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Phantom Activation Volumes[1]

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Activation volumes $\Delta V^{\neq}$ are presently obtained by measuring the effect of pressure on the reaction rate. It is implicitly assumed that the entire response of rate to pressure is volume related, i.e., acceleration by high pressure reveals shrinkage as the reactants progress to the transition state, and vice versa. However, we now demonstrate that high pressure accelerates some bond-making reactions in an additional, nonvolume-related way, through its elevation of solvent viscosity. Diels−Alder reactions, 1,3-dipolar cycloadditions, and Claisen rearrangements are accelerated by rising viscosity and are therefore subject to viscosity-associated acceleration at raised pressures. This gives rise to the illusion of volume shrinkage along the reaction coordinate toward the transition state, which we term phantom $\Delta V^{\neq}$. Thus the true $\Delta V^{\neq}$ for these reactions, while negative, is less negative than previously believed. Corrections in $\Delta V^{\neq}$, calculated from experimental rate−viscosity plots, range up to 61%.

Introduction

Volume of activation ($\Delta V^{\neq}$) is an important criterion of mechanism. It is defined in eq 1.[2]

$$\Delta V^{\neq} = -RT\delta \log k/\delta P \qquad (1)$$

Acceleration of a reaction by pressure means that the size of the transition state (TS), which includes not only the reacting atoms but also the surrounding solvent molecules, is smaller than that of the reactants, and retardation by pressure means that the TS is larger. For example, bond-forming reactions are usually pressure accelerated and bond-breaking reactions pressure retarded. However, sometimes solvent effects are overwhelming, as with solvolyses which, although bond breaking, are pressure accelerated owing to electrostriction of solvent by the polar TS.[3] Another example of solvent domination is the Diels−Alder reaction. It is pressure accelerated as expected for a bond-forming reaction, but the $\Delta V^{\neq}$ term arises predominantly from reduction of empty space in the solvent around the TS rather than from bond formation itself.[4]

It is generally assumed that the $\Delta V^{\neq}$ calculated from eq 1 is truly a volume term. It is our thesis, however, that $\Delta V^{\neq}$ is only partially volume derived, because increasing the pressure can induce kinetic effects that do not arise from volume changes. We propose the term "phantom activation volumes" for pressure-induced rate changes that are defined as $\Delta V^{\neq}$ in eq 1 but are in fact not volume related. It is important to identify them because otherwise erroneous decisions regarding mechanism might be made.

How can phantom $\Delta V^{\neq}$s come about? One way is to have a reaction whose rate is sensitive to solvent polarity. Since solvent polarities rise modestly with pressure,[5] such a reaction would respond to increased pressure by a rate increase caused, not by shrinkage in the TS, but instead by the pressure-induced increase in solvent polarity.[6] Of course, superimposed on this rate increase might be another arising from electrostriction (a true volume effect), and yet another positive or negative rate effect resulting from bond formation or breakage. Thus only a portion of the $\Delta V^{\neq}$ calculated from eq 1 would be a true $\Delta V^{\neq}$, and the other portion a phantom (negative) $\Delta V^{\neq}$.

Another possible source of phantom $\Delta V^{\neq}$s is the fact that solvent viscosity usually rises geometrically with pressure.[7] The isomerization of azo compounds and benzylideneaniline, which is normally slowed by high viscosity, shows a strong pressure-induced retardation that arises, not from shrinkage, but instead from the increase in viscosity created by the pressure.[8] In this instance the viscosity-derived phantom $\Delta V^{\neq}$ is positive, but a negative one can also come about (vide infra).[1,9,10]

The Diels−Alder Reaction

Diels−Alder (DA) reactions are strongly accelerated by pressure, with $\Delta V^{\neq}$s ranging from −30 to −50 cm^3/mol, which are not far from the overall reaction volumes, ΔV_R ($\Delta V_R = V_{\text{product}} - V_{\text{reactant}}$).[11] This has been interpreted in terms of a concerted mechanism, in which both new bonds are forming in the TS, rather than a stepwise-diradical mechanism, where only one new bond is created in the rate-determining step (RDS). We have opposed this interpretation[4] on two grounds: (1) a concerted mechanism would have an early TS while a diradical one would have a late RDS TS, so that contraction arising from bond formation would not differ much between them; (2) shrinkage at the TS is so dominated by reduction in empty space among solvent molecules that bond formation hardly affects $\Delta V^{\neq}$ at all. Thus drawing mechanistic conclusions is not warranted.

Another controversy involves the observation that $|\Delta V^{\neq}|$ often equals or even exceeds $|\Delta V_R|$,[11] i.e., the TS is apparently no larger, and may even be *smaller*, than the cycloadduct. This has been attributed to secondary orbital interactions, but they can be ruled out since often endo and exo cycloadditions do not differ in their $\Delta V^{\neq}$s.[12] It is impossible for a partially formed bond to be the same size, not to mention smaller than a fully formed bond.

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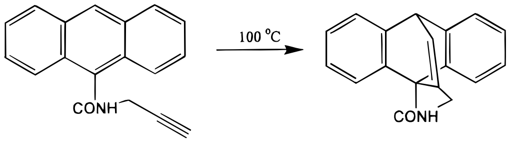
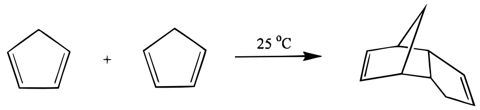
We propose that this apparent paradox can be resolved by recognizing that a significant portion of the published negative $\Delta V^{\neq}$s for DA reactions are phantom volumes. This is because solvent viscosities rise geometrically with increasing pressure[7] and DA rates rise with increasing viscosity.[13] This was first recognized more than a decade ago,[10] in an intramolecular case (reaction 1) run in mono-, di-, tri-, and tetraglyme. Unfortunately, quantitative calculation of phantom volumes was not then possible owing to our somewhat imprecise viscometry, and also to the fact that viscosity data had been obtained on pure solvents but not the reaction mixtures themselves.



(1)

We now report accurate viscosities of the actual reaction mixtures,[1] enabling the determination of a reliable figure for the rate−viscosity relationship. This relationship is also reported for another DA, the dimerization of cyclopentadiene (reaction 2), run in pure linear saturated hydrocarbon solvents.[1] For this intermolecular case, as viscosity rises to about 1−2 centipoise (cP), the rate first rises and then falls. The rise stems from viscosity-induced acceleration in the collision-controlled regime. Only the rising portion of the curve concerns us here, because it is over this viscosity range that published activation volumes were determined. These rate−viscosity relationships permit the determination of what the accelerations would have been in the pressure-accelerated DA cases had they arisen solely from viscosity effects. The viscosity-induced acceleration turns out to be a significant factor of the observed pressure-induced acceleration. Clearly, therefore, *all published $\Delta V^{\neq}$s are too negative* for DA and other reactions that are viscosity-accelerated, e.g., Claisen rearrangements (vide infra).



(2)

The simplest relationship between rate and viscosity is a linear one, which was permitted by the 1981 data and tentatively adopted at that time for parsimony's sake.[10] However, after examination of all the DA pressure−rate data[11] for which the solvents' pressure−viscosity relationships were available,[7] it became apparent that only a log−log relationship (natural log, base e) gives straight lines (Figures 1 and 2).[14] We therefore calculated the phantom volumes in this manner. Since $\Delta V^{\neq}$s are usually based on the initial slopes, i.e., at atmospheric pressure,[2] the calculations involve little if any extrapolation beyond our range of measured viscosities.

Figures 1 and 2 depict typical cases from the literature of rates vs viscosities. The viscosities were obtained by replacing the pressures in the original kinetic papers by the corresponding viscosities at those pressures for those solvents.[7] Of course only a portion of the slopes arises from viscosity effects, but the linearity is very good up to ∼1 − ∼2 cP,[15] and this is significant because the viscosity-induced rate effects are a large portion of the total.

Table 1 provides literature data for a large selection of intermolecular DA cases for which both apparent $\Delta V^{\neq}$s and solvent viscosities under pressure are available, along with the phantom $\Delta V^{\neq}$s. These are defined as that portion of the apparent $\Delta V^{\neq}$s that arise from viscosity rather than from true volume effects. Viscosity effects on the rates were taken as the initial slope (0.96) of log reduced rate vs log reduced viscosity for reaction 2, and of necessity were all assigned the same slope, clearly an approximation. Phantom activation volumes, then, are the ratios of 0.96 and each total slope, times apparent $\Delta V^{\neq}$s. In all entries the closest temperature between reaction and literature pressure−viscosity data was used (≤ 20 °C); extrapolation or interpolation with respect to temperature was not done. Phantom $\Delta V^{\neq}$s are as high as 61% and true $\Delta V^{\neq}$s as small as 39% of the apparent $\Delta V^{\neq}$s, a significant correction (but vide infra).

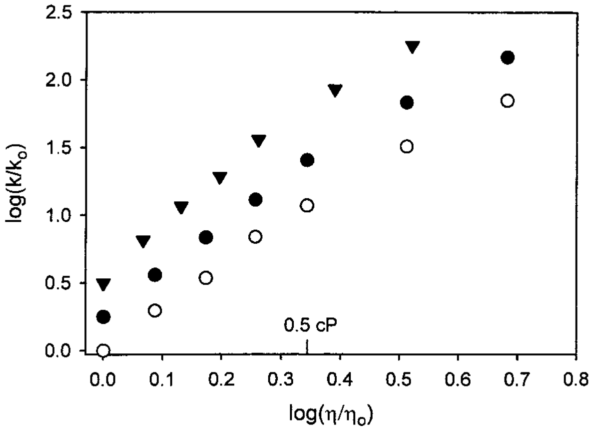


Figure 1. Log reduced viscosity vs log reduced rate of selected literature data. Data taken from Table 1 entry 4 (symbol ▼), entry 5 (symbol ●), entry 17 (symbol ○).

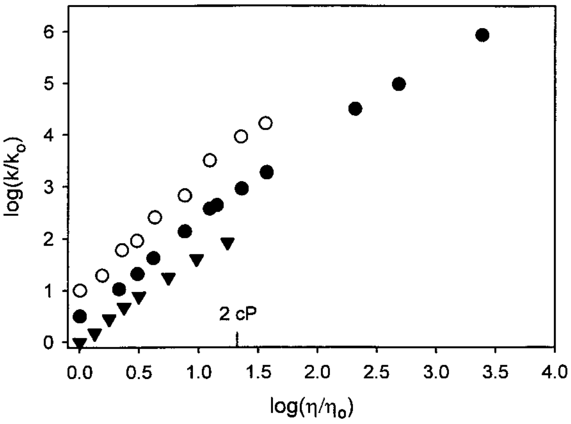


Figure 2. Log reduced viscosity vs log reduced rate of selected literature data. Data taken from Table 1 entry 10 (symbol ○), entry 13 (symbol ●), entry 19 (symbol ▼).

Table 2 provides three literature examples of intramolecular DA cases for which both apparent $\Delta V^{\neq}$s and solvent viscosities under pressure are available, along with the phantom $\Delta V^{\neq}$s which were calculated as before, but using the slope (0.204) of the intramolecular DA (reaction 1) in ref 1. In all entries the closest temperature between reaction and literature pressure−viscosity data was used. In entries 2 and 3 which were done at 153 °C in hexane, we were limited to use pressure−viscosity data of 75 °C.[16] Phantom $\Delta V^{\neq}$s are as high as 21%, and true $\Delta V^{\neq}$s as small as 79% of the apparent $\Delta V^{\neq}$s, a significant correction (but vide infra).

We recognize, of course, that the use of only a single case each of intermolecular and intramolecular DA to calculate phantom activation volumes for all cases is only semiquantitatively justified, since rate−viscosity slopes undoubtedly vary from case to case. Thus only approximate conclusions can be drawn, except for CPD itself.

TABLE 1: Diels-Alder Cycloadditions

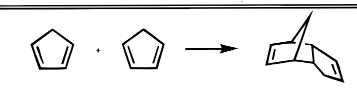
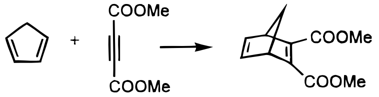
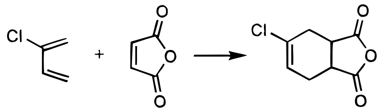
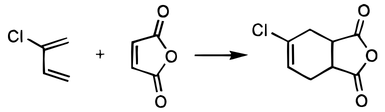
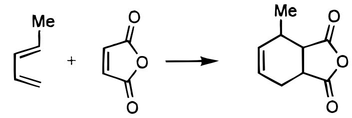
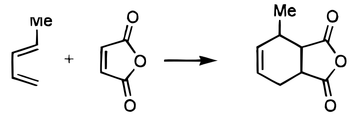
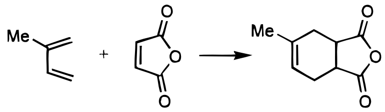
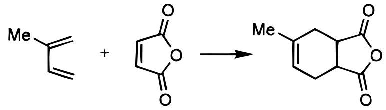
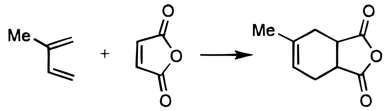
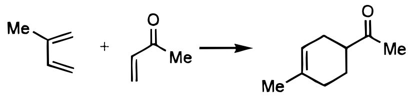
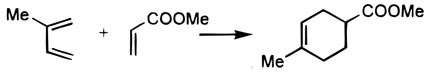
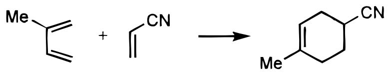
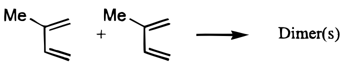
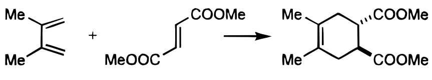
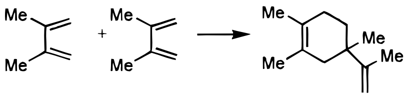
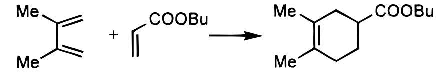
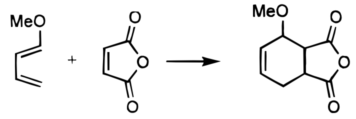
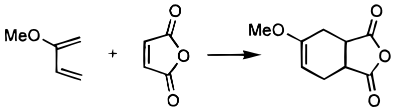
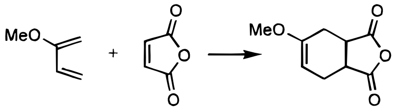
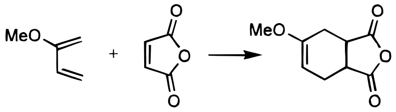
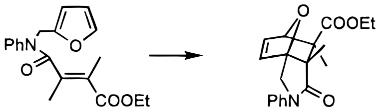
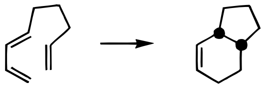
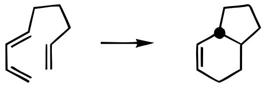
Entry	Reaction	Solvent (Temp, °C)	$\Delta V^{\neq}$ (mL/mol)	ΔV_R (mL/mol)	k-η Slope	Volume Correction (%)	Corrected $\Delta V^{\neq}$ (mL/mol)	Ref.
1		*n*-BuCl (30)[a]	-22.3	-32.2	1.56	61	-9.6	18
2		EtOAc (10)[b]	-30.2	-33.9	1.56	61	-13.0	19
3		*n*-BuCl (65)[c]	-51.1		2.51	38	-36.8	20
4		Acetone (65)[d]	-48.6		3.36	29	-35.0	20
5		*n*-BuCl (35)[e]	-48.9		2.85	34	-33.3	20
6		Acetone (35)[b]	-47.3		2.97	32	-33.1	20
7		*n*-BuCl (35)[e]	-38.0		2.47	39	-24.3	21
8		EtOAc (35)[b]	-37.4	-36.8	1.95	49	-20.2	21
9		Acetone (35)[b]	-39.0	-35.9	2.47	39	-25.0	21
10		*n*-BuBr (21)[b]	-36.9	-37.1	2.16	44	-21.4	22
11		*n*-BuBr (21)[b]	-30.8	-36.9	1.88	51	-16.0	22
12		*n*-BuBr (21)[b]	-33.1	-37.0	1.84	52	-16.9	22
13		nBuBr (40)[b]	-35.3	-41.9	1.83	52	-18.0	23
14		*n*-BuCl (40)[e]	-32.9	-37.2	2.25	43	-19.7	24
15		*n*-BuBr (70)[d]	-33	-44.9	1.57	61	-14.2	25

TABLE 1: Diels-Alder Cycloadditions

Entry	Reaction	Solvent (Temp, °C)	$\Delta V^{\neq}$ (mL/mol)	ΔV_R (mL/mol)	k-η Slope	Volume Correction (%)	Corrected $\Delta V^{\neq}$ (mL/mol)	Ref.
16		PhBr (80)[d]	-22.6		1.85	52	-11.5	18
17		*n*-BuCl (35)[e]	-45.4	-35.5	2.74	35	-30.4	20
18		*n*-BuCl (50)[e]	-36.7		2.26	42	-22.0	20
19		CCl_4 (50)[b]	-37.6		1.57	61	-16.2	20
20		Acetone (50)[b]	-36.2	-34.1	2.50	38	-23.2	20

[a] Pressure−viscosity data of 25 °C was used (ref 7c). [b] Pressure−viscosity data of 30 °C was used (ref 7a). [c] Pressure−viscosity data of 60 °C was used (ref 7c). [d] Pressure−viscosity data of 75 °C was used (ref 7a). [e] Pressure−viscosity data of 40 °C was used (ref 7c).

TABLE 2: Intramolecular Diels−Alder Reactions

Entry	Reaction	Solvent (Temp, °C)	$\Delta V^{\neq}$ (mL/mol)	ΔV_R (mL/mol)	k-η Slope	Volume Correction (%)	Corrected $\Delta V^{\neq}$ (mL/mol)	Ref.
1		EtOH (25)[a]	-25	-23	2.34	8.7	-22.8	26
2		Hexane (153)[b]	-24.8	-32	0.98	21	-19.6	27
3		Hexane (153)[b]	-24.8	-28.5	0.98	21	-19.6	27

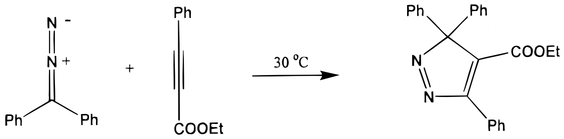
[a] Pressure−viscosity data of 30 °C was used (ref 7a). [b] Pressure−viscosity data of 75 °C was used (ref 7a).

Although it is clear that phantom $\Delta V^{\neq}$s are indeed large, the corrections do not settle the mechanistic controversy of concerted vs diradical, but they do resolve the paradox of TSs seeming impossibly small. In all the cases where ($|\Delta V_R| - |\Delta V^{\neq}|$) appeared to be <0, TSs are now seen to be comfortably larger than the cycloadducts.

It is noteworthy that the rate−viscosity slope of the intramolecular DA[10] is much less than that of the intermolecular one.[1] This is expected, because translation of the two reacting atoms relative to each other is much more severely restricted in the intramolecular case, and therefore the rate has less to gain by way of soft collisions[17] when the viscosity is increased.

1,3-Dipolar Cycloadditions

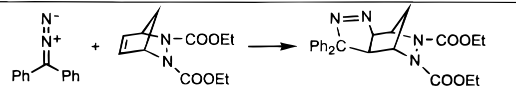
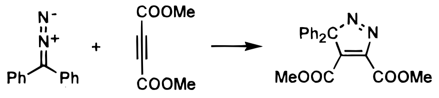
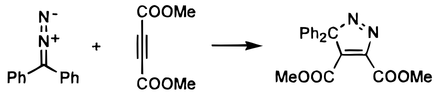
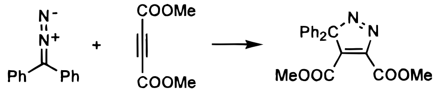
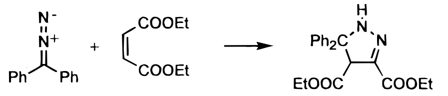
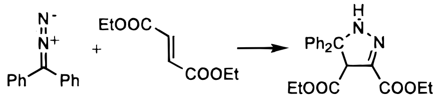
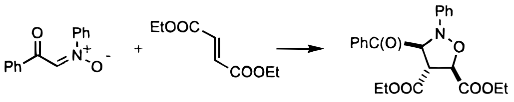
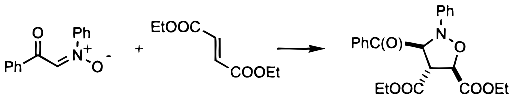
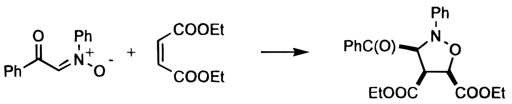
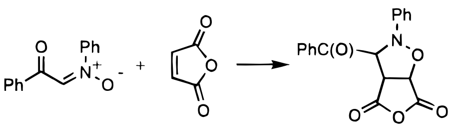
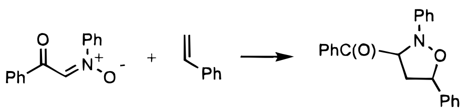
1,3-Dipolar cycloadditions are well known to be pressure accelerated.[11] We examined a typical case, diphenyldiazomethane and ethyl phenylpropiolate (reaction 3), in pure saturated hydrocarbon solvents.[1] Like the intermolecular DA it is viscosity accelerated up to about 1−2 cP.[1] Its rising slope of log reduced rate vs log reduced viscosity is 0.277. We attribute this relatively low value to electrostatic preassociation of the somewhat polar reactants, which reduces the translational component of their relative motion in the same manner that intramolecularity does (vide supra).



(3)

Comparison of our observed slope with those of log *k* vs log *η* (viscosity, cP) for several examples from the literature (Table 3) again shows that a significant portion of the apparent $\Delta V^{\neq}$ is actually viscosity related and that the true $\Delta V^{\neq}$s, with the

TABLE 3: Dipolar Cycloadditions

Entry	Reaction	Solvent [°C]	ΔV≠ [mL/mol]	ΔV_R [mL/mol]	k-η Slope	Volume orrection (%)	Corrected ΔV≠ [mL/mol]	Ref.
1		Toluene 25[a]	-30		1.67	17	-25	[28]
2		Hexane 25[a]	-24	-35	1.16	24	-18	[29]
3		Toluene 25[a]	-23	-27	1.23	23	-18	[29]
4		PhCl 25[a]	-18	-26	1.22	23	-14	[29]
5		PhCl 25[a]	-24	-25	1.48	19	-20	[29]
6		PhCl 25[a]	-21	-27	1.47	19	-17	[29]
7		Toluene 25[a]	-21.7	-22.7	1.09	25	-16.2	[30]
8		$CHCl_3$ 25[a]	-24.3		1.61	17	-20.1	[30]
9		Toluene 25[a]	-19.9		1.07	26	-14.7	[30]
10		Toluene 25[a]	-22.9		1.19	23	-17.6	[30]
11		Toluene 25[a]	-19.5		1.05	26	-14.4	[30]

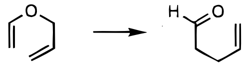
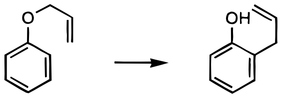
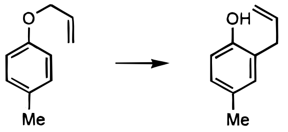
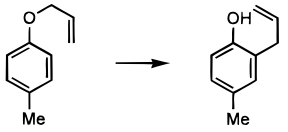
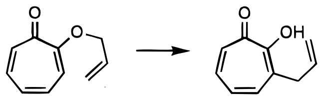
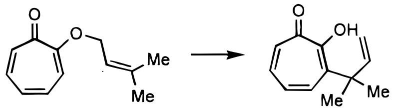
[a] Pressure−viscosity data of 30 °C was used (ref 7a).

same caveat as before, are quite a bit lower than the apparent ones. In all entries the closest temperature between reaction and literature pressure−viscosity data was used (≤ 5 °C), and extrapolation or interpolation with respect to temperature was not done. The viscosity−rate slope used in the calculations is 0.277.[1] Phantom volumes are as high as 26% and true volumes as low as 74% of apparent $\Delta V^{\neq}$.

The Claisen Rearrangement

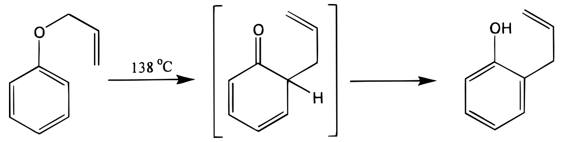
The Claisen rearrangement, like the DA, is pressure accelerated,[31,32] but unlike the DA there is little or no reduction in size during traversal of the reaction coordinate, since the molar volume of allylacetaldehyde is only 6 mL/mol (5.4%) smaller than allyl vinyl ether. This is not surprising since the numbers

TABLE 4: Claisen Rearrangement

Entry	Reaction	Solvent (Temp, °C)	$\Delta V^{\neq}$ (mL/mol)	ΔV_R (mL/mol)	k-η Slope	Volume Correction (%)	Corrected $\Delta V^{\neq}$ (mL/mol)	Ref.
1				-5.6[a]				
2				-4.8[a]				
3		Cyclohexane (176)[b]	-18	-5.7[a]	0.97	12	-16	31
4		Benzene (186)[b]	-18	-5.7[a]	0.65	17	-15	31
5		Ph-Br (120)[b]	-10.6		0.77	14	-9.1	33
6		Ph-Br (120)[b]	-18		0.40	28	-13	33

[a] Molar volumes were determined by using the median density of a compound at 20 °C from Beilstein Crossfire (v 3.1). [b] Pressure−viscosity data of 75 °C was used (ref 7a).

of bonds and rings do not change. Thus, in contrast to the DA and 1,3-dipolar cycloadditions, the Claisen rearrangement products gain almost no thermodynamic advantage over reactants when the pressure rises.[34] There is then only one volume-related kinetic factor here instead of two, namely the fact that the cyclic TS has one more ring than the reactant. It therefore has a higher packing fraction,[4] and the ensemble (TS + solvent) is smaller than (reactant + solvent), giving rise to a negative $\Delta V^{\neq}$.



(4)

We have previously published on the rate increase with viscosity of phenyl allyl ether to 2-allylphenol in hydrocarbon solvents (reaction 4).[10] From our data in Table 4, again with improved viscometry on the actual reaction mixtures[17b] (rate−viscosity slope is 0.118), we are able to calculate the (negative) phantom $\Delta V^{\neq}$ of reported Claisen rearrangements. In all entries the closest temperature between reaction and literature pressure−viscosity data was used. In entries 3 through 6, which were done significantly higher than 75 °C, we were limited to pressure−viscosity data of 75 °C.[16] Once again there is a significant correction of the apparent $\Delta V^{\neq}$s,[1] which are lowered by 12−28%. Relatively low rate−viscosity slopes are seen for this intramolecular reaction as with the intramolecular DA (vide supra) presumably for the same reason.

Conclusion

All activation volumes for Diels−Alder reactions, 1,3-dipolar cycloadditions, and Claisen rearrangements include a significant portion that is not volume related, which we term phantom activation volumes. They arise from the fact that these reactions are accelerated by rising solvent viscosity. Activation volumes are determined by running the reactions at increasing pressure, which unavoidably causes increasing viscosity also, and this imposes a previously unsuspected influence on rates that is independent of volume. The corrections we have calculated, while subject to uncertainty because rate−viscosity slopes were experimentally determined for only a few of the many cases in Tables 1−4, cannot but be significant fractions of the total activation volumes.

It seems likely that many other reactions will exhibit viscosity acceleration. Some other possible examples that we ascribe to viscosity acceleration have previously appeared in the literature,[35] and we have plans to go further afield. In all of these cases it will be necessary to make corrections for phantom volumes for each reaction if possible in order to uncover the true $\Delta V^{\neq}$s, if meaningful mechanistic conclusions are to be drawn.

References and Notes

(1) Vibrational Activation, Paper VI. Paper V: Acceleration of Bimolecular Reactions by Solvent Viscosity. Swiss, K. A.; Firestone, R. A. *J. Phys. Chem. A* **1999**, *103*, 5369−5372. The Supporting Material contains kinetic data and viscosities of the individual reaction mixtures for reactions 1, 2, and 3.

(2) le Noble, W. J. in *Organic High-Pressure Chemistry*; le Noble, W. J., Ed.; Elsevier: Amsterdam, 1988, p 5.

(3) le Noble, W. J. in *Organic High-Pressure Chemistry*; le Noble, W. J., Ed.; Elsevier: Amsterdam, 1988, p 238.

(4) Firestone, R. A.; Smith, G. M. *Chem. Ber.* **1989**, *122*, 1089, and references therein.

(5) Schornack, L. G.; Eckert, C. A. *J. Phys. Chem.* **1970**, *74*, 3014. le Noble, W. J.; Kelm, H. *Angew. Chem., Int. Ed. Engl.* **1980**, *19*, 841.

(6) Prof. A. De Meijere, personal communication. Lansky, A., Dissertation, Gottingen, Germany, 1992, p 42.

(7) As an example, for pure EtOAc at 3871 atm (relative to 1 atm), a 558% increase in η was found (ref 7a). (a) Bridgman, P. W. *Proc. Am. Acad. Arts Sci.* **1926**, *61*, 57. (b) Bridgman, P. W. *The Physics of High Pressure*; G. Bell & Sons: London, 1949. (c) for *n*-BuCl only see Kuss, E. *Z. Angew. Phys.* **1955**, *7*, 372. (d) for *n*-heptane only see Ducoulombier, D.; Zhou, H.; Boned, C.; Peyrelasse, J.; Saint-Guirons, H.; Xans, P. *J. Phys. Chem.* **1986**, *90*, 1692.

(8) Asano, T.; Matsuo, K.; Sumi, H. *Bull. Chem. Soc. Jpn.* **1997**, *70*, 239. For other cogent comments on pressure, viscosity, and activation volumes see Troe, J. *J. Phys. Chem.* **1986**, *90*, 357−365.

(9) One might say that even a viscosity-induced high-pressure kinetic effect is volume derived because whether or not a reaction is taking place, there is reduction in the solvent's void space under pressure which may underlie that increase in viscosity. However, viscosity is an independent parameter of rate even at normal pressure,[1,8,10] and therefore must be so at elevated pressures as well.

(10) Firestone, R. A.; Vitale, M. A. *J. Org. Chem.* **1981**, *46*, 2160.

(11) Asano, T.; le Noble, W. J. *Chem. Rev.* **1978**, *78*, 407.

(12) Seguchi, K.; Sera, A.; Maruyama, K. *Tetrahedron Lett.* **1973**, 1585.

(13) This was first established with an intramolecular DA (ref 10) in somewhat polar solvents, but we have recently obtained quantitatively similar results with an intermolecular, entirely nonpolar case (ref 1).

(14) A log−log relationship has been previously reported, see Oh-oka, H.; Iwaki, M.; Itoh, S. *Biochem.* **1997**, *36*, 9267.

(15) The log−log plots (Figures 1 and 2) exhibit a downward deviation from linearity as viscosities rise above approximately 1−2 cP, right where the downturn occurs with reactions 1 and 2 (ref 1). In Figures 1 and 2 the downturn is less because it is superimposed upon continuation of a pressure-induced rate increase arising from true volume effects.

(16) The absolute viscosities will be considerably different between wide temperature ranges; however, the reduced viscosities are less subject to temperature differences. Bridgman reports (ref 7a) for *n*-hexane: η(30 °C)/η(75 °C) ratios: 1 atm, 1.574; 968 atm, 1.449; 1936 atm, 1.521; 3872 atm, 1.633.

(17) (a) Firestone, R. A.; Christensen, B. G. *Tetrahedron Lett.* **1973**, 389. (b) Swiss, K. A.; Firestone, R. A. *J. Org. Chem.* **1999**, *64*, 2158.

(18) Walling, C.; Schugar, H. J. *J. Am. Chem. Soc.* **1963**, *85*, 607.

(19) Grieger, R. A.; Eckert, C. A. *J. Am. Chem. Soc.* **1970**, *92*, 7149.

(20) McCabe, J. R.; Eckert, C. A. *Ind. Eng. Chem. Fundam.* **1974**, *13*, 168.

(21) Grieger, R. A.; Eckert, C. A. *Trans. Faraday Soc.* **1970**, *66*, 2579.

(22) Brun, C.; Jenner, G. *Tetrahedron* **1972**, *28*, 3113.

(23) Rimmelin, J.; Jenner, G. *Tetrahedron* **1974**, *30*, 3081.

(24) Seguchi, K.; Sera, A.; Maruyama, K. *Bull. Chem. Soc. Jpn.* **1974**, *47*, 2242.

(25) Jenner, G.; Rimmelin, J. *Tetrahedron Lett.* **1980**, 3039.

(26) Isaacs, N. S.; Van der Beeke, P. *Tetrahedron Lett.* **1982**, 2147.

(27) Diedrich, M. K.; Klaerner, F−G. *J. Am. Chem. Soc.* **1998**, *120*, 6212.

(28) Isaacs, N. S.; Rannala, E. *J. Chem. Soc., Perkin Trans. 2* **1975**, 1555.

(29) Swieton, G.; von Jouanne, J.; Kelm, H.; Huisgen, R. *J. Org. Chem.* **1983**, *48*, 1035.

(30) Yoshimura, Y.; Osugi, J.; Nakahara, M. *J. Am. Chem. Soc.* **1983**, *105*, 5414.

(31) Brower, K. R. *J. Am. Chem. Soc.* **1961**, *83*, 4370.

(32) Walling, C.; Naiman, M. *J. Am. Chem. Soc.* **1962**, *84*, 2628.

(33) Sugiyama, S.; Mori, A.; Kato, N.; Takeshita, H. *Bull. Chem. Soc. Jpn.* **1989**, *62*, 1143.

(34) The assumption of negligible net change in size during the rearrangement of phenyl allyl ether is justified even though the isolated product 2-allylphenol is 4.8 mL/mol (3.5%) smaller, because the primary rearrangement product is actually 5-allyl-2,4-cyclohexadienone (shown in reaction 4) which, unlike 2-allylphenol, is probably not aggregated by hydrogen bonds and is therefore probably about as dense as allyl phenyl ether.

(35) (a) Sternbach, D. D.; Rossana, D. M. *Tetrahedron Lett.* **1982**, 303. (b) Dolbier, W. R., Jr.; Seabury, M. J. *J. Am. Chem. Soc.* **1987**, *109*, 4393. (c) Dunams, T.; Hoekstra, W.; Pentaleri M.; Liotta, D. *Tetrahedron Lett.* **1988**, 3745. (d) Dolbier, W. R., Jr.; Weaver, S. L. *J. Org. Chem.* **1990**, *55*, 711. (e) Metz, P.; Seng D.; Frohlich R.; Wibbeling, B. *Synlett* **1996**, 741. (f) Valgimigli, L.; Ingold K. U.; Lusztyk, J. *J. Org. Chem.* **1996**, *61*, 7947.