THE EFFECT OF INITIAL MIXTURE TEMPERATURE ON BURNING VELOCITY

Sir:

The effect of initial mixture temperature on the burning velocity of ethylene–air mixtures has been determined as an extension of the work done with propane–air mixtures.[1] The burning velocities were based on the inner edge of the shadow cast by the Bunsen cone and were computed by a total area method.[1] The data were smoothed by a method in which flame surfaces were related to heights, similar to the graphical method of Bollinger and Williams.[2] The maximum burning velocity was obtained with mixtures containing approximately 7.5% ethylene by volume for each of the initial temperatures, which ranged from 34 to 344°. These results are plotted in Fig. 1, along with curves predicted on a relative basis by the thermal theory as presented by Semenov[3] and by the square root law of Tanford and Pease.[4] These theoretical equations were used in the same manner as they were used for propane,[1] except that the temperature dependence of the Tanford and Pease equation has been re-evaluated and certain inconsistencies[5] have been eliminated. The corrected expression for predicting the relative effect of temperature on flame speed by the square root law is

$$u_f \propto \sqrt{(\Sigma k_1 p_1 D_{i,r}) T_o{}^2 T_f{}^{-1.33}} \qquad (1)$$

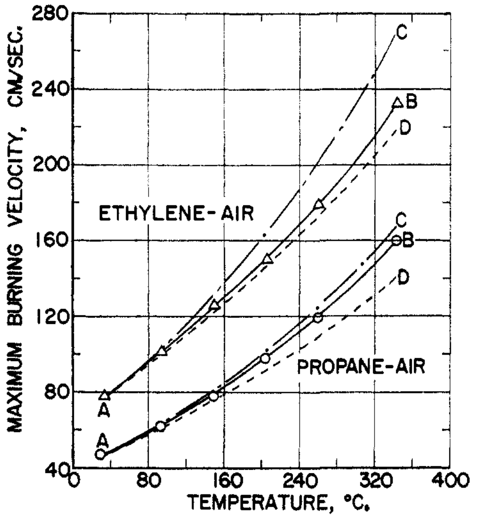

Fig. 1.—Effect of temperature on burning velocity: AB, experimental; AC, thermal theory; AD, square root law, $\Sigma_i p_i D_{i,r} = 6.5\, p_H + p_{OH} + p_O$.

The temperature dependence of k_i, the rate constant for interaction between active particles and combustible material, has been neglected in plotting the curves in Fig. 1. Constant values of k_i were used by Tanford[6] for variations in flame temperature of the order of magnitude considered in this paper. If the temperature dependence of k_i is represented by $\exp(-E/RT_f)$, the temperature effect on flame speed due to k_i will be small, since the activation energies for active-radical–fuel-molecule reactions are small. For example, using activation energies for reactions between hydrogen atoms and hydrocarbon molecules of 8.6 kcal./g.-mole for propane[7] and 2.6 kcal./g.-mole for ethylene[8] would raise the square-root law curves for the two gases by 10 cm./sec. and 4 cm./sec. at 344°, respectively. From another viewpoint, the activation energies which are required to give the best fits to the experimental curves (geometric mean values of computations for 5 temperature increments) are 9.9 and 8.3 kcal. for propane and ethylene, respectively.

The author gratefully acknowledges the suggestion of Dr. Robert N. Pease that the Tanford-Pease equation should be extended by substituting an Arrhenius expression for the rate constant.[9]

---

[1] Dugger, THIS JOURNAL, **72**, 5271 (1950).

[2] Bollinger and Williams, Natl. Adv. Comm. Aeron. Report No. 932, 1949.

[3] Semenov, *Progress of Phys. Sci. (USSR)*, **24**, No. 4 (1940). Translation appears as Natl. Adv. Comm. Aeron. Technical Memorandum No. 1026, 1942. The thermal theory curves are based on the equation for bimolecular reactions, since the same equation predicts that flame velocity should be independent of pressure, as was found for ethylene by Linnett according to Gaydon and Wolfhard, *Fuel*, **29**, 15 (1950), and for propane by Huebner and Wolfhard, Volkenrode Monograph. Die Verbrennung in Grossen Hoehen, Issued by Technical Information Bureau for Chief Scientist, Ministry of Supply, 1946.

[4] Tanford and Pease, *J. Chem. Phys.*, **15**, 861 (1947).

[5] Tanford and Pease assumed diffusion coefficients to be proportional to T^2 rather than $T^{1.67}$ as used by the author and had converted diffusion coefficients at combustion zone temperature to those at room temperature on this basis. Further, they assumed the mean combustion zone temperature to be proportional to flame temperature, T_f, hence concentrations should be proportional to T_f, rather than the arithmetic average $(T_o + T_f)/2$ used by the author.

[6] Tanford, "Third Symposium on Combustion, Flame and Explosion Phenomena," The Williams and Wilkins Company, Baltimore, Md., 1949, pp. 140–146.

[7] Steacie, *J. Chem. Phys.*, **16**, 361 (1948).

[8] Computed from data of Melville and Robb, *Proc. Roy. Soc. (London)*, **A196**, 445 (1949), assuming a steric factor of 0.1.

[9] Dr. Hartwell F. Calcote's helpful criticism is also acknowledged.

LEWIS FLIGHT PROPULSION LABORATORY
GORDON L. DUGGER
NATIONAL ADVISORY COMMITTEE FOR AERONAUTICS
CLEVELAND, OHIO

RECEIVED FEBRUARY 26, 1951

TESTS WITH LYXOFLAVIN FOR VITAMIN ACTIVITY

Sir:

L-Lyxoflavin has been examined biologically for vitamin activity. In a rat assay for unidentified vitamins in liver and other source materials, L-lyxoflavin has shown growth-promoting or vitamin-like activity.

In 1949, Pallares and Garza[1] isolated from human heart myocardium a pentose-flavin, which was found to be identical with synthetic L-lyxoflavin. Lyxoflavin differs from riboflavin only in the configuration of the groups about C_4 of the pentose side chain. The significance of the presence of lyxoflavin in human heart muscle was not discussed[1], but another paper[2] considers it as "stereoisomerism of vitamin B_2."

α-Ribazole[3] (1-α-D-ribofuranosido-5,6-dimethylbenzimidazole) is a unit of the vitamin B_{12} molecule. Vitamins B_2 and B_{12} and lyxoflavin contain a 1,2-diamino-4,5-dimethylbenzene moiety linked through nitrogen to a pentose. It seemed *a priori* that lyxoflavin might be a new member of the vitamin B complex with a biological role different from

---

[1] Pallares and Garza, *Arch. Biochem.*, **22**, 63 (1949).

[2] Pallares and Garza, *Arch. Inst. Cardio. Mex.*, **19**, 735 (1949).

[3] Brink, Holly, Shunk, Peel, Cahill and Folkers, THIS JOURNAL, **72**, 1866 (1950).