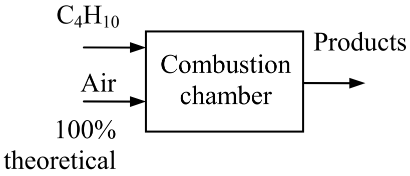
15-18 n-Octane is burned with stoichiometric amount of air. The mass fraction of each product, the mass of water in the products and the mass fraction of each reactant are to be determined.

Assumptions **1** Combustion is complete. **2** The combustion products contain CO_2, H_2O, O_2, and N_2 only.

Properties The molar masses of C, H_2, O_2 and air are 12 kg/kmol, 2 kg/kmol, 32 kg/kmol, and 29 kg/kmol, respectively (Table A-1).

Analysis The reaction equation for 100% theoretical air is



$$C_8H_{18} + a_{th}[O_2 + 3.76N_2] \longrightarrow B\,CO_2 + D\,H_2O + E\,N_2$$

where a_{th} is the stoichiometric coefficient for air. The coefficient a_{th} and other coefficients are to be determined from the mass balances

Carbon balance: $B = 8$

Hydrogen balance: $2D = 18 \longrightarrow D = 9$

Oxygen balance: $2a_{th} = 2B + D \longrightarrow a_{th} = 0.5(2\times 8 + 9) = 12.5$

Nitrogen balance: $a_{th} \times 3.76 = E \longrightarrow E = 12.5 \times 3.76 = 47$

Substituting, the balanced reaction equation is

$$C_8H_{18} + 12.5[O_2 + 3.76N_2] \longrightarrow 8\,CO_2 + 9\,H_2O + 47\,N_2$$

The mass of each product and the total mass are

$$m_{CO2} = N_{CO2}M_{CO2} = (8\text{ kmol})(44\text{ kg/kmol}) = 352\text{ kg}$$
$$m_{H2O} = N_{H2O}M_{H2O} = (9\text{ kmol})(18\text{ kg/kmol}) = 162\text{ kg}$$
$$m_{N2} = N_{N2}M_{N2} = (47\text{ kmol})(28\text{ kg/kmol}) = 1316\text{ kg}$$
$$m_{total} = m_{CO2} + m_{N2} + m_{H2O} = 352 + 162 + 1316 = 1830\text{ kg}$$

Then the mass fractions are

$$\text{mf}_{CO2} = \frac{m_{CO2}}{m_{total}} = \frac{352\text{ kg}}{1830\text{ kg}} = \mathbf{0.1923}$$
$$\text{mf}_{H2O} = \frac{m_{H2O}}{m_{total}} = \frac{162\text{ kg}}{1830\text{ kg}} = \mathbf{0.0885}$$
$$\text{mf}_{N2} = \frac{m_{N2}}{m_{total}} = \frac{1316\text{ kg}}{1830\text{ kg}} = \mathbf{0.7191}$$

The mass of water per unit mass of fuel burned is

$$\frac{m_{H2O}}{m_{C8H18}} = \frac{(9\times 18)\text{ kg}}{(1\times 114)\text{ kg}} = \mathbf{1.421\,kg\ H_2O/kg\ C_8H_{18}}$$

The mass of each reactant and the total mass are

$$m_{C8H18} = N_{C8H18}M_{C8H18} = (1\text{ kmol})(114\text{ kg/kmol}) = 114\text{ kg}$$
$$m_{air} = N_{air}M_{air} = (12.5\times 4.76\text{ kmol})(29\text{ kg/kmol}) = 1725.5\text{ kg}$$
$$m_{total} = m_{C8H18} + m_{air} = 114 + 1725.5 = 1839.5\text{ kg}$$

Then the mass fractions of reactants are

$$\text{mf}_{C8H18} = \frac{m_{C8H18}}{m_{total}} = \frac{114\text{ kg}}{1839.5\text{ kg}} = \mathbf{0.0620}$$
$$\text{mf}_{air} = \frac{m_{air}}{m_{total}} = \frac{1725.5\text{ kg}}{1839.5\text{ kg}} = \mathbf{0.9380}$$

15-27 Butane C_4H_{10} is burned with 200 percent theoretical air. The kmol of water that needs to be sprayed into the combustion chamber per kmol of fuel is to be determined.

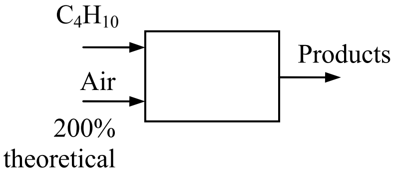
Assumptions **1** Combustion is complete. **2** The combustion products contain CO_2, H_2O, O_2, and N_2 only.

Properties The molar masses of C, H_2, O_2 and air are 12 kg/kmol, 2 kg/kmol, 32 kg/kmol, and 29 kg/kmol, respectively (Table A-1).

Analysis The reaction equation for 200% theoretical air without the additional water is

$$\mathrm{C_4H_{10}} + 2a_{\mathrm{th}}[\mathrm{O_2} + 3.76\mathrm{N_2}] \longrightarrow B\,\mathrm{CO_2} + D\,\mathrm{H_2O} + E\,\mathrm{O_2} + F\,\mathrm{N_2}$$

where a_{th} is the stoichiometric coefficient for air. We have automatically accounted for the 100% excess air by using the factor $2a_{\mathrm{th}}$ instead of a_{th} for air. The coefficient a_{th} and other coefficients are to be determined from the mass balances



Carbon balance: $B = 4$

Hydrogen balance: $2D = 10 \longrightarrow D = 5$

Oxygen balance: $2 \times 2a_{\mathrm{th}} = 2B + D + 2E$

$a_{\mathrm{th}} = E$

Nitrogen balance: $2a_{\mathrm{th}} \times 3.76 = F$

Solving the above equations, we find the coefficients (E = 6.5, F = 48.88, and a_{th} = 6.5) and write the balanced reaction equation as

$$\mathrm{C_4H_{10}} + 13[\mathrm{O_2} + 3.76\mathrm{N_2}] \longrightarrow 4\,\mathrm{CO_2} + 5\,\mathrm{H_2O} + 6.5\,\mathrm{O_2} + 48.88\,\mathrm{N_2}$$

With the additional water sprayed into the combustion chamber, the balanced reaction equation is

$$\mathrm{C_4H_{10}} + 13[\mathrm{O_2} + 3.76\mathrm{N_2}] + N_v\,\mathrm{H_2O} \longrightarrow 4\,\mathrm{CO_2} + (5 + N_v)\,\mathrm{H_2O} + 6.5\,\mathrm{O_2} + 48.88\,\mathrm{N_2}$$

The partial pressure of water in the saturated product mixture at the dew point is

$$P_{v,\mathrm{prod}} = P_{\mathrm{sat@60°C}} = 19.95\,\mathrm{kPa}$$

The vapor mole fraction is

$$y_v = \frac{P_{v,\mathrm{prod}}}{P_{\mathrm{prod}}} = \frac{19.95\,\mathrm{kPa}}{100\,\mathrm{kPa}} = 0.1995$$

The amount of water that needs to be sprayed into the combustion chamber can be determined from

$$y_v = \frac{N_{\mathrm{water}}}{N_{\mathrm{total,product}}} \longrightarrow 0.1995 = \frac{5 + N_v}{4 + 5 + N_v + 6.5 + 48.88} \longrightarrow N_v = \mathbf{9.796\ kmol}$$

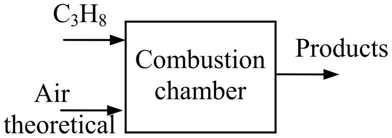
15-51 The higher and lower heating values of liquid propane are to be determined and compared to the listed values.

Assumptions **1** Combustion is complete. **2** The combustion products contain CO_2, H_2O, and N_2. **3** Combustion gases are ideal gases.

Properties The molar masses of C, O_2, H_2, and air are 12, 32, 2, and 29 kg/kmol, respectively (Table A-1).

Analysis The combustion reaction with stoichiometric air is

$$C_3H_8(l) + 5(O_2 + 3.76N_2) \longrightarrow 3CO_2 + 4H_2O + 18.8N_2$$



Both the reactants and the products are taken to be at the standard reference state of 25°C and 1 atm for the calculation of heating values. The heat transfer for this process is equal to enthalpy of combustion. Note that N_2 and O_2 are stable elements, and thus their enthalpy of formation is zero. Then,

$$q = h_C = H_P - H_R = \sum N_P \overline{h}^{\circ}_{f,P} - \sum N_R \overline{h}^{\circ}_{f,R} = \left(N\overline{h}^{\circ}_f\right)_{\mathrm{CO2}} + \left(N\overline{h}^{\circ}_f\right)_{\mathrm{H2O}} - \left(N\overline{h}^{\circ}_f\right)_{\mathrm{C3H8}}$$

The $\overline{h}^{\circ}_f$ of liquid propane is obtained by adding $\overline{h}_{fg}$ of propane at 25°C to $\overline{h}^{\circ}_f$ of gas propane (103,850 + 44.097 × 335 = 118,620 kJ/kmol). For the HHV, the water in the products is taken to be liquid. Then,

$$\begin{aligned} h_C &= (3\ \text{kmol})(-393{,}520\ \text{kJ/kmol}) + (4\ \text{kmol})(-285{,}830\ \text{kJ/kmol}) - (1\ \text{kmol})(-118{,}620\ \text{kJ/kmol}) \\ &= -2{,}205{,}260\ \text{kJ/kmol propane} \end{aligned}$$

The HHV of the liquid propane is

$$\text{HHV} = \frac{-h_C}{M_m} = \frac{2{,}205{,}260\ \text{kJ/kmol}\ C_3H_8}{44.097\ \text{kg/kmol}\ C_3H_8} = \mathbf{50{,}010\ kJ/kg\ C_3H_8}$$

The listed value from Table A-27 is **50,330 kJ/kg**. For the LHV, the water in the products is taken to be vapor. Then,

$$\begin{aligned} h_C &= (3\ \text{kmol})(-393{,}520\ \text{kJ/kmol}) + (4\ \text{kmol})(-241{,}820\ \text{kJ/kmol}) - (1\ \text{kmol})(-118{,}620\ \text{kJ/kmol}) \\ &= -2{,}029{,}220\ \text{kJ/kmol propane} \end{aligned}$$

The LHV of the propane is then

$$\text{LHV} = \frac{-h_C}{M_m} = \frac{2{,}029{,}220\ \text{kJ/kmol}\ C_3H_8}{44.097\ \text{kg/kmol}\ C_3H_8} = \mathbf{46{,}020\ kJ/kg\ C_3H_8}$$

The listed value from Table A-27 is **46,340 kJ/kg**. The calculated and listed values are practically identical.

15-66 A mixture of propane and methane is burned with theoretical air. The balanced chemical reaction is to be written, and the amount of water vapor condensed and the the required air flow rate for a given heat transfer rate are to be determined.

Assumptions **1** Combustion is complete. **2** The combustion products contain CO_2, CO, H_2O, O_2, and N_2 only.

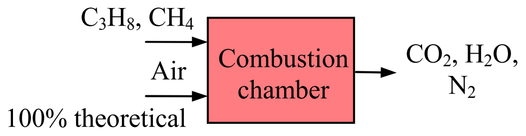
Properties The molar masses of C, H_2, O_2, N_2 and air are 12 kg/kmol, 2 kg/kmol, 32 kg/kmol, 28 kg/kmol, and 29 kg/kmol, respectively (Table A-1).

Analysis (a) The balanced reaction equation for stoichiometric air is

$$0.4\,C_3H_8 + 0.6\,CH_4 + a_{th}\left[O_2 + 3.76N_2\right] \longrightarrow 1.8\,CO_2 + 2.8\,H_2O + a_{th} \times 3.76\,N_2$$

The stoicihiometric coefficient a_{th} is determined from an O_2 balance:

$$a_{th} = 1.8 + 1.4 = 3.2$$



Substituting,

$$0.4\,C_3H_8 + 0.6\,CH_4 + 3.2\left[O_2 + 3.76N_2\right] \longrightarrow 1.8\,CO_2 + 2.8\,H_2O + 12.032\,N_2$$

(b) The partial pressure of water vapor is

$$P_v = \frac{N_{H2O}}{N_{total}} P_{total} = \frac{2.8}{1.8 + 2.8 + 12.032}(100\text{ kPa}) = \frac{2.8\text{ kmol}}{16.632\text{ kmol}}(100\text{ kPa}) = 16.84\text{ kPa}$$

The dew point temperature of the product gases is the saturation temperature of water at this pressure:

$$T_{dp} = T_{sat@16.84\text{ kPa}} = 56.2°\text{C} \quad \text{(Table A-5)}$$

Since the temperature of the product gases are at 398 K (125°C), there will be no condensation of water vapor.

(c) The heat transfer for this combustion process is determined from the energy balance $E_{in} - E_{out} = \Delta E_{system}$ applied on the combustion chamber with $W = 0$. It reduces to

$$-Q_{out} = \sum N_P\left(\bar{h}_f^\circ + \bar{h} - \bar{h}^\circ\right)_P - \sum N_R\left(\bar{h}_f^\circ + \bar{h} - \bar{h}^\circ\right)_R$$

The products are at 125 °C, and the enthalpy of products can be expressed as

$$\left(\bar{h} - \bar{h}^\circ\right) = \bar{c}_p \Delta T$$

where $\Delta T = 125 - 25 = 100°\text{C} = 100\text{ K}$. Then, using the values given in the table,

$$\begin{aligned} -Q_{out} &= (1.8)(-393{,}520 + 41.16 \times 100) + (2.8)(-241{,}820 + 34.28 \times 100) + (12.032)(0 + 29.27 \times 100) \\ &\quad - (0.4)(-103{,}850) - (0.6)(-74{,}850) \\ &= -1{,}246{,}760\text{ kJ/kmol fuel} \end{aligned}$$

or $\quad Q_{out} = 1{,}246{,}760\text{ kJ/kmol fuel}$

For a heat transfer rate of 97,000 kJ/h, the molar flow rate of fuel is

$$\dot{N}_{fuel} = \frac{\dot{Q}_{out}}{Q_{out}} = \frac{97{,}000\text{ kJ/h}}{1{,}246{,}760\text{ kJ/kmol fuel}} = 0.07780\text{ kmol fuel/h}$$

The molar mass of the fuel mixture is

$$M_{fuel} = 0.4 \times 44 + 0.6 \times 16 = 27.2\text{ kg/kmol}$$

The mass flow rate of fuel is

$$\dot{m}_{fuel} = \dot{N}_{fuel} M_{fuel} = (0.07780\text{ kmol/h})(27.2\text{ kg/kmol}) = 2.116\text{ kg/h}$$

The air-fuel ratio is

$$\text{AF} = \frac{m_{air}}{m_{fuel}} = \frac{(3.2 \times 4.76 \times 29)\text{ kg}}{(0.4 \times 44 + 0.6 \times 16)\text{ kg}} = 16.24\text{ kg air/kg fuel}$$

The mass flow rate of air is then

$$\dot{m}_{air} = \dot{m}_{fuel}\text{AF} = (2.116\text{ kg/h})16.24) = \mathbf{34.4\text{ kg/h}}$$

15-73 A high efficiency gas furnace burns gaseous propane C_3H_8 with 140 percent theoretical air. The volume flow rate of water condensed from the product gases is to be determined.

Assumptions **1** Combustion is complete. **2** The combustion products contain CO_2, H_2O, O_2, and N_2 only.

Properties The molar masses of C, H_2, O_2 and air are 12 kg/kmol, 2 kg/kmol, 32 kg/kmol, and 29 kg/kmol, respectively (Table A-1).

Analysis The reaction equation for 40% excess air (140% theoretical air) is

$$C_3H_8 + 1.4a_{th}[O_2 + 3.76N_2] \longrightarrow B\,CO_2 + D\,H_2O + E\,O_2 + F\,N_2$$

where a_{th} is the stoichiometric coefficient for air. We have automatically accounted for the 40% excess air by using the factor $1.4a_{th}$ instead of a_{th} for air. The coefficient a_{th} and other coefficients are to be determined from the mass balances

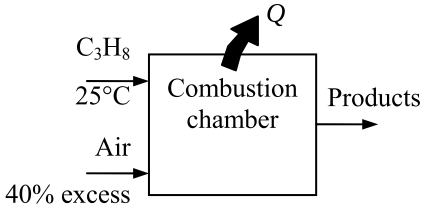
Carbon balance: $B = 3$

Hydrogen balance: $2D = 8 \longrightarrow D = 4$

Oxygen balance: $2\times 1.4a_{th} = 2B + D + 2E$

$0.4a_{th} = E$

Nitrogen balance: $1.4a_{th} \times 3.76 = F$



Solving the above equations, we find the coefficients (E = 2, F = 26.32, and a_{th} = 5) and write the balanced reaction equation as

$$C_3H_8 + 7[O_2 + 3.76N_2] \longrightarrow 3\,CO_2 + 4\,H_2O + 2\,O_2 + 26.32\,N_2$$

The partial pressure of water in the saturated product mixture at the dew point is

$$P_{v,\text{prod}} = P_{\text{sat@40°C}} = 7.3851\,\text{kPa}$$

The vapor mole fraction is

$$y_v = \frac{P_{v,\text{prod}}}{P_{\text{prod}}} = \frac{7.3851\,\text{kPa}}{100\,\text{kPa}} = 0.07385$$

The kmoles of water condensed is determined from

$$y_v = \frac{N_{\text{water}}}{N_{\text{total,product}}} \longrightarrow 0.07385 = \frac{4 - N_w}{3 + 4 - N_w + 2 + 26.32} \longrightarrow N_w = 1.503\,\text{kmol}$$

The steady-flow energy balance is expressed as

$$\dot{N}_{\text{fuel}} H_R = \dot{Q}_{\text{fuel}} + \dot{N}_{\text{fuel}} H_P$$

where $$\dot{Q}_{\text{fuel}} = \frac{\dot{Q}_{\text{out}}}{\eta_{\text{furnace}}} = \frac{31{,}650\,\text{kJ/h}}{0.96} = 32{,}969\,\text{kJ/h}$$

$$H_R = \bar{h}^o_{f\,\text{fuel@25°C}} + 7\bar{h}_{\text{O2@25°C}} + 26.32\bar{h}_{\text{N2@25°C}}$$
$$= (-103{,}847\,\text{kJ/kmol}) + 7(0) + 26.32(0) = -103{,}847\,\text{kJ/kmol}$$

$$H_P = 3\bar{h}_{\text{CO2@25°C}} + 4\bar{h}_{\text{H2O@25°C}} + 2\bar{h}_{\text{O2@25°C}} + 26.32\bar{h}_{\text{N2@25°C}} + N_w(\bar{h}^o_{f\,\text{H2O(liq)}})$$
$$= 3(-393{,}520\,\text{kJ/kmol}) + 4(-241{,}820\,\text{kJ/kmol}) + 2(0) + 26.32(0) + 1.503(-285{,}830\,\text{kJ/kmol})$$
$$= -2.577\times 10^6\ \text{kJ/kmol}$$

Substituting into the energy balance equation,

$$\dot{N}_{\text{fuel}} H_R = \dot{Q}_{\text{fuel}} + \dot{N}_{\text{fuel}} H_P$$

$$\dot{N}_{\text{fuel}}(-103{,}847\,\text{kJ/kmol}) = 32{,}969\,\text{kJ/h} + \dot{N}_{\text{fuel}}(-2.577\times 10^6\ \text{kJ/kmol}) \longrightarrow \dot{N}_{\text{fuel}} = 0.01333\,\text{kmol/h}$$

The molar and mass flow rates of the liquid water are

$$\dot{N}_w = N_w \dot{N}_{\text{fuel}} = (1.503\ \text{kmol/kmol fuel})(0.01333\ \text{kmol fuel/h}) = 0.02003\ \text{kmol/h}$$

$$\dot{m}_w = \dot{N}_w M_w = (0.02003\ \text{kmol/h})(18\ \text{kg/kmol}) = 0.3608\ \text{kg/h}$$

The volume flow rate of liquid water is

$$\dot{V}_w = (v_{f\,@25°C})\dot{m}_w = (0.001003\ \text{m}^3/\text{kg})(0.3608\ \text{kg/h}) = 0.0003619\ \text{m}^3/\text{h} = \mathbf{8.7\ L/day}$$

15-80 Propane is burned with stoichiometric and 50 percent excess air. The adiabatic flame temperature is to be determined for both cases.

Assumptions **1** Steady operating conditions exist. **2** Air and combustion gases are ideal gases. **3** Kinetic and potential energies are negligible. **4** There are no work interactions. **5** The combustion chamber is adiabatic.

Analysis Under steady-flow conditions the energy balance $E_{\text{in}} - E_{\text{out}} = \Delta E_{\text{system}}$ applied on the combustion chamber with $Q = W = 0$ reduces to

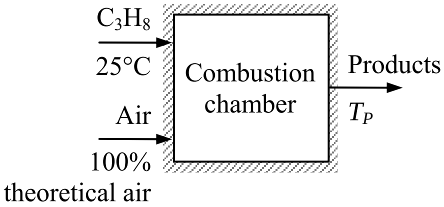
$$\sum N_P\left(\bar{h}_f^\circ + \bar{h} - \bar{h}^\circ\right)_P = \sum N_R\left(\bar{h}_f^\circ + \bar{h} - \bar{h}^\circ\right)_R \longrightarrow \sum N_P\left(\bar{h}_f^\circ + \bar{h} - \bar{h}^\circ\right)_P = \sum N_R \bar{h}_{f,R}^\circ$$

since all the reactants are at the standard reference temperature of 25°C. Then, for the stoicihiometric air

$$C_3H_8 + a_{\text{th}}(O_2 + 3.76\,N_2) \longrightarrow 3\,CO_2 + 4\,H_2O + a_{\text{th}} \times 3.76\,N_2$$

where a_{th} is the stoichiometric coefficient and is determined from the O_2 balance,

$$a_{\text{th}} = 3 + 2 = 5$$



Thus,

$$C_3H_8 + 5\,(O_2 + 3.76\,N_2) \longrightarrow 3\,CO_2 + 4\,H_2O + 18.8\,N_2$$

From the tables,

Substance	$\bar{h}_f^\circ$ **kJ/kmol**	$\bar{h}_{298K}$ **kJ/kmol**
C_3H_8 (*g*)	-103,850	---
O_2	0	8682
N_2	0	8669
H_2O (*g*)	-241,820	9904
CO_2	-393,520	9364

Thus,

$$(3)\left(-393{,}520 + \bar{h}_{CO2} - 9364\right) + (4)\left(-241{,}820 + \bar{h}_{H2O} - 9904\right) + (18.8)\left(0 + \bar{h}_{N2} - 8669\right) = (1)\left(-103{,}850\right) + 0 + 0$$

It yields $\quad 3\bar{h}_{CO2} + 4\bar{h}_{H2O} + 18.8\bar{h}_{N2} = 2{,}274{,}680 \text{ kJ}$

The adiabatic flame temperature is obtained from a trial and error solution. A first guess is obtained by dividing the right-hand side of the equation by the total number of moles, which yields 2,274,680/(3 + 4 + 18.8) = 88,166 kJ/kmol. This enthalpy value corresponds to about 2650 K for N_2. Noting that the majority of the moles are N_2, T_P will be close to 2650 K, but somewhat under it because of the higher specific heat of H_2O.

At 2500 K:

$$\begin{aligned} 3\bar{h}_{CO2} + 4\bar{h}_{H2O} + 18.8\bar{h}_{N2} &= 3 \times 131{,}290 + 4 \times 108{,}868 + 18.8 \times 82{,}981 \\ &= 2{,}389{,}380 \text{ kJ} \ \text{(Higher than 2,274,680 kJ)} \end{aligned}$$

At 2450 K:

$$\begin{aligned} 3\bar{h}_{CO2} + 4\bar{h}_{H2O} + 18.8\bar{h}_{N2} &= 3 \times 128{,}219 + 4 \times 106{,}183 + 18.8 \times 81{,}149 \\ &= 2{,}334{,}990 \text{ kJ} \ \text{(Higher than 2,274,680 kJ)} \end{aligned}$$

At 2400 K:

$$\begin{aligned} 3\bar{h}_{CO2} + 4\bar{h}_{H2O} + 18.8\bar{h}_{N2} &= 3 \times 125{,}152 + 4 \times 103{,}508 + 18.8 \times 79{,}320 \\ &= 2{,}280{,}704 \text{ kJ} \ \text{(Higher than 2,274,680 kJ)} \end{aligned}$$

At 2350 K:

$$3\bar{h}_{CO2} + 4\bar{h}_{H2O} + 18.8\bar{h}_{N2} = 3\times 122{,}091 + 4\times 100{,}846 + 18.8\times 77{,}496$$
$$= 2{,}226{,}580 \text{ kJ (Lower than 2,274,680 kJ)}$$

By interpolation of the two results,

T_P = **2394 K = 2121°C**

When propane is burned with 50% excess air, the reaction equation may be written as

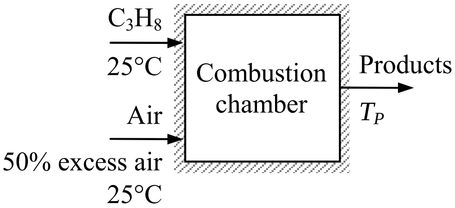
$$C_3H_8 + 1.5\times a_{th}(O_2 + 3.76N_2) \longrightarrow 3CO_2 + 4H_2O + 0.5\times a_{th}O_2 + 1.5\times a_{th}\times 3.76N_2$$

where a_{th} is the stoichiometric coefficient and is determined from the O_2 balance,

$$1.5a_{th} = 3 + 2 + 0.5a_{th} \longrightarrow a_{th} = 5$$

Thus,

$$C_3H_8 + 7.5(O_2 + 3.76N_2) \longrightarrow 3CO_2 + 4H_2O + 2.5O_2 + 28.2N_2$$



Using the values in the table,

$$(3)\left(-393{,}520 + \bar{h}_{CO2} - 9364\right) + (4)\left(-241{,}820 + \bar{h}_{H2O} - 9904\right) + (2.5)\left(0 + \bar{h}_{O2} - 8682\right)$$
$$+ (28.2)\left(0 + \bar{h}_{N2} - 8669\right) = (1)\left(-103{,}850\right) + 0 + 0$$

It yields $3\bar{h}_{CO2} + 4\bar{h}_{H2O} + 2.5\bar{h}_{O2} + 28.2\bar{h}_{N2} = 2{,}377{,}870$ kJ

The adiabatic flame temperature is obtained from a trial and error solution. A first guess is obtained by dividing the right-hand side of the equation by the total number of moles, which yields 2,377,870/(3+4+2.5+28.2) = 63,073 kJ/kmol. This enthalpy value corresponds to about 1960 K for N_2. Noting that the majority of the moles are N_2, T_P will be close to 1960 K, but somewhat under it because of the higher specific heat of H_2O.

At 1800 K:

$$3\bar{h}_{CO2} + 4\bar{h}_{H2O} + 2.5\bar{h}_{O2} + 28.2\bar{h}_{N2} = 3\times 88{,}806 + 4\times 72{,}513 + 2.5\times 60{,}371 + 28.2\times 57{,}651$$
$$= 2{,}333{,}160 \text{ kJ (Lower than 2,377,870 kJ)}$$

At 1840 K:

$$3\bar{h}_{CO2} + 4\bar{h}_{H2O} + 2.5\bar{h}_{O2} + 28.2\bar{h}_{N2} = 3\times 91{,}196 + 4\times 74{,}506 + 2.5\times 61{,}866 + 28.2\times 59{,}075$$
$$= 2{,}392{,}190 \text{ kJ (Higher than 2,377,870 kJ)}$$

By interpolation,

T_P = **1830 K = 1557°C**

15-89 Ethylene gas is burned steadily with 20 percent excess air. The temperature of products, the entropy generation, and the exergy destruction (or irreversibility) are to be determined.

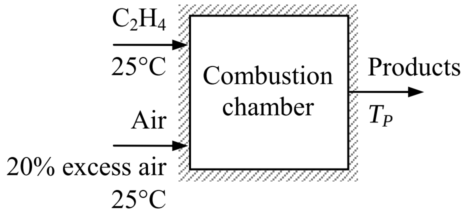
Assumptions **1** Combustion is complete. **2** Steady operating conditions exist. **3** Air and the combustion gases are ideal gases. **4** Changes in kinetic and potential energies are negligible.

Analysis (*a*) The fuel is burned completely with the excess air, and thus the products will contain only CO_2, H_2O, N_2, and some free O_2. Considering 1 kmol of C_2H_4, the combustion equation can be written as

$$\mathrm{C_2H_4}(\mathrm{g}) + 1.2a_{\mathrm{th}}(\mathrm{O_2} + 3.76\mathrm{N_2}) \longrightarrow 2\mathrm{CO_2} + 2\mathrm{H_2O} + 0.2a_{\mathrm{th}}\mathrm{O_2} + (1.2)(3.76)a_{\mathrm{th}}\mathrm{N_2}$$

where a_{th} is the stoichiometric coefficient and is determined from the O_2 balance,

$$1.2a_{\mathrm{th}} = 2 + 1 + 0.2a_{\mathrm{th}} \longrightarrow a_{\mathrm{th}} = 3$$



Thus,

$$\mathrm{C_2H_4}(g) + 3.6(\mathrm{O_2} + 3.76\mathrm{N_2}) \longrightarrow 2\mathrm{CO_2} + 2\mathrm{H_2O} + 0.6\mathrm{O_2} + 13.54\mathrm{N_2}$$

Under steady-flow conditions, the exit temperature of the product gases can be determined from the steady-flow energy equation, which reduces to

$$\sum N_P\left(\bar{h}_f^\circ + \bar{h} - \bar{h}^\circ\right)_P = \sum N_R \bar{h}_{f,R}^\circ = \left(N\bar{h}_f^\circ\right)_{\mathrm{C_2H_4}}$$

since all the reactants are at the standard reference state, and for O_2 and N_2. From the tables,

Substance	$\bar{h}_f^\circ$ kJ/kmol	$\bar{h}_{298\ K}$ kJ/kmol
C_2H_4 (*g*)	52,280	---
O_2	0	8682
N_2	0	8669
H_2O (*g*)	-241,820	9904
CO_2	-393,520	9364

Substituting,

$$\begin{aligned}(2)(-393{,}520 + \bar{h}_{\mathrm{CO_2}} - 9364) + (2)(-241{,}820 + \bar{h}_{\mathrm{H_2O}} - 9904)\\ + (0.6)(0 + \bar{h}_{\mathrm{O_2}} - 8682) + (13.54)(0 + \bar{h}_{\mathrm{N_2}} - 8669) = (1)(52{,}280)\end{aligned}$$

or, $$2\bar{h}_{\mathrm{CO_2}} + 2\bar{h}_{\mathrm{H_2O}} + 0.6\bar{h}_{\mathrm{O_2}} + 13.54\bar{h}_{\mathrm{N_2}} = 1{,}484{,}083\ \mathrm{kJ}$$

By trial and error,

$$T_P = \mathbf{2269.6\ K}$$

(*b*) The entropy generation during this adiabatic process is determined from

$$S_{\mathrm{gen}} = S_P - S_R = \sum N_P \bar{s}_P - \sum N_R \bar{s}_R$$

The C_2H_4 is at 25°C and 1 atm, and thus its absolute entropy is 219.83 kJ/kmol·K (Table A-26). The entropy values listed in the ideal gas tables are for 1 atm pressure. Both the air and the product gases are at a total pressure of 1 atm, but the entropies are to be calculated at the partial pressure of the components which is equal to $P_i = y_i P_{\mathrm{total}}$, where y_i is the mole fraction of component i . Also,

$$S_i = N_i \bar{s}_i(T, P_i) = N_i\left(\bar{s}_i^\circ(T, P_0) - R_u \ln(y_i P_m)\right)$$

The entropy calculations can be presented in tabular form as

	$\mathbf{N_i}$	$\mathbf{y_i}$	$\bar{\mathbf{s}}_{\mathbf{i}}^{\circ}(\mathbf{T,1atm})$	$\mathbf{R_u ln(y_i P_m)}$	$\mathbf{N_i \bar{s}_i}$
C_2H_4	1	1.00	219.83	---	219.83
O_2	3.6	0.21	205.14	-12.98	784.87
N_2	13.54	0.79	191.61	-1.96	2620.94
					S_R = 3625.64 kJ/K
CO_2	2	0.1103	316.881	-18.329	670.42
H_2O	2	0.1103	271.134	-18.329	578.93
O_2	0.6	0.0331	273.467	-28.336	181.08
N_2	13.54	0.7464	256.541	-2.432	3506.49
					S_P = 4936.92 kJ/K

Thus,

$$S_{\text{gen}} = S_P - S_R = 4936.92 - 3625.64 = \mathbf{1311.28\ kJ/kmol \cdot K}$$

and

(*c*) $$X_{\text{destroyed}} = T_0 S_{\text{gen}} = (298\text{ K})(1311.28\text{ kJ/kmol}\cdot\text{K C}_2\text{H}_4) = \mathbf{390{,}760\ \ kJ}\ \ (\text{per kmol C}_2\text{H}_4)$$

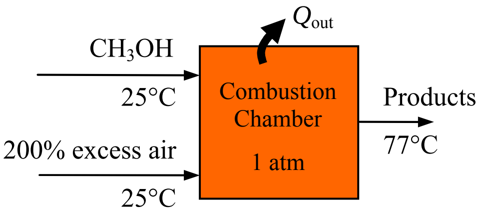
15-95 Methyl alcohol is burned steadily with 200 percent excess air in an automobile engine. The maximum amount of work that can be produced by this engine is to be determined.

Assumptions **1** Combustion is complete. **2** Steady operating conditions exist. **3** Air and the combustion gases are ideal gases. **4** Changes in kinetic and potential energies are negligible.

Analysis The fuel is burned completely with the excess air, and thus the products will contain only CO_2, H_2O, N_2, and some free O_2. Considering 1 kmol CH_3OH the combustion equation can be written as

$$\mathrm{CH_3OH} + 3a_{\mathrm{th}}(\mathrm{O_2} + 3.76\mathrm{N_2}) \longrightarrow \mathrm{CO_2} + 2\mathrm{H_2O} + 2a_{\mathrm{th}}\mathrm{O_2} + 3a_{\mathrm{th}} \times 3.76\mathrm{N_2}$$

where a_{th} is the stoichiometric coefficient and is determined from the O_2 balance,



$$0.5 + 3a_{\mathrm{th}} = 1 + 1 + 2a_{\mathrm{th}} \longrightarrow a_{\mathrm{th}} = 1.5$$

Thus,

$$\mathrm{CH_3OH} + 4.5(\mathrm{O_2} + 3.76\mathrm{N_2}) \longrightarrow \mathrm{CO_2} + 2\mathrm{H_2O} + 3\mathrm{O_2} + 16.92\mathrm{N_2}$$

Under steady-flow conditions the energy balance $E_{\mathrm{in}} - E_{\mathrm{out}} = \Delta E_{\mathrm{system}}$ applied on the combustion chamber with $W = 0$ reduces to

$$-Q_{\mathrm{out}} = \sum N_P\left(\bar{h}_f^\circ + \bar{h} - \bar{h}^\circ\right)_P - \sum N_R\left(\bar{h}_f^\circ + \bar{h} - \bar{h}^\circ\right)_R$$

Assuming the air and the combustion products to be ideal gases, we have $h = h(T)$. From the tables,

Substance	$\bar{h}_f^\circ$ **kJ/kmol**	$\bar{h}_{298\mathrm{K}}$ **kJ/kmol**	$\bar{h}_{350\mathrm{K}}$ **kJ/kmol**
CH_3OH	-200,670	---	---
O_2	0	8682	10,213
N_2	0	8669	10,180
H_2O (*g*)	-241,820	9904	11,652
CO_2	-393,520	9364	11,351

Thus,

$$\begin{aligned} -Q_{\mathrm{out}} &= (1)(-393{,}520 + 11{,}351 - 9364) + (2)(-241{,}820 + 11{,}652 - 9904) \\ &\quad + (3)(0 + 10{,}213 - 8682) + (16.92)(0 + 10{,}180 - 8669) - (1)(-200{,}670) \\ &= -663{,}550 \text{ kJ/kmol of fuel} \end{aligned}$$

The entropy generation during this process is determined from

$$S_{\mathrm{gen}} = S_P - S_R + \frac{Q_{\mathrm{out}}}{T_{\mathrm{surr}}} = \sum N_P \bar{s}_P - \sum N_R \bar{s}_R + \frac{Q_{\mathrm{out}}}{T_{\mathrm{surr}}}$$

The entropy values listed in the ideal gas tables are for 1 atm pressure. Both the air and the product gases are at a total pressure of 1 atm, but the entropies are to be calculated at the partial pressure of the components which is equal to $P_{\mathrm{i}} = y_{\mathrm{i}} P_{\mathrm{total}}$, where y_{i} is the mole fraction of component *i*. Then,

$$S_i = N_i \bar{s}_i(T, P_i) = N_i\left(\bar{s}_i^\circ(T, P_0) - R_u \ln(y_i P_m)\right)$$

The entropy calculations can be presented in tabular form as

	N_i	y_i	$\bar{s}_i^\circ(T,1\text{atm})$	$R_u \ln(y_i P_m)$	$N_i \bar{s}_i$
CH_3OH	1	---	239.70	---	239.70
O_2	4.5	0.21	205.04	-12.98	981.09
N_2	16.92	0.79	191.61	-1.960	3275.20
					S_R = 4496 kJ/K
CO_2	1	0.0436	219.831	-26.05	245.88
H_2O (*g*)	2	0.0873	194.125	-20.27	428.79
O_2	3	0.1309	209.765	-16.91	680.03
N_2	16.92	0.7382	196.173	-2.52	3361.89
					S_P = 4717 kJ/K

Thus,

$$S_{\text{gen}} = S_P - S_R + \frac{Q_{\text{out}}}{T_{\text{surr}}} = 4717 - 4496 + \frac{663{,}550}{298} = 2448 \text{ kJ/K (per kmol fuel)}$$

The maximum work is equal to the exergy destruction

$$W_{\max} = X_{\text{dest}} = T_0 S_{\text{gen}} = (298)(2448\,\text{kJ/K}) = 729{,}400\,\text{kJ/K (per kmol fuel)}$$

Per unit mass basis,

$$W_{\max} = \frac{729{,}400\,\text{kJ/K}\cdot\text{kmol}}{32\,\text{kg/kmol}} = \mathbf{22{,}794\,kJ/kg\ fuel}$$