Heat of Hydrogenation of a Cis Imine. An Experimental and Theoretical Study

Kenneth B. Wiberg,* David Y. Nakaji, and Kathleen M. Morgan

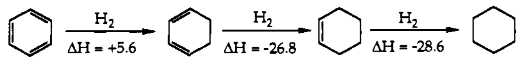
Contribution from the Department of Chemistry, Yale University, New Haven, Connecticut 06511

Received December 7, 1992

Abstract: The heat of hydrogenation of 1-azacyclopentene was determined via a measurement of the heat of hydrogenation of its trimer, and of the dissociation energy of the trimer. The observed ΔH_{hydrog} was −15.9 ± 0.3 kcal/mol and is considerably smaller than that for cyclopentene (−26.9 ± 0.1 kcal/mol). The data were combined with the results of theoretical calculations (MP3/6-311++G**//RHF/6-31G*) to give information on the heats of formation, heats of hydrogenation, and strain energies of a number of cyclic and acyclic imines. The "resonance energy" of pyridine was estimated based on these data.

1. Introduction

A detailed analysis of the thermochemically derived "resonance energy" of heterosubstituted benzenes requires a knowledge of the energy changes that occur in each step of the hydrogenation to the saturated counterpart. Thus, for benzene, the available data[1] allow the calculation of the energies for the following steps:



The clear difference between the first step and the subsequent ones provides the strongest evidence for the usual stability of benzene. The stabilization energy of benzene is then given by the difference between the first and last step, corrected for the stabilization energy of the diene (1.8 kcal/mol) or 36 kcal/mol. A similar study of pyridine would require information on the heat of reduction of the cis imine 1-azacyclohexene (**1**). Information of the reduction of imines is available only for some heavily substituted trans-imines,[2] and no cis-imine has been studied.

In order to gain information concerning the heats of hydrogenation of compounds such as **1**, we have carried out an experimental study making use of reaction calorimetry; we have also carried out a theoretical study of **1** and a number of related compounds.

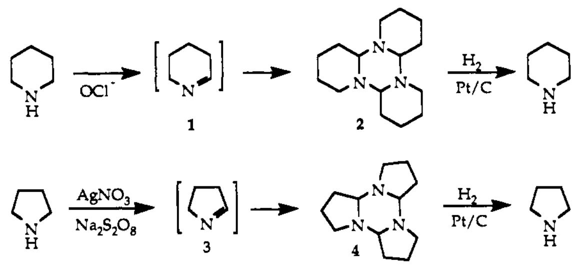
2. Results of the Calorimetric Study

It is possible to prepare **1** by the oxidation of piperidine with hypochlorite,[3] but it immediately trimerizes to form **2**. If it were possible to measure the heat of reduction of **2** to piperidine, and measure the equilibrium constant between **1** and **2**, the heat of reduction of **1** could be calculated. The hydrogenation of **2** does occur in acetic acid solution with platinum as the catalyst to form piperidine. The heat of reduction was determined to be −30.2 ± 0.5 kcal/mol (Table I). However, attempts to measure the thermally induced equilibrium between **1** and **2** failed, and no significant vapor pressure was observed for **2** at temperatures as high as 220 °C.

Table I. Heats of Hydrogenation and Solution in AcOH

compound	ΔH_{rxn} (cal/mol)	no. of runs
1-azacyclohexene trimer (**2**)	−81 550 ± 493[a]	5
pyrroline trimer (**4**)	−78 160 ± 63	7
piperidine	−17 120 ± 47	5
pyrrolidine	−17 710 ± 43	5

[a] The uncertainties are given as 2s̄.



It is possible to prepare 1-pyrroline (**3**) by catalytic oxidation of pyrrolidine with silver ion.[4] It also rapidly trimerizes to form **4**. The trimer does undergo hydrogenation to pyrrolidine in acetic acid with platinum as the catalyst, and the heat of hydrogenation was measured. The heat of solution of pyrrolidine in acetic acid also was determined. The combination of these data gives the heat of hydrogenation of liquid trimer to liquid pyrrolidine as −25.0 ± 0.1 kcal/mol (Table I). In this case it was possible to measure the equilibrium constant between **3**(g) and **4**(l) by measuring the vapor pressure of **3** and **4** as a function of temperature (Table II). It has been shown that **4** does not exist in the vapor phase.[5] The equilibrium constant is then $K = P^3/\chi_T$, where P is the vapor pressure of the monomer and χ_T is the mole fraction of the trimer in the liquid phase. The latter was determined by examining the NMR spectra of neat trimer at various temperatures (Table III). A plot of $3 \ln P - \ln \chi_T$ against $1/T$ gave a linear relationship whose slope was $-\Delta H/R$ at the mean temperature. ΔH at 328.5 K thus derived was 48.1 ± 0.3 kcal/mol.[6] The heat of vaporization is temperature dependent and can be corrected to 298 K using the heat capacities of the trimer and monomer. The C_p of the liquid trimer was estimated to be 102.7[7] and that for 3 mol of the monomer as a gas was calculated[8] to be 55.3 cal/mol·K using the scaled HF/6-31G* vibrational frequencies. The correction is 1.44 kcal/mol giving ΔH_r = 49.6 ± 0.5[6] kcal/mol at 298 K. The combination of the

---

[1] Conant, J. B.; Kistiakowsky, G. B. *Chem. Rev.* **1937**, *20*, 181.

[2] Häfelinger, G.; Steinmann, L. *Angew. Chem., Int. Ed. Engl.* **1977**, *16*, 47.

[3] Schöpf, C.; Komzak, A.; Braun, F.; Jacobi, E. *Annalen* **1947**, *559*, 7, 24.

[4] Nomura, Y.; Ogawa, K.; Takeuchi, Y.; Tomoda, S. *Chem. Lett.* **1977**, 693.

[5] Edwards, G. B.; Yamanouchi, K.; Kuchitsu, K.; Sugie, M.; Takeo, H.; Matsumura, C.; Ogawa, K.; Takeuchi, Y. *J. Mol. Spectrosc.* **1985**, *111*, 301.

[6] The estimated uncertainty in heats of vaporization calculated in this manner is 0.1 kcal/mol or less. This was multiplied by 3, corresponding to the cubic dependence on P, to give the estimated uncertainty. The uncertainty in the correction to 298 K is at most ±0.2 kcal/mol, and the final ΔH was given a conservative uncertainty of ±0.5 kcal/mol.

[7] The liquid-phase heat capacity was estimated using a group contribution method: Reid, R. C.; Prausnitz, J. M.; Sherwood, T. K. *The Properties of Gases and Liquids*, 3rd ed.; McGraw-Hill, New York, 1977; p 152.

[8] Janz, G. J. *Estimation of Thermochemical Properties of Organic Compounds*; Academic Press: New York, 1958; p 20.

0002-7863/93/1515-3527$04.00/0

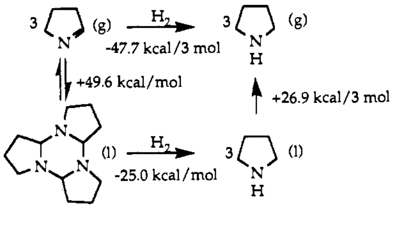
Table II. Vapor Pressure Data

P (mm Hg)	χ_T	T (K)
18.55	0.916	322.21
31.65	0.890	329.20
47.53	0.866	334.69
19.73	0.915	322.55
25.38	0.903	325.98
62.13	0.846	338.82
38.85	0.879	331.81
13.80	0.928	318.24

Table III. Mole Fraction of Trimer in the Liquid Phase

T (K)	χ_T
316.0	0.934
326.7	0.902
336.7	0.855
346.8	0.802

hydrogenation and equilibrium data along with the heat of vaporization of pyrrolidine[9] gives ΔH of hydrogenation of **3**(g) to pyrrolidine(g) as -15.9 ± 0.3 kcal/mol. The gas-phase heat of formation of pyrrolidine[9] is -0.80 ± 0.15 kcal/mol. The gas-phase heat of formation of **3** is then $+15.1 \pm 0.3$ kcal/mol. The calorimetric results are summarized in Table I, and the heat of vaporization data are given in Tables II and III.



As a check on the internal consistency of the data, the ΔH for conversion of **4**(l) to **3**(l) was calculated from the liquid-phase data via a plot of ln $(\chi_M{}^3/\chi_T)$ versus $1/T$. The slope was -12.39 giving $\Delta H = +24.6$. Combined with the data for the conversion of **4**(l) to **3**(g) leads to ΔH_v of 25.0 for 3 mol of **3**(l). This is a reasonable value in view of the observed value for 3 mol of pyrrolidine (26.9 kcal/mol) and the usually slightly lower value for unsaturated compounds as compared to their saturated analogues.

3. Experimental Section

3.1. Vapor Pressure Measurements. The vaporization enthalpy of pyrroline trimer (**4**) was determined using an apparatus similar to one described by McLaughlin and Tamres.[10] It consisted of a sample tube, connected to a mercury-filled U-tube. Two Teflon stopcocks allowed the system to be sealed or opened to vacuum through an outlet. The whole assembly was constructed within an 88-mm O-ring joint, allowing the apparatus to be placed into a glass jacket and maintained at a constant temperature with a refluxing solvent. The reflux temperature could be varied by a manostat controlling the pressure within the jacket surrounding the apparatus. In practice it was found that two manostats in series were required to maintain adequate control of the reflux temperature of the solvent. The temperature was monitored with a Hewlett-Packard quartz thermometer. It was possible to maintain the temperature to ±0.05° over the course of a measurement. The difference in height of the mercury was determined with a cathetometer, which was accurate to within ±0.05 mm.

To do a run the U-tube was filled with mercury, being careful not to allow any gas bubbles to be trapped between the mercury and the glass. The apparatus was then pumped down and gently rocked back and forth to remove all the air in the mercury. The sample was introduced by vacuum transfer or syringe. In the later case it was extremely important to submit the sample to several freeze–pump–thaw cycles to remove all traces of dissolved gas. The manostats were then set on a low pressure and the solvent was heated to reflux. Vapor pressure readings were taken every 10 min until the observed pressure did not change from the previous reading. The temperature of the refluxing solvent was increased by increasing the pressure in the jacket. The procedure was repeated until the limit of the U-tube was reached (~10 cm). The procedure was then reversed and the temperature was decreased after each measurement. If the same PVT curve was obtained for heating as for cooling and if the cell returned to the original starting pressure, then it was assumed that no significant decomposition occurred during the run.

The mole fraction of trimer in the liquid phase was determined by variable-temperature NMR spectroscopy using a sealed sample. The temperature was calibrated using ethylene glycol.[11] The NMR spectrum had bands characteristic of the monomer (δ 7.15 (s, 1 H), 3.81 (m, 2 H), 2.49 (m, 2 H), and 1.8 (m, 2 H)) and of the trimer (δ 2.98 (m, 6 H), 2.30 (m, 3 H), 1.8 (m, 12 H)). The ratio of the appropriate bands gave the mole fraction of the trimer. The receiver delay was set to 10 s to ensure proper relaxation of all components. The data were fit to $\chi_T = a + bT + cT^2$, where $a = -3.8209$, $b = 3.2699 \times 10^{-2}$, and $c = -5.5856 \times 10^{-5}$. The χ_T for the temperature at which the vapor pressure was determined was estimated from this equation.

3.2. Calorimetry. The calorimeter used in this study was an extensively remodeled LKB 8700-1 isoperibol reaction calorimeter.[12,13] Data were recorded and analyzed automatically by a microprocessor.[14] A new cell and head assembly was designed for hydrogenations. The cell consisted of a 64-mm diameter tube of approximately 250-mL capacity with a glass post for breaking the ampoule and a 6-mm port on the side for introducing solvent to the cell. On the exterior of the cell there were two hooks for securing the cell to the head. The cell was fit into the head via an O-ring vacuum fitting. The head was entirely constructed of stainless steel, except for the cap, which was Plexiglas. There were five ports through the head, leading to the interior of the cell: thermometer well, heater well, hydrogen inlet, breaker shaft, and stirrer shaft. The stirrer shaft was connected to a 600-rpm synchronous motor via a magnetic coupling, the hydrogen inlet consisted of a $^1/_8$" stainless steel tube silver soldered to the head, and the rest were connected with Cajon™ Ultra-Torr quick connects. An additional inlet led through the head to the outside of the cell so that cooling gas could be blown over the cell. The head and cell were clamped into a brass canister which was then immersed in a constant-temperature bath. The temperature of the bath could be held to ±0.0003° during the course of a run.

In a typical run the cell was charged with catalyst (PtO_2) and an ampoule containing the weighed sample was placed in the ampoule holder. The cell was then evacuated to ~0.03 Torr and filled with acetic acid by canula. Hydrogen gas was bubbled through the solvent via a needle through the side port of the cell for 3 min. Hydrogen was connected to the inlet port and blown through the cell for an additional 2 min. The cell was then sealed, filled to 5-cm positive pressure of hydrogen, and placed in the canister. The canister was placed in the bath and the cell was cooled to below the bath temperature. The starting temperature was chosen such that the reaction would bring the cell to above its equilibrium temperature. The temperature was automatically measured every 10 s. Data collection was begun at a temperature below the starting temperature, such that 30–50 points would be collected before the ampoule was broken, about 0.01–0.02°. When the starting temperature was reached, the ampoule was broken. Temperature data were collected for enough time to ensure that the reaction was complete and the cell was returning toward its equilibrium temperature. For heats of solution, 25 min was adequate. The hydrogenations were generally much slower, requiring between 50 min and 2.5 h. The cell was cooled and a calibration run was done over the same temperature range using the electric heater. This gave the heat capacity of the cell.

The measured heat of reaction corresponds to liquid trimer going to the reduced monomer in solution. The heat of solution of the reduced monomer was measured in the same fashion, and the combination of the two measurements gave the heat of reduction of the liquid trimer to the liquid reduced monomer.

---

(9) Pedley, J. B.; Naylor, R. D.; Kirby, S. P. *Thermochemical Data of Organic Compounds*, 2nd ed.; Chapman and Hall: London, 1986.

(10) McLaughlin, D. E.; Tamres, M. *J. Am. Chem. Soc.* **1960**, *82*, 5618.

(11) Amman, C.; Meier, P.; Merbach, A. E. *J. Magn. Reson.* **1982**, *46*, 319.

(12) LKB 8700-1 Precision Calorimetry System; LKB Produkter, A. B., Stockholm-Bromma I, Sweden.

(13) Sunner, S.; Wadsö, I. *Acta Chem. Scand.* **1959**, *13*, 97. A more detailed description of the calorimeter is given in the Ph.D. thesis of D.Y.N., Yale, 1992.

(14) Wiberg, K. B.; Squires, R. R. *J. Chem. Thermodyn.* **1979**, *11*, 773. Martin, E. Ph.D. Thesis, Yale University, 1984, Chapters 2 and 3. Crocker, L. S. Ph.D. Thesis, Yale University, 1990 Chapter 2.

Table IV. Calculated Energies[a]

compound	HF/6-31G*	MP3/6-31G**	MP3/6-311++G**	ZPE	$\Delta H_f{}^\circ{}_{298}$(g)
H_2	−1.126 83	−1.163 16	−1.166 26	5.9	0.0
ethene	−78.031 72	−78.339 98	−78.371 27	30.7	12.5 ± 0.1
formaldimine	−94.028 46	−94.355 99	−94.397 36	24.3	20.6 ± 2.0[b]
formaldehyde	−113.866 33	−114.188 78	−114.249 52	16.3	−26.0 ± 0.1[c]
propene	−117.071 47	−117.536 73	−117.582 49	47.8	4.8 ± 0.2
(*E*)-1-azapropene	−133.073 84	−133.557 80	−133.613 89	41.4	
$CH_3N{=}CH_2$	−133.061 49	−133.543 54	−133.598 29	41.2	
acetaldehyde	−152.915 97	−153.394 18	−153.469 13	33.6	−39.7 ± 0.1
(*Z*)-2-aza-2-butene	−172.099 12	−172.738 35	−172.807 16	58.1	
(*E*)-2-aza-2-butene	−172.105 83	−172.744 74	−172.813 99	58.1	
pyrroline (**3**)	−209.978 72	−210.743 75	−210.825 59	63.6	15.1 ± 0.3[d]
1-azacyclohexene (**1**)	−249.014 82	−249.937 80	−250.033 81	81.2	
ethane	−79.288 76	−79.572 76	−79.603 80	44.6	−20.0 ± 0.1
methylamine	−95.209 83	−95.573 87	−95.618 89	38.6	−5.5 ± 0.1
methanol	−115.035 42	−115.399 08	−115.464 03	31.0	−48.1 ± 0.1
propane	−118.263 65	−118.765 28	−118.811 12	61.9	−25.0 ± 0.1
ethylamine	−134.247 61	−134.769 61	−134.828 88	55.8	−11.3 ± 0.2
$(CH_3)_2NH$	−134.238 85	−134.757 45	−134.815 51	55.7	−4.4 ± 0.2
ethanol	−154.075 74	−154.596 64	−154.675 90	48.2	−56.2 ± 0.1
$CH_3NHCH_2CH_3$	−173.276 49	−173.952 92	−174.025 61	72.8	−10.9 ± 0.5[e]
n-butane	−157.298 41	−157.957 83		79.1	−30.0 ± 0.2
n-pentane	−196.333 10	−197.150 37		96.3	−35.1 ± 0.2
pyrrolidine	−211.145 33	−211.947 34	−212.032 68	78.3	−0.8 ± 0.2
piperidine	−250.188 71	−251.149 75	−251.249 12	96.0	−11.3 ± 0.2

[a] Total energies are given in hartrees; ZPE and ΔH_f are given in kcal/mol. The ZPE were obtained from HF/6-31G* calculations and scaled by 0.8934. The observed ΔH_f's are taken ref 9. [b] Based on G2 calculations (ref 23). [c] There are two apparently precise values for the heat of formation of formaldehyde: −26.0 ± 0.1 kcal/mol and −28.7 ± 0.1 kcal/mol (ref 9). The former appears to be preferred. [d] This work. [e] The ΔH_f of methylethylamine was taken as the average of that for dimethylamine and diethylamine.

4. Calculations

Ab initio calculations provide a useful tool for examining compounds when there are limited experimental data for comparison. For the current study all structures were optimized at the RHF/6-31G* level of theory[15] using the Gaussian 90 series of programs[16] (Tables IV and V). Vibrational frequencies were calculated at the same level of theory and scaled by 0.8934 to obtain zero-point energies, so that energies of reaction could be converted to enthalpies. Electron correlation was included up to the MP3[17] level using both the 6-31G** and 6-311++G** basis sets. The latter basis is valence triple-ζ with diffuse and polarization functions added at both heavy atoms and hydrogen. We have found previously[18] that fairly reliable relative energies can be obtained at the MP3/6-31G** level with a careful choice of isodesmic reactions and that MP3/6-311++G** is generally reliable for predicting isodesmic reactions to within ±2 kcal of experiment.

Both path angles were derived from the theoretical calculations and are defined as the angle between the paths of maximum electron density connecting the nuclei.[19] For all molecules without symmetry the bond path angles will deviate from the geometric angles (defined by the line connecting the atomic centers); i.e., there will be some degree of bond bending. The amount of bond bending is representative of the magnitude of the forces distorting an atom away from its preferred geometry.[20]

5. Discussion

Any experimental determination is subject to a measure of uncertainty. While careful control of the experimental conditions can often assure good precision, the accuracy can still be questionable. Heats of formation are only widely known for hydrocarbons and to a lesser extent for oxygenated compounds and amines. For imines, heat of formation data are nearly nonexistent, and only a few gas-phase determinations have been made. This means that there is no framework of data available from different laboratories to determine whether a particular measurement is reasonable. For example, the heat of formation of the parent imine, formaldimine, has been the subject of several studies.[21] The published heats of formation range from 16.5 ± 2 kcal/mol to 32.2 ± 3 kcal/mol. Obviously the true value falls outside several of the authors' error bars. In the absence of experimental values with which to compare, ab initio calculation can serve the role of determining whether a particular measurement is "in the ballpark".

For formaldimine and a few of the other small imines, the heats of formation that exist in the literature are determined by schemes such as the following:

$$\Delta H_f{}^\circ(H_2C{=}NH) = PA - \Delta H_f{}^\circ(H^+) + \Delta H_f{}^\circ(H_2C{=}NH_2{}^+) \quad (1)$$

The heat of formation of a proton is well known, but the proton affinity and the heat of formation of the protonated species are not easily determined and are very large numbers. Relatively small errors in these lead to large uncertainties in the heat of formation of the neutral species. Fortunately, molecules of this size are amenable to very good theoretical treatment. Smith, Pople, Curtiss, and Radom[22] have determined the heat of formation of formaldimine through use of G2 theory[23] which is effectively a calculation at the QCISD(T)/6-311+G(3df,p)//MP2/6-31G(d) + ZPE level of theory. G2 predicts atomization energies, ionization energies, electron affinities, and proton

---

(15) Hariharan, P. C.; Pople, J. A. *Theor. Chim. Acta* **1973**, *28*, 213. Franci, M. M.; Pietro, W. J.; Hehre, W. J.; Binkley, J. S.; Gordon, M. S.; Defrees, D. J.; Pople, J. A. *J. Chem. Phys.* **1982**, *77*, 3654.

(16) Frisch, M. J.; Head-Gordon, M.; Trucks, G. W.; Foresman, J. B.; Schlegel, H. B.; Raghavachari, K.; Robb, M.; Binkley, J. S.; Gonzalez, C.; Defrees, D. J.; Fox, D. J.; Whiteside, R. A.; Seeger, R.; Melius, C. F.; Baker, J.; Martin, R. L.; Kahn, L. R.; Stewart, J. J. P.; Topiol, S.; Pople, J. A. *Gaussian 90*, Revision F; Gaussian Inc.: Pittsburgh, PA, 1990.

(17) Møller, C.; Plesset, M. S. *Phys. Rev.* **1934**, *46*, 618. Pople, J. A.; Binkley, J. S.; Seeger, R. *Int. J. Quantum. Chem. Symp.* **1986**, *10*, 1.

(18) Wiberg, K. B.; Nakaji, D.; Breneman, C. M. *J. Am. Chem. Soc.* **1989**, *111*, 4178.

(19) Bader, R. F. W. *Acc. Chem. Res.* **1985**, *9*, 18. Bader, R. F. W. *Atoms in Molecules. A Quantum Theory*; Clarendon Press: Oxford, 1990.

(20) Bader, R. F. W.; Essén, H. *J. Chem. Phys.* **1984**, *80*, 1943. Bader, R. F. W.; Tang, T.-H.; Tal, Y. Biegler-König, F. W. *J. Am. Chem. Soc.* **1982**, *104*, 940.

(21) Defrees, D. J.; Hehre, W. J. *J. Phys. Chem.* **1978**, *82*, 391. Tarasenko, N. A.; Tishenko, A. A.; Zaikin, V. G.; Volkova, V. V.; Gusel'nikov, L. E. *Izv. Akad. Nauk. SSSR, Ser. Khim.* **1986**, 2397. Grela, M. A.; Colussi, A. J. *Int. J. Chem. Kinet.* **1988**, *20*, 713. Lias, S. G.; Bartmess, J. E.; Kiebman, J. F.; Holmes, J. L.; Levin, R. D.; Mallard, W. G. *J. Phys. Chem. Ref. Data* **1988**, *17*, Suppl 1. Peerboom, R. A. L.; Ingemann, S.; Nibbering, N. M. M.; Liebman, J. F. *J. Chem. Soc., Perkin Trans. 2* **1990**, *2*, 1825.

(22) Smith, B. J.; Pople, J. A.; Curtiss, L. A.; Radom, L. *Aust. J. Chem.* **1992**, *45*, 285.

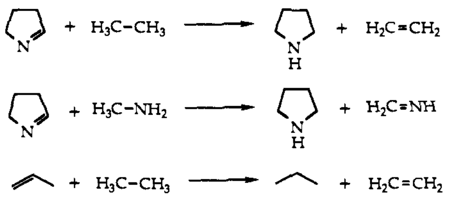
(23) Curtiss, L. A.; Raghavachari, K.; Trucks, G. W.; Pople, J. A. *J. Chem. Phys.* **1991**, *94*, 7221.

Table V. Calculated (RHF/6-31G*) and Observed Geometries

	parameter	calcd	exptl[a]
formaldimine	r(C=N)	1.2505	1.273(4)
	r(N—H)	1.0061	1.020(20)
	r(C—H_c)	1.0840	1.092(20)
	r(C—H_t)	1.079	1.092(20)
	∠(HNC)	111.45	110.5(15)
	∠(NCH_{cis})	124.69	125.1(20)
	∠(NCH_{trans})	119.21	117.9(20)
N-methylformaldimine	r(C=N)	1.2466	1.30(2)
	r(N—C)	1.4439	1.44 (assumed)
	∠(CNC)	118.57	116.9(15)
(*E*)-1-azapropene	r(N=C)	1.2527	1.273 (assumed)
	r(C—C)	1.4996	1.510(10)
	r(H—N)	1.0056	1.021 (assumed)
	∠(CCN)	121.98	121(1)
	∠(HNC)	111.42	110.4 (assumed)
(*Z*)-1-azapropene	r(N=C)	1.2522	1.273 (assumed)
	r(C—C)	1.5049	1.525(10)
	r(H—N)	1.0076	1.021 (assumed)
	∠(CCN)	127.75	126(1)
	∠(HNC)	111.53	110.4 (assumed)
(*E*)-2-aza-2-butene	r(C—N)	1.4431	1.43(2)
	r(N=C)	1.2483	1.30 (assumed)
	r(C—C)	1.5003	1.50(2)
	∠(CNC)	118.78	118(1)
	∠(NCC)	122.42	121.5(10)
(*Z*)-2-aza-2-butene	r(C—N)	1.4432	
	r(N=C)	1.2530	
	r(C—C)	1.5097	
	∠(CNC)	123.27	
	∠(NCC)	131.19	
pyrroline	r(N=C_2)	1.2517	1.306(2)
	r(C_2—C_3)	1.5109	
	r(C_3—C_4)	1.5386	
	r(C_4—C_5)	1.5443	
	r(C_5—N)	1.461	
	∠(C_5NC_2)	109.33	
	∠(NC_2C_3)	116.71	
	∠($C_2C_3C_4$)	101.01	
	∠($C_3C_4C_5$)	103.16	
	∠(C_4C_5N)	106.73	
	τ($NC_2C_3C_4$)	−12.83	
	τ($C_2C_3C_4C_5$)	41.45	
	τ($C_3C_4C_5N$)	−58.15	
1-azacyclohexene	r(N=C_2)	1.2513[b]	
	r(C_2—C_3)	1.5122	
	r(C_3—C_4)	1.5280	
	r(C_4—C_5)	1.5266	
	r(C_5—C_6)	1.5279	
	r(C_6—N)	1.4547	
	∠(CNC)	120.11	
	∠(NC_2C_3)	127.54	
	∠($C_2C_3C_4$)	111.77	
	∠($C_3C_4C_5$)	108.98	
	∠($C_4C_5C_6$)	109.88	
	∠(C_5C_6N)	115.30	

[a] The observed values were taken from *Landolt-Bornstein*; Springer-Verlag: Berlin; Vol. II/7 (1976) and Vol. II/15 (1987). [b] The structure of 1-azacyclohexene has recently been studied at the MP2/6-31G* level: Bachrach, S. M.; Liu, M. Tetrahedron *Lett.* **1992**, *33*, 6771.

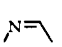
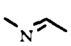
affinities to within ±2 kcal/mol of experiment. Their predicted $\Delta H_f°_{298}$(g) for the parent imine is 20.6 ± 2 kcal/mol. This serves as a starting point with which we can now compare our experimental results, and the results of our more modest calculations.



	Exp.	Calc[a]	
	16.6±0.4	16.8	(2)
	10.2±2.0	9.5	(3)
	2.7±0.3	2.6	(4)

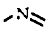
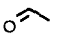
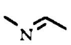
[a]MP3/6-311++G**//RHF/6-31G*+ZPE.

Table VI. Hydrogen Transfer Energies from Ethane[a] R_1—X=CHR_2 + H_3C—CH_3 → R_1—$XHCH_2R_2$ + H_2C=CH_2

compound	HF/ 6-31G*	MP3/ 6-31G**	MP3/ 6-311++G**	$\Delta H_{rxn}°$(g)
ethene	0.0	0.0	0.0	0.0
H_2C=NH	10.2	9.7	7.3	6.4 ± 2.0
H_2C=O	18.3	14.9	12.1	10.4 ± 0.2
propene	3.2	3.6	2.6	2.7 ± 0.2
	15.1	13.7	11.5	
CH_3N=CH_2	12.9	12.4	10.2	
	24.1	19.7	16.9	16.0 ± 0.2
	13.1	12.2	9.6	
	17.4	16.2	13.9	
pyrroline	19.9	19.1	16.8	16.6 ± 0.3
1-azacyclohexene	15.4	14.0	11.7	

[a] The experimental data were taken from ref 9.

Table VII. Heats of Hydrogenation[a] R_1—X=CHR_2 + H_2 → R_1—$XHCH_2R_2$

compound	HF/ 6-31G*	MP3/ 6-31G**	MP3/ 6-311++G**	$\Delta H_{rxn}°$(g)
ethene	−36.1	−35.8	−33.6	−32.5 ± 0.2
imine	−25.8	−25.9	−26.3	−26.1 ± 2.0
propene	−32.8	−32.8	−30.9	−29.8 ± 0.2
formaldehyde	−17.7	−20.8	−21.5	−22.1 ± 0.2
	−21.0	−22.0	−22.1	
	−23.1	−23.2	−23.4	
	−12.0	−16.0	−16.7	−16.5 ± 0.2
	−22.9	−23.5	−23.9	
	−18.7	−19.5	−19.7	
pyrroline	−16.2	−16.6	−16.8	−15.9 ± 0.3
1-azacyclohexene	−20.6	−21.7	−21.9	

[a] The experimental data were taken from ref 9.

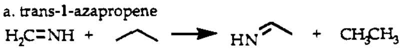
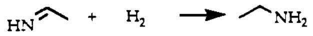
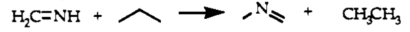
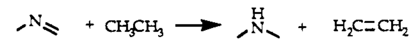
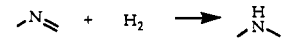
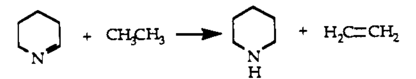
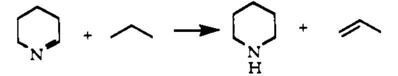
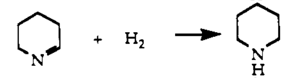
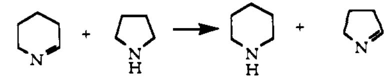
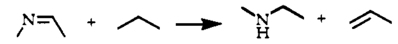
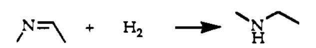
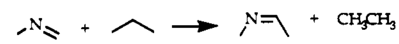
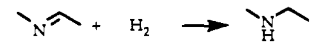
Our study of pyrroline, **3**, is the first in which the heat of hydrogenation of a cis-imine has been measured, and the first imine whose gas-phase heat of formation has been determined without going through the ions. As can be seen from eqs 2–4, the calculated enthalpies for the selected isodesmic reactions agree quite well with experiment.

6. Estimating the Heats of Formation of Imines

The good agreement between the experimental and theoretical values means we can estimate the heats of formation of some of the imines. In particular, we want to know the heat of formation of **1**. It is necessary to know the heat of hydrogenation of **1**, if the resonance energy of pyridine or any of the other azines is to be estimated. From the data given in Tables VI and VII, we can see that the MP3/6-311++G** reproduce hydrogen-transfer energies with an average error 0.7 ± 0.5 kcal/mol and hydrogenation enthalpies with an average error of −0.5 ± 0.4 kcal/mol. The MP3/6-31G** values are a little worse, being 2.8 ± 0.8 kcal/mol and −0.8 ± 1.4 kcal/mol, respectively. Though the number of data points is limited, it is important to note that the errors in the MP3/6-311++G** hydrogen-transfer energies are comparable to the uncertainty of the experimental values. The MP3/6-31G** calculations show a larger deviation from experiment. They are systematically too endothermic for the hydrogen transfers and too exothermic for the hydrogenations. With this in mind, we can now estimate the heats of formation of the imines.

Table VIII shows isodesmic reactions where only one heat of formation is not known. The $\Delta H_f°$(est) value is heat of formation of this species estimated using the best calculated value. Only

Table VIII. Isodesmic Reactions for Estimating Heats of Formation[a]

	RHF	MP3[a]	MP3[b]	$\Delta H°_f$[c]
a. trans-1-azapropene				
	-6.8	-6.0	-6.0	9.6
	15.1	13.7	11.5	9.7
	11.9	10.8	8.9	9.6
	-21.0	-22.0	-22.1	10.8
b. 2-azapropene				
	0.8	2.7	3.6	19.2
	12.9	12.4	10.2	17.9
	9.7	9.6	7.6	17.8
	-23.1	-23.2	-23.4	19.0
c. 1-azacyclohexene				
	15.4	14.0	11.7	9.5
	12.2	11.1	9.1	10.1
	-20.6	-21.7	-21.9	10.6
	-4.5	-5.1	-5.1	9.7
d. Z-2-aza-2-butene				
	13.1	12.2	9.6	12.0
	9.9	9.4	7.0	11.9
	-22.9	-23.5	-23.9	13.0
	-2.1	-1.8	-1.4	12.0
	5.4	6.9	8.2	12.8
e. E-2-aza-2-butene				
	17.4	16.2	13.9	7.7
	14.1	13.4	11.3	7.6
	-18.7	-19.5	-19.7	8.8
	-6.3	-5.8	-5.9	7.5
	1.2	2.9	3.9	8.5

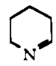
[a] The experimental heats of formation were taken from ref 9, and are listed in Table I. [b] MP3/6-31G**. [c] MP3/6-311++G**. Six Cartesian d functions were used. [d] Estimated $\Delta H_f°$ of unknown using the MP3/6-311++G** level for the enthalpy of reaction.

the $\Delta H_f°$ estimated from the hydrogenation reactions differs significantly from the others. As noted previously, the calculated hydrogenation reactions tend to be too exothermic by 0.5 kcal/mol at the MP3/6-311++G** level and 0.8 kcal/mol at the MP3/6-31G** level. Taking this into account brings this estimate into line with the others. The estimated heats of formation should be accurate to about ±1 kcal/mol for the 6-311++G** values and ±2 kcal/mol for the 6-31G** values. Table IX summarizes these results.

7. Effects of Nitrogen Substitution in Five- and Six-Membered Rings

We now have sufficient data to compare the effects of nitrogen substitution on the properties of the five- and six-membered rings.

Table IX. Calculated Heats of Formation of Imines, kcal/mol

compound	$\Delta H_f°_{298}$
$H_2C=NH$	20.6
	9.6
	18.3
	12.2
	7.8
	15.1[a]
	9.5

[a] Experimental value, ±0.3 kcal/mol.

Table X. Estimated Strain Energies of Five- and Six-Membered Rings

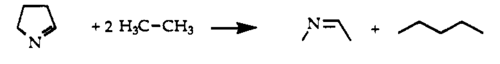
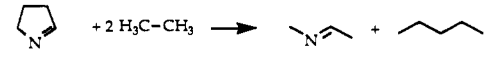
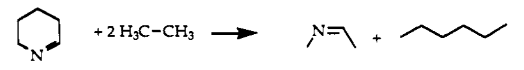
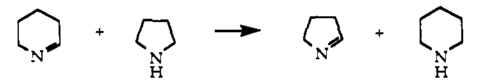
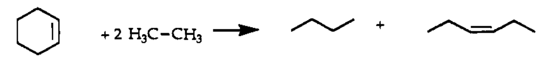
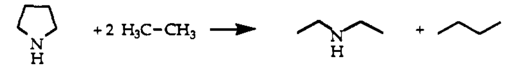
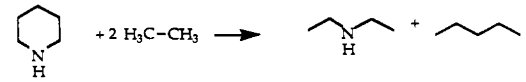
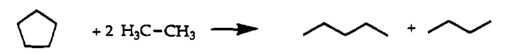
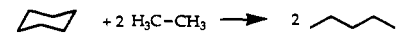
	ΔH°
	+2.0
	-2.4
	+2.8
	-4.9
	-4.7
	-5.7
	-0.2
	-1.9
	-6.5
	-1.1
	-6.8
	-0.7

Table X shows isodesmic reactions where the ring is opened to unstrained acyclic products. The negatives of these energies are the strain energies of the cyclic compounds. For the saturated compounds the effect of nitrogen substitution is negligible. Both pyrrolidine and cyclopentane have about 6.5 kcal/mol of strain, and piperidine and cyclohexane are essentially unstrained. The situation is quite different when unsaturation is added to the ring. Cyclopentene is less strained relative to an acyclic alkene than cyclopentane is relative to an acyclic alkane. Cyclohexene is unstrained relative to a cis alkene. For the imines, the isodesmic reactions converting **1** and **3** to acyclic compounds are *endothermic* relative to a cis imine. The cis imine is less stable than the trans imine by 4.4 kcal/mol.

Examination of the bond path angles about the imine nitrogen in Table XI offers an explanation of the destabilization of the cis imine. The bond path angles about nitrogen for the N-substituted imines are all 106° except for (*Z*)-2-aza-2-butene. For this compound the bond path angle has widened considerably to 109.6°

Table XI. Comparison of Geometric and Bond Path Angles

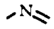
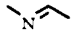
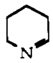
compound	angle:	CNC(H)	bpa	NCC(H)	bpa
$H_2C{=}NH$		111.45	102.90	124.69	122.84
		111.42	103.40	121.98	124.74
		118.57	106.1		
		118.78	106.38	122.42	125.59
		123.27	109.56	131.19	126.25
		113.39	106.86	116.71	116.67
		120.11	106.94	127.54	124.04

Table XII. MM3 Calculations

	$\Delta H_f°$ obs	MM3 calc	MM3[a] Strain E.
cyclopentene	8.2	9.1	5.0
cyclohexene	−1.2	−1.2	−0.4
cyclopentane	−18.4	−18.9	5.7
pyrrolidine	−0.8	−0.9	6.8
cyclohexane	−29.5	−30.0	−0.4
piperidine	−11.3	−11.8	0.8
pyrroline trimer		33.2	11.5
pyrroline trimer 1 axial		33.6	11.9
pyrroline trimer 2 axial		38.0	16.3
1-azacyclohexene trimer all eq		6.3	−0.6
1-azacyclohexene trimer 1 axial		5.1	−1.8
1-azacyclohexene trimer 1 boat		9.7	2.8
1-azacyclohexene trimer 2 axial		14.2	7.3
cyclopentene trimer all eq		−34.3	16.6
cyclopentene trimer 1 axial		−33.5	17.4
cyclohexene trimer all eq		−65.8	−0.1
cyclohexene trimer 1 axial		−63.5	2.2

[a] Based on unstrained models using Franklin's group equivalents (CH_2 = −4.926, CH = −1.09, cis HC=CH = +18.88, NH = −12.0, N = −23.1 kcal/mol). Franklin, J. L. *Ind. Eng. Chem.* **1949**, *41*, 1070.

and the geometric angle is even larger at 123.3°. Obviously the steric interactions in this molecule are the most severe, although all the imines show a fairly large degree of bond bending. Pyrroline shows the least distortion about the nitrogen, and this is consistent with the fact that it shows the greatest stabilization relative to its reference compounds.

The difference between the cycloalkyl series and the azacycloalkyl series is perhaps best seen in the hydrogen-transfer reactions between the five- and six-membered rings. In the cycloalkyl series, the reaction of cyclohexene is exothermic by −1.9 kcal/mol, whereas the reaction of 1-azacyclohexene is exothermic by −4.9 kcal/mol. This results from the small difference in strain between pyrroline and 1-azacyclohexene, on one hand, and the larger difference between cyclopentene and cylcohexane.

8. Trimers

Trimer **2** has a greater thermal stability than trimer **4**. We can see from the isodesmic reaction transferring hydrogen between **1** and pyrrolidine to give **3** and piperidine that it is favorable by 4.9 kcal/mol to have the double bond in the five-membered ring. This is due in large part to the greater strain energy in pyrrolidine relative to piperidine. There may also be a difference in strain energy between the trimers. It is not practical to study them by ab initio calculations, but it is reasonable to examine them using the molecular mechanics program MM3 (Table XII).

MM3 has been parameterized for amines[24] and alkanes[25] among other compounds. Trimer **2** has two low-energy conformations. The lowest energy conformation has one of the six-membered rings twisted axially relative to the central ring. The second conformation has all the rings in equatorial positions. An NMR study[26] determined that the conformation with one ring axial must be favored by at least 0.6 kcal/mol. MM3 finds that this conformation is favored by 1.2 kcal/mol, in good agreement with the experiment. The calculated strain energy of the conformation is −1.8 kcal/mol.

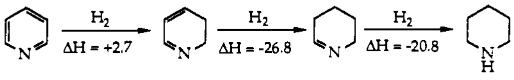
No experimental data exist on the conformation of the pyrroline trimer in solution. As with **2** there are two low-energy conformations, one where all the rings are equatorial and one with one ring axial. MM3 predicts the two conformations to be of nearly identical energy. The all-equatorial conformation is slightly favored, but only by 0.4 kcal/mol. The strain energy of the lower energy conformation is 11.5 kcal/mol.

The analogous hydrocarbons have strain energies of 16.6 versus 11.5 (**4**) and −0.1 versus −1.8 (**2**). In both cases the trimeric amines are less strained than their alkane counterparts. The estimated heats of trimerization of 3 mol of pyrroline and 1-azacyclohexene in the gas phase are −12 kcal/mol and −23 kcal/mol, respectively. These data combined with the experimental energy of **4**(l) going to **3**(g) give a ΔH_v for **4** of 37.6 kcal/mol. Heats of vaporization increase by about 1.1 kcal/mol per methylene group, leading to an estimated ΔH_v for **2** of 41 kcal/mol. Then, the ΔH for the conversion of **2**(l) to **1**(g) would be 64 kcal/mol. In view of this large value, it is not suprising that no significant amount of **1** was found when **2** was heated.

The entropy change in the trimerization of **1** and **3** in the gas phase may be estimated to be about −105 and −108 eu, respectively,[27] giving $-T\delta S$ at 298 K as 31 and 32 kcal/mol. The free-energy changes for the trimerizations in the gas phase at 298 K would then be +8 kcal/mol for **1** and +20 kcal/mol for **3**. It can be seen that the trimers would not be present in significant amount in the gas phase.

9. "Resonance Energy" of Pyridine

These data allow us to make an estimate of the heat change for each step in the hydrogenation of pyridine. The overall



$\Delta H_{\mathrm{hydrog}}$ is −44.9 ± 0.2 kcal/mol.[9] The ΔH for the last step is given by the data in Tables I and IX as −20.8 ± 1.0 kcal/mol. The ΔH for the second step might reasonably be taken as about the same as that for 1,3-cyclohexadiene. The heat change for the first step is then +2.7 kcal/mol with an uncertainty of about 2 kcal/mol. Thus, a comparison with the first reduction step for benzene indicates that pyridine has a slightly lower "resonance energy" than benzene, in good accord with some other estimates.[18]

10. Conclusions

We have determined the heat of hydrogenation of the cis-imine **3**. Using ab initio calculations, we have been able to predict the heats of formation of several other imines, including **1**, which serves as a reference compound for predicting the resonance energy of pyridine and the other azines. The data indicate that the "resonance energy" of pyridine is about 2 kcal/mol less than that of benzene.

Acknowledgment is made to the donors of the Petroleum Research Fund, administered by the American Chemical Society, for the support of this research.

---

(24) Schmitz, L. R.; Allinger, N. L. *J. Am. Chem. Soc.* **1990**, *112*, 8307.

(25) Allinger, N. L.; Yuh, Y. H.; Lii, J. H. *J. Am. Chem. Soc.* **1989**, *111*, 8551. Cf.: Allinger, N. L.; Li, F.; Yan, L. *J. Comput. Chem.* **1990**, *11*, 848, 868.

(26) Kessler, H.; Möhrle, H.; Zimmermann, G. *J. Org. Chem.* **1977**, *42*, 66.

(27) Using MM3, the entropies are estimated to be: **1**, 76 eu; **2**, 122 eu; **3**, 73 eu; and **4**, 111 eu.