q_i(\beta) + q_i(\gamma) \quad (5)$$

In order that an element may be included in the scheme, it is necessary to obtain values for the electronegativity and polarizability of that element, in whichever hybridization state is appropriate.

In previous work, the electronegativities used were taken from the values given by Gasteiger and Marsili [10], which were based on the orbital electronegativities compiled by Hinze and Jaffe [13] on the Mulliken scale.

One problem with these electronegativities is that although the individual  $s$  and  $p$  hybridised orbital electronegativities can be derived, the hybridisation (e.g. of Cl in a C–Cl bond) is open to question and as a result so is the electronegativity.

It seemed only logical, in view of the good correlation of charge vs. chemical shift for the first-row atoms considered, to obtain the electronegativities of Cl, Br and I directly from the appropriate proton chemical shifts of the MeX compounds. Indeed this is merely adopting the well-known electronegativity scale of Allred and Rochow [14], which is based on the MeX chemical shifts, and relating it to the Mulliken scale. The electronegativities so derived are given in Table 8.

Also in the previous scheme, perhaps partly to compensate for the smaller values of the electronegativities used previously, the  $\beta$  and  $\gamma$  effects of these atoms were enhanced by a common factor. The necessity for this enhanced  $\beta$  effect can now be tested against the available data set. This was done and we now find that this enhanced  $\beta$  effect is not required for the  $\beta$  effect on hydrogen, thus

retaining the charge vs. chemical shift correlation, but does produce more accurate dipole moments when implemented for all other  $\beta$  atoms. This minor coding change has been implemented for all the results to be discussed. A small modification has also been made to the  $\gamma$  effect as it was noted from the  $\delta$  values for the n-BuX compounds that the  $\beta(\text{S.C.S.})/\gamma(\text{S.C.S.})$  for C, N, O and F is about 10/1, while for Cl, Br and I it is nearer to 5/1. These two values replace the original one in Eq. 4.

The parameterisation method also follows the previous pattern, in attempting to produce parameters which provide the best fit between experimental and calculated dipole moments. As far as possible the dipole moments and accurate gas-phase molecular geometries are taken from microwave studies. However, other techniques have been considered when this information was unavailable.

The parameter fitting was carried out via a nonlinear least-squares fit using the partial derivatives of the dipole moments, with respect to the parameters. The resulting values of the parameters are then tested in the CHARGE3 program.

## RESULTS AND DISCUSSION

Each group of compounds will be discussed separately. This reflects the order in which the parameterisation was carried out. Where possible, the results will be compared to those from the most recent parameterisation of the Gasteiger variable electronegativity scheme [11], and reference will also be made to our own attempts to include variable electronegativity in the  $\alpha$  effect.

### *Fluoro and chloro compounds*

At an early stage in the development of these calculations [2], attempts were made to include some kind of charge dependence on the electronegativity. This follows the earlier ideas of Iczkowski and Margrave [15] who showed that the value of an atom's electronegativity will vary according to the degree of polarity of its surroundings. This idea was taken up and applied by Huheey, who first defined the concept in his charge scheme [16]. The early attempts in this laboratory proved to give poorer results than the nonvariable scheme.

Also, a second investigation was made of the possibility of including variable electronegativity. In Table 1 a comparison is presented of the new parameterisation with dipole moments obtained using variable electronegativity. Our original intention was to use a quadratic expression to describe the variation of electronegativity with charge, but after consideration of the ideas of Iczkowski and Margrave [15] it appeared that a cubic expression might provide a better description. Consider an electronegative atom like fluorine: as more negative charge gets transferred to the fluorine, the electronegativity should decrease; when this variation takes place using a quadratic function then the fluorine electronegativity decreases in a constant way. If one considers the situation where fluorine has a charge of 0.9 electrons, it seems more reasonable that the fluorine would still be in a position to readily attract the final 0.1 electrons, as it can easily support an integral negative charge. For this reason a cubic curve seemed more reasonable. There is also the added advantage that an additional barrier to charge loss is also present for electronegative atoms.

A parameterisation was attempted for the fluoro compounds using both variable electronegativity functions, however, in neither case was it possible to obtain as close agreement with experiment as was possible using the original method [2]. Two interesting features did arise from this

TABLE I  
OBSERVED AND CALCULATED DIPOLE MOMENTS FOR FLUOROALKANES

Compound	Obs <sup>a</sup>	Calc	Calc <sup>b</sup>	Calc <sup>c</sup>	Gast <sup>d</sup>
HF	1.8	1.81	—	—	—
CH <sub>3</sub> F	1.855	1.85	1.78	1.88	1.85
CH <sub>2</sub> F <sub>2</sub>	1.96	1.96	1.89	2.03	1.86
CHF <sub>3</sub>	1.645	1.64	1.57	1.62	1.50
CH <sub>3</sub> CH <sub>2</sub> F	1.96	1.95	1.90	1.90	2.02
CH <sub>3</sub> CHF <sub>2</sub>	2.30	2.26	2.27	2.24	2.44
CH <sub>3</sub> CF <sub>3</sub>	2.32	2.26	2.40	2.38	2.61
CF <sub>3</sub> CH <sub>2</sub> F	—	2.14	—	—	—
CF <sub>3</sub> CHF <sub>2</sub>	1.54	1.68	—	—	1.64
CH <sub>2</sub> FCH <sub>2</sub> F (g)	2.67	2.81	2.73	2.91	—
CHF <sub>2</sub> CHF <sub>2</sub> (g)	2.9	2.78	—	—	—
CHF <sub>2</sub> CH <sub>2</sub> F (g)	—	1.80	—	—	—
CHF <sub>2</sub> CH <sub>2</sub> F (t)	—	3.41	—	—	—
nPrF (g)	1.90	1.93	1.89	1.90	—
nPrF (t)	2.05	1.97	1.97	1.90	—
iPrF	1.96,2.15	2.03	—	—	2.14
tBuF	1.96,2.15	2.10	—	—	2.22
cHexylF (eq)	2.11	2.09	2.13	1.95	—
cHexylF (ax)	1.81	2.03	2.01	1.95	—

<sup>a</sup> Dipole moments from Ref. 19 and geometries from Ref. 20.

<sup>b</sup> Electronegativity variation using quadratic function.

<sup>c</sup> Electronegativity variation using cubic function.

<sup>d</sup> Results from Gasteiger scheme [11].

work: when the charges for a longer chain alkyl fluoride are calculated it is found that the charges on the carbon atoms show a very good reflection of the inductive effect due to fluorine, i.e. there is a gradual increase in the charge as one moves down the chain from the halogen. This contradicts quite categorically the idea of charge alternation which resulted from the Mulliken population analysis of CNDO wavefunctions [17], and is in agreement with our previous calculations. More importantly, it was found that the variable electronegativity method does converge before electronegativity equalisation takes place. This shows that the damping factor ( $1/2^n$ ), where  $n$  is the number of the iteration, is not in fact necessary and may be a constraint which modifies the final answer. In the latest parameterisation [11] this factor has been increased to 0.6.

The actual charges obtained by variable electronegativity methods showed that very little charge transfer takes place beyond the one-bond effect, as opposed to the gradual redistribution of charge which should take place in consecutive iterations. Taking ethyl fluoride as an example, it can be seen that the calculated dipole moment is due almost entirely to the large C–F bond dipole, the rest of the molecule being effectively uncharged. The charge on the  $\alpha$ -H's is almost identical to that of the  $\beta$ -H's, indicating that there is very little interaction between atoms separated by more than one bond. The parameters used to obtain the results in Table I were markedly different to those used in the previous scheme and this is reflected in the above observations. In order that the dipole moments were reproduced it was necessary to decrease the size of ACF, i.e. increase the

polarity of the C–F bond. This compensates for the lack of charge transmission over more than one bond but results in a very strongly polar C–F bond. When the cubic term was applied, this effect was further enlarged, and in the case of cyclohexyl fluoride the dipole moment is independent of conformation. Although Table 1 shows reasonable agreement between the observed dipole moments and those calculated using variable electronegativity it was decided to discontinue this approach as the calculated charges appear to be very insensitive to chemical environment.

In the earlier papers [2,3] describing the theory behind the charge scheme only three AXY parameters were used in the parameterisation. Although these parameters produced very good agreement with experiment for haloalkanes, a number of discrepancies appeared when the scheme was applied to nitrogen and oxygen compounds.

As very little data exist on the dipole moments of alkanes it is not easy to find a suitable C–H parameter. During the fluoro parameterisation both ACH and ACF were varied and it was shown that the best value of ACH should be around 20–21. ACH was therefore set equal to 20.0 and the rest of the parameterisation proceeded from this point. In a recent paper Miller [18] has revised the earlier estimate of the fluorine polarisability from 0.697 to 0.863. We use this latter value henceforth.

The results in Table 1 show essentially complete agreement between calculated and experimental dipole moments. This agreement is even better than that of the original parameterisation [2]. The Gasteiger scheme [11], on the other hand, has been calibrated against fluoromethane and consequently a very close fit is obtained, but the dipole moments for difluoro and trifluoro methane are considerably lower than experiment, i.e. the C–F dipole is too large. A similar situation is also found for fluoroethanes.

Similar excellent agreement is also found for the chloro compounds (see Table 2). As expected, the more polarisable chlorine atoms do not retain their charge to such an extent as the fluorine atoms, and as the substitution pattern increases, the charge on the chlorine decreases.

A more general feature of the fluoro and chloro alkanes which is well reproduced is the effect of conformation upon the dipole moments of halo-cyclohexanes. In both cases the larger dipole moment of the equatorial mono-substituted form is found, and this may be considered as resulting solely from the conformation of the ring, with respect to the dominant C–X dipole; in each case the charges are identical as in the present scheme the charges calculated are invariant of conformation in wholly saturated systems.

#### *Bromo, iodo and mixed-halo compounds*

In the earlier parameterisation [2] of this charge method one 'A' parameter was chosen to represent bonding interactions with second-row or greater elements. Using the Gasteiger and Marsili [10] values of the bromine and iodine electronegativities,  $E_{\text{Br}}$  10.87,  $E_{\text{I}}$  9.99, it was found that the A parameter produced dipole moments which were too low. The electronegativity values obtained from the charge vs. chemical shift correlation are 12.7 and 11.2, on the Mulliken scale.

In this parameterisation these higher electronegativities were used and generally better agreement was obtained for the bromo compounds with the higher  $E_{\text{Br}}$ . A similar situation was found for iodine. The results in Tables 3 and 4 show good agreement between observed and calculated values; again better than that found by Gasteiger [11]. The Gasteiger scheme seems to produce C–X dipoles which are too large, in an attempt to fit the mono-halo cases and the resulting dipole moments for the multi-substituted cases are in error.

TABLE 2  
OBSERVED AND CALCULATED DIPOLE MOMENTS FOR CHLOROALKANES

Compound	Obs <sup>a</sup>	Calc	Gast <sup>c</sup>
HCl	1.08	1.09	—
CH <sub>3</sub> Cl	1.89	1.89	1.73
CH <sub>2</sub> Cl <sub>2</sub>	1.62	1.62	1.63
CHCl <sub>3</sub>	1.04	1.09	1.16
CH <sub>3</sub> CH <sub>2</sub> Cl	2.00	1.97	1.86
CH <sub>3</sub> CHCl <sub>2</sub>	1.92	2.02	2.06
CH <sub>3</sub> CCl <sub>3</sub>	1.76	1.93	2.01
CCl <sub>3</sub> CH <sub>2</sub> Cl	1.45	1.81	1.80
CCl <sub>3</sub> CHCl <sub>2</sub>	1.10	1.23	1.23
nPrCl (g)	1.96	1.95	—
nPrCl (l)		2.03	—
iPrCl	2.07	2.06	1.96
tBuCl	2.15	2.11	2.04
cHexylCl (eq)	2.20 <sup>b</sup>	2.08	—
cHexylCl (ax)	2.09 <sup>b</sup>	2.00	—
1,1 cHexylCl <sub>2</sub>	2.48	2.31	2.44
cis 1,2 cHexylCl <sub>2</sub>	3.12	3.22	3.14
cis 1,4 cHexylCl <sub>2</sub>	2.89	2.78	2.69
tr 1a,2a cHexylCl <sub>2</sub>	1.20	1.00	—
tr 1e,2e cHexylCl	3.30	3.42	—

<sup>a</sup> Dipole moments from Ref. 19 and geometries from Ref. 20.

<sup>b</sup> Dipole moment of corresponding 4-tert-butyl-1-chlorocyclohexane.

<sup>c</sup> Results from Gasteiger scheme [11].

The individual halo-alkane reparameterisation shows a significant improvement in the agreement between the observed and calculated dipole moments and a further significant improvement was found for the mixed-halo compounds. In Table 5 the results are presented together with those from the original parameterisation [2]. It is important to note that these compounds were not actually included in the parameterisation process, however, some of the systematic discrepancies present in the original scheme have been reduced. In almost every case the original parameterisation produced results which were larger than the experimental values, however, this has been altered by reducing the polarity of the C–X bonds.

#### Nitrogen compounds

Following the original results for haloalkanes [2], attempts were made to test the program for simple amine compounds, that is to say, compounds where resonance effects would not play a significant part. This work was done as a prelude to a modification of the scheme to handle resonance systems [3]. The original nitrogen parameterisation was based upon the simple scheme, i.e. one parameter to deal with all bonds between first-row elements. The electronegativity of nitrogen was taken to be that of the  $te^2tetete$  configuration [3,13] and reasonable agreement was shown between the observed and calculated dipole moments. There were, however, systematic errors when dealing with compounds with N–H bonds, i.e. the dipole moment of NH<sub>3</sub> was too low, likewise

TABLE 3  
OBSERVED AND CALCULATED DIPOLE MOMENTS FOR BROMOALKANES

Compound	Obs <sup>a</sup>	Calc	Gast <sup>d</sup>
HBr	0.82	0.82	—
CH <sub>3</sub> Br	1.80	1.86	1.71
CH <sub>2</sub> Br <sub>2</sub>	1.43	1.50	1.59
CHBr <sub>3</sub>	1.00	0.95	1.14
CH <sub>3</sub> CH <sub>2</sub> Br	2.03	1.92	1.84
CH <sub>3</sub> CHBr <sub>2</sub>	1.92	1.84	2.02
CH <sub>3</sub> CBr <sub>3</sub>	—	1.67	—
nPrBr (g)	2.03	1.89	—
nPrBr (t)	—	1.96	—
iPrBr	2.1	1.97	1.94
tBuBr	2.2	2.01	2.02
cHexylBr (eq)	2.25 <sup>b</sup>	2.02	—
cHexylBr (ax)	2.19 <sup>b</sup>	1.90	—
1,1 cHexylBr <sub>2</sub>	2.44	2.15	2.39
cis 1,2 cHexylBr <sub>2</sub>	3.12	3.04	3.08
cis 1,4 cHexylBr <sub>2</sub>	2.91	2.69	2.64
tr 1a,2a cHexylBr <sub>2</sub>	1.20 <sup>c</sup>	0.91	—
tr 1e,2e cHexylBr <sub>2</sub>	3.30 <sup>c</sup>	3.21	—

<sup>a</sup> Dipole moments from Ref. 19 and geometries from Ref. 20.

<sup>b</sup> Dipole moment of corresponding 4-tert-butyl-1-bromocyclohexane.

<sup>c</sup> Dipole moment of corresponding 4-tert-butyl-1,2-dibromocyclohexane.

<sup>d</sup> Results from Gasteiger scheme [11].

TABLE 4  
OBSERVED AND CALCULATED DIPOLE MOMENTS FOR IODOALKANES

Compound	Obs <sup>a</sup>	Calc	Gast <sup>b</sup>
HI	0.44	0.44	—
CH <sub>3</sub> I	1.62	1.71	1.52
CH <sub>2</sub> I <sub>2</sub>	1.10	1.29	1.37
CHI <sub>3</sub>	0.85	0.75	0.89
CH <sub>3</sub> CH <sub>2</sub> I	1.77	1.71	1.62
CH <sub>3</sub> CHI <sub>2</sub>	—	1.52	1.70
CH <sub>3</sub> CI <sub>3</sub>	—	1.29	—
nPrI (g)	1.85	1.68	—
nPrI (t)	—	1.74	1.70
iPrI	2.00	1.72	1.76
tBuI	2.13	1.73	—
cHexylI (eq)	2.00	1.74	—
cHexylI (ax)	—	1.65	—
cis 1,4 cHexylI <sub>2</sub>	2.45	2.36	—

<sup>a</sup> Dipole moments from Ref. 19 and geometries from Ref. 20.

<sup>b</sup> Results from Gasteiger scheme [11].



TABLE 5  
OBSERVED AND CALCULATED DIPOLE MOMENTS FOR MIXED-HALO COMPOUNDS

Compound	Obs <sup>a</sup>	Calc	Calc <sup>b</sup>
Chloro-fluoro			
CF <sub>3</sub> Cl	0.50	0.46	0.55
CF <sub>2</sub> Cl <sub>2</sub>	0.53	0.49	0.57
CFCl <sub>3</sub>	0.46	0.48	0.55
CHF <sub>2</sub> Cl	1.43	1.49	1.55
CHFC1 <sub>2</sub>	1.30	1.35	1.43
CH <sub>2</sub> FCl	1.82	1.82	1.85
CF <sub>3</sub> CH <sub>2</sub> Cl	—	2.09	—
CF <sub>3</sub> CF <sub>2</sub> Cl	0.52	0.43	0.51
CH <sub>3</sub> CF <sub>2</sub> Cl	2.14	2.16	2.24
CH <sub>2</sub> FCH <sub>2</sub> Cl (g) 60°	2.72	2.89	2.80
Fluoro-bromo			
CF <sub>3</sub> Br	0.65	0.75	0.79
CF <sub>2</sub> Br <sub>2</sub>	0.66	0.84	0.88
CFBr <sub>3</sub>	0.58	0.72	0.75
CHF <sub>2</sub> Br	1.31	1.50	1.55
CHBrBr <sub>2</sub>	—	1.28	—
Chloro-bromo			
CCl <sub>3</sub> Br	0.21	0.30	0.24
CCl <sub>2</sub> Br <sub>2</sub>	0.25	0.35	0.28
CClBr <sub>3</sub>	0.20	0.30	0.24
CHCl <sub>2</sub> Br	1.31	1.07	1.15
CH <sub>2</sub> ClBr	1.66	1.56	1.64
Fluoro-iodo			
CF <sub>3</sub> I	1.0	1.24	1.29

<sup>a</sup> Dipole moments from Ref. 19 and geometries from Ref. 20.

<sup>b</sup> Calculated dipole moments from Ref. 2.

that of CH<sub>3</sub>NH<sub>2</sub>, but trimethylamine was calculated in good agreement with experiment [3].

In the present reparameterisation a larger data set of saturated nitrogen compounds was used (Table 6), and the two integrals involving ANH and ACN, were optimised for the best agreement using this data set. Note that this produces a set of 13 equations in the two unknowns, so that the solutions are overdetermined.

The results of this optimisation are shown in Table 6, together with the dipole moments calculated on the original scheme, and the optimised values of ANH and ACN in Table 9.

The agreement with the observed dipole moments is now impressive and well worthy of comment. The dipole moments of ammonia and the simple methylamines are now well reproduced, in contrast to those of Ref. 2, and thus there is no need to invoke, at this stage, any lone-pair dipole for amines.

The dipole moments of the cyclic amines are also well reproduced, and of particular note is the agreement for the calculated and observed axial and equatorial N–H conformers of piperidine.

TABLE 6  
OBSERVED AND CALCULATED DIPOLE MOMENTS FOR NITROGEN COMPOUNDS

Compound	Obs <sup>a</sup>	Calc <sup>b</sup>	Calc <sup>c</sup>
NH <sub>3</sub>	1.468	1.50	1.16
CH <sub>3</sub> NH <sub>2</sub>	1.33	1.24	0.96
(CH <sub>3</sub> ) <sub>2</sub> NH	1.01	1.04	0.88
(CH <sub>3</sub> ) <sub>3</sub> N	0.63	0.85	0.85
<i>N</i> -Methyl aziridine <sup>d</sup>	1.236	1.27	1.27
<i>N</i> -Methyl piperidine	0.78	0.68	0.67
Piperidine <sup>e</sup> eq	0.82	0.84	0.68
ax	1.19	1.14	0.98
<i>N</i> -Methyl morpholine <sup>f</sup>	1.21	1.24	1.25
Morpholine <sup>g</sup> eq	1.71	1.70	1.41
ax		1.76	1.47
Ethylene imine <sup>h</sup>	1.89	1.41	1.26

<sup>a</sup> Dipole moments from Ref. 19 and geometries from Ref. 20, unless otherwise stated.

<sup>b</sup> This work.

<sup>c</sup> From Ref. 3.

<sup>d</sup> D.M. and geometry from Ref. 22.

<sup>e</sup> D.M. from Ref. 21.

<sup>f</sup> D.M. and geometry from Ref. 23.

<sup>g</sup> Geometry from Ref. 24.

<sup>h</sup> D.M. and geometry from Ref. 25.

Parkin et al. [21] noted that 'on simple bond moment arguments these values should be identical'. Here these differences come directly from the charge scheme, precisely analogous to the differing dipole moments of the axial and equatorial cyclohexyl halides.

In the analogous morpholine case, only the N–H equatorial conformer was observed in the gas phase [24]. Note also that the *N*-methyl derivatives of piperidine and morpholine exist entirely in the methyl equatorial conformation.

The only anomalous result in this table is that of ethyleneimine (but not, note, the *N*-methyl analogue). This is not altogether surprising, as the hybridisation of the nitrogen in this molecule, as a consequence of the ring strain, differs appreciably from that of an unstrained amine, and this would affect the electronegativity of the nitrogen atom. It would be possible to correct for this in the CHARGE scheme but we have not considered it necessary at the moment.

### Oxygen compounds

In the original scheme the oxygen compounds were treated in a similar way to the nitrogen compounds; a standard first-row parameter was used [3] and this was sufficient to give better agreement with experiment than is found when using the CNDO/2 method [17]. It was stated earlier that the electronegativity of an orbital shows a linear dependence upon the amount of *s*-character which is present [13], and this may play an important role in the case of alcohols and ethers, where there is considerable variation of the hybridisation state of the oxygen. It was decided, however, to use a single value for the oxygen electronegativity as an angle dependence of the electronegativity would add unnecessary complexity to the scheme.

For this reparameterisation the two independent integrals for the CO and OH bonds were optimised over a more extensive data set (Table 7). The parameterisation process took the same

form as that of the nitrogen compounds. First, ether-type compounds were subjected to a rough fitting to provide an approximate value for ACO. With this value it was then possible to proceed with a more refined fitting of both parameters over the complete data set. The results for the oxygen data set are given in Table 7.

The calculated results show a very good agreement with experiment, with two notable exceptions. The first concerns the cyclic ether, 1,3,5 trioxan, where the calculated dipole moment is higher than the observed one by ca. 0.2 D. It may be possible to attribute the difference to the anomeric effect [32] where donation of lone-pair electron density to the  $\sigma^*(\text{C-H})$  orbitals can take place. This would result in a decrease in the overall charge on the oxygen and would lead to a reduction in the size of the molecular dipole moment. The second major discrepancy concerns 7-oxabicyclo[2.2.1]heptane. Again the dipole moment is calculated to be higher than experiment. In this compound the COC angle is only  $95.3^\circ$  [30], and this increase in the  $p$ -character may lead to a greatly reduced value of the oxygen electronegativity, below that used here. This may result in a more polar C–O bond than is necessary and hence the calculated dipole moment will be too high. It should be pointed out that the dimethyl epoxide compound, which has an even smaller bond angle at the oxygen, is calculated to be in good agreement with experiment.

## CONCLUSIONS

The charge scheme described in Ref. 2 has been reparameterised to obtain a more general series of parameters to describe the bonding interactions in saturated organic molecules. There now

TABLE 7  
OBSERVED AND CALCULATED DIPOLE MOMENTS FOR OXYGEN COMPOUNDS

Compound	Obs <sup>a</sup>	Calc	
		This work	Ref. 3
H <sub>2</sub> O	1.846	1.87	1.68
CH <sub>3</sub> OH	1.662	1.62	1.59
CH <sub>3</sub> OCH <sub>3</sub>	1.30	1.39	1.54
<i>t</i> -BuOH	1.67	1.64	1.64
Tetrahydropyran <sup>b</sup>	1.60	1.61	1.77
1,3 Dioxan	2.14	2.04	2.27
1,3,5 Trioxan	2.08	2.26	2.50
<i>tr</i> -Ethanol <sup>c</sup>	1.44, 1.58	1.48	
2-Methoxy ethanol <sup>d</sup>	2.36	2.28	
2,3 Epoxybutane <sup>e</sup>	2.03	1.91	
7-Oxabicycloheptane <sup>f</sup>	1.621	1.99	
Trioxaadamantane <sup>g</sup>	3.01	3.19	
Ethylene oxide	1.88	1.81	

<sup>a</sup> Dipole moments from Ref. 19 and geometries from Ref. 20, unless otherwise stated.

<sup>b</sup> Geometry from Ref. 26.

<sup>c</sup> D.M. from Ref. 27.

<sup>d</sup> D.M. and geometry from Ref. 28.

<sup>e</sup> D.M. and geometry from Ref. 29.

<sup>f</sup> D.M. and geometry from Ref. 30.

<sup>g</sup> D.M. and geometry from Ref. 31.

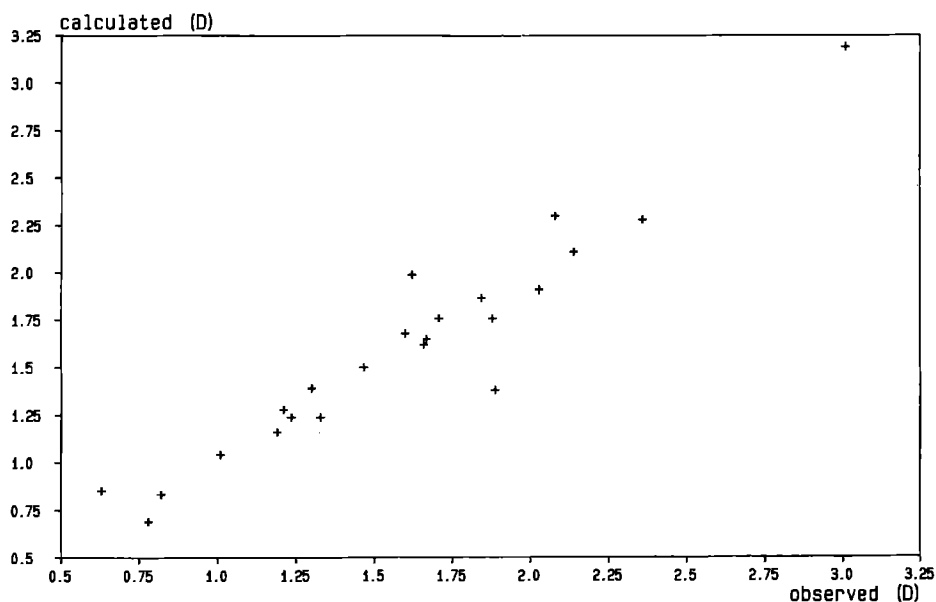


Fig. 1. Observed vs. calculated dipole moments for nitrogen and oxygen compounds.

exists a series of parameters for individual bonding pairs, each of which has been obtained by fitting calculated dipole moments to the experimental values.

The new parameterisation has been shown to give better values than those of the original scheme for the individual haloalkanes. The mixed-halo compounds also show better agreement than had been previously found. There are some problems associated with the bromo and iodo compounds, as it was not possible to obtain a very close fit to the experimental values. This could, however, be due to the quality of the experimental data which cannot be considered as uniformly reliable as that for the other haloalkanes. A case in point is 1,1-diiodo ethane, for which the calculated dipole moment is almost 1 D too small. There seems to be very little reason for this discrepancy.

TABLE 8  
ELECTRONEGATIVITIES AND POLARISABILITIES

	E	P
H	7.17	0.628
C	7.98	1.056
N	11.54	1.085
O	14.20	0.912
F	16.96	0.863
Cl	13.6	1.518
Br	12.7	1.885
I	11.2	2.423

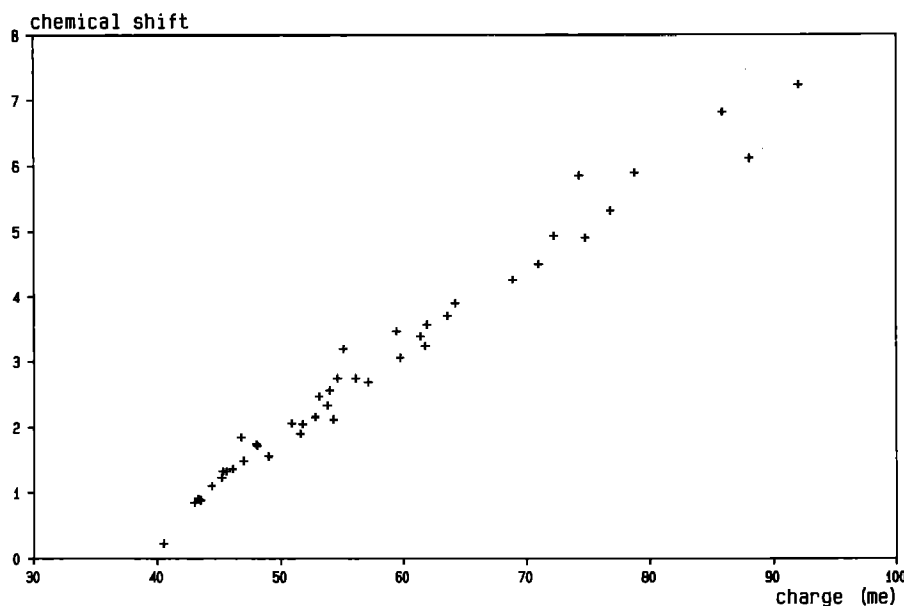


Fig. 2. Correlation of  $^1\text{H}$  chemical shifts (ppm) with partial charges (me). A least-squares fit gives the equation,  $\delta = 1.317q_{\text{H}} - 4.729$ , correlation coefficient = 0.994.

In contrast, the new parameterisation for the oxygen and nitrogen compounds shows a much improved reproduction of the dipole moment values. These are plotted in Fig. 1. It is interesting to note that it is possible to obtain a good fit for the series ammonia through to trimethylamine, which suggests that lone-pair dipoles or hyperconjugative interactions between the lone pair and the methyl groups do not need to be invoked at this stage. In Fig. 2 the  $^1\text{H}$  chemical shifts for a number of the compounds are plotted against their partial atomic charges. This near-linear plot indicates that the charges could be of considerable value in  $^1\text{H}$  shift predictions using the equation given in the figure legend.

In comparison to other methods the model performs particularly well; the reproduction of dipole moments is very much better than the CNDO/2 program [10], which has also received some

TABLE 9  
BOND PARAMETERS

	New AXY	Previous AXY		New AXY	Previous AXY
ACH	20.0	23.80	ACBr	21.6	21.00
ACN	43.2	40.40	ACI	21.6	21.00
ACO	48.2	40.40	ANH	16.0	23.80
ACF	40.0	40.40	AOH	21.4	23.80
ACS	21.6	21.00	AHF	23.8	23.80
ACCl	30.5	21.00			

parameterisation. The haloalkane parameterisation also shows better agreement than the empirical scheme of Gasteiger [11].

In conclusion, the CHARGE program [2,3] has been subjected to a major reparameterisation of the  $\sigma$  scheme and shows good agreement between observed and calculated dipole moments. The required atomic parameters, electronegativities and polarisabilities, and the bond pair parameters are presented in Tables 8 and 9.

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