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Heats of Formation of Organic Compounds by a Simple Calculation

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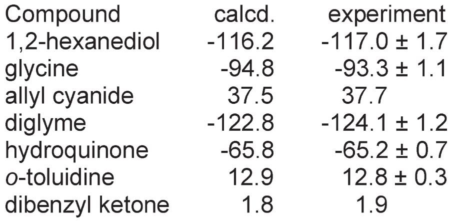

Compound	calcd.	experiment
1,2-hexanediol	-116.2	-117.0 ± 1.7
glycine	-94.8	-93.3 ± 1.1
allyl cyanide	37.5	37.7
diglyme	-122.8	-124.1 ± 1.2
hydroquinone	-65.8	-65.2 ± 0.7
o-toluidine	12.9	12.8 ± 0.3
dibenzyl ketone	1.8	1.9

Heats (enthalpies) of formation of organic compounds are obtained by adding a constant for a functional group to the heat of formation of its hydrocarbon precursor obtained by counting hydrogens, as previously described. A single molecule is used to obtain the contribution of each functional group. Unlike all other empirical schemes, *no global fitting is used*. The calculation is done with pencil and paper or a hand calculator. The accuracy of the results is demonstrated by their agreement with more than 360 species of experimentally known heats of formation with a mean average deviation of 0.75 kcal mol^{-1}, very nearly equal to the mean experimental uncertainty. Alcohols, hydroperoxides, ethers, amines, ketones, aldehydes, carboxylic acids, esters, chlorides, phenols, anilines, aryl ethers, nitriles, amides, and alkyl radicals are used as examples of the method. This success allows the calculation of experimentally unknown heats of formation. The method is limited to compounds that are not subject to strain or effects of conjugation. In such cases, the experimental heat of formation of the hydrocarbon precursor may be used. The method can be extended to functional groups not treated here.

Introduction

Empirical schemes exist for estimating gas phase heats (enthalpies) of formation of organic molecules and of free radicals, and they generally succeed because of regularities present in such compounds and of the fact that many of their parameters are found by global fitting to a large number of molecules.[1] Ab initio calculations also are being widely employed to the same end.[2] In this work, we present what we feel is the simplest empirical approach. It is more limited than other approaches in that it does not apply to compounds exhibiting "special" effects, such as strained compounds or effects of resonance. However, our approach is not based on global fitting. The procedure is performed with pencil and paper or a hand calculator. Thus this approach brings into sharp focus the existing regularities that may not be evident in more widely applicable but more complex approaches. Regularities in the heats of combustion of normal alkanes have been known for a long time.[1h]

We have devised a simple group additivity approach (eq 1) for calculating heats of formation, $\Delta_f H$, of hydrocarbons based on counting the number of hydrogens of each type: n_1 of primary (1°) aliphatic hydrogens (H's bonded to sp^3 carbon), n_2 of 2° aliphatic H's, n_3 of 3° aliphatic H's, n_4 of vinyl H's, n_5 of H's missing from C=C groups, n_6 is the number of terminal alkynes including acetylene, n_7 is the number of internal alkynes, n_8 is the number of H's missing

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from benzene rings, and n_9 is the number of benzene rings.[3]

$$\Delta_f H = -3.345n_1 - 2.49n_2 - 1.59n_3 + 3.135n_4 + 5.60n_5 + 54.42n_6 + 55.31n_7 + 2.17n_8 + 19.82n_9,\ \text{kcal mol}^{-1}\ \text{at 298 K} \quad (1)$$

The procedure is applicable to all species that are not subject to "special" effects. It is not based on global fitting to a large variety of known compounds, as is done by well-known approaches such as Benson's group additivity[1a,b] or by Allinger's successful molecular mechanics[4] approach that establishes transferable structural regularities by examining a large variety of molecules with known properties. The coefficients of each term in eq 1 were obtained from a typical molecule in each case: c_1 from 2,2-dimethylpropane, c_2 from butane, c_3 from 2,4-dimethylpentane, c_4 from ethylene, c_5 from *trans*-2-butene, c_6 from 1-butyne, c_7 from 2-butyne and 2-pentyne, c_8 from toluene, and c_9 from benzene.

Some examples of applying eq 1 to obtain results equal to experimental values within experimental error include the following: ethane, $n_1 = 6$, all other $n_i = 0$, calculated $\Delta_f H = -20.07$ kcal mol^{-1}; cyclohexane, $n_2 = 12$, all other $n_i = 0$, calculated $\Delta_f H = -29.9$; isobutane, $n_1 = 9$, $n_3 = 1$, all other $n_i = 0$, and $\Delta_f H = -31.7$; propylene, $n_1 = 3$, $n_4 = 3$, $n_5 = 1$, all other $n_i = 0$, calculated $\Delta_f H = 4.97$; acetylene, $n_6 = 1$, all other $n_i = 0$, and $\Delta_f H = 54.42$; 1-octyne, $n_1 = 3$, $n_2 = 10$, $n_6 = 1$, all other $n_i = 0$, calculated $\Delta_f H = 19.5$; *n*-butylbenzene, $n_1 = 3$, $n_2 = 6$, $n_8 = 1$, $n_9 = 1$, all other $n_i = 0$, calculated $\Delta_f H = -2.99$; *trans*-decalin, $n_2 = 16$, $n_3 = 2$, all other $n_i = 0$, and $\Delta_f H = -43.0$; triphenylmethane, $n_3 = 1$, $n_8 = 3$, $n_9 = 3$, all other $n_i = 0$, calculated $\Delta_f H = 64.4$. For over 100 other hydrocarbons, $\Delta_f H$ calculated by eq 1 is equal to experimental values, within experimental error, for species not subject to strain or effects of resonance.[3] We have not attributed any particular rationalizations for the success of this approach. For molecules subject to special effects, the difference between the calculated and experimental value provides a measure of the magnitude of the effect.

New schemes for estimating heats of formation of hydrocarbons that are not subject to special effects and of their radicals have appeared recently, and they have caused lively debate about the rationalizations offered.[5] It turns out that successful schemes can be shown to map onto each other.[5g] Hence, the simplest should be the best for establishing regularities in heats of formation not only of hydrocarbons but also of a wide variety of other classes of organic compounds.

We report here an extension of the approach of eq 1 for estimating heats of formation of hydrocarbon radicals and of organic compounds other than hydrocarbons. When $\Delta_f H$ of one hydrocarbon derivative compound is known, homologues also can be estimated by extending eq 1. For example, if the hydrocarbon precursor of acetone is propane and $\Delta_f H$[acetone] is known, $\Delta_f H$ of other ketones can be obtained from $\Delta_f H$ of the precursor hydrocarbon calculated by eq 1 plus a stabilization constant that is equal to the difference in $\Delta_f H$ of propane and acetone. Common functional groups are treated: alkyl radicals, alcohols, hydroperoxides, ethers, amines, ketones, aldehydes, carboxylic acids, esters, chlorides, phenols, anilines, aryl ethers, nitriles, amides, and their combinations in the same molecule. Sufficient examples are provided to establish the broad applicability of this approach and the regularities present.

Results and Discussion

Carbon-Centered Radicals. Gas phase heats of formation of 1, 2, and 3° alkyl radicals can be obtained accurately by eqs 2−4, where R denotes an alkyl group. These relationships are valid for hydrocarbon radicals that are not resonance-stabilized, for example, not allylic, propargylic, or benzylic. Heats of formation of the hydrocarbons can be obtained from eq 1 or from experimental values in the literature. The constants in eqs 2−4 are based on 1-propyl radical, on 3-pentyl radical, and on *tert*-butyl radical, respectively.

$$\Delta_f H[\mathrm{RCH_2^\bullet}] = \Delta_f H[\mathrm{R_2CH_3}] + 48.9\ \text{kcal mol}^{-1} \quad (2)$$

$$\Delta_f H[\mathrm{R_2CH^\bullet}] = \Delta_f H[\mathrm{R_2CH_2}] + 46.5\ \text{kcal mol}^{-1} \quad (3)$$

$$\Delta_f H[\mathrm{R_3C^\bullet}] = \Delta_f H[\mathrm{R_3CH}] + 44.1\ \text{kcal mol}^{-1} \quad (4)$$

Table 1 shows some applications of eq 1 and of eqs 2−4 for 25 alkyl radicals, for which experimental values of heats of formation are available. Five more are given as examples for calculating experimentally unknown values.

Cyclization of 5-hexen-1-yl radicals to cyclohexyl or cyclopentylmethyl radicals (Scheme 1) is a reaction of the type used as a "radical clock". Rearrangements involving similar species have been studied as part of synthetic approaches for forming C−C bonds. In this context, Houk et al. have performed ab initio calculations (G3MP2B3) of the thermochemistries of the cyclizations.[9] Path a was found to be exothermic by −20.4 kcal mol^{-1} and by path b by −6.2, with an uncertainty estimated as ±1 kcal mol^{-1}.

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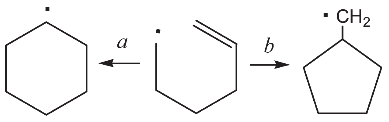
TABLE 1. Precursor Alkane, Radical, Heat of Formation of Alkane by eq 1, Heat of Formation of the Radical by eqs 2−4, and Experimental Heat of Formation of the Radical, kcal mol^{-1} at 298 K

RH	R$^{\bullet}$	calcd $\Delta_f H$[RH]	calcd $\Delta_f H$[R$^{\bullet}$]	literature $\Delta_f H$[R$^{\bullet}$]^a
CH_3CH_3	$CH_3CH_2^{\bullet}$	−20.0	28.9	28.4 ± 0.5
$CH_3CH_2CH_3$	$CH_3CH_2CH_2^{\bullet}$	−25.0	23.9	23.9 ± 0.5
$CH_3CH_2CH_3$	$(CH_3)_2CH^{\bullet}$	−25.0	21.5	22.0 ± 0.5, 21.0 ± 0.7^b
$CH_3(CH_2)_2CH_3$	$CH_3(CH_2)_2CH_2^{\bullet}$	−30.0	18.9	18.6^b
$CH_3(CH_2)_2CH_3$	$CH_3(CH_2)_2C{\cdot}HCH_3$	−30.0	16.5	16.0 ± 0.5, 16.2 ± 0.5^b
$(CH_3)_2CHCH_3$	$(CH_3)_2CHCH_2^{\bullet}$	−32.1	16.8	17.0 ± 0.5, 16.7 ± 1.0^b
$(CH_3)_2CHCH_3$	$(CH_3)_3C^{\bullet}$	−32.1	12.0	11.0 ± 0.7, 11.5 ± 0.7^b
$CH_3(CH_2)_3CH_3$	$CH_3(CH_2)_3CH_2^{\bullet}$	−35.1	13.8	13.0
$CH_3(CH_2)_3CH_3$	$CH_3(CH_2)_2C{\cdot}HCH_3$	−35.1	11.4	12.0
$CH_3(CH_2)_3CH_3$	$CH_3CH_2C{\cdot}HCH_2CH_3$	−35.1	11.4	11.2^b
$(CH_3)_3CCH_3$	$(CH_3)_3CCH_2^{\bullet}$	−40.1	8.8	8.7 ± 2, 8.0^c
$(CH_3)_2CHCH_2CH_3$	$(CH_3)_2C{\cdot}CH_2CH_3$	−36.7	7.4	6.7 ± 0.7, 7.0^b
$CH_3(CH_2)_4CH_3$	$CH_3(CH_2)_4CH_2^{\bullet}$	−39.9	9.0	8.0^b
$CH_3(CH_2)_4CH_3$	$CH_3(CH_2)_3C{\cdot}HCH_3$	−39.9	6.6	7.0^b
$(CH_3CH_2)_2CHCH_3$	$(CH_3CH_2)_2C{\cdot}CH_3$	−41.0	3.1	3.4^b
$(CH_3)_2CHCH_2CH_2CH_3$	$(CH_3)_2C{\cdot}CH_2CH_2CH_3$	−41.7	2.4	0.8 ± 2^b
$CH_3CH(CH_3)CH(CH_3)CH_3$	$CH_3C{\cdot}(CH_3)CH(CH_3)CH_3$	−42.5	1.6	0.7 ± 2.4^b
$CH_3(CH_2)_5CH_3$	$CH_3(CH_2)_5CH_2^{\bullet}$	−45.0	3.9	4^c
$CH_3(CH_2)_5CH_3$	$CH_3(CH_2)_4C{\cdot}HCH_3$	−45.0	1.6	2^c
$CH_2{=}CHCH_2CH_3$	$CH_2{=}CHCH_2CH_2^{\bullet}$	0.0	48.9	46^b
$CH_2{=}CH(CH_2)_2CH_3$	$CH_2{=}CH(CH_2)_2CH_2^{\bullet}$	−5.0	43.9	42.9^b
$CH_2{=}CHCH(CH_3)_2$	$CH_2{=}CHCH(CH_3)CH_2^{\bullet}$	−6.7	42.2	42.3^c
$C_6H_5CH_2CH_3$	$C_6H_5CH_2CH_2^{\bullet}$	7.0	55.9	55.9^b
$C_6H_5(CH_2)_3CH_3$	$C_6H_5(CH_2)_3CH_2^{\bullet}$	−3.0	45.9	45.9^b
$C_6H_5(CH_2)_3CH_3$	$C_6H_5CH_2C{\cdot}HCH_2CH_3$	−3.0	43.5	44.1^b
cyclo-$C_5H_9CH_3$^{d}	cyclo-$C_5H_9CH_2^{\bullet}$	−25.3	23.6	
$CH_2{=}CH(CH_2)_3CH_3$	$CH_2{=}CH(CH_2)_3CH_2^{\bullet}$	−10.0	38.9	
cyclo-$C_6H_{11}CH_3$	cyclo-$C_6H_{11}CH_2^{\bullet}$	−38.2	5.9	
$CH_3C{\equiv}C(CH_2)_4CH_3$	$CH_3C{\equiv}C(CH_2)_4CH_2^{\bullet}$	15.3	64.2	
$C_6H_5(CH_2)_2CH_3$	$C_6H_5CH_2CH_2CH_2^{\bullet}$	2.0	50.9	

[a] From ref 6, unless indicated otherwise. [b] From ref 7. [c] From ref 8. [d] Because the precursor alkane is strained, the experimental[6] value of $\Delta_f H$[RH] is used.

SCHEME 1



The experimental heat of formation of the cyclohexyl radical is known, 18 ± 1.5 kcal mol^{-1}, but those of 5-hexen-1-yl and cyclopentylmethyl are not. Our approach can be used to estimate their values. All experimental values cited throughout the text are from ref 6, unless specified otherwise. By eqs 1 and 2, $\Delta_f H$[5-hexen-1-yl] = $\Delta_f H$[1-hexene] + 48.9 = −10.0 + 48.9 = 38.9. Because cyclopentane is strained and eq 1 is not applicable, the experimental $\Delta_f H$[methylcyclopentane] = −25.3 is used and, by eq 2, $\Delta_f H$[cyclopentylmethyl radical] = −25.3 + 48.9 = 23.6. Hence, the exothermicity of path a is 18.0 − 38.9 = −20.9 kcal mol^{-1} and of path b is 18.0 − 23.6 = −5.6. Both of our estimates are in excellent agreement with the high level ab initio estimates.

Estimates by group additivity values (GAV) based on CBS-QB3 calculations have been made for a very large variety of carbon radicals, including vinyl, alkynyl, aromatic radicals, and conjugated and strained varieties, etc. In all, 59 group additivity values were used.[1c] For the 13 in common with Table 1, GAV results yield $\Delta_f H$[R$^{\bullet}$] consistently somewhat greater than experimental values by about 1.5 kcal mol^{-1}. This pattern persists for many additional carbon radicals that are not subject to special effects. Some examples are as follows for our approach and for the CBS-QB3/GAV results in parentheses: 3-buten-1-yl, 48.9 (50.5); 4-pentyn-1-yl, 84.2 (85.9); 5-penten-1-yl, 43.9 (45.6) 5-hexen-1-yl, 38.9 (40.7); 2-methyl-3-buten-1-yl, 42.2 (43.8); 3-methyl-4-penten-2-yl, 34.9 (36.2); 2-phenylethyl, 55.9 (57.4); 3-phenylpropyl, 50.9 (52.5).

The successful agreement with experimental values demonstrated in Table 1 indicates that accurate estimates of $\Delta_f H[R^{\bullet}]$ can be obtained for alkyl radicals that are not subject to effects of strain or resonance. This is the same limitation on two more elaborate recent schemes.[5a,e]

The same approach of eqs 2−4 is now applied to radicals that are subject to special effects. For $^{\bullet}CH_2CH_2^{\bullet}$, the precursor alkane is ethane, $\Delta_f H[CH_3CH_3]$ = −20.0 kcal mol^{-1}. By eq 2 for two 1° alkyl radicals, $\Delta_f H[^{\bullet}CH_2CH_2^{\bullet}]$ = −20.0 + 2 × 48.9 = 77.8 kcal mol^{-1} versus experimental $\Delta_f H[CH_2{=}CH_2]$ = 12.54. The difference, 77.8 − 12.54 = 65.3 kcal mol^{-1}, is the often quoted strength of the π bond of ethylene based on the energy of 65.0 kcal mol^{-1} required for the *cis*/*trans* isomerization of 1,2-dideuterioethylene.[10]

For the benzyl radical, $\Delta_f H[C_6H_5CH_2^{\bullet}] = \Delta_f H[C_6H_5CH_3]$ + 48.9 = (12 ± 0.3) + 48.9 = 60.9 ± 0.3 kcal mol^{-1}. Subtraction of the known[11] conjugation stabilization of the benzyl radical yields 60.9 − 11.7 = 49.2 ± 0.3, which is its experimental heat of formation, 49.5 ± 1.

Cyclohexane has been reported to be a strain-free molecule.[3] The enthalpy of reaction for breaking one C−C bond in cyclohexane to yield the diradical hexane-1,6-diyl can be calculated from eqs 1 and 2. By eq 1, $\Delta_f H$[cyclohexane] = −2.490 × 12 = −29.88 kcal mol^{-1} versus experimental −29.78

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TABLE 2. Primary Alcohols, $\Delta_fH[\mathrm{RCH_3}]$ by Equation 1, Calculated $\Delta_fH[\mathrm{RCH_2OH}] = \Delta_fH[\mathrm{RCH_3}] - (36.0 \pm 0.5)$, and Experimental $\Delta_fH[\mathrm{RCH_2OH}]$, kcal mol^{-1} at 298 K

alcohol	calcd $\Delta_fH[\mathrm{RCH_3}]$	calcd $\Delta_fH[\mathrm{RCH_2OH}]$	experimental[a] $\Delta_fH[\mathrm{RCH_2OH}]$
ethanol	−20.0	−56.0	−56.0 ± 0.5
1,2-ethanediol	−20.0	−92.0	−94.3, −92.7, −93.3, −92.6
1-propanol	−25.0	−61.0	−61.2 ± 0.5
1,3-propanediol	−25.0	−97.0	−97.6 ± 1.2, −93.7 ± 0.8, −96.9[b]
allyl alcohol	5.0	−31.0	−29.6 ± 0.4, −30.7 ± 0.7[c]
propargyl alcohol	−44.5	8.5	11[b]
1-butanol	−30.0	−66.0	−66.0 ± 1.0
2-methyl-1-propanol	−31.7	−67.7	−67.8 ± 0.2
1,4-butanediol	−30.0	−102.0	−101.8 ± 1.4, 102.0 ± 0.6
2-methyl-2-propen-1-ol	−2.6	−38.6	−38[b]
2-buten-1-ol	−2.6	−38.6	−37[b]
1-pentanol	−35.0	−71.0	−71.0 ± 1.0
2-methyl-1-butanol	−36.7	−72.7	−72.2 ± 0.4
3-methyl-1-butanol	−36.7	−72.7	−71.9 ± 0.4, −72.0 ± 0.4
neopentyl alcohol	−40.1	−76.1	−76.1[b]
4-pentene-1-ol	−5.0	−41.0	−41.0[b]
1,5-pentanediol	−35.0	−107.0	−105.6 ± 1.4, −107.3 ± 0.8
1-hexanol	−40.0	−76.0	−75.0 ± 2.0
1,6-hexanediol	−40.0	−112.0	−109.8 ± 0.5, −110.4 ± 1.1
1-heptanol	−45.0	−81.0	−82 ± 1, −80.2 ± 0.4[b]
benzyl alcohol	11.9	−24.1	−22.6 ± 0.7
1-octanol	−50.0	−86.0	−85.0 ± 1.0
2-ethyl-1-hexanol	−51.6	−87.6	−87.8 ± 0.5
1-nonanol	−54.9	−90.9	−90.0 ± 2.0
1-decanol	−59.9	−95.9	−94.0 ± 3.0
1-dodecanol	−69.9	−105.9	−104.3 ± 0.2
2-phenylethanol	7.0	−29.0	−28.8[b]
3-buten-1-ol	−0.0	−36.0	
2-ethyl-1-butanol	−46.6	−82.6	
cyclohexylmethanol	−36.5	−72.5	
1-undecanol	−64.9	−100.9	

[a] From ref 6, unless indicated otherwise. [b] From ref 8. [c] From ref 12.

and $\Delta_fH[n\text{-hexane}] = -3.345 \times 6 - 2.490 \times 8 = -39.99$ versus experimental −39.94. By eq 2, $\Delta_fH[{}^\bullet\mathrm{CH_2(CH_2)_4CH_2}^\bullet] = \Delta_fH[n\text{-hexane}] + 2 \times 48.9 = 57.81$. Hence the enthalpy of reaction for cyclo-$\mathrm{C_6H_{12}} \rightarrow {}^\bullet\mathrm{CH_2(CH_2)_4CH_2}^\bullet$ is given by $\Delta H_{\mathrm{rxn}} = 57.81 - (-29.88) = 87.69$, which is the calculated bond dissociation energy, BDE, of a C−C bond in cyclohexane. This can be compared to the experimental value for the unstrained central bond of hexane, $\mathrm{BDE[CH_3CH_2CH_2{-}CH_2CH_2CH_3]} = 87.74 \pm 0.71$. The two values are equal, indicating no strain in cyclohexane and confirming the accuracy of the value of 48.9 of eq 2.

Alcohols. Equation 1 is also used successfully for estimating heats of formation of homologues of non-hydrocarbon molecules by relating them to their precursor hydrocarbon of the same carbon skeleton plus a constant, as shown below for alcohols and other functional groups in the following subsections. In comparing our calculated values to experimental data, it should be kept in mind that cited experimental uncertainties normally denote precision of the measurements, not necessarily accuracy.

For primary alcohols, the most accurately determined value is for ethanol, $\Delta_fH[\mathrm{CH_3CH_2OH}] = -56.0 \pm 0.5$ kcal mol^{-1}. From this as a basis, we obtain the difference between the experimental heat of formation of the alcohol and the heat of formation of the precursor hydrocarbon, the latter calculated by eq 1: $\Delta_fH[\mathrm{CH_3CH_2OH}] - \Delta_fH[\mathrm{CH_3CH_3}] = (-56.0 \pm 0.5) - (-20.0) = -36.0 \pm 0.5$. The difference of replacing a 1° H of the hydrocarbon by OH remains constant for all primary alcohols not subject to intramolecular hydrogen bonding, as shown in Table 2, where the reported uncertainty of the calculated $\Delta_fH[\mathrm{RCH_2OH}]$ is ±0.5 kcal mol^{-1}. Table 2 also contains values of four alcohols for which experimental Δ_fH is not available. By counting hydrogens for eq 1 and adding a stabilization constant of −36.0 kcal mol^{-1}, heats of formation of primary alcohols are obtained accurately.

For secondary alcohols, with isopropanol as a basis, the difference between the heat of formation of the alcohol and that of the precursor hydrocarbon is $\Delta_fH[\mathrm{CH_3CH(OH)CH_3}] - \Delta_fH[\mathrm{CH_3CH_2CH_3}] = (-65.2 \pm 0.2) - (-25.0) = -40.2 \pm 0.2$ kcal mol^{-1}. The difference in Δ_fH brought about by replacing a 2° H by OH remains constant for other secondary alcohols, as shown in Table 3. Lack of an experimental entry indicates apparently unavailable experimental heats of formation. In this and subsequent tables, it is demonstrated that the effect of two different functional groups is additive, as in 1,2-propanediol, unless doing so introduces a special effect (strain, resonance, or intramolecular hydrogen bonding).

There is a significant discrepancy in Table 3 between the reported experimental value of −87.8 kcal mol^{-1} for (*E*)-4-methylcyclohexanol versus our calculated −76.7. With both substituent groups on cyclohexane being equatorial and not causing any significant strain, our calculated value should be accurate, as it is for cyclohexanol and for methylcyclohexane.

For tertiary alcohols, there are not many experimental heats of formation available. A constant pattern appears to hold: $\Delta_fH[\mathrm{R_3COH}] = \Delta_fH[\mathrm{R_3CH}] - 43.0$ kcal mol^{-1}. With $\Delta_fH[\mathrm{R_3CH}]$ calculated by eq 1, the following $\Delta_fH[\mathrm{R_3COH}]$

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(12) From the heat of hydrogenation, −30.5 kcal mol^{-1}, of allyl alcohol to 1-propanol: Dolliver, M. A.; Gresham, T. L.; Kistiakowsky, G. B.; Smith, E. A.; Vaughan, W. E. *J. Am. Chem. Soc.* **1938**, *60*, 440.

TABLE 3. Secondary Alcohols, $\Delta_f H$[RR′CH$_2$] by Equation 1, Calculated $\Delta_f H$[RR′CHOH] = $\Delta_f H$[RR′CH$_2$] − (40.20 ± 0.2), and Experimental $\Delta_f H$[RR′CHOH], kcal mol^{-1} at 298 K

alcohol	calcd $\Delta_f H$[RR′CH$_2$]	calcd $\Delta_f H$[RR′CHOH]	experimental[a] $\Delta_f H$[RR′CHOH]
2-propanol	−25.0	−65.2	−65.2 ± 0.2
1,2-propanediol[b]	−25.0	−101.2	−102.7 ± 1.0, −100.7 ± 0.8
1,2,3-propanetriol[b]	−25.0	−137.2	−138.1 ± 0.3, −137.9
2-butanol	−30.0	−70.2	−70.1 ± 0.2
1,3-butanediol[b]	−30.0	−106.2	−103.7 ± 0.5, −106.9
3-butyn-2-ol	39.4	−0.8	2,[c] (−0.1 estimated)[c]
2-pentanol	−35.0	−75.2	−74.7 ± 0.2
3-pentanol	−35.0	−75.2	−75.8, −75.2, −75.4 ± 0.2[c]
1-pentene-3-ol	−5.0	−45.2	−45[c]
4-pentene-2-ol	−5.0	−45.2	−43[c]
3-methyl-2-butanol	−36.7	−76.9	−75.6 ± 0.4, −74.8
1,2-pentanediol	−35.0	−111.2	−111.2 ± 1.8
2-hexanol	−40.0	−80.2	−79.7 ± 0.5
3-hexanol	−40.0	−80.2	−79.5 ± 0.5, −79[c]
4-methyl-2-pentanol	−41.7	−81.9	−82.2 ± 0.5
3-methyl-2-pentanol	−41.7	−81.9	−80.9[c]
2-methyl-3-pentanol	−41.7	−81.9	−81.0 ± 0.5
1,2-hexanediol[b]	−40.0	−116.2	−117.0 ± 1.7
cyclohexanol	−29.9	−70.1	−69 ± 2
2-heptanol	−45.0	−85.2	−84.9 ± 0.5
3-heptanol	−45.0	−85.2	−84.9 ± 0.5
4-heptanol	−45.0	−85.2	−84.8 ± 0.5
(*E*)-4-methylcyclohexanol	−36.5	−76.7	−87.8[d]
cyclobutanol[e]	6.8	−33.4	−34.6, −32[c]
cyclopentanol[e,f]	−18.3	−58.5	−58.1, −58.0 ± 0.4
7-norbornanol[e,g]	−13.1	−53.3	−52.0 ± 2.0
2-adamantanol[e,h]	−32.1	−72.3	−71.5 ± 1.1
2-octanol	−50.0	−90.2	
1-phenylethanol	7.0	−33.2	
2-nonanol	−54.9	−95.1	
1-phenyl-2-propanol	2.0	−38.2	
diphenylmethanol	39.0	−1.2	

[a] From ref 6, unless indicated otherwise. [b] Also subtracting 36.0 for each primary alcohol stabilization. [c] From ref 8. [d] From ref 13. [e] For strained hydrocarbons the experimental heat of formation is used. [f] Experimental $\Delta_f H$[cyclobutane] quoted in ref 14. [g] Bicyclo[2.2.1]heptan-7-ol. [h] Tricyclo-[3.3.1.1^{3,7}]decan-2-ol.

values are obtained (experimental values in parentheses): *tert*-butyl alcohol, −74.7 (−74.7 ± 0.2, −74.9 ± 0.3); 2-methyl-2-butanol, −78.1 (−78.7[8]); 2-methyl-3-buten-2-ol, −49.7 (estimated −50[8]); 3-methyl-3-pentanol, −84.7 (−83.2 ± 0.1[8]). For three precursor hydrocarbons suspected of strain, the experimental heat of formation must be used: 2,3-dimethyl-2-butanol, −42.5 − 43.0 = −85.5 (−85.3 ± 0.1[8]); 1-methylcyclopentanol, −25.3 − 43.0 = −68.3 (−67.9 ± 0.3[8]); 1-adamantanol (tricyclo[3.3.1.1^{3,7}]decan-1-ol), −32.1 − 43.0 = −75.1 (−74.3 ± 0.8).

Equation 1 is not applicable to methane. Methane is unique among hydrocarbons in that it does not contain any C−C bonds and its heat of formation does not decrease by 4.95 kcal mol^{-1} on going to ethane, the next alkane in the series, as is the case of all linear hydrocarbons all the way from ethane to octadecane. However, from the experimental $\Delta_f H[CH_3OH] = 49 \pm 3$ kcal mol^{-1} and $\Delta_f H[CH_4] = -17.9$ kcal mol^{-1}, the difference is −31.1. The increasing stabilization constants of methyl, primary, secondary, and tertiary alcohols are −31.1, −36.0, −40.2, and −43.0 kcal mol^{-1}, relative to their precursor hydrocarbons. These differences are due to the different electronegativities, χ, of the groups $\chi[CH_3] = 2.525$, $\chi[RCH_2] = 2.462$, $\chi[R_2CH] = 2.411$, and $\chi[R_3C] = 2.378$, all bonded to a hydroxyl group with $\chi[OH] = 3.500$.[11] Figure 1 shows a plot of the stabilization constants of the alcohols versus the dipole contribution of strengthening the C−O bond of the alcohol. The dipole contribution is $23(\Delta\chi)^2$ kcal mol^{-1}. The plot has a correlation coefficient of 1.000. We are not aware of this type of correlation ever having been noted before as a function of $\Delta_f H$.

Hydroperoxides. Peroxides are of current interest for their involvement in combustion, in autoxidation, in biological oxidation of unsaturated lipids, etc.[15] $\Delta_f H$ of 1, 2, and 3° hydroperoxides can be estimated from $\Delta_f H$ of the precursor hydrocarbon of the same carbon skeleton by accounting for stabilizations of −18.3, −22.3, and −25.3 kcal mol^{-1}, respectively, as shown in Table 4. The table includes estimates of four examples of apparently unknown $\Delta_f H$. Two are of the type involved in peroxidations of unsaturated lipids, a secondary allylic hydroperoxide as a model for monounsaturated lipids, and another secondary with a CH$_2$ being doubly allylic as a model for similarly di-unsaturated lipids.

Because $\Delta_f H[CH_4]$ cannot be obtained from eq 1, its experimental value of −17.9 kcal mol^{-1} compared to experimental $\Delta_f H[CH_3OOH] = -31.2$ yields a stabilization of −13.3. A plot of stabilization energies for methyl, 1, 2, and

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(13) From the compilation available at http://courses.chem.indiana.edu/c360/documents/thermodynamicdata.pdf.

(14) National Institute of Standards and Technology, Computational Chemistry Comparison and Benchmark DataBase, Release 15, February 2010, NIST Standard Reference Database 101 (http://cccbdb.nist.gov/).

(15) For some examples, see: (a) Chatgilialoglu, C.; Ferreri, C. *Acc. Chem. Res.* **2005**, *38*, 441–448. (b) Sebbar, N.; Bozzelli, J. W.; Bockhorn, H. *J. Phys. Chem. A* **2004**, *108*, 8353–8366. (c) Sebbar, N.; Bockhorn, H.; Bozzelli, J. W. *Phys. Chem. Chem. Phys.* **2002**, *4*, 3691–3703. (d) Simmie, J. M.; Black, G.; Curran, H. J.; Hinde, J. P. *J. Phys. Chem. A* **2008**, *112*, 5010–5016. (e) Stevens, W. R.; Walker, S. H.; Shuman, N. S.; Baer, T. *J. Phys. Chem. A* **2010**, *114*, 804–810.

3° hydroperoxides versus $23(\Delta\chi)^2$ with $\chi[\mathrm{OOH}] = 3.334$,[11] as in Figure 1, yields a straight line: stabilization $= -1.878 \times 23(\Delta\chi)^2 + 19.06$; correlation coefficient = 0.9997.

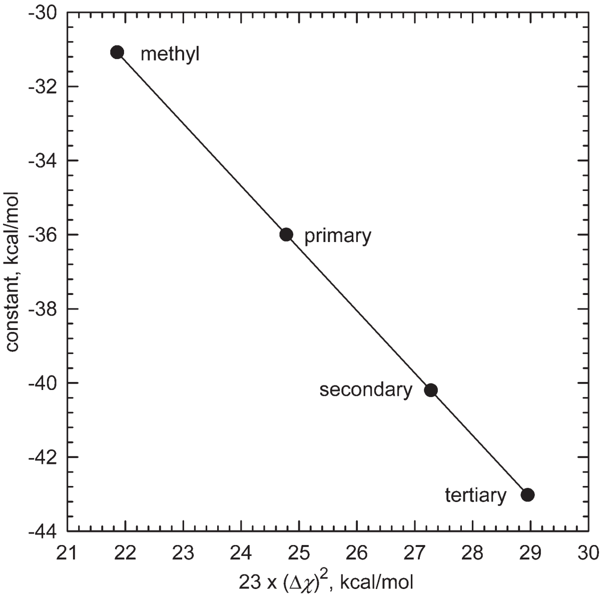

FIGURE 1. Stabilization constants of the alcohols versus $23(\Delta\chi)^2$ of the C−OH dipole contribution. Stabilization constant $= -1.684 \times 23(\Delta\chi)^2 + 5.722$; correlation coefficient = 0.9999997.

Ethers show regularities similar to those of alcohols. The insertion of an oxygen atom into a hydrocarbon R−R′ to form ether R−O−R′ stabilizes the precursor molecule in proportion to the magnitude of the bond dipole effects, as previously illustrated in Figure 1. For one $\mathrm{O{-}CH_3}$ bond, this stabilization is −12.0 kcal mol^{-1}. For two such bonds, as in dimethyl ether, twice −12.0. For one $\mathrm{O{-}CH_2R}$ bond, it is −14.7. For one $\mathrm{O{-}CHR_2}$ bond, it is −16.4. For one $\mathrm{O{-}CR_3}$ bond, it is −16.7. Some examples follow with $\Delta_f H$[R−R′] calculated by eq 1 and the experimental values in parentheses. For dimethyl ether, $\Delta_f H[\mathrm{CH_3OCH_3}] = \Delta_f H[\mathrm{CH_3CH_3}] - 12.0 \times 2 = -20.0 - 24.0 = -44.0$ (−44.0 ± 0.1). For ethyl methyl ether, $\Delta_f H[\mathrm{CH_3CH_2OCH_3}] = -25.0 - 14.7 - 12.0 = -51.7$ (−51.7 ± 0.2). For methyl propyl ether, $\Delta_f H[\mathrm{CH_3OCH_2CH_2CH_3}] = -30.0 - 12.0 - 14.7 = -56.7$ (−56.9[2g]). For diethyl ether, $\Delta_f H[\mathrm{CH_3CH_2OCH_2CH_3}] = -30.0 - 14.7 \times 2 = -59.4$ (−60.4 ± 0.5). For methyl isopropyl ether, $\Delta_f H[\mathrm{CH_3OCH(CH_3)_2}] = -31.7 - 12.0 - 16.4 = -60.1$ (−60.2 ± 0.2). For butyl methyl ether, $\Delta_f H[\mathrm{CH_3OCH_2CH_2CH_2CH_3}] = -35.0 - 12.0 - 14.7 = -61.7$ (−61.7[2g]). For diisopropyl ether, $\Delta_f H[\mathrm{(CH_3)_2CHOCH(CH_3)_2}] = -43.3 - 16.4 \times 2 = -76.1$ (−76.3[2g]). For di-*sec*-butyl ether, $\Delta_f H(\mathrm{CH_3CH_2CH(CH_3))_2O}] = -53.3 - 16.4 \times 2 = -86.1$ (−86.3 ± 0.5). For the gasoline additive *tert*-butyl methyl ether, $\Delta_f H[\mathrm{(CH_3)_3COCH_3}] = -40.1 - 16.7 - 12.0 = -68.8$ (−68.1, −67.7 ± 0.3, and −67.5 ± 0.5), indicating small, if any, steric effect. For *tert*-butyl isopropyl ether, $\Delta_f H[\mathrm{(CH_3)_3COCH(CH_3)_2}] = -51.7 - 16.4 - 16.6 = -84.8$ (−85.5[2g]) For di-*n*-propyl ether, $\Delta_f H[\mathrm{(CH_3CH_2CH_2)_2O}] = -40.0 - 14.7 \times 2 = -64.4$ (−65.1[2g]). For di-*n*-butyl ether, $\Delta_f H[\mathrm{(CH_3CH_2CH_2CH_2)_2O}] = -50.0 - 14.7 \times 2 = -79.4$ (−79.8 ± 0.5). For 1,2-dimethoxyethane, $\Delta_f H[\mathrm{CH_3OCH_2CH_2OCH_3}] = -30.0 - 12.0 \times 2 - 14.7 \times 2 = -83.4$ (−81.9 ± 0.2, −93.1 ± 0.5 and −81.4 ± 0.2). For 1,2-ethoxyethane, $\Delta_f H[\mathrm{CH_3CH_2OCH_2CH_2OCH_2CH_3}] = -40.0 - 14.7 \times 4 = -98.8$ (−98.1 and −97.6 ± 1.0). For diglyme (diethylene glycol dimethyl ether), $\Delta_f H[\mathrm{(CH_3OCH_2CH_2)_2O}] = \Delta_f H[\mathrm{CH_3(CH_2)_4CH_3}] = -40.0 - 12.0 \times 2 - 14.7 \times 4 = -122.8$ (−124.6, −122.3, and −125.4). For dibenzyl ether, $\Delta_f H[\mathrm{(C_6H_5CH_2)_2O}] = 34.0 - 14.7 \times 2 = 4.6$ (4.7[8]). For 2-methoxyethanol, by also taking account of stabilization by a 1° alcohol function, $\Delta_f H[\mathrm{CH_3OCH_2CH_2OH}] = -25.0 - 12.0 - 14.7 - 36.0 = -87.7$ (−90.1 ± 1.9). For 3-methoxypropanol, with stabilization by a 1° alcohol, $\Delta_f H[\mathrm{CH_3OCH_2CH_2CH_2OH}] = -30.0 - 12.0 - 14.7 - 36.0 = -92.7$ (93[8]). There is steric hindrance in di-*tert*-butyl ether, as indicated by calculated $\Delta_f H = -93.6$ versus experimental −86.5 and −87.0 ± 0.2.[8]

Amines. For amines, replacing a 1° H with $\mathrm{NH_2}$ destabilizes the molecule by +8.7 kcal mol^{-1} on the basis of $\Delta_f H$ $[\mathrm{CH_3CH_3}] = -20.0$ kcal mol^{-1} and $\Delta_f H$ $[\mathrm{CH_3CH_2NH_2}] = -11.3$. Twenty-one examples for which $\Delta_f H$ values are available and four new estimates are shown in Table 5.

Heats of formation of alkylamines with the $\mathrm{NH_2}$ group on 2° carbon (replacing a 2° H by $\mathrm{NH_2}$) are obtained accurately by $\Delta_f H[\mathrm{RR'CHNH_2}] = \Delta_f H[\mathrm{RR'CH_2}] + 5.0$ kcal mol^{-1},

TABLE 4. **Hydroperoxides, $\Delta_f H$[hydrocarbon] by Equation 1, Calculated $\Delta_f H$[ROOH] = $\Delta_f H$[hydrocarbon] + Stabilization,[a] and Experimental $\Delta_f H$[ROOH], kcal mol^{-1} at 298 K**

hydroperoxide	calcd $\Delta_f H$[RH]	calcd $\Delta_f H$[ROOH]	experimental[b] $\Delta_f H$[[ROOH]
$\mathrm{CH_3CH_2OOH}$	−20.0	−38.3	−50.0, −39.5 ± 0.7,[c] 38.8[d]
$\mathrm{CH_3CH_2CH_2OOH}$	−25.0	−43.3	−60.0, −44.1,[e] −43.8[f]
$\mathrm{CH_2{=}CHCH_2OOH}$	5.0	−13.3	−13.6[g]
$\mathrm{CH_3CH_2CH_2CH_2OOH}$	−30.0	−48.3	−49,[h] −48.4[f]
$\mathrm{CH_3(CH_2)_3CH_2OOH}$	−35.0	−53.3	−52[h]
$\mathrm{(CH_3)_2CHCH_2OOH}$	−30.1	−48.4	−50.2[f]
$\mathrm{(CH_3)_2CHOOH}$	−25.1	−47.4	−47.1, −48.1,[d] −47.9[f]
$\mathrm{CH_3CH_2CH(CH_3)OOH}$	−30.0	−52.3	−52.7[f]
cyclo-$\mathrm{C_6H_{11}OOH}$	−29.9	−52.2	−51.4
$\mathrm{(CH_3)_3COOH}$	−30.1	−55.4	−56.1, −57.2,[f] −55.7 ± 0.6[i]
$\mathrm{C_6H_5C(CH_3)_2OOH}$[j]	3.9	−21.4	−18.8 ± 1.6
$\mathrm{CH_3CH{=}CHCH_2OOH}$	−2.6	−20.9	
$\mathrm{CH_3CH{=}CHCH(CH_3)OOH}$	−41.0	−63.3	
$\mathrm{(CH_2{=}CH)_2CHOOH}$	25.0	2.7	
$\mathrm{C_6H_5CH_2OOH}$	11.9	−6.4	

[a] Stabilization = −18.3 for 1° hydroperoxides, −22.3 for 2°, and −25.3 for 3°. [b] Experimental values from ref 6, unless indicated otherwise. [c] From ref 16. [d] From ref 17, theoretical CBS-Q. [e] Selected value, ref 15b. [f] From ref 15d. [g] From ref 15c. [h] From ref 8. [i] From ref 15e, CBS-Q calculation. [j] The precursor isopropylbenzene is subject to steric strain. Its experimental $\Delta_f H = 3.9 \pm 1.1$ was used, and the calculated value has the same uncertainty.

TABLE 5. Alkylamines (NH_2 on 1° Carbon), $\Delta_fH[RCH_3]$ by Equation 1, Calculated $\Delta_fH[R_2CH_2NH_2] = \Delta_fH[RCH_3] + 8.7$, and Experimental $\Delta_fH[RCH_2NH_2]$, kcal mol^{-1} at 298 K

amine	calcd $\Delta_fH[RCH_3]$	calcd $\Delta_fH[RCH_2NH_2]$	experimental[a] $\Delta_fH[RCH_2NH_2]$
ethaneamine	−20.0	−11.3	−13.8, −11.3,[b] −11.4[c]
2-aminoethanol[d]	−20.0	−47.3	−48,[b] −48.2[c]
propanamine	−25.0	−16.3	−16.7 ± 0.2, −25.5
allylamine	5.0	13.7	11.0,[b] estimated 13.1[b]
3-aminopropan-1-ol[d]	−25.0	−52.3	−52[b]
2-Me-2-propen-1-amine	−2.6	6.1	5[b]
butanamine	−30.0	−21.3	−22.7, −35.7
2-methylpropanamine	−31.7	−23.0	−23.6 ± 0.1
2-methyl-2-propenamine	−2.6	6.1	5.0[b]
pentanamine	−35.0	−26.3	−23.0,[b] −27.3[e]
3-methylbutanamine	−36.7	−28.0	−28.2[b]
neopentylamine	−40.1	−31.4	−31[b]
hexanamine	−40.0	−31.3	−46.3, −32[b]
heptanamine	−45.0	−36.3	−51.6
octanamine	−50.0	−41.3	−41.5 ± 0.3, −41[b]
decacanamine	−59.9	−51.2	−51[b]
benzylamine	11.9	20.6	21.0 ± 0.7, 20.0 ± 0.7[b]
2-phenylethanamine	7.0	15.7	16[b]
3-phenylpropanamine	−2.0	10.7	11.0[b]
1,2-ethanediamine	−20.0	−2.6	−4.1 ± 0.2, −4.3 ± 0.5[b]
1,6-hexanediamine	−40.0	−22.4	−23.3[b]
aminoacetone[f]	−25.0	−43.5	
1,4-butanediamine	−30.0	−12.6	
6-aminohexan-1-ol[d]	−40.0	−67.2	
nonanamine	−54.9	−46.2	

[a] Experimental values from ref 6, unless indicated otherwise. [b] From ref 8. [c] From ref 18a. [d] Also subtracting 36 for a primary alcohol stabilization. [e] Ab initio G3 calculation, ref 2g. [f] Also subtracting 27.2 for stabilization by one ketone function (see below, Table 7).

where $\Delta_fH[RR'CH_2]$ is calculated by eq 1, and the destabilization constant is derived from isopropylamine: $\Delta_fH[CH_3CH(NH_2)CH_3] - \Delta_fH[CH_3CH_2CH_3] = -20.0 - (-25.0) = 5.0$. Some examples of calculated heats of formation are (experimental values in parentheses): isopropylamine, −20.0 kcal mol^{-1} (−20.0); 2-butanamine, −25.0 (−25.1); cyclohexylamine, −24.9 (−23.6, −25.1 ± 0.3); amphetamine (1-phenyl-2-aminopropane), 7.0 (7.0[8]); 2-amino-1-butanol, −61.0 (−61.5[8]). For cyclopentylamine, the cyclopentane ring is strained by 6.6 kcal mol^{-1} and eq 1 is not applicable.[3] Using the experimental $\Delta_fH[\text{cyclo-}C_5H_{10}] = -13.3$ and adding the destabilization of 5.0 yields $\Delta_fH[\text{cyclo-}C_5H_9NH_2] = -13.3$ kcal mol^{-1} versus experimental −13.1. The gas phase heat of formation of norephedrine (1-phenyl-2-aminopropan-1-ol) is not known. The heat of formation of the precursor hydrocarbon propylbenzene is 2.0 kcal mol^{-1} by eq 1, and we obtain $\Delta_fH[\text{norephedrine}] = -29.0$ by subtracting 36 for stabilization by the primary alcohol (Table 2) and adding 5.0 for destabilization by the NH_2 group on 2° carbon.

For the NH_2 group connected to a 3° carbon, the experimental value of *tert*-butylamine is $\Delta_fH[(CH_3)_3CNH_2] = -28.8 \pm 0.2$ kcal mol^{-1}. The difference from the heat of formation of the precursor hydrocarbon, isobutane (−31.7), is +2.9. For NH_2 connected to a 3° carbon, one more value appears to be available with reported[18] $\Delta_fH[(CH_3)_2CH(NH_2)CH_2OH] = -64.9$ kcal mol^{-1}. From $\Delta_fH[(CH_3)_3CH] = -31.7$, we obtain −64.8 by adding 2.9 for destabilization by the amino group and subtracting 36.0 for stabilization by the 1° alcohol, in excellent agreement.

For the unique methylamine, using experimental values, $\Delta_fH[CH_3NH_2] - \Delta_fH[CH_4] = -5.6 - (-17.9) = 12.3$ as the destabilizing value of the NH_2 group on methane. The destabilizations induced by the NH_2 group connected to methyl, 1, 2, and 3° carbons are 12.3, 8.7, 5.0, and 2.9. A plot of these values versus $23(\Delta\chi)^2$, with $\chi[NH_2] = 3.071$,[11] in Figure 2 gives a straight line as also observed with the alcohols.

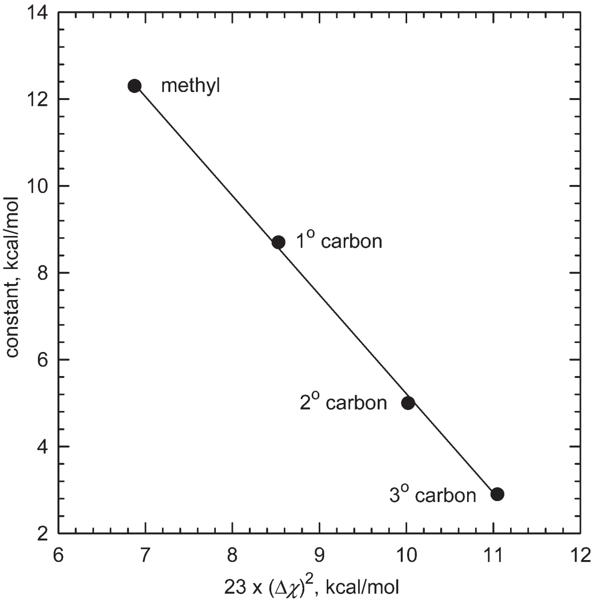

FIGURE 2. Destabilization constants versus $23(\Delta\chi)^2$ of the NH_2 group connected to methyl, 1, 2, and 3° carbon versus $23(\Delta\chi)^2$ of the $C-NH_2$ dipole contribution. Destabilization constant = $-2.281 \times 23(\Delta\chi)^2 + 28.02$; correlation coefficient = 0.99949.

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TABLE 6. Tertiary Amines. Contributions of $\Delta_f H$[hydrocarbon groups] by Equation 1, Calculated $\Delta_f H[R_1R_2R_3N] = \Sigma(\Delta_f H$[hydrocarbon groups]$) + 22.9$, and Literature Values of $\Delta_f H[R_1R_2R_3N]$, kcal mol^{-1} at 298 K

tertiary amine	$\Sigma(\Delta_f H)[R_1R_2R_3]$	calcd $\Delta_f H[R_1R_2R_3N]$	literature[a] $\Delta_f H[R_1R_2R_3N]$
$(CH_3)_3N$	-30.1	-7.2	-5.7 ± 0.2, -7.4, -6.9[b]
$(CH_3)_2NCH_2CH_3$	-35.1	-12.2	-11.6[b]
$CH_3N(CH_2CH_3)_2$	-40.1	-17.2	-17,[c] -19.4[d]
$(CH_3)_2NCH(CH_3)_2$	-41.7	-18.8	-18,[c] -21.5[d]
$(CH_3)_2N(CH_2)_3NH_2$[e]	-31.4	-8.5	-8.3 ± 1, -8.4, -10.7
$(CH_3CH_2)_3N$	-45.0	-22.1	-22.1 ± 0.1, -22.4,[b] -22.1[c]
$(CH_3)_2N(CH_2)_3CH_3$	-45.0	-22.1	-20,[c] -22.6[d]
$(CH_3)_2NCH_2CH(CH_3)_2$	-46.7	-23.8	-24.8,[c] -24.8[d]
$CH_3CH_2CH_2N(CH_2CH_3)_2$	-50.0	-27.1	-27,[c] -30.9[d]
$(CH_3CH_2CH_2)_3N$	-60.0	-37.1	-38.4 ± 0.2, -37.6[b]
$(CH_3CH_2CH_2CH_2)_3N$	-74.9	-52.0	-53 ± 0.3,[c] -55.5[d]
$C_6H_5N(CH_3)_2$[f]	0.7	23.6	24.0 ± 0.8, 21.1, 24,[c] 21.8[d]
$(CH_3)_2NCH_2C_6H_5$	-3.1	19.8	20,[c] 19.4[d]
p-$CH_3C_6H_4N(CH_3)_2$[g]	-7.2	15.7	16.5 ± 1.8, 17,[c] 13.9[d]
m-$CH_3C_6H_4N(CH_3)_2$[h]	-7.2	15.7	17.4 ± 1.7
$C_6H_5N(CH_2CH_3)_2$[i]	-9.2	13.7	14.8 ± 1.8, 9.5,[c] 12.7[d]
$(HOCH_2CH_2)_3N$[j]	-153.0	-130.1	-133.4 ± 0.7, -133.3[d]
cyclo-$C_6H_{11}N(CH_3)_2$	-46.6	-23.7	-25,[c] -25.8[d]
$(CH_3)_2NCH_2COCH_3$[k]	-67.3	-44.4	-43.0 ± 0.7
$((CH_3)_2NCH_2)_2$[l]	-50.1	-4.3	-4.7[c]
$(CH_3)_2NCH_2CO_2CH_3$[m]	-12.2	-90.2	-88.5 ± 0.2
$(CH_3)_2NCH_2CH_2CH_3$	-40.1	-17.2	
$(CH_3CH_2)_2NCH(CH_3)_2$	-51.7	-28.8	
$(CH_2{=}CHCH_2)_3N$	30.1	53.0	
$(C_6H_5CH_2)_3N$	17.0	39.9	

[a] Experimental values from ref 6, unless indicated otherwise. [b] By G3 calculation, ref 2g. [c] From ref 8. [d] By Benson's group additivity method in ref 8. [e] Calculated for *N*,*N*-dimethylpropylamine $\Sigma(\Delta_f H) = -40.1$, plus 8.7 for the destabilization of one NH_2 (Table 6) $= -31.4$. [f] From the experimental $\Delta_f H$[aniline] $= 20.8$ (Table 12) plus the contribution of 6 methyl hydrogens (-3.345×6) yields 0.7. [g] From the calculated $\Delta_f H$[p-toluidine] $= 12.9$ (Table 12) plus the contribution of 6 methyl hydrogens (-3.345×6). [h] From $\Delta_f H$[m-toluidine] $= 12.9$ (Table 12), plus the contribution of six 1° H's (-3.345×6) yields -7.2. [i] From $\Delta_f H$[aniline] $= 20.8$ (Table12) plus the contribution of six 1° H's (-3.345×6) plus the contribution of four 2° H's (-2.490×4) yields -9.2. [j] From the calculated $\Delta_f H$[triethylamine] $= -22.1$ plus the stabilization of three 1° alcohol groups (-36.0×3) yields -153.0. [k] From *N*,*N*-dimethylpropylamine $\Sigma(\Delta_f H) = -40.1$, minus 27.2 for stabilization of the ketone (Table 5) yields -63.7. [l] *N*,*N*,*N*′,*N*′-Tetramethyl-1,2-diaminoethane; $\Sigma(\Delta_f H[R_1R_2R_3])$ from twelve 1° H's and four 2° H's yields -50.1. [m] The precursor is $(CH_3)_2NCH_2{-}CH_3$ minus 78 for insertion of CO_2 to form the methyl ester (Table 10).

Tertiary amines are treated in Table 6. They follow a regular pattern, and their heats of formation are given by the contributions of the three hydrocarbon groups attached to nitrogen plus a destabilization value of 22.9 kcal mol^{-1}, derived from triethylamine. The number of methyl hydrogens are counted as primary, the number of methylene hydrogens as secondary, and the number of methine hydrogens as tertiary. For example, for *N*,*N*-dimethylethylamine, the two methyl groups by eq 1 ($n_1 = 3 + 3$) contribute -20.1 kcal mol^{-1} and the ethyl group ($n_1 = 3$, $n_2 = 2$) contributes -15.0. To their sum is added 22.9 to yield $\Delta_f H[CH_3CH_2N(CH_3)_2] = -12.2$ kcal mol^{-1} versus -11.6 by an ab initio calculation by G3.[2g] Results are provided in Table 6 for compounds for which there are literature values of $\Delta_f H$ and some other estimates. Agreement with available experimental values is within the order of expected experimental uncertainties.

The destabilization constant for 3° amines does not appear to depend on the type of alkyl groups present, unlike the case of alcohols, 1° amines, ethers, and chlorides. There are reasons that lead to this result. (a) Uncertainties in experimental $\Delta_f H$ are often greater in 3° amines, and small variations will not be apparent. (b) The electronegativity of R_2N groups would be substantially lower than $\chi[NH_2] = 3.071$[11] because the electron-donating alkyl groups make the R_2N group more capable of supporting a partial positive charge. This decreases the $23(\Delta\chi)^2$ term for bonds to carbon and decreases the magnitude of the dipole effect. (c) Reliable $\Delta_f H$ values do not appear to be available for *tert*-butyl groups in 3° amines. The dipole effect would be greatest with a *tert*-butyl group and might have been more manifest. Available experimental values in Table 6 predominantly pertain to tertiary amines of methyl and 1° alkyl groups.

Table 6 also incorporates values obtained by other approaches for direct comparison with this work. It provides examples of high level ab initio calculations (G3)[2g] for tertiary amines for values in common with this work (footnote b). The maximum deviation in the four such cases is 0.6 kcal mol^{-1}. Even though global fitting is not made here, the empirical character of this work has similarities with Benson's group additivities because successful schemes map onto each other, as mentioned in the Introduction. Table 6 incorporates some such values (footnote d) for comparison. It is seen that occasionally there are differences in the $\Delta_f H$ values obtained, for example, *N*,*N*-dimethylethylamine, -17.0 (this work), -19.4 (Benson's GA),[19] and -17 (experimental); *N*,*N*-dimethyl-*p*-toluidine, 15.7 (this work), 13.9 (Benson's GA), 16.5 ± 1.8 (experimental). There are disparities with other functional groups, for example, benzophenone in Table 7, 11.9 (this work), 14.6 (Benson's GA), 11.8 (experimental) and glycine in Table 9, -94.9 (this work), -90.9 (Benson's GA), -93.3 ± 1.1 (experimental). The Benson approach does not seem to have been parametrized for some species that are treated successfully here such as acrylamide and benzamide (see below), methyl vinyl ketone, etc.[19]

Ketones. The difference in heats of formation remains constant for a hydrocarbon and a ketone of the same carbon

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(19) Stein, S. E.; Brown, R. L. Structures and Properties Group Additivity Model as implemented in ref 6 under "Group Additivity Based Estimates" (retrieved June 18, 2010).

TABLE 7. Ketones, $\Delta_f H[RR'CH_2]$ by Equation 1, Calculated $\Delta_f H[RR'C{=}O] = \Delta_f H[RR'CH_2] - 27.2$, and Experimental $\Delta_f H[RR'C{=}O]$, kcal mol^{-1} at 298 K

ketone	calcd $\Delta_f H[RR'CH_2]$	calcd $\Delta_f H[RR'C{=}O]$	experimental[a] $\Delta_f H[RR'C{=}O]$
acetone	−25.1	−52.3	−52.2 ± 0.1, −51.9 ± 0.1, −52.0 ± 0.2
butanone	−30.0	−57.2	−57.0 ± 0.2
2-pentanone	−35.0	−62.2	−61.9 ± 0.3
3-pentanone	−30.0	−62.2	−60.6 ± 0.2, −61.7 ± 0.2, −62.3 ± 0.4
2,4-pentanedione	−35.0	−89.4	−91.9 ± 0.3, −89.9 ± 0.5, −90.4[b]
3-methyl-2-butanone	−36.7	−63.9	−62.8 ± 0.2
2-hexanone	−40.0	−67.2	−66.9 ± 0.3
3-hexanone	−40.0	−67.2	−66.5 ± 0.2
2-methyl-3-pentanone	−41.7	−68.9	−68.4 ± 0.2
3-methyl-2-pentanone	−41.7	−68.9	−67.9 ± 0.3
4-methyl-2-pentanone	−41.7	−68.9	−69.6 ± 0.4
4-heptanone	−45.0	−72.2	−71.3 ± 0.3
2,4-dimethyl-3-pentanone	−48.3	−75.5	−74.4 ± 0.3
4,4-dimethyl-2-pentanone	−50.1	−77.3	−76.6 ± 0.5
3-methyl-2-hexanone	−46.6	−73.8	−73[c]
5-methyl-2-hexanone	−46.6	−73.8	−74[c]
2-octanone	−50.0	−77.2	−80.8 ± 0.8, −77.0[c]
4-octanone	−50.0	−77.2	−83.5 ± 0.7, −77.0[c]
acetophenone	7.0	−20.2	−20.7 ± 0.4
4-methylacetophenone	−0.9	−28.1	−28[c]
2-nonanone	−54.9	−82.1	−81.5 ± 0.4
5-nonanone	−54.9	−82.1	−82.4 ± 0.3
2,6-dimethyl-4-heptanone	−58.3	−85.5	−85.5 ± 0.3
1-phenyl-1-propanone	2.0	−25.2	−26[c]
1-phenyl-2-propanone	2.0	−25.2	−22.6, −23.6 ± 0.3[c]
6-undecanone	−64.9	−92.1	−92.6 ± 0.5
2-dodecanone	−69.9	−97.1	−96.6 ± 0.6
benzophenone	39.0	11.8	11.9 ± 0.7
1-phenyl-1,3-butanedione	−3.0	−57.4	−58.3 ± 0.7
dibenzyl ketone	−29.0	1.8	1.9[c]
1-cyclopropylethanone[d]	−1.0	−28.2	−27.6 ± 0.3, −28[c]
$C_6H_5COCH_2COC_6H_5$[e]	29.0	−25.4	−26.1 ± 0.5
cyclododecanone	−59.7	−86.9	−85.9 ± 0.7, −83.4 ± 0.6
1,3-dihydroxyacetone[f]	−25.1	−124.3	−125 ± 1[g]
3-hydroxy-2-butanone[h]	−25.1	−92.5	
4-penten-2-one	−5.0	−32.2	

[a] Experimental $\Delta_f H$ from ref 6, unless indicated otherwise. [b] From ref 20a, keto form. [c] From ref 8. [d] For strained methyl cyclopropyl ketone, $\Delta_f H$(experimental) of the precursor ethylcyclopropane is used. [e] 1,3-Diphenyl-1,3-propanedione. [f] Also −72.0 for two 1° alcohols. [g] $\Delta_f H$(experimental) from ref 20b. [h] Also −40.2 for the 2° alcohol stabilization.

skeleton in the absence of strain or conjugation effects. With acetone as a basis, the difference between the heat of formation of the ketone and that of the precursor hydrocarbon is $\Delta_f H[CH_3C(O)CH_3] - \Delta_f H[CH_3CH_2CH_3] = -52.2 - (-25.0) = -27.2$ kcal mol^{-1} and, in general, $\Delta_f H[RC(O)R'] = \Delta_f H[RCH_2R'] - 27.2$. This is not a unique description of the regularities present: $\Delta_f H[RC(O)R'] = \Delta_f H[R{-}R'] - 32.2$ yields identical results. Table 7 shows the results, where $\Delta_f H[RR'CH_2]$ is obtained by eq 1, unless indicated otherwise.

Unlike the alcohols, there is no difference based on whether the carbonyl group is connected to a methyl, 1°, or 2° carbon. This is due to the small electronegativity difference between the carbonyl carbon and other carbon groups.[11] The calculated $\Delta_f H$ of *tert*-butyl methyl ketone is more negative than the experimental value by 3 kcal mol^{-1}, and this is the strain of the compound. Similarly, the strain is about 9 kcal mol^{-1} in di-*tert*-butyl ketone.

The entries for acetophenone, benzophenone, and 1-phenyl-1,3-butanedione in Table 7 indicate that there is no significant stabilization by conjugation of the α carbonyl with the benzene ring or stabilization by formation of the enol form of the 1,3-dione in the gas phase. The "3d mol" structure of benzophenone[6] indicates that the carbonyl is not coplanar with the rings, and no evidence of conjugation has been found previously in acetophenone.[11] In general, there seems to be a small effect of conjugation with a double bond, as indicated by comparing calculated heats of formation of α,β-unsaturated ketones to experimental values: 3-buten-2-one, calculated $\Delta_f H = -27.2$ kcal mol^{-1} versus experimental −27.4 ± 2.6; 4-methyl-3-penten-2-one, calculated −42.4 versus experimental −42.6 ± 0.2 and −44.1 ± 0.6; 1-penten-3-one, calculated −32.2 versus experimental −33.0[8] and a theoretical value of −32.6;[21] 3-methyl-3-hepten-2-one, calculated −54.8 versus experimental −54;[8] doubly conjugated 1,4-pentadien-3-one, calculated −2.2 versus experimental −0.8;[8] doubly conjugated 2-methylpenta-1,4-dien-3-one, calculated −9.7 versus theoretical −10.5.[21] For the extended conjugation in 4-phenyl-3-buten-2-one, there is substantial conjugation stabilization of −9.2 ± 0.9 shown by calculated $\Delta_f H = -2.3$ versus experimental −11.5 ± 0.9, and this is consistent with the known ease of dehydration at room temperature of the product of aldol condensation of acetone and benzaldehyde, 4-hydroxy-4-phenyl-2-butanone, to give 4-phenyl-3-buten-2-one.

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(20) (a) Temprado, M.; Roux, M. V.; Umnahanant, P.; Zhao, H.; Chickos, J. S. *J. Phys. Chem. B* **2005**, *109*, 12590–12595. (b) Dorofeeva, O. V.; Vogt, N.; Vogt, J.; Popik, M. V.; Rykov, A. N.; Vilkov, L. V. *J. Phys. Chem. A* **2007**, *111*, 6434–6442.

(21) Morales, G.; Martinez, R. *J. Phys. Chem. A* **2009**, *113*, 8683–8703. Ab initio CBS-Q calculation.

The heat of reaction (ΔH_{rxn}) for the base-catalyzed aldol condensation of acetone to diacetone alcohol (4-methyl-2-pentanon-4-ol) is easily calculated. $\Delta_f H$[diacetone alcohol] is not known accurately, with two disparate values available:[6]

TABLE 8. Aldehydes, $\Delta_f H$[RCH$_3$] by Equation 1, Calculated $\Delta_f H$[R$_2$CHO] = $\Delta_f H$[RCH$_3$] − 20.0, and Experimental $\Delta_f H$[RCHO], kcal mol^{-1} at 298 K

aldehyde	calcd $\Delta_f H$[RCH$_3$]	calcd $\Delta_f H$[RCHO]	experimental[a] $\Delta_f H$[RCHO]
acetaldehyde	−20.0	−40.0	−40.8 ± 0.4
propanal	−25.0	−45.0	−45.1 ± 0.2, −44.8 ± 0.4[b]
butanal	−30.0	−50.0	−50.6 ± 0.2, −48.9 ± 0.3
2-methylpropanal	−31.7	−51.7	−51.6 ± 0.4, −51.9[c]
pentanal	−35.0	−55.0	−55.1,[b] −54.6[c]
3-methylbutanal	−36.7	−56.7	−57.0[c]
2,2-dimethylpropanal	−40.1	−60.1	−59.2[c]
hexanal	−40.0	−60.0	−59.0,[b] −59.5[c]
3-methylpentanal	−41.7	−61.7	−62.0[b]
3,3-dimethylbutanal	−45.1	−65.1	−64.0,[b] −64.8[c]
heptanal	−45.0	−65.0	−63.1 ± 1.0
octanal	−50.0	−70.0	−69.8 ± 0.7
2-ethylhexanal	−51.6	−71.6	−71.6
cyclohexanecarbaldehyde	−36.5	−56.5	−56[b]
benzaldehyde	11.9	−8.1	−8.8 ± 0.7, −8.9 ± 1.0, 8.0
phenylacetaldehyde	7.0	−13.0	−13[b]
2-phenylpropanal	2.0	−18.0	−17[b]
terephthaldehyde[d]	4.1	−35.9	−37.6 ± 2
butanedial	−30.0	−70.0	
hexanedial	−40.0	−80.0	
4-methylbenzaldehyde	4.1	−15.9	

[a] Experimental values from ref 6, unless indicated otherwise. [b] From ref 8. [c] From ref 23, MM3 calculation. [d] 1,4-Benzenedicarboxaldehyde.

TABLE 9. Carboxylic Acids, $\Delta_f H$[RCH$_3$] by Equation 1, Calculated $\Delta_f H$[RCO$_2$H] = $\Delta_f H$[RCH$_3$] − 83.5, and Experimental $\Delta_f H$[RCO$_2$H], kcal mol^{-1} at 298 K

carboxylic acid	calcd $\Delta_f H$[RCH$_3$]	calcd $\Delta_f H$[RCO$_2$H]	experimental[a] $\Delta_f H$[[RCO$_2$H]
ethanoic	−20.0	−103.5	−103.5 ± 0.6, −103.3 ± 0.1[b]
propanoic	−25.0	−108.5	−108.9 ± 0.5, −107.0 ± 0.5[b]
butanoic	−30.0	−113.5	−113.7 ± 1, −113.0 ± 1[b]
4-aminobutanoic[c]	−30.0	−104.8	−105.0 ± 0.5
pentanoic	−35.0	−118.5	−117 ± 4, −117.0 ± 0.5[b]
3-methylbutanoic	−36.7	−120.2	−120.1 ± 1.6, −123 ± 1[b]
pivalic acid	−40.1	−123.6	−122[b]
hexanoic	−40.0	−123.5	−122.1 ± 1, −122.8 ± 0.4[b]
6-aminohexanoic[c]	−40.0	−114.8	−115.3 ± 0.7
benzoic	−11.9	−71.6	−70.7 ± 1.0, −70.3[b]
heptanoic	−45.0	−128.5	−128.2, −128.9 ± 0.4
octanoic	−50.0	−133.5	−132.5 ± 0.3, 133.0 ± 0.3
2-ethylhexanoic	−50.0	−133.5	−133.7 ± 0.5
2-phenylacetic	7.0	−76.5	−76[b]
nonanoic	−54.9	−138.4	−137.7 ± 0.2, −138.5 ± 0.2
decanoic	−59.9	−143.4	−149.2 ± 1.2
hexadecanoic	−89.8	−173.3	−174.5 ± 1.3
3-methylbenzoic	4.1	−79.4	−78.4 ± 0.3, −78.7
4-methylbenzoic	4.1	−79.4	−79.0 ± 0.4, −79, −77.6
3,4-dimethylbenzoic	−3.8	−87.3	−86.6 ± 0.4
3,5-dimethylbenzoic	−3.8	−87.3	−87.1 ± 0.4
3-ethylbenzoic	−0.9	−84.4	−82.9 ± 0.4
4-ethylbenzoic	−0.9	−84.4	−85.0 ± 0.5
propanedioic[d]	−25.0	−192.0	−188, −181,−193.9[b]
glycine[e]	−20.0	−94.8	−93.3 ± 1.1
D-alanine[f]	−25.0	−103.5	−113.3, −102.4
(CH$_3$)$_2$CH(OH)CO$_2$H[g]	−31.7	−155.4	−155[b]
(*E*)-H$_2$NC$_7$H$_{12}$CO$_2$H[h]	−43.2	−118.0	−119.5 ± 1.2[h]
cyclopropanecarboxylic[i]	5.5	−78.0	78[b]
3-oxo-butanoic[j]	−30.0	−140.7	
adipic[k]	−40.0	−207.0	
cyclohexanecarboxylic	−36.5	−120.0	
terephthalic	4.1	−162.9	

[a] Experimental values from ref 6, unless indicated otherwise. [b] From ref 8. [c] Also adding 8.7 for the amino destabilization. [d] Malonic acid. The first two experimental values are from reported[6] two different $\Delta_f H$ values of the solid plus the heat of sublimation of the acid. [e] Also adding 8.7 for destabilization of replacing a 1° H by NH$_2$. [f] The precursor is propane and 5.0 is added for replacing a 2° H by NH$_2$ (see above). Two experimental values are reported[6] for the solid, −113.3 and −102.4, plus the heat of sublimation of +31.7. [g] Also subtracting 40.2 for the 2° alcohol function of 2-hydroxy-2-methylpropanoic acid. [h] (*E*)-4-Aminomethylcyclohexanecarboxylic acid. The precursor is (*E*)-1,4-dimethylcyclohexane, $\Delta_f H$ = −43.2 by eq 1, also adding 8.7 for the amino function (Table 5). Experimental value from ref 18b. [i] The precursor is the strained methylcyclopropane, and its experimental $\Delta_f H$ = 5.5 is used. [j] Acetoacetic acid (keto form), also subtracting 27.2 for the ketone stabilization. Our MP2/6-31G(d,p) calculation did not indicate stabilization by intramolecular hydrogen bonding. [k] Hexanedioic.

TABLE 10. Esters. Contribution of Δ_fH[R−R′] by Equation 1, Calculated Δ_fH[RCO_2R'] = Δ_fH[R−R′] + Stabilization,[a] and Experimental Values of Δ_fH[RCO_2R'], kcal mol^{-1} at 298 K

esters	calcd Δ_fH[R−R′]	calcd Δ_fH[RCO_2R']	cxperimental[b] Δ_fH [RCO_2R']
$CH_3CO_2CH_3$	−20.0	−98.0	−98.0
$CH_3CO_2CH_2CH_3$	−25.0	−106.2	−106.5 ± 0.2, 106.3 ± 0.1, 106.1
$CH_3CO_2CH_2CH_2CH_3$	−30.0	−111.2	−109.0[c]
$CH_3CO_2CH_2CH{=}CH_2$	−0.0	−81.2	−83[c]
$CH_3CO_2CH(CH_3)_2$	−31.7	−114.2	−117.0 ± 0.9, −115.1 ± 0.2[c]
$CH_3CO_2(CH_2)_3CH_3$	−35.1	−116.3	−135.3, −116.1[c]
$CH_3CO_2C(CH_3)_3$	−40.1	−123.4	−123.4 ± 0.3
$ClCH_2CO_2CH_3$[d]	−26.9	−104.9	−106.1 ± 2.4, −100[c]
$HOCH_2CO_2CH_3$[e]	−56.0	−134.0	−133.1 ± 1.5
$CH_3CO_2CH_2CH_2OCH_3$[f]	−56.7	−137.9	−136.9 ± 2.1
$(CH_3)_2NCH_2CO_2CH_3$[g]	−12.2	−90.2	−88.5 ± 0.2
$(CH_3)_2NCH_2CO_2CH_2CH_3$[h]	−17.2	−98.4	−97.3 ± 0.3
$CH_3CH_2CO_2CH_2CH_3$	−30.0	−111.2	−111.5 ± 0.1, 110.8 ± 0.2, 111[c]
$CH_3(CH_2)_2CO_2CH_3$	−30.0	−108.0	−112.7 ± 0.4, −116.0, −108[c]
$CH_3(CH_2)_2CO_2CH(CH_3)CH_2CH_3$	−46.4	−129.1	−130.3 ± 1.0[i]
$CH_3(CH_2)_3CO_2CH_3$	−35.0	−113.0	−112.7 ± 0.4, −116.0 ± 0.4
$CH_2{=}CH(CH_2)_2CO_2CH_2CH_3$	−10.0	−91.2	−92.1 ± 0.6
$CH_2{=}CH(CH_2)_2CO_2CH(CH_3)_2$	−16.6	−99.1	−99.7 ± 0.5
$CH_3CH{=}CHCH_2CO_2CH_2CH_3$[j]	−12.6	−93.8	−93.2 ± 0.7
$CH_3CH{=}CHCH_2CO_2CH(CH_3)_2$[j]	−19.2	−101.7	−101.6 ± 0.6
$CH_3CH_2CH(CH_3)CO_2CH_2CH_3$	−41.7	−122.9	−123.4 ± 0.3, −124.9 ± 2.1
$(CH_3)_3CCO_2CH_3$	−40.1	−118.1	−118.2, −122.8 ± 1.8, −117.0[c]
$(CH_3)_3CCO_2CH_2CH_3$	−45.1	−126.3	−125.6 ± 0.3, −128.1 ± 2.0[i]
$CH_3(CH_2)_4CO_2CH_3$	−40.0	−118.0	−118.0 ± 0.4
$CH_3(CH_2)_5CO_2CH_2CH_3$	−50.0	−131.2	−131.0 ± 0.5
$C_6H_5CO_2CH_3$	11.9	−66.1	−64.4 ± 1.2, 71.7, 69 ± 2[c]
p-$CH_3C_6H_4CO_2CH_3$	4.1	−73.9	−74.7 ± 1.6, −77[c]
$CH_3(CH_2)_8CO_2CH_2CH_3$	−64.9	−146.1	−146.3 ± 0.4
$CH_3C(O)CH_2CO_2CH_2CH_3$[k]	−62.2	−143.4	−140.1 ± 0.4, −133.4, −121.3
$CH_3CH_2OC(O)CH_2CO_2CH_2CH_3$[l]	−35.0	−197.4	−190.1, −220.3 ± 1.0
$CH_3CH_2CO_2(CH_2)_2CH_3$	−35.0	−116.2	−116.5[m]
$CH_3CH_2CO_2(CH_2)_3CH_3$	−40.0	−121.2	−123.4[n]
$CH_3OC(O)C_6H_4CO_2CH_3$[o]	4.1	−151.9	−144.7,[o] −150.1 ± 0.2[o]
$CH_3CO_2C_6H_5$	11.9	−66.1	−66.8 ± 0.3, −67.0 ± 0.5
$C_6H_5CO_2C_6H_5$	44.0	−34.0	34.0 ± 0.6, 36.3 ± 1.1, 34 ± 0.7[c]
$CH_3CO_2CH{=}CH_2$[p]	5.0	−73.0	−75.0, −73.8 ± 2.4
$CH_3CO_2CH_2C_6H_5$	7.0	−74.2	
$CH_3OC(O)(CH_2)_4CO_2CH_3$	−40.0	−196.0	
$C_6H_5CO_2CH_2C_6H_5$	39.0	−42.2	

[a] Stabilization = −78.0 kcal mol^{-1} for esters of methanol, phenol, and vinyl alcohol; −81.2 for esters of 1° alcohols; −82.5 for esters of 2° alcohols; and −83.3 for esters of 3° alcohols. [b] Experimental values from ref 6, unless indicated otherwise. [c] From ref 8. [d] The precursor is $ClCH_2-CH_3$ with Δ_fH[R−R′] = −26.9 from Table 11. [e] The precursor is $HOCH_2-CH_3$ with Δ_fH[R−R′] = −56.0 from Table 2. [f] The precursor is $CH_3-CH_2CH_2OCH_3$ with Δ_fH [R−R′] = −56.7 (see text). [g] The precursor is $(CH_3)_2NCH_2-CH_3$ with Δ_fH[R−R′] = −12.2 calculated for *N*,*N*-dimethylethylamine in Table 6. [h] The precursor is $(CH_3)_2NCH_2-CH_2CH_3$ with Δ_fH[R−R′] = −17.2 calculated for *N*,*N*-dimethylpropylamine in Table 6. [i] From ref 4f. [j] The experimental and calculated values are for the *trans*-isomer. [k] Acetoacetic ester (keto form) and the precursor is $CH_3C(O)CH_2-CH_2CH_3$ with Δ_fH[2-pentanone] = −62.2 from Table 7. [l] Malonic ester and insertion of two CO_2 units into $CH_3CH_2-CH_2-CH_2CH_3$. [m] Δ_fH[RCO_2R'] of the liquid is reported[6] as −126.9, the heat of vaporization is 10.4, and we estimate their sum of −116.5 for the gas phase. [n] Δ_fH[RCO_2R'] of the liquid is estimated[6] as −132.7, the heat of vaporization is 9.3, and we estimate their sum of −123.4 for the gas phase. [o] Δ_fH[R−R′] is calculated by eq 1 for $CH_3-C_6H_4-CH_3$. Δ_fH of solid dimethyl terephthalate is reported[6] as −169.7 or −175.1 ± 0.2 to which we add the heat of sublimation of 25.0. [p] Vinyl esters appear to have the same stabilization of −78.0 as phenyl esters.

−116.2 kcal mol^{-1} by the semiempirical PM3 calculation and −129.1 kcal mol^{-1} by the static bomb calorimeter technique. Its precursor hydrocarbon is 2-methylpentane, $\Delta_fH = -41.7$ by eq 1. Subtracting the stabilization of the ketone (−27.2) and of the 3° alcohol (−43.0), we obtain Δ_fH[diacetone alcohol] = −111.9 kcal mol^{-1}, with the known Δ_fH[acetone] = −52.2, $\Delta H_{rxn} = -111.9 - 2(-52.2) = -7.5$ kcal mol^{-1} for the gas phase reaction and neglecting possible intramolecular H-bonding. A liquid phase value of $\Delta H_{rxn} = -8.2$ has been reported (no solvent, insoluble $Ba(OH)_2$ catalyst),[22] which may indicate some small effect of H-bonding.

For 2,3-butanedione, there is a large discrepancy between the reported experimental heat of formation and the one calculated by our approach, −78.1 versus −84.3 kcal mol^{-1}, respectively. The compound is less stable than calculated, which indicates not only lack of conjugation stabilization by its alternating double-single-double bonds with parallel p orbitals, but destabilization (see below). We also performed a G3 model chemistry calculation and obtained −79.1 kcal mol^{-1} for the most stable *s-trans* conformer of 2,3-butanedione, thus confirming the experimental value. Benzil (dibenzoyl), another α-diketone, also shows destabilization by 7 kcal mol^{-1}, $\Delta_fH = -13.0 \pm 0.7$ experimental versus −20.4 calculated.

Aldehydes. The difference in heats of formation remains constant for a hydrocarbon and an aldehyde of the same carbon skeleton in the absence of strain or resonance effects.

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[22] Davis, G. L.; Burrows, G. H. *J. Am. Chem. Soc.* **1936**, *58*, 311–312.

[23] Allinger, N. L.; Chen, K.; Rahman, M.; Pathiaseril, A. *J. Am. Chem. Soc.* **1991**, *113*, 4505–4517. MM3 calculation.

This is demonstrated in Table 8, where $\Delta_f H[RCHO] = \Delta_f H[RCH_3] - 20.0$ kcal mol^{-1} for propanal. $\Delta_f H[RCH_3]$ is obtained by eq 1, unless indicated otherwise.

For α,*β*-unsaturated aldehydes, there is greater conjugation stabilization than for α,*β*-unsaturated ketones. For acrolein, calculated $\Delta_f H[CH_2{=}CHCHO] = -14.5$ kcal mol^{-1} versus experimental −18.0, which indicates conjugation stabilization of −3.5 kcal mol^{-1}. For $\Delta_f H$[*trans*-2-butenal], the calculated value is −22.4 kcal mol^{-1} versus −25.7 experimental,[8] for stabilization of −3.2. For $\Delta_f H$[*trans*-2-pentenal], the calculated value is −27.2 kcal mol^{-1} versus −30.0 experimental[8] for stabilization of −2.8. For benzaldehyde (Table 8), the calculated and experimental values are in agreement, indicating no significant conjugation stabilization with the ring.

With ethanedial (glyoxal), the hydrocarbon precursor is ethane and calculated $\Delta_f H[HCOCHO] = -20.0 - 2 \times (20.0) = -60.0$ kcal mol^{-1} is substantially more negative than the experimental value of −50.7, as was found with 2,3-butanedione. It appears that repulsive interactions between the two adjacent carbons of the carbonyls decreases the stability of α-dicarbonyls by 6.2 kcal mol^{-1} in 2,3-butanedione and by 9.3 kcal mol^{-1} in ethanedial due to the substantial partial positive charge on carbonyl carbons. Again, for 2-oxopropanal with precursor propane, the calculated $\Delta_f H[CH_3COCHO] = -25 - 27.2 - 20 = -72.2$ is more negative than the experimental −64.8 ± 1.2, for a difference of about 7.4 kcal mol^{-1}. In aldol condensations, aldehydes are more reactive than ketones because the second electron-donating alkyl group in the latter diminishes the partial positive charge of their carbonyl carbon. Thus, ethanedial would be expected to have the largest such destabilizing effect, 2,3-butanedione the smallest, and 2-oxopropanal to be in between. In fact, this is what the calculated values demonstrate quantitatively. The unusually weak $C(sp^2){-}C(sp^2)$ central bonds of 2,3-butanedione and of glyoxal, 72.9 and 71.5 kcal mol^{-1}, respectively,[6] also demonstrate this effect.

Carboxylic Acids. Regularities of carboxylic acids are shown in Table 9. On the basis of the experimental value of acetic acid, $\Delta_f H[CH_3CO_2H] = -103.5 \pm 0.6$ kcal mol^{-1} and $\Delta_f H[CH_3CH_3] = -20.0$ by eq 1, the stabilization energy of replacing CH_3 by CO_2H is −83.5 ± 0.6. In general, $\Delta_f H[RCO_2H] = \Delta_f H[RCH_3] - 83.5$. Again, this is not a unique description of the regularities: the same results are obtained by $\Delta_f H[RCO_2H] = \Delta_f H[RH] - 88.5$.

Benzoic acids fit the same pattern of other carboxylic acids, which indicates that there is no significant conjugation stabilization of the carbonyl with the ring as also seen with

TABLE 11. Chlorides. $\Delta_f H$[hydrocarbon] by Equation 1, Calculated $\Delta_f H[RCl] = \Delta_f H$[hydrocarbon] + Stabilization,[a] and Experimental $\Delta_f H[RCl]$, kcal mol^{-1} at 298 K

chloride	calcd $\Delta_f H[RH]$	calcd $\Delta_f H[RCl]$	experimental[b] $\Delta_f H[[RCl]$
CH_3CH_2Cl	−20.0	−26.9	−26 ± 2, −26.8 ± 0.1[c]
$CH_3CH_2CH_2Cl$	−25.0	−31.9	−31.7 ± 0.2, 31.4 ± 0.2
$ClCH_2CH_2CH_2Cl$	−25.0	−38.8	−38 ± 2[c]
$CH_3(CH_2)_2CH_2Cl$	−30.0	−36.9	−37.0 ± 0.3, −36.9 ± 0.2[c]
$ClCH_2CH_2CH_2CH_2Cl$	−30.0	−43.8	−43.8 ± 0.5
$CH_3(CH_2)_3CH_2Cl$	−35.0	−41.9	−41.9 ± 0.3
$CH_3(CH_2)_6CH_2Cl$	−50.0	−56.9	−57.1 ± 0.5
$CH_3(CH_2)_{10}CH_2Cl$	−69.9	−76.8	−77.0 ± 0.6
$CH_3(CH_2)_{16}CH_2Cl$	−99.8	−106.7	−106.6 ± 0.7
$CH_2{=}CHCH_2Cl$	5.0	−1.9	−1.3, −1.3 ± 0.6[c]
$CH_2{=}CHCH_2CH_2Cl$	0.0	−6.9	−6.9
trans-$CH_3CH{=}CHCH_2Cl$	−2.6	−9.5	−10.3
$C_6H_5CH_2Cl$	11.9	5.0	4.5 ± 0.8
$(CH_3)_2CHCl$	−25.0	−34.8	−34 ± 1
$CH_3CH_2CH(CH_3)Cl$	−30.0	−39.8	−39.8 ± 0.2
$CH_3CHClCH_2CH_2Cl$	−30.0	−46.7	−46.6 ± 0.5
cyclo-$C_6H_{11}Cl$	−29.9	−39.7	−39.8 ± 0.5, −39.2 ± 0.2
$(CH_3)_3CCl$	−31.7	−43.0	−43.0
$CH_3CH_2C(CH_3)_2Cl$	−36.7	−48.0	−48.2[d]
$HOCH_2CH_2Cl$[e]	−56.0	−62.9	−62[c]
CH_3OCH_2Cl[f]	−44.0	−50.9	−51[c]
$ClCH_2CHO$[g]	−40.0	−46.9	−47 ± 4[c]
C_6H_5Cl	19.8	13.0	13.0, 13.0 ± 0.2
p-ClC_6H_4Cl	19.8	6.1	5.9, 5.9[c]
m-ClC_6H_4Cl	19.8	6.1	6.7
1,3,5-$C_6H_3Cl_3$	19.8	−0.8	−0.6 ± 0.3, −3.2, 0.0[c]
p-$CH_3C_6H_4Cl$	11.9	5.1	4.0[c]
p-$ClC_6H_4CO_2H$[h]	−71.6	−78.4	−76.6 ± 0.2
p-ClC_6H_4CHO[i]	−8.1	−14.9	−16[c]
2,2′-$C_6H_5C_6H_5Cl$[j]	44.0	30.2	30.6 ± 1.1
$CH_2{=}CHCl$[k]	12.5	5.7	5.3 ± 0.7, 7.0, 5.0, 8.4 ± 0.3
$HC{\equiv}CCH_2Cl$	44.4	37.5	
trans-1,4-$ClC_6H_{10}Cl$[l]	−29.9	−49.5	
p-$CH_3C_6H_4CH_2Cl$	4.1	−2.8	
$C_6H_5C(CH_3)_2Cl$	0.3	−11.0	

[a] Stabilization = −6.9 for 1° alkyl and aryl chlorides; −9.8 for 2° alkyl chlorides; and −11.3 for 3° alkyl chlorides. [b] Experimental values from ref 6, unless indicated otherwise. [c] From ref 8. [d] Obtained from the liquid $\Delta_f H = -56.2$ plus the heat of vaporization, $\Delta H_{vap} = 8.0$. [e] Precursor is ethanol, $\Delta_f H = -56.0$. [f] Precursor is dimethyl ether, $\Delta_f H = -44.0$. [g] Precursor is acetaldehyde, $\Delta_f H = -40.0$. [h] Precursor is benzoic acid (Table 9). [i] Precursor is benzaldehyde (Table 8). [j] 2,2′-Dichlorobiphenyl. [k] Precursor is ethylene, and the same stabilization of −6.9 of aryl appears applicable. [l] *trans*-1,4-Dichlorocyclohexane (both chlorines equatorial).

acetophenone and benzaldehyde. Some very small conjugation stabilization may be present in α,β-unsaturated carboxylic acids. For acrylic acid, calculated $\Delta_fH[CH_2{=}CHCO_2H] = -78.5$ kcal mol^{-1} versus experimental −79.0 ± 1, −80.5 ± 0.6, and −77.6; calculated for *trans*-2-butenoic acid −86.1 kcal mol^{-1} versus experimental −88.1 ± 0.3; calculated for *trans*-2-pentenoic acid −91.1 versus experimental −90 ± 2;[8] calculated for *trans*-cinnamic acid $\Delta_fH[C_6H_5\text{-}CH{=}CHCO_2H] = -54.0$ versus experimental −54.[8] Destabilization due to repulsion is also evident in the α-dicarbonyls of oxalic acid with calculated $\Delta_fH[HOC(O)CO_2H] = -187.0$ kcal mol^{-1} versus one reported experimental value of −175.0 ± 0.7. Pyruvic acid also shows repulsion between the α-dicarbonyl groups with calculated $\Delta_fH[CH_3C(O)\text{-}CO_2H] = -135.7$ versus experimental[8] −131.0 kcal mol^{-1}.

Esters. For esters, there are several published heats of formation in disagreement with each other, and the approach used here can help establish the more reliable values. Heats of formation of esters are obtained by calculating the effect of inserting a CO_2 group into a hydrocarbon R−R′ to form R−CO_2−R′: $\Delta_fH[RCO_2R'] = \Delta_fH[R{-}R'] - 78.0$ kcal mol^{-1} for methyl esters, $\Delta_fH[R{-}R'] - 81.2$ for esters of 1° alcohols, $\Delta_fH[R{-}R'] - 82.5$ for esters of 2° alcohols, and $\Delta_fH[R{-}R'] - 83.3$ for esters of 3° alcohols. Each of these stabilization values was obtained from a reliable heat of formation of a single ester: methyl acetate, ethyl acetate, isopropyl pentanoate, and *tert*-butyl acetate. Insertion of the CO_2 unit into the hydrocarbon R−R′ leads to the above different stabilizations of the O−R′ bond, as discussed above for a similar effect on alcohols, ethers, etc., in the context of Figures 1 and 2. The precursor R−R′ is stabilized by −78.0 kcal mol^{-1} in phenyl and vinyl esters, as found for methyl esters. The success in matching most of the many cases presented in Table 10 indicates that reported values in significant disagreement with those calculated here are suspect.

TABLE 12. Phenols, Anilines, and Aromatic Ethers. $\Delta_fH[Ar{-}H]$ by Equation 1 or from Previous Tables. Calculated $\Delta_fH[Ar{-}OH] = \Delta_fH[Ar{-}H] - 42.8$. $\Delta_fH[Ar{-}NH_2] = \Delta_fH[Ar{-}H] + 1.0$; $\Delta_fH[Ar{-}O{-}R] = \Delta_fH[Ar{-}R] - 16.0$, and Experimental Values, kcal mol^{-1} at 298 K

aryl-X	calcd $\Delta_fH[Ar{-}H]$	calcd $\Delta_fH[Ar{-}X]^b$	experimental^a $\Delta_fH[ArX]^b$
phenol	19.8	−23.0	−23.0 ± 0.1
p-cresol	11.9	−30.9	−30.0 ± 0.4, −29.9[c]
m-cresol	11.9	−30.9	−31.9 ± 0.3, −31.6 ± 0.3, −28.0
o-cresol	11.9	−30.9	−30.7 ± 0.2, −30.0[c]
hydroquinone	19.8	−65.8	−66.2 ± 0.3, −65.0, −64.3
resorcinol	19.8	−65.8	−68.0 ± 0.3, −65.7, −63.4
pyrocatechol	19.8	−65.8	−65.7 ± 0.3, −63.9 ± 0.5, −64.9
3,5-dimethylphenol	4.1	−38.7	−38.6 ± 0.2
p-phenylphenol[d]	43.9	1.1	0.0, 8.5
3,5-dichlorophenol[e]	6.0	−36.6	−35.4 ± 0.4
2,4-dichlorophenol[e]	6.0	−36.6	−37.4 ± 0.5
p-*tert*-butylphenol[f]	−5.4	−48.2	−44.5, −51.3, −41.7
p-hydroxybenzoic acid[g]	−71.6	−114.4	−117.7 ± 0.5, −114.8
vinyl alcohol[h]	12.5	−30.3	−30.6 ± 2.0
aniline	19.8	20.8	20.8 ± 0.2, 19.7, 19.9
p-toluidine[i]	12.0	13.0	10.0, 14.0, 13[c]
m-toluidine[i]	12.0	13.0	14.6, 21.0, 13.0 ± 0.4[c]
o-toluidine[i]	12.0	13.0	12.7, 13.1, 12.3, 13.0 ± 0.2[c]
o-phenylaniline[d]	43.9	44.9	44.8
p-chloroaniline[j]	13.0	14.0	13.7, 13[c]
2,6-dichloroaniline[e]	6.0	7.2	7[c]
p-aminobenzoic acid[g]	−71.6	−70.6	−70.2 ± 1.0
m-aminobenzoic acid[g]	−71.6	−70.6	−69.2 ± 0.9
o-animobenzoic acid[g]	−71.6	−70.6	−70.8 ± 0.4
p-aminophenol[g]	−23.0	−22.0	−21.6 ± 0.3, −19.5 ± 0.4
m-aminophenol[g]	19.8	−22.0	−23.6 ± 0.4, −21.4 ± 0.4
vinylamine[k]	12.5	13.5	14[c]
anisole[l]	12.0	−16.0	−16.2 ± 0.3
ethyl phenyl ether[m]	7.0	−23.7	−24.3 ± 0.3
methyl vinyl ether[n]	5.0	−23.0	−24 ± 2[c]
divinyl ether[o]	30.0	−2.0	−3.0 ± 0.2
phenyl vinyl ether[p]	37.0	5.0	5.5 ± 0.5
diphenyl ether[d,q]	43.9	12.0	12.3 ± 0.6
p-hydroxymethylphenol[r]	−30.9	−66.9	
p-hydroxybenzaldehyde[s]	−8.1	−50.9	
p-aminobenzaldehyde[s]	−8.1	−7.1	
1,4-benzenediamine	19.8	21.8	

[a] Experimental values from ref 6, unless indicated otherwise. [b] X is either OH, NH_2, or OR. [c] From ref 8. [d] The precursor is biphenyl and $\Delta_fH[C_6H_5C_6H_5] = 43.9$ is calculated by eq 1 ($n_8 = 2$, $n_9 = 2$) because the rings are not coplanar and there is little conjugation stabilization (experimental $\Delta_fH = 43.1 \pm 0.7^6$ and 43.6 ± 0.3^8). [e] The precursor is *m*-dichlorobenze (Table 11). [f] *tert*-Butylbenzene is strained, and its experimental value of Δ_fH is used. [g] The precursor is benzoic acid (Table 9). [h] The precursor is ethylene, and the same stabilization of −42.8 of phenol appears applicable. [i] The precursor is toluene. [j] The precursor is chlorobenzene. [k] The precursor is ethylene, and the same destabilization of +1.0 of aniline appears applicable. [l] The precursor is toluene, and 12.0 is subtracted for the CH_3−O ether function (see treatment of alkyl ethers) and 16.0 is subtracted for the O−Ar function. [m] The precursor is ethylbenzene, and 14.7 is subtracted for the CH_3CH_2−O ether function and 16.0 for O−Ar. [n] The precursor is propylene, and 12.0 is subtracted for the CH_3−O function and 16.0 for the O−CH=CH_2 function. [o] The precursor is 1,3-butadiene by eq 1 (without conjugation stabilization) minus twice 16.0. [p] The precursor is styrene by eq 1 (without conjugation stabilization), and 16.0 is subtracted twice. [q] From biphenyl, 16.0 is subtracted twice. [r] The precursor is *p*-cresol (this Table) minus 36.0 for the 1 alcohol function. [s] The precursor is *p*-methylphenol (this Table) minus 36.0 for the 1° alcohol function. [t] The precursor is benzaldehyde (Table 8).

Dimethyl oxalate (dimethyl ethanedioate) also shows repulsion of 6.5 kcal mol^{-1} between the carbons of the adjacent carbonyl functions. Our calculated value is $\Delta_f H = -176.0$ kcal mol^{-1} versus experimental -169.5 ± 0.1. α,β-Unsaturated esters may show a small extent of conjugation of about 1 kcal mol^{-1}, with experimental values often showing somewhat greater stability than calculated by assuming no special effect of conjugation stabilization. For example, for methyl methacrylate calculated $\Delta_f H[CH_2{=}C(CH_3)CO_2CH_3] = -80.6$ versus experimental -83.3, -81.8, and -79.1.

Chlorides are treated in Table 11. Replacement of a hydrogen atom by chlorine stabilizes the precursor molecule by -6.9 kcal mol^{-1} for 1°, -9.8 for 2°, and -11.3 for 3° alkyl chlorides. Replacement of an aromatic hydrogen stabilizes the precursor by -6.9, as with 1° alkyl. When two chlorines are bonded to the same carbon, the stabilization is not additive but is somewhat reduced, as expected; *vic*-dichlorides show the same effect.

Phenols, Anilines, and Aryl Ethers. These compounds are subject to a significant special effect of resonance by strong electron donation by oxygen and nitrogen into the ring, as is evident from their high reactivity in electrophilic aromatic substitutions. Regularities of compounds having OH, NH_2, or OR connected to benzene rings are shown in Table 12 for some available experimental values and for some estimates of unknown ones. Replacing an aromatic H with OH stabilizes the precursor Ar$-$H by -42.8 kcal mol^{-1} (based on phenol) and with NH_2 destabilizes by $+1.0$ (based on aniline), for phenyl and vinyl ethers formed by insertion of O into Ar$-$R, $\Delta_f H[ArOR] = \Delta_f H[Ar{-}R] - 16.0$ for each Ar$-$O linkage.

For salicylic acid, assuming no special effect of intramolecular hydrogen bonding, the calculated $\Delta_f H[o\text{-}HOC_6H_4\text{-}CO_2H]$ is -114.4 kcal mol^{-1} versus experimental -118.5 ± 0.3, -117.9, and -116.9, indicating stabilization of about 3 to 4 kcal mol^{-1} by internal hydrogen bonding.

Nitriles. For nitriles of known experimental heats of formation, $\Delta_f H[R{-}C{\equiv}N] = \Delta_f H[R{-}CH_3] + 37.5$ kcal mol^{-1}. Calculating $\Delta_f H[R{-}CH_3]$ by eq 1 yields results that are accurate to within the uncertainty of available experimental values for acetonitrile, propanenitrile, allyl cyanide, butanenitrile, 3-butenenitrile, 2-methylpropanenitrile, pentanenitrile, 3-pentenenitrile, 2,2-dimethylpropanenitrile, 2-methylbutanenitrile, 1,6-hexanedinitrile, heptanenitrile, octanenitrile, decanenitrile, and undecanenitrile. By also taking account of the carbonyl group stabilization, $\Delta_f H$[α-cyanoacetophenone] and $\Delta_f H$[β-cyanopropiophenone] are calculated as 17.3 and 7.3 kcal mol^{-1}, respectively, versus experimental 16.8 and 7.2 ± 0.1. In acrylonitrile, there does not appear to be any significant thermodynamic stabilization by conjugation because the calculated $\Delta_f H[CH_2{=}CHC{\equiv}N] = 42.5$ kcal mol^{-1} matches experimental values of 41.3 and 43.0.

Amides. For amides, $\Delta_f H[R{-}C(O)NH_2] = \Delta_f H[R{-}CH_3] - 36.5$ kcal mol^{-1} with $\Delta_f H[R{-}CH_3]$ calculated by eq 1. Accurate results are obtained for acetamide, propanamide, butanamide, 2-methylpropanamide, and 2,2-dimethylpropanamide. For acrylamide, methacrylamide, benzamide, and terephthalamide, the calculated and experimental values agree and indicate no significant thermodynamic stabilization by conjugation.

TABLE 13. Summary of Functional Group Stabilization/Destabilization Effects: Functional Group (RX), Precursor Hydrocarbon and Effect (RH + E), kcal mol^{-1} at 298 K

RX	RH + E
$RCH_2{-}OH$[a]	$RCH_3 - 36.0$
$RR'CH{-}OH$[b]	$RR'CH_2 - 40.2$
$RR'R''C{-}OH$[b]	$RR'R''CH - 43.0$
$RCH_2{-}OOH$[a]	$RCH_3 - 18.3$
$RR'CH{-}OOH$[b]	$RR'CH_2 - 22.3$
$RR'R''C{-}OOH$[b]	$RR'R''CH - 25.3$
$RO{-}CH_3$[c]	$RCH_3 - 12.0$
$RO{-}CH_2R'$[d]	$RCH_2R' - 14.7$
$RO{-}CHR'R''$[d]	$RCHR'R'' - 16.4$
$RO{-}CR'R''R'''$[d]	$RCR'R''R''' - 16.7$
$RCH_2{-}NH_2$[a]	$RCH_3 + 8.7$
$RR'CH{-}NH_2$	$RR'CH_2 + 5.0$
$RR'R''C{-}NH_2$[b]	$RR'R''CH + 2.9$
$RR'R''N$	$\Sigma(\Delta_f H[R_1R_2R_3]) + 22.9$
$RR'C{=}O$[b]	$RR'CH_2 - 27.2$
$RCH{=}O$[c]	$RCH_3 - 20.0$
$R{-}CO_2H$[a]	$RCH_3 - 83.5$
$R{-}CO_2{-}CH_3$[a,e]	$RCH_3 - 78.0$
$R{-}CO_2{-}CH_2R'$[b]	$RR'CH_2 - 81.2$
$R{-}CO_2{-}CHRR'$	$RCHRR' - 82.5$
$R{-}CO_2{-}\ CR'R''R'''$[b]	$RCR'R''R''' - 83.3$
$RCH_2{-}Cl$[a]	$RCH_3 - 6.9$
$RR'CH{-}Cl$[b]	$RR'CH_2 - 9.8$
$RR'R''C{-}Cl$[b]	$RR'R''CH - 11.3$
$Ar{-}Cl$	$Ar{-}H\ {-6.9}$
$Ar{-}OH$[f]	$ArH - 42.8$
$Ar{-}NH_2$[f]	$ArH + 1.0$
$Ar{-}OR$[d]	$ArR - 16.0$
$R{-}CN$	$RCH_3 + 37.5$
$RC(O)NH_2$[a]	$RCH_3 - 36.5$

[a] Also R = Ar or vinyl. [b] Also R = Ar. [c] Also R and R′ = Ar. [d] For each of such ether linkage. [e] Also esters of phenol and vinyl alcohol. [f] Also Ar = vinyl.

Summary of Constants. A summary of the single value applicable to each functional group's stabilization or destabilization is given in Table 13. The values given are additive for multifunctional compounds that are free of special effects.

The same approach of relating heats of formation of other classes of organic compounds to those of their hydrocarbon precursors should be applicable to functional groups not treated here, such as thiols, bromides, peroxides, etc.

Conclusions

The regularity of $\Delta_f H$ of organic compounds enables calculation of their values by a simple procedure for many functional groups. For compounds free of strain or resonance effects, $\Delta_f H$ of the hydrocarbon precursor is obtained by counting hydrogens as previously described and then adding a constant that reflects the effect on $\Delta_f H$ of introducing the functional group(s). This provides accurate estimates of experimentally unavailable values and selection of the most likely accurate value among disparate experimental reports. The contribution of functional groups X to $\Delta_f H$ correlates well with Pauling's $23(\Delta\chi)^2$ dipole contribution to BDE[R$-$X]. The approach used should be applicable to functional groups not treated here.

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