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Thermochemistry of Ethyl 3-Oxobutanoate Revisited: Observance of a Non-Zero Enthalpy of Mixing between Tautomers and Its Effects on Enthalpies of Formation

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The enthalpies of formation of pure liquid and gas-phase ethyl 3-oxobutanoate and ethyl *Z*-3-hydroxy-2-butenoate are examined in the light of some recent NMR studies on the enthalpy differences between gas-phase enthalpies of the two tautomers. Correlation-gas chromatography experiments are used to evaluate the vaporization enthalpies of the pure tautomers. Values of (61.6 ± 2.2) and (54.7 ± 1.3) kJ·mol^{-1} at T = 298.15 K are measured for pure ethyl 3-oxobutanoate and ethyl *Z*-3-hydroxy-2-butenoate, respectively. These values can be contrasted to a value of (54.2 ± 1.0) kJ·mol^{-1} when the vaporization enthalpy is measured as a mixture of tautomers. The difference is attributed to an endothermic enthalpy of mixing term of (7.1 ± 2.1) kJ·mol^{-1}, which destabilizes the mixture relative to the pure components. Using the existing liquid-phase enthalpy differences measured by NMR, the endothermic enthalpy of mixing term and the condensed phase enthalpy of formation of the equilibrium mixture, new enthalpies of formation for ethyl 3-oxobutanoate and ethyl *Z*-3-hydroxy-2-butenoate in both the gas, $\Delta_f H^\circ_m$(g) = (−578.2 ± 3.3), (−593.0 ± 2.7) kJ·mol^{-1}, respectively, and liquid phases, $\Delta_f H^\circ_m$(l) = (−639.8 ± 2.4), (−647.7 ± 2.4) kJ·mol^{-1}, respectively, are calculated. A comparison of the difference in gas phase enthalpy of formation of the two tautomers reproduces the enthalpy difference measured previously by gas-phase NMR studies.

Introduction

Similar to 1,3-dicarbonyl compounds, ethyl acetoacetate can act as a bidentate ligand and form complexes with metal ions.[1] Its enthalpy of formation is a crucial property in the evaluation of the enthalpies of formation of these complexes. Unlike most organic compounds that exist predominately in one form, two tautomeric forms, ethyl 3-oxobutanoate and ethyl *Z*-3-hydroxy-2-butenoate, coexist with the former predominating at equilibrium. As with 2,4-pentanedione, the presence of this equilibrium complicates the evaluation of the thermodynamic properties of the two pure constituents.[2] This paper examines some of these properties in light of some recent findings in the 2,4-pentanedione tautomeric system and reports some additional experimental results that elucidate the thermochemistry of the processes involved in this equilibrium. To avoid any confusion, the term ethyl acetoacetate will be used when referring to the equilibrium mixture, and the terms ethyl 3-oxobutanoate and ethyl *Z*-3-hydroxy-2-butenoate will be used to refer to each respective tautomer.

The enthalpy of formation of ethyl acetoacetate in the liquid phase, $\Delta_f H^\circ_m$(l, 298.15 K) = (−640.4 ± 1.1) kJ·mol^{-1}, has been reported recently.[1] The measurement is a redetermination of a value reported by Kharasch some time ago and is some 13 kJ·mol^{-1} more negative than previously reported.[3]

Several values for the enthalpy of vaporization of ethyl acetoacetate have also been reported.[1,4,5] These values are summarized in Table 1. Good agreement is found for most of the measurement if the experimental data are adjusted to T = 298.15 K. Equation 1 was used for this purpose. The term $C_p(\mathrm{l})_\mathrm{m}$ represents the heat capacity of the liquid that was estimated by group additivity.[6] Since the heat capacity of the liquid enol-ester and keto-ester estimate slightly differently, a weighted average (see footnote a of Table 1) was used in the adjustments:

$$\Delta_l^g H^\circ_m(298.15\ \mathrm{K})/\mathrm{kJ{\cdot}mol^{-1}} = \Delta_l^g H_m(T_m) + [(10.58 + 0.26C_p(\mathrm{l})_m/\mathrm{J{\cdot}mol^{-1}{\cdot}K^{-1}})(T_m/\mathrm{K} - 298.15\ \mathrm{K})/1000 \quad (1)$$

The composition of ethyl acetoacetate at equilibrium as a neat liquid at T = 298.15 K has been measured several times.[7−11] The results are summarized in Table 2. Agreement between different measurements is good. A value of 7.4 % has been used by Ribeiro da Silva et al.[1] for the composition of the enol-ester at equilibrium in the pure liquid, and we have retained this value for use in our calculations. Although the amount of enol-ester in the pure liquid is small, approximately equal amounts of both tautomers are present at equilibrium in the gas phase.[7] Ribeiro da Silva et al.[1] added the vaporization enthalpy of the equilibrium mixture, (54.2 ± 1.0) kJ·mol^{-1}, to the enthalpy of formation they measured, and reported a value of (−586.2±1.5) kJ·mol^{-1} for the enthalpy of formation of the gas-phase equilibrium mixture (0.507 keto, 0.493 enol). Using available condensed phase enthalpies of enolization,[8] they estimated a gas-phase enolization enthalpy of (−8 ± 2) kJ·mol^{-1}. Together with the known gas-phase composition of keto to enol form, 50.7 to 49.3 %, respectively, they arrived at an estimate of the enthalpy of formation of ethyl 3-oxobutanoate, $\Delta_f H^\circ_m$(g, 298.15 K) = (−582.2 ± 2.5) kJ·mol^{-1}.

Folkendt et al.[9] have reported NMR measurements of the temperature dependence of the equilibrium constant

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Table 1. Summary of Vaporization Enthalpies Measured for Ethyl Acetoacetate[a]

$\Delta_l^g H_m(T_m)$ kJ·mol^{-1}	T_m K	$\Delta C_p(l)_m \Delta T$ kJ·mol^{-1}	$\Delta_l^g H_m^\circ$(298.15 K) kJ·mol^{-1}	method	reference
54.2 ± 1.0	298.15		54.2 ± 1.0	drop calorimetry	1
52.5	316	1.35	53.8		4
55.0					5

[a] A value of 249.7 J·mol^{-1}·K^{-1} was used for $C_p(l)_m$ (enol−ester, mol fraction: 0.074, $C_p(l)_m$ = 267.5); (keto-ester, mol fraction: 0.926, $C_p(l)_m$ = 248.3 J·mol^{-1}·K^{-1}).

Table 2. Summary of Equilibrium Composition and Enthalpy Differences Measured between Ethyl 3-Oxobutanoate and *Z*-Ethyl 3-Hydroxy-2-butenoate in the Liquid and Gas Phase[a]

% enol liquid phase	$\Delta H_{\text{keto/enol}}(T_m)$ kJ mol^{-1} liquid phase	% enol gas phase	$\Delta H_{\text{keto/enol}}(T_m)_{\text{gas}}$ kJ mol^{-1}	T_m K	reference
7.4				293.15	11, 7
8.0				298.15	8
7.39 ± 0.2		49.3			7
6.8	0.75 ± 0.83	46.6	13.3 ± 1.0	298.15	9
		45.3−46.9			10

[a] Enthalpy differences measured by the temperature dependence of the equilibrium constant.

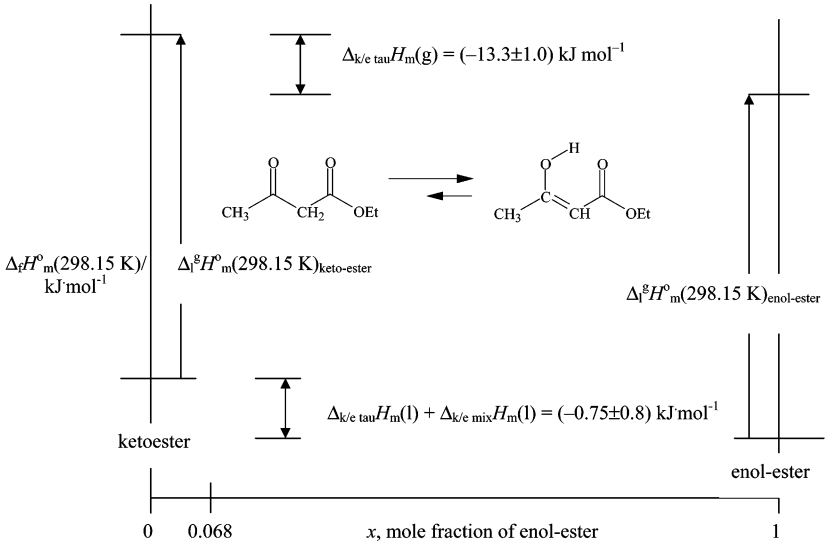

Figure 1. Thermochemical scheme resulting from the NMR studies of Folkendt et al.[9]

for the neat liquid in both the condensed and gas phases. From their measurements they were able to calculate the enthalpy differences between the two tautomers. The results of Folkendt et al.[9] on ethyl acetoacetate are summarized in Figure 1. Enthalpy differences between the keto-ester and enol-ester of (−0.75 ± 0.84) kJ·mol^{-1} and (−13.3 ± 1.0) kJ·mol^{-1} were obtained for the neat liquid and the gas phase, respectively, with the enol-ester enthalpically the more stable in the gas phase. As noted above, the thermodynamic equilibrium, however, favors the keto-ester in the condensed phase. Despite perhaps a small enthalpic advantage in the condensed phase, the enol-ester suffers from an entropic disadvantage in both phases.

It is significant to point out that the enthalpic difference measured by Folkendt et al.[9] between the enol-ester and the keto-ester, by the temperature dependence of the equilibrium constant of the liquid phase, is actually a sum of the enthalpy of enolization, $\Delta_{\text{ketoester/enol}}H_m(l)$, and the enthalpy of intermolecular interactions between the enol-ester and keto-ester forms. The enthalpy of intermolecular interaction between the two tautomers will be referred to as the enthalpy of mixing, $\Delta_{k/e}H_{\text{mix}}$. The fact that the enthalpic difference in the gas phase differs from that in the condensed phase suggests that the vaporization enthalpies of the two tautomers differ, that there is an nonzero enthalpy of mixing term in the liquid-phase influencing the enthalpic differences observed, or both. Without prior knowledge of the vaporization enthalpy of both pure tautomers, it is not possible to distinguish between these three possibilities.

Correlation-gas chromatography is an ideal method for determining the vaporization enthalpy of a pure material even though the material of interest may be present in a mixture. In the case of a dynamic equilibrium, all that is required for measurement is that the components in the mixture not equilibrate during the chromatography. In the absence of equilibration, the vaporization enthalpy of the pure material is obtained. A study of the vaporization enthalpy of the two tautomers, measured as pure liquids, is reported below.

Experimental Section

The compounds studied were all obtained from commercial vendors in high chemical purity (98+ %). The gas chromatograph used to measure retention times was an HP 5980 series II instrument equipped with a split−splitless capillary injection port and a FID detector. A split ratio of approximately 50:1 was used. A 30 m SPB-5 capillary column was used for the analyses. The retention

Table 3. Retention Times (min) of Standards and the Tautomers of Ethyl Acetyl Acetate

Run 1

	T/K						
	333.3	338.4	343.4	348.5	353.6	358.6	363.7
methane	0.395	0.395	0.401	0.406	0.410	0.415	0.418
2,2,4,4-tetramethylcyclobutanedione	5.334	4.425	3.725	3.165	2.721	2.367	2.065
benzoquinone	6.471	5.352	4.488	3.795	3.248	2.811	2.439
2,5-hexanedione	7.038	5.774	4.793	4.016	3.4	2.914	2.509
ethyl 3-oxobutanoate	8.37	6.777	5.553	4.595	3.843	3.253	2.774
1,4-cyclohexanedione	23.983	18.781	14.95	12.02	9.766	8.025	6.637

Run 2

	T/K					
	323.2	328.2	333.2	338.3	343.3	348.4
methane	0.395	0.402	0.475	0.408	0.415	0.419
2,4-pentanedione	4.463	3.726	3.125	2.653	2.275	1.982
methyl 3-oxobutanoate	7.284	6.01	4.846	4.04	3.394	2.847
benzoquinone	10.078	8.215	6.754	5.554	4.64	3.915
2,5-hexanedione	11.239	9.18	7.298	6.02	4.997	4.115
ethyl 3-oxobutanoate	13.748	11.095	8.676	7.072	5.794	4.705
1,4-cyclohexanedione	41.73	32.5	24.908	19.653	15.671	12.398

Run 3

	T/K						
	343.5	348.5	353.6	358.6	363.8	368.8	374
methane	0.401	0.401	0.411	0.411	0.416	0.416	0.427
methyl *Z*-3-hydroxy-2-butenoate	2.243	2.243	1.737	1.737	1.392	1.392	1.184
ethyl *Z*-3-hydroxy-2-butenoate	3.879	3.879	2.822	2.822	2.13	2.13	1.702
ethyl 3-hydroxybutanoate	4.97	4.97	3.54	3.54	2.6	2.6	2.035
ethyl 3-hydroxyhexanoate	20.427	20.427	13.06	13.06	8.614	8.614	6.017
Z-2,2,6,6-tetramethyl-5-hydroxy-4-heptene-3-one	22.209	22.209	14.167	14.167	9.33	9.33	6.484
o-hydroxyacetophenone	25.365	25.365	16.451	16.451	10.95	10.95	7.725

Run 4

	T/K						
	373.8	368.8	363.7	359	353.5	348.4	343.4
methane	0.43	0.425	0.421	0.416	0.41	0.406	0.401
ethyl *Z*-3-hydroxy-2-butenoate	1.694	1.899	2.142	2.45	2.836	3.3	3.89
ethyl 3-hydroxybutanoate	2.007	2.274	2.605	3.02	3.546	4.174	4.978
ethyl 3-hydroxyhexanoate	6	7.179	8.748	10.739	13.358	16.565	20.87
o-hydroxyacetophenone	7.645	9.122	11.016	13.437	16.585	20.457	25.575
methyl *o*-hydroxybenzoate	9.22	11.1	13.445	16.506	20.469	25.513	32.15

times were recorded to three significant figures following the decimal point on an HP 3356 series II integrator. The solvent used was either methylene chloride or methanol. At temperatures below 100 °C, the solvent was retained on the column. Methane was bubbled into the solution just prior to injection and was used as the non-retained reference. The retention time of methane was used to determine the dead volume of the column. Adjusted retention times (t_a) were calculated by subtracting the retention time of methane from the retention time recorded for each analyte. At the higher temperatures the retention time of the solvent was used as the non-retained reference. Column temperatures were controlled by the instrument and monitored using a Fluke 51 K/J thermometer. GC−MS experiments were performed on an HP 5988 A instrument. The methods and procedures used to determine the enthalpy of vaporization and vapor pressures have been previously described in the literature.[12] Retention times are recorded in Table 3.

The vaporization enthalpies of the compounds used as standards for ethyl 3-oxobutanoate and ethyl *Z*-3-hydroxy-2-butenoate are summarized in Table 4, Sections A and B.[2,4,14−21] Vaporization enthalpies were adjusted to T = 298.15 K if necessary using eq 1. Values for $C_p(\mathrm{l})_\mathrm{m}$ in eq 1 were estimated.[6] The vaporization enthalpies of two of the standards used, 2,2,4,4-tetramethyl-1,3-cyclobutanedione and 1,4-cyclohexanedione, were calculated by subtracting the fusion enthalpy adjusted to T = 298.15 K from the experimental sublimation enthalpy also adjusted to T = 298.15 K (eq 2). Equation 3, shown to provide reliable temperature adjustments,[13] was used to adjust fusion enthalpies to T = 298.15 K:

$$\Delta_\mathrm{l}{}^\mathrm{g}H^\circ_\mathrm{m}(298.15\ \mathrm{K})/\mathrm{kJ\cdot mol^{-1}} = \Delta_\mathrm{cr}{}^\mathrm{g}H^\circ_\mathrm{m}(298.15\ \mathrm{K}) - \Delta_\mathrm{cr}{}^\mathrm{l}H^\circ_\mathrm{m}(298.15\ \mathrm{K}) \quad (2)$$

$$\Delta_\mathrm{cr}{}^\mathrm{l}H^\circ_\mathrm{m}(298.15\ \mathrm{K}) = \Delta_\mathrm{cr}{}^\mathrm{l}H^\circ_\mathrm{m}(T_\mathrm{fus}) + [0.15C_p(\mathrm{cr})_\mathrm{m} - 0.26C_p(\mathrm{l})_\mathrm{m} - 9.83\ \mathrm{J\cdot K^{-1}\cdot mol^{-1}}](T_\mathrm{fus} - 298.15\ \mathrm{K})/1000 \quad (3)$$

The vaporization enthalpy of 2,2,4,4-tetramethylcyclobutanedione, $\Delta_\mathrm{l}{}^\mathrm{g}H^\circ_\mathrm{m}(298.15\ \mathrm{K}) = (54.2 \pm 0.3)\ \mathrm{kJ\cdot mol^{-1}}$, obtained in this manner has recently been reported.[2] Two discordant sublimation enthalpies at T = 298.15 K have been reported for 1,4-cyclohexanedione, both from reputable laboratories. A value of 84.2 kJ·mol^{-1} was measured by mass effusion−torsion effusion,[15] while a value of (75 ± 1) kJ·mol^{-1} was obtained by "vacuum drop microcalorimetry".[16] The phase change enthalpies of 1,4-cyclohexanedione have also been reported {($\Delta_\mathrm{tr}H_\mathrm{m}$ (322.2 K) =

Table 4. Summary of Literature Values of the Vaporization Enthalpy Standards

Section A: Reference Compounds for Keto-esters

	$\Delta_l^gH_m(T_m)$ kJ· mol^{-1}	range of measurement K	T_m K	$C_p(l)_m$[a] J·mol^{-1}·K^{-1}	$\Delta_l^gH_m^\circ(298\ K)$[b] kJ·mol^{-1}	mean value $\Delta_l^gH_m^\circ(298\ K)$ kJ· mol^{-1}	ref
$C_5H_8O_2$ 2,4-pentanedione (100 %)					51.2 ± 2.2		2
$C_5H_8O_3$ methyl 3-oxobutanoate	49.7	353−384	369	216.4	54.4		14
$C_6H_4O_2$ benzoquinone	47.8	388−402	395	180	53.4		4
$C_6H_8O_2$ 1,4-cyclohexanedione					68.7[c]		15
					59.4[d]		16
$C_6H_{10}O_2$ 2,5-hexanedione	50.1	386−416	401	236.6	57.5		4
$C_8H_{12}O_2$ 2,2,4,4-tetramethylcyclobutanedione					54.3		16, 2
					54.1	54.2 ± 0.3	17, 2

Section B: Reference Compounds for Enol-esters

	$\Delta_l^gH_m(T_m)$ kJ· mol^{-1}	range of measurement K	T_m K	$C_p(l)_m$[a] J·mol^{-1}·K^{-1}	$\Delta_l^gH_m^\circ(298\ K)$[b] kJ·mol^{-1}	mean value $\Delta_l^gH_m^\circ(298\ K)$ kJ· mol^{-1}	ref
$C_6H_{12}O_3$ ethyl 3-hydroxybutanoate	55.9		298		55.9	55.9 ± 0.6	19
$C_8H_8O_2$ o-hydroxyacetophenone	59.6		298			59.6 ± 0.6	2
$C_8H_8O_3$ methyl *o*-hydroxybenzoate	59.9	327−357	342	269	63.4		4
	58.7	329−359	344		62.4	62.0 ± 1.8	4
	56.9	288−318	303		57.3[d]		4
	55.8	327−383	355		60.3		20
$C_8H_{16}O_3$ ethyl 3-hydroxyhexanoate	61.9		298		61.9	61.9 ± 0.6	19
2,2,6,6-tetramethyl-5-hydroxyhept-4-ene-3-one (98 %)	61.9 ± 2.0		298		61.9 ± 2.0	60.7 ± 1.5	2
	59.5 ± 0.2		298		59.5 ± 0.2		21

[a] Heat capacities were estimated by group additivity. [b] Vaporization enthalpies were adjusted to 298.15 K using eq 1. [c] See Experimental Section. [d] Value not used.

6.15; $\Delta_{tr}H_m$ (339.2 K) = 0.96; $\Delta_{cr}{}^lH_m$ (348.2 K) = 10.0) kJ·mol^{-1}}.[22] Adjusting the total phase change enthalpy to T = 298.15 K using eq 3 and estimated liquid and solid heat capacities of (196.4 and 167) J·mol^{-1}·K^{-1},[6] respectively, resulted in a fusion enthalpy of 15.6 kJ·mol^{-1} at T = 298.15 K. The predicted vaporization enthalpies for the two sublimation entries, according to eq 2, resulted in $\Delta_l^gH_m^\circ$(298.15 K) values of (68.7 and 59.4) kJ·mol^{-1}, respectively. As illustrated and discussed below, the values measured by torsion effusion−mass effusion gave the best correlation with the vaporization enthalpies of other dicarbonyl-containing compounds.

Results and Discussion

Correlation-gas chromatography is an ideal method for determining the vaporization enthalpy of a pure material even though the material of interest may be present in a mixture. This has recently been demonstrated in determination of the vaporization enthalpy of acetoin which exists in equilibrium with a series of stereoisomeric dimers[19] and acetylacetone,[2] which exists as an equilibrium mixture of two tautomers. All that is required is a demonstration that the rate of equilibration is slow relative to the time scale of the chromatography. Interconversion between the two tautomers of ethyl acetoacetate is relatively slow. A gas chromatographic trace of ethyl acetylacetate (solvent not shown) is illustrated in Figure 2. The keto-ester form predominates at equilibrium. The trace illustrates the enol-ester form, which is the sharp peak at approximately 360 time units, a sharp peak for the major tautomer at approximately 480 time units and a broad rather featureless absorption between the two pure tautomers. The sharpness of both of these peaks confirms that tautomerism is relatively slow on the time scale of the chromatogram. The relative intensity of the peaks is in agreement with the GC−MS studies of Allegretti et al.[23] on the composition of the equilibrium mixture in the gas phase. The broad relatively featureless contribution to the relative intensity between the two peaks is due to a portion of the ethyl acetoacetate that has tautomerized during the chromatography. This behavior has been observed and studied previously.[23] We have confirmed this interpretation with our own GC−MS studies. The mass spectra obtained for the two sharp peaks and at intermediate times were consistent with ethyl acetoacetate {*m*/*z* 130 parent ion (both), major fragments at *m*/*z* 88 (keto-ester), 85 (enol-ester), and 43 (both}. Differences in the fragmentation pattern and fragment intensity were observed in the two sharp peaks; these were consistent with those previously described and assigned for the two tautomers.[23]

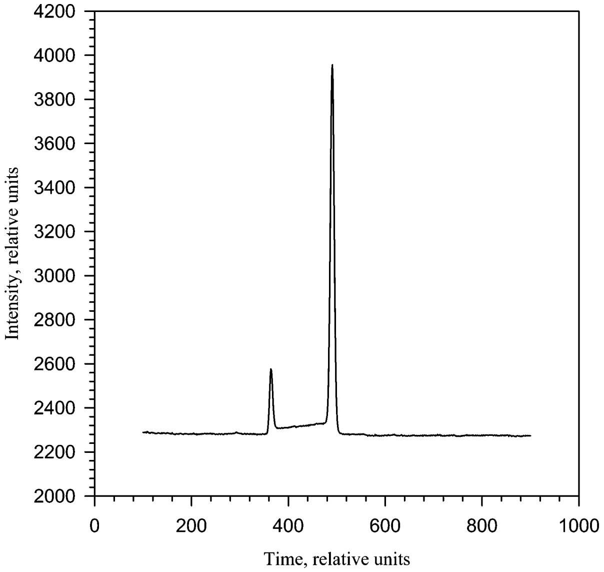

Figure 2. Gas chromatogram of ethyl acetoacetate.

Correlation-gas chromatography is not a direct method for determining vaporization enthalpies. What is determined directly is the enthalpy of transfer from the condensed phase of the column to the gas phase, $\Delta_{sln}{}^gH_m(T_m)$.

Table 5. Enthalpies of Transfer and Vaporization Enthalpies Obtained for Ethyl Acetoacetate

compound	slope	intercept	$\Delta_{\mathrm{sln}}{}^{\mathrm{g}}H_{\mathrm{m}}(T_{\mathrm{m}})$ kJ·mol^{-1}	$\Delta_{\mathrm{l}}{}^{\mathrm{g}}H^{\circ}_{\mathrm{m}}$(298.15 K) kJ·mol^{-1} (lit)	$\Delta_{\mathrm{l}}{}^{\mathrm{g}}H^{\circ}_{\mathrm{m}}$(298.15 K) kJ·mol^{-1} (calcd)
Run 1					
2,2,4,4-tetramethylcyclobutanedione	−4368.3	11.517	36.32	54.2	53.7 ± 1.0
benzoquinone	−4379.4	11.343	36.41	53.4	53.9 ± 1.0
2,5-hexanedione	−4605.3	11.929	38.29	57.5	57.5 ± 1.0
ethyl 3-oxobutanoate	−4861.3	12.515	40.41		61.6 ± 1.0
1,4-cyclohexanedione	−5306.8	12.772	44.12	68.7	68.7 ± 1.0
$\Delta_{\mathrm{l}}{}^{\mathrm{g}}H^{\circ}_{\mathrm{m}}(298.15\ \mathrm{K})/\mathrm{kJ\cdot mol^{-1}} = (1.92 \pm 0.077)\cdot\Delta_{\mathrm{sln}}{}^{\mathrm{g}}H^{\circ}_{\mathrm{m}}(348\ \mathrm{K}) - (16.01 \pm 0.49)$ $r^2 = 0.9968$ (4)					
Run 2					
2,4-pentanedione	−4271.8	11.82	35.51	51.2	49.9 ± 2.2
methyl 3-oxobutanoate	−4658.1	12.49	38.73	54.4	55.5 ± 2.2
1,4-benzoquinone	−4556	11.83	37.88	53.4	54.0 ± 2.2
2,5-hexanedione	−4804.7	12.48	39.94	57.5	57.6 ± 2.2
ethyl 3-oxobutanoate	−5074.7	13.11	42.19		61.6 ± 2.2
1,4-cyclohexanedione	−5529.6	13.39	45.97	68.7	68.2 ± 2.2
$\Delta_{\mathrm{l}}{}^{\mathrm{g}}H^{\circ}_{\mathrm{m}}(298.15\ \mathrm{K})/\mathrm{kJ\cdot mol^{-1}} = (1.75 \pm 0.14)\cdot\Delta_{\mathrm{sln}}{}^{\mathrm{g}}H_{\mathrm{m}}(336\ \mathrm{K}) - (12.21 \pm 1.1)$ $r^2 = 0.981$ (5)					
Run 3					
methyl *Z*-3-hydroxy-2-butenoate	−3493.3	10.042	29.04		52.3 ± 1.3
ethyl *Z*-3-hydroxy-2-butenoate	−3921	10.708	32.60		54.9 ± 1.3
ethyl 3-hydroxybutanoate	−4089.8	10.952	34.00	55.9	55.9 ± 1.3
ethyl 3-hydroxyhexanoate	−4970.9	12.157	41.33	61.9	61.2 ± 1.3
Z-2,2,6,6-tetramethyl-5-hydroxy-4-heptene-3-one	−4984.9	12.113	41.44	61.9	61.3 ± 1.3
o-hydroxyacetophenone	−4798.8	11.413	39.90	60.0	60.2 ± 1.3
$\Delta_{\mathrm{l}}{}^{\mathrm{g}}H^{\circ}_{\mathrm{m}}(298.15\ \mathrm{K})/\mathrm{kJ\cdot mol^{-1}} = (0.725 \pm 0.108)\cdot\Delta_{\mathrm{sln}}{}^{\mathrm{g}}H_{\mathrm{m}}(343\ \mathrm{K}) + (31.22 \pm 0.65)$ $r^2 = 0.9852$ (6)					
Run 4					
ethyl *Z*-3-hydroxy-2-butenoate	−4278.1	11.214	35.57		54.5 ± 1.2
ethyl 3-hydroxybutanoate	−4497.5	11.58	37.39	55.9	55.8 ± 1.2
ethyl 3-hydroxyhexanoate	−5497.5	12.993	45.70	61.9	62.0 ± 1.2
o-hydroxacetophenone	−5274	12.137	43.85	60.0	60.6 ± 1.2
methyl *o*-hydroxybenzoate	−5412	12.309	44.99	62.0	61.4 ± 1.2
$\Delta_{\mathrm{l}}{}^{\mathrm{g}}H^{\circ}_{\mathrm{m}}(298.15\ \mathrm{K})/\mathrm{kJ\cdot mol^{-1}} = (0.739 \pm 0.088)\cdot\Delta_{\mathrm{sln}}{}^{\mathrm{g}}H_{\mathrm{m}}(359\ \mathrm{K}) + (28.19 \pm 0.58)$ $r^2 = 0.9724$ (7)					

A value of $\Delta_{\mathrm{sln}}{}^{\mathrm{g}}H_{\mathrm{m}}(T_{\mathrm{m}})$ is obtained by plotting $\ln(1/t_{\mathrm{a}})$, where t_{a} represents the time each analyte spends on the column, against reciprocal temperature, K^{-1}. The reciprocal of t_{a} is proportional to the vapor pressure of each analyte on the stationary phase of the column at a particular temperature. Once evaluated, enthalpies of transfer measured experimentally are correlated to known vaporization enthalpies measured by some other means. The vaporization enthalpy of the target is obtained from the correlation equation. Selection of the reference compounds is crucial in the evaluation. Generally, reference compounds are chosen with the same type and number of functional groups as the target compound. Some flexibility is possible in certain cases where it has been demonstrated empirically that functional group substitution still provides suitable correlations. For example, good correlations have previously been obtained when an ester group is substituted for a ketone as in this study.[24] A rationale for this has recently been reported.[25]

The vaporization enthalpies of only a limited number of enol-esters are available as reference compounds for ethyl *Z*-3-hydroxy-2-butenoate. Many enolic esters exist in equilibrium with their keto-ester form. Literature vaporization enthalpies measured on these mixtures suffer from the same ambiguities as those associated with ethyl acetoacetate. Phenolic esters such as methyl *o*-hydroxybenzoate are good models, but in order to obtain a spread of retention times, several hydroxyketones and hydroxyesters have also been used to model the enol-ester tautomer.

Reference compounds for the keto-ester, ethyl 3-oxobutanoate, are also problematic. In this instance it has been necessary to use diketones as models for the keto-ester. As a test of whether diketones would serve as reliable models for keto-esters, the vaporization enthalpy of methyl 3-oxobutanoate was also included in the mixtures studied. Depending on the solvent used, methyl 3-oxobutanoate also exists in equilibrium with an enol form, methyl *Z*-3-hydroxy-2-butenoate.[10] In the pure liquid, the keto-ester form predominates to the extent of 99+ %.[8] A value for $\Delta_{\mathrm{l}}{}^{\mathrm{g}}H^{\circ}_{\mathrm{m}}(369.1\ \mathrm{K})$ of 49.7 kJ·mol^{-1} was calculated from the reported vapor pressure of pure methyl acetoacetate as a function of temperature (353.2−385 K).[14] Adjusting the vaporization enthalpy to $T = 298.15$ K using eq 1 and an estimated liquid heat capacity of 216.4 J·mol^{-1}·K^{-1} for the pure keto-ester, results in a $\Delta_{\mathrm{l}}{}^{\mathrm{g}}H^{\circ}_{\mathrm{m}}(298.15\ \mathrm{K})$ value of 54.4 kJ·mol^{-1}. This compound was included as a standard to observe the quality of the correlation obtained when using diketones as standards for keto-esters. As a methylene chloride solution, we could also detect a small amount of methyl *Z*-3-hydroxy-2-butenoate in equilibrium with its keto-ester form. Under these conditions, it was also possible to evaluate the vaporization enthalpy of enol-ester form as well.

A series of two mixtures for each tautomer were examined using different standards for each tautomer. A summary of the vaporization enthalpies of the standards used are listed in Table 4. Plots of $\ln(1/t_{\mathrm{a}})$ versus $1/T$ using the retention times given in Table 3 resulted in linear correlations with correlation coefficients ($R^2 > 0.99$). The slopes and intercepts obtained in these plots for ethyl 3-oxobutanoate, ethyl *Z*-3-hydroxy-2-butenoate, and the standards are summarized in Table 5, runs 1−4. Plotting the vaporization enthalpies of the standards versus the enthalpies of transfer from solution to the gas phase, $\Delta_{\mathrm{sln}}{}^{\mathrm{g}}H_{\mathrm{m}}(T_{\mathrm{m}})$, resulted in correlation eqs 4 to 7 given at the bottom of each run in Table 5. Also included in these tables

Table 6. Summary of Standard Molar Enthalpies at T = 298.15 K of the Two Ethyl Acetoacetate Tautomers[a]

compound	$\Delta_f H^\circ_m(l)$ kJ·mol^{-1}	$\Delta_l^g H^\circ_m(298.15\ K)$ kJ·mol^{-1}	$\Delta_f H^\circ_m(g)$ kJ·mol^{-1}	$\Delta_{k/e\ tau} H_m(g)$ kJ·mol^{-1}
ethyl 3-oxobutanoate	−639.8 ± 2.4	61.6 ± 2.2	−578.2 ± 3.3 [−586.2 ± 1.5]	
ethyl *Z*-3-hydroxy-2-butenoate	−647.7 ± 2.4	54.7 ± 1.3	−593.0 ± 2.7	−14.8 ± 3.7 [−13.3 ± 1.0]

[a] $\Delta_f H^\circ_m$(T = 298.15 K, l, 7.4 % enol-ester and 92.6 % keto-ester) = (−640.4 ± 1.1) kJ·mol^{-1}. Values in brackets are the previous literature values.

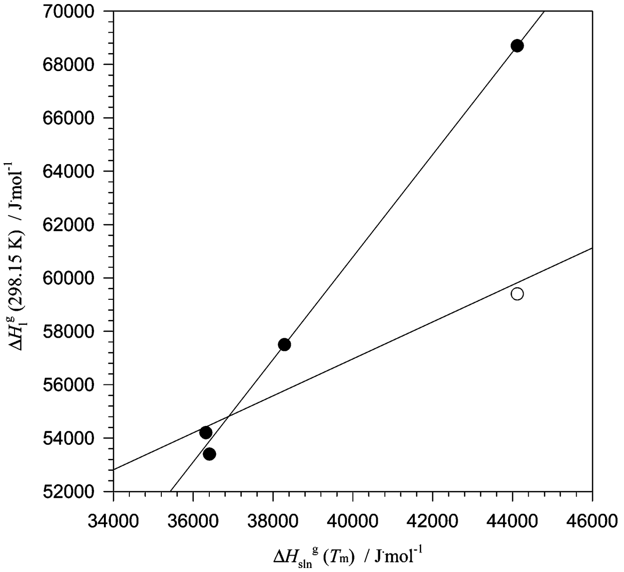

Figure 3. Comparison of the correlations obtained by plotting literature vaporization enthalpies vs enthalpies of transfer by linear regression using the two literature values reported for the sublimation enthalpy of 1,4-cyclohexanedione (solid vs empty circles).

are the vaporization enthalpies of methyl 3-oxobutanoate and methyl *Z*-3-hydroxy-2-butenoate obtained from the correlations.

As discussed in the Experimental Section, two vaporization enthalpies for 1,4-cyclohexanedione can be calculated from available sublimation enthalpy data. Figure 3 illustrates a typical correlation obtained using the two vaporization enthalpies calculated for 1,4-cyclohexanedione. This correlation was obtained using the data in run 1. It is clear form this plot that the vaporization enthalpy (68 700 J·mol^{-1}) obtained using the sublimation enthalpy measured by torsion effusion−mass effusion correlates best with the other data; this value was used in subsequent correlations.

The test to see how well diketones can serve as substitutes for keto-esters is illustrated in Table 5, run 2. Although methyl 3-ketobutanoate is used as a standard in the correlation equation, the value calculated for this substance from the correlation equation clearly demonstrates that the ester does fit quite well. If the vaporization enthalpy of methyl 3-ketobutanoate was omitted from the correlation and treated as an unknown, a value of (55.8 ± 2.2) kJ·mol^{-1} was calculated for the vaporization enthalpy of this compound. This compares to a calculated value of (55.5 ± 2.0) kJ·mol^{-1} when it is included in the correlation.

Ethyl acetoacetate is similar to acetylacetone in that the enol-ester form of the tautomer is more volatile than the keto-ester. Unlike acetylacetone, the vaporization enthalpies of the two tautomers are found to differ somewhat. For ethyl 3-oxobutanoate, a vaporization enthalpy of (61.6 ± 2.2) kJ·mol^{-1} was obtained from runs 1 and 2; a vaporization enthalpy of (54.7 ± 1.3) kJ·mol^{-1} was obtained for ethyl *Z*-3-hydroxy-2-butenoate in runs 3 and 4. The uncertainty in these values represents two standard deviation of the largest uncertainty associated with the intercept of the correlation equation, eqs 4 to 7.

Conclusions

Given the vaporization enthalpies of both individual tautomers, it is now possible to calculate an ideal vaporization enthalpy of the equilibrium mixture without the enthalpy of mixing term, $\Delta_l^g H^\circ_m(298.15\ K)_{ideal\ mix}$ (eq 8), and compare it to the vaporization enthalpy measured on the equilibrium mixture, $\Delta_l^g H^\circ_m(298.15K)_{mix}$ = (54.0 ± 0.6) kJ·mol^{-1}(average of the first two entries, Table 1).

$$\Delta_l^g H^\circ_m(298.15\ \mathrm{K})_{ideal\ mix} = 0.074 \cdot (54.7 \pm 1.3) + 0.926 \cdot (61.6 \pm 2.2) = (61.1 \pm 2.0)\ \mathrm{kJ \cdot mol^{-1}} \quad (8)$$

The difference, (7.1 ± 2.1) kJ·mol^{-1}, should correspond to the enthalpy of mixing, $\Delta_{k/e\ mix} H_m(l)$. Given that $\Delta_{k/e\ tau} H_m(l) + \Delta_{k/e\ mix} H_m(l) = (-0.75 \pm 0.8)$ from the temperature dependence of the equilibrium constant,[9] the enthalpy difference between the two pure tautomers is now calculated as (−7.85 ± 2.3) kJ·mol^{-1}. From the experimental enthalpy of formation of liquid ethyl acetoacetate (−640.4 ± 1.1 kJ mol^{-1}), the enthalpy of tautomerism without mixing and the equilibrium concentration of the liquid (7.4 % enol-ester), enthalpies of formation of the liquid enol-ester ($\Delta_f H^\circ_m(l)_{enol-ester}$) and keto-ester ($\Delta_f H^\circ_m(l)_{keto-ester}$) can be calculated (eqs 9 and 10).

$$\Delta_f H^\circ_m(l)_{enol-ester} - \Delta_f H^\circ_m(l)_{keto\ ester} = (-7.85 \pm 2.3)\ \mathrm{kJ \cdot mol^{-1}} \quad (9)$$

$$0.074 \cdot \Delta_f H^\circ_m(l)_{enol-ester} + 0.926 \cdot \Delta_f H^\circ_m(l)_{keto-ester} = (-640 \pm 1.1)\ \mathrm{kJ \cdot mol^{-1}} \quad (10)$$

The results are shown in Table 6. Addition of the corresponding vaporization enthalpy of each respective tautomer results in gas-phase enthalpies of formation of (−593.0 ± 2.7) and (−578.2 ± 3.3) kJ·mol^{-1}. The gas-phase difference between the two tautomers now calculates as (−14.8 ± 3.7) kJ·mol^{-1}; this value can be compared to the gas-phase NMR value of (−13.3 ± 1.0) kJ·mol^{-1} reported by Folkendt et al.[9]

A positive enthalpy of mixing between between alcohols and ketones is frequently observed.[26] An enthalpy of mixing value of (7.6 ± 0.6) kJ·mol^{-1} calculated in the tautomerization of acetylacetone,[2] which exists primarily as the enol, can be compared to the value obtained in this study, (7.1 ± 2.1) kJ·mol^{-1}, in which the keto-ester tautomer is favored at equilibrium. In both ethyl acetoacetate and in acetylacetone,[2] the pure liquid enol forms are enthalpically favored over the pure keto forms by (−7.85 ± 2.3) and (−18.9 ± 0.7) kJ·mol^{-1}, respectively. The keto form is favored entropically in both cases.[9] Differences in the magnitude of the positive enthalpy of mixing can explain the differences in equilibrium keto/enol concentrations observed between the methyl and ethyl esters of acetoacetic acid. A positive enthalpy of mixing favors the predominate tautomer; in methyl acetoacetate this occurs to the virtual exclusion of the other. It can also explain the sensitivity of the equilibrium mixture to solvent as is observed in this and related systems. If the solute−solvent energetics are less unfavorable than the energetics associated with tautomer interaction, then the equilibrium mixture observed in solution will vary both with solvent and concentration reflecting the enthalpic differences of the pure tautomers as measured in studies of acetylacetone and ethyl acetoacetate. Whether similar enthalpies of mixing influence other tautomeric systems remains to be determined.

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