RATE OF REACTION
BETWEEN MOLECULAR HYDROGEN
AND MOLECULAR OXYGEN

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16. Abstract

The shock tube data of Jachimowski and Houghton have been rigorously analyzed to obtain rate constants for the candidate initiation reactions $H_2 + O_2 \rightarrow H + HO_2$, $k_{01} = 1.9 \times 10^{13}$ exp(-24 100/T); $H_2 + O_2 \rightarrow H_2O + O$, $k_{02} = 4.1 \times 10^{13}$ exp(-25 400/T); and $H_2 + O_2 \rightarrow OH + OH$, $k_{03} = 2.3 \times 10^{13}$ exp(-25 200/T). (Rate constants in cm$^3$ mole$^{-1}$ sec$^{-1}$, temperature in degrees kelvin.) Reaction (01) is probably not the initiation process because the activation energy obtained is less than the endothermicity and because the derived rates greatly exceed values inferred in the literature from the reverse of reaction (01). Reactions (02) and (03) remain as possibilities, with reaction (02) slightly favored on the basis of steric and statistical considerations. The solution of the differential equations is presented in detail to show how the kinetics of other ignition systems may be solved.
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SUMMARY

The shock tube data of Jachimowski and Houghton have been rigorously analyzed to obtain rate constants for the candidate initiation reactions: 

\[ \text{H}_2 + \text{O}_2 \xrightarrow{k_1} \text{H} + \text{HO}_2, \]

\[ k_{01} = 1.9 \times 10^{13} \exp(-24 100/T); \]

\[ \text{H}_2 + \text{O}_2 \xrightarrow{k_{02}} \text{H}_2\text{O} + \text{O}, \]

\[ k_{02} = 4.1 \times 10^{13} \exp(-25 400/T); \]

and 

\[ \text{H}_2 + \text{O}_2 \xrightarrow{k_{03}} \text{OH} + \text{OH}, \]

\[ k_{03} = 2.3 \times 10^{13} \exp(-25 200/T). \]

(Rate constants in \( \text{cm}^3 \text{ mole}^{-1} \text{ sec}^{-1} \), temperature in degrees kelvin.) Reaction (01) is probably not the initiation process because the activation energy obtained is less than the endothermicity and because the derived rates greatly exceed values inferred in the literature from the reverse of reaction (01). Reactions (02) and (03) remain as possibilities, with reaction (02) slightly favored on the basis of steric and statistical considerations. The solution of the differential equations is presented in detail to show how the kinetics of other ignition systems may be solved.

INTRODUCTION

When a mixture containing hydrogen and oxygen is subjected to a temperature and pressure pulse in a shock tube, small concentrations of atoms and free radicals are first formed from molecular hydrogen and oxygen. Reactions which have been suggested include (ref. 1)

\[ \text{H}_2 + \text{O}_2 \xrightarrow{k_{01}} \text{H} + \text{HO}_2 \quad \Delta H_{298}^o/R = 28700 \text{ K} \quad \text{(01)} \]

\[ \text{H}_2 + \text{O}_2 \xrightarrow{k_{02}} \text{H}_2\text{O} + \text{O} \quad \Delta H_{298}^o/R = 900 \text{ K} \quad \text{(02)} \]

\[ \text{H}_2 + \text{O}_2 \xrightarrow{k_{03}} \text{OH} + \text{OH} \quad \Delta H_{298}^o/R = 9400 \text{ K} \quad \text{(03)} \]
The dissociation processes

\[ \text{H}_2 + M \rightarrow 2\text{H} + M \]

\[ \text{O}_2 + M \rightarrow 2\text{O} + M \]

are sufficiently endothermic that they are unimportant except, possibly, at very high temperatures.

After small concentrations of H, OH, or O are formed from reactions (01), (02), or (03), the atom and radical concentrations grow exponentially via the well-known branched-chain scheme

\[ \text{OH} + \text{H}_2 \xrightarrow{k_1} \text{H}_2\text{O} + \text{H} \]  \hspace{1cm} (I)

\[ \text{H} + \text{O}_2 \xrightarrow{k_2} \text{OH} + \text{O} \]  \hspace{1cm} (II)

\[ \text{O} + \text{H}_2 \xrightarrow{k_3} \text{OH} + \text{H} \]  \hspace{1cm} (III)

At low temperatures and high pressures the chainbreaking reaction

\[ \text{H} + \text{O}_2 + M \xrightarrow{k_4} \text{HO}_2 + M \]  \hspace{1cm} (IV)

must also be considered.

Recently Jachimowski and Houghton (ref. 2) studied the hydrogen-oxygen initiation process behind incident shocks. They analyzed their data to obtain rate constants assuming initiation by reaction (01) or (03). Their analysis used an approximate and intuitive formulation proposed to explain hydrogen-oxygen ignition delays (ref. 3).

In this report the data of Jachimowski and Houghton are reexamined, using a rigorous formulation of the initiation and chain-branching kinetics. Rate constants are obtained by assuming initiation by reaction (01), (02), or (03). The values of rate constants indicate the more likely initiation processes. The solution of the differential equations is presented in detail to provide a guide as to how the kinetics of other similar ignition systems may be solved.

THEORETICAL CONSIDERATIONS

The differential equations governing the growth of radical concentrations during the induction period are as follows (ref. 3):
\[
\frac{d[H]}{dt} = - (\nu_2 + \nu_4)[H] + \nu_1[OH] + \nu_3[O] + i_1
\]  
(1)

\[
\frac{d[OH]}{dt} = \nu_2[H] - \nu_1[OH] + \nu_3[O] + 2i_3
\]  
(2)

\[
\frac{d[O]}{dt} = \nu_2[H] - \nu_3[O] + i_2
\]  
(3)

Here \([H], [OH],\) and \([O]\) are the concentrations of hydrogen atoms, hydroxyl radicals, and oxygen atoms. Also, \(\nu_1 = k_1[H_2], \nu_2 = k_2[O_2], \nu_3 = k_3[H_2], \nu_4 = k_4[O_2][M],\) and \(i_1 = k_{01}[H_2][O_2], i_2 = k_{02}[H_2][O_2], i_3 = k_{03}[H_2][O_2],\) where \([H_2]\) and \([O_2]\) are the concentrations of molecular hydrogen and oxygen, \([M]\) is the total gas concentration, and the \(k's\) are the specific reaction rate constants for reactions (01) to (03) and (I) to (IV).

During the induction period the concentrations of \(H, OH,\) and \(O\) build up rapidly, while the concentrations of \(H_2\) and \(O_2\) are scarcely depleted. Hence, the \(\nu's\) and \(i's\) in equations (1) to (3) may be taken as constants.

The initiation rates \(i_1, i_2,\) and \(i_3\) can be eliminated from the differential equations by introducing new variables \(C_H = [H] + a_H, C_{OH} = [OH] + a_{OH},\) and \(C_O = [O] + a_O,\) where \(a_H, a_{OH},\) and \(a_O\) are constants. If these new variables are substituted into equations (1) to (3), the initiation rates are eliminated by equating the sums of the constant terms to zero:

\[
(\nu_2 + \nu_4)a_H = - \nu_1a_{OH} - \nu_3a_O + i_1 = 0
\]  
(4)

\[
-\nu_2a_H + \nu_1a_{OH} - \nu_3a_O + 2i_3 = 0
\]  
(5)

\[
-\nu_2a_H + \nu_3a_O + i_2 = 0
\]  
(6)

The new differential equations are

\[
\frac{dC_H}{dt} = - (\nu_2 + \nu_4)C_H + \nu_1C_{OH} + \nu_3C_O
\]  
(7)

\[
\frac{dC_{OH}}{dt} = \nu_2C_H - \nu_1C_{OH} + \nu_3C_O
\]  
(8)

\[
\frac{dC_O}{dt} = \nu_2C_H - \nu_3C_O
\]  
(9)
A particular solution to equations (7) to (9) is (refs. 3 and 4)

\[ C_i = A_i \exp(\lambda t), \quad i = \text{H, OH, O} \quad (10) \]

Substitution of equations (10) into equations (7) to (9) yields the relations

\[ -(\nu_2 + \nu_4 + \lambda)A_H + \nu_1 A_{\text{OH}} + \nu_3 A_O = 0 \quad (11) \]

\[ \nu_2 A_H - (\nu_1 + \lambda)A_{\text{OH}} + \nu_3 A_O = 0 \quad (12) \]

\[ \nu_2 A_H - (\nu_3 + \lambda)A_O = 0 \quad (13) \]

A nontrivial solution requires that the determinant of the coefficients of the \( A_i \) be zero, which leads to the cubic equation

\[ \lambda^3 + (\nu_1 + \nu_2 + \nu_3 + \nu_4)\lambda^2 + [\nu_1 \nu_3 + \nu_1 \nu_4 + \nu_3 \nu_4]\lambda - \nu_1 \nu_3 (2\nu_2 - \nu_4) = 0 \quad (14) \]

In the chain-branching region \((2\nu_2 > \nu_4)\) Descartes' rule of signs indicates that, if the roots of equation (14) are real, there are one positive and two negative roots. It can be shown that the roots are always real provided \(k_1 > 2k_3\), which is true for temperatures below about 2280 K, and hence true for the data of Jachimowski and Houghton (ref. 2), which are in the range 1100 to 1900 K.

The general solution to equations (7) to (9) is

\[ C_i = A_{1,i} \exp(\lambda_1 t) + A_{2,i} \exp(\lambda_2 t) + A_{3,i} \exp(\lambda_3 t) \quad i = \text{H, OH, O} \quad (15) \]

where \(\lambda_1, \lambda_2, \lambda_3\) are the roots of equation (14). We are specifically interested in the growth of hydroxyl concentration, which was the quantity monitored in reference 2. If we designate \(\lambda_1\) as the positive root and \(\lambda_2, \lambda_3\) as negative roots, then

\[ [\text{OH}] = A_{1,\text{OH}} \exp(\lambda_1 t) + A_{2,\text{OH}} \exp(\lambda_2 t) + A_{3,\text{OH}} \exp(\lambda_3 t) - a_{\text{OH}} \]

\[ \equiv A_{1,\text{OH}} \exp(\lambda_1 t) \quad (16) \]

The approximation shown is a good one, except very early in the reaction, because \(a_{\text{OH}}\) is small and the negative exponentials soon die out.

We will now evaluate the coefficient \(A_{1,\text{OH}}\). Initially, at \(t = 0\), the hydroxyl radical concentration is zero, and equation (16) gives

\[ [\text{OH}]_{t=0} = 0 = A_{1,\text{OH}} + A_{2,\text{OH}} + A_{3,\text{OH}} - a_{\text{OH}} \quad (17) \]
One can write similar expressions for hydrogen and oxygen atoms and then use any two of equations (11) to (13) to eliminate the $A_H$ and $A_O$ in favor of $A_{OH}$. The same result can be obtained more easily by successive differentiation of equations (2) and (16):

$$\left( \frac{d[OH]}{dt} \right)_{t=0} = \lambda_1 A_{1,OH} + \lambda_2 A_{2,OH} + \lambda_3 A_{3,OH} = 2i_3 \quad (18)$$

and

$$\frac{d^2[OH]}{dt^2}_{t=0} = \lambda_1 A_{1,OH} + \lambda_2 A_{2,OH} + \lambda_3 A_{3,OH}$$

$$= \nu_2 \left( \frac{d[H]}{dt} \right)_{t=0} - \nu_1 \left( \frac{d[OH]}{dt} \right)_{t=0} + \nu_3 \left( \frac{d[O]}{dt} \right)_{t=0}$$

$$= \nu_2 i_1 - 2\nu_1 i_3 + \nu_3 i_2 \quad (19)$$

Equations (17) to (19) are solved to obtain

$$A_{1,OH} = \frac{\lambda_2 \lambda_3 a_{OH} - 2i_3(\lambda_2 + \lambda_3) + \nu_2 i_1 + \nu_3 i_2 - 2\nu_1 i_3}{(\lambda_2 - \lambda_1)(\lambda_3 - \lambda_1)} \quad (20)$$

where the constant

$$a_{OH} = \frac{[2\nu_2(\nu_1 + \nu_2) + \nu_4(\nu_1 + 2\nu_3)]}{\nu_1(2\nu_2 - \nu_4)} \quad (21)$$

is obtained by solving equations (4) to (6). Expressions for $A_{2,OH}$ and $A_{3,OH}$ can be obtained by permitting the indices on $\lambda_1$, $\lambda_2$, and $\lambda_3$ in equation (20).

Thus, equations (16), (20), and (21) describe the growth of hydroxyl concentration until such time as the effects of depletion of molecular hydrogen and oxygen are important or until the temperature rises due to atom and radical recombination processes.

**ANALYSIS OF EXPERIMENTAL INDUCTION TIMES**

The induction times reported by Jachimowski and Houghton (ref. 2) correspond to the time at which the hydroxyl concentration has risen to $10^{-10}$ mole per cubic centimeter. They also report experimental values of the growth constant. Thus, one can
obtain experimental values of $A_1, OH$ from equation (16)

$$A_1, OH = 10^{-10} \exp(-\lambda \tau) \quad (22)$$

where $\lambda$ is the experimental growth constant and $\tau$ is the induction time. Rate constants for initiation by reactions (01), (02), and (03) were obtained from equation (20), assuming that only one of the initiation reactions was occurring:

$$k_{01} = \frac{A_1, OH}{[H_2][O_2]} (\lambda_2 - \lambda_1)(\lambda_3 - \lambda_1) \left[ \frac{2\nu_2 \lambda_2 \lambda_3}{\nu_1(2\nu_2 - \nu_4) + \nu_2} \right]^{-1} \quad (23)$$

$$k_{02} = \frac{A_1, OH}{[H_2][O_2]} (\lambda_2 - \lambda_1)(\lambda_3 - \lambda_1) \left[ \frac{(2\nu_2 + \nu_4)\lambda_2 \lambda_3}{\nu_1(2\nu_2 - \nu_4) + \nu_3} \right]^{-1} \quad (24)$$

$$k_{03} = \frac{2A_1, OH}{[H_2][O_2]} (\lambda_2 - \lambda_1)(\lambda_3 - \lambda_1) \left[ \frac{\nu_4 \lambda_2 \lambda_3}{\nu_1(2\nu_2 - \nu_4)} - \lambda_2 - \lambda_3 - \nu_1 \right]^{-1} \quad (25)$$

In these calculations the rate constants for reactions I, II, and III were taken from reference 5. The rate constants for reaction IV was taken from reference 6. And $\lambda_1$, $\lambda_2$, and $\lambda_3$ were obtained by solution of equation (14).

These rate constants were least-squares fitted to the Arrhenius equation

$$k = A \exp\left(-\frac{E}{RT}\right) \quad (26)$$

Results are summarized in table I, where the Arrhenius equation parameters and their standard deviations are presented.

**DISCUSSION OF RESULTS**

In this section the three candidate initiation reactions will be discussed in turn, with an indication as to which are the most likely initiation processes.

The reaction

$$H_2 + O_2 \rightarrow H + HO_2 \quad (01)$$

is a simple abstraction or two center reaction involving the breaking and formation of
one bond. Hence it is the most likely candidate based on steric considerations. However, as Jachimowski and Houghton (ref. 2) have already observed, the activation energy, 24 100 K, is substantially less than the endothermicity of 28 700 K. This amounts to more than three standard deviations in the activation energy and would seem to eliminate reaction (01), barring a temperature dependent systematic error in the experimental data.

Further, in reference 7 the rate of reaction (01) is reported to be

$$k_{01} = 5.5 \times 10^{13} \exp(-29 100/T)$$

(27)

in the range 290 to 800 K, with an uncertainty of a factor of 2.5. Equation (27) predicts a rate at 1100 K which is less than one-thirtieth of the rate at that temperature given by the constants in table I. Thus it is very unlikely that reaction (01) is the initiation process in these mixtures.

The reaction

$$\text{H}_2 + \text{O}_2 \rightarrow \text{H}_2\text{O} + \text{O}$$

(02)

might be termed a three center reaction inasmuch as new bonds are formed among three atoms. This would seem sterically more probable than the four center reaction

$$\text{H}_2 + \text{O}_2 \rightarrow \text{OH} + \text{OH}$$

(03)

In addition, reaction (02) is less endothermic than reaction (03), although in either case the endothermicities are substantially smaller than the activation energies. Finally, the standard deviations in ln k, ln A, and E/R are slightly smaller for reaction (02) than for either reaction (01) or reaction (03). Initiation in these mixtures may be due to either reaction (02) or reaction (03) (or both), with reaction (02) the slightly more probable candidate.

Lewis Research Center,
National Aeronautics and Space Administration,
Cleveland, Ohio, November 10, 1972, 502-04.
REFERENCES


<table>
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<tr>
<th>Reaction</th>
<th>$A$, cm$^3$ mole$^{-1}$ sec$^{-1}$</th>
<th>Activation energy, E/R, K</th>
<th>Standard deviations of -</th>
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<tr>
<td>$H_2 + O_2 \xrightarrow{k_{01}} H + HO_2$</td>
<td>$1.9 \times 10^{13}$</td>
<td>24</td>
<td>100</td>
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<tr>
<td>$H_2 + O_2 \xrightarrow{k_{02}} H_2O + O$</td>
<td>$4.1 \times 10^{13}$</td>
<td>25</td>
<td>400</td>
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<tr>
<td>$H_2 + O_2 \xrightarrow{k_{03}} OH + OH$</td>
<td>$2.3 \times 10^{13}$</td>
<td>25</td>
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