Solutions Manual for

Thermodynamics: An Engineering Approach

Seventh Edition

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Chapter 15

CHEMICAL REACTIONS

Fuels and Combustion

15-1C Nitrogen, in general, does not react with other chemical species during a combustion process but its presence affects the outcome of the process because nitrogen absorbs a large proportion of the heat released during the chemical process.

15-2C Moisture, in general, does not react chemically with any of the species present in the combustion chamber, but it absorbs some of the energy released during combustion, and it raises the dew point temperature of the combustion gases.

15-3C The number of atoms are preserved during a chemical reaction, but the total mole numbers are not.

15-4C Air-fuel ratio is the ratio of the mass of air to the mass of fuel during a combustion process. Fuel-air ratio is the inverse of the air-fuel ratio.

15-5C No. Because the molar mass of the fuel and the molar mass of the air, in general, are different.

15-6C The dew-point temperature of the product gases is the temperature at which the water vapor in the product gases starts to condense as the gases are cooled at constant pressure. It is the saturation temperature corresponding to the vapor pressure of the product gases.

15-7 Sulfur is burned with oxygen to form sulfur dioxide. The minimum mass of oxygen required and the mass of sulfur dioxide in the products are to be determined when 1 kg of sulfur is burned.

Properties The molar masses of sulfur and oxygen are 32.06 kg/kmol and 32.00 kg/kmol, respectively (Table A-1).

Analysis The chemical reaction is given by

$$\mathrm{S} + \mathrm{O}_2 \longrightarrow \mathrm{SO}_2$$

$$\boxed{\mathrm{S} + \mathrm{O}_2 \longrightarrow \mathrm{SO}_2}$$

Hence, 1kmol of oxygen is required to burn 1 kmol of sulfur which produces 1 kmol of sulfur dioxide whose molecular weight is

$$M_{\mathrm{SO2}} = M_{\mathrm{S}} + M_{\mathrm{O2}} = 32.06 + 32.00 = 64.06\ \text{kg/kmol}$$

Then,

$$\frac{m_{\mathrm{O2}}}{m_{\mathrm{S}}} = \frac{N_{\mathrm{O2}} M_{\mathrm{O2}}}{N_{\mathrm{S}} M_{\mathrm{S}}} = \frac{(1\ \text{kmol})(32\ \text{kg/kmol})}{(1\ \text{kmol})(32.06\ \text{kg/kmol})} = \mathbf{0.998\ kg\ O_2/kg\ S}$$

and

$$\frac{m_{\mathrm{SO2}}}{m_{\mathrm{S}}} = \frac{N_{\mathrm{SO2}} M_{\mathrm{SO2}}}{N_{\mathrm{S}} M_{\mathrm{S}}} = \frac{(1\ \text{kmol})(64.06\ \text{kg/kmol})}{(1\ \text{kmol})(32.06\ \text{kg/kmol})} = \mathbf{1.998\ kg\ SO_2/kg\ S}$$

15-8E Methane is burned with diatomic oxygen. The mass of water vapor in the products is to be determined when 1 lbm of methane is burned.

Properties The molar masses of CH_4, O_2, CO_2, and H_2O are 16, 32, 44, and 18 lbm/lbmol, respectively (Table A-1E).

Analysis The chemical reaction is given by

$$\mathrm{CH_4 + 2O_2 \longrightarrow CO_2 + 2H_2O}$$

$$\mathrm{CH_4 + 2O_2} \\ \mathrm{\rightarrow CO_2 + 2H_2O}$$

Hence, for each lbmol of methane burned, 2 lbmol of water vapor are formed. Then,

$$\frac{m_{\mathrm{H2O}}}{m_{\mathrm{CH4}}} = \frac{N_{\mathrm{H2O}} M_{\mathrm{H2O}}}{N_{\mathrm{CH4}} M_{\mathrm{CH4}}} = \frac{(2\ \mathrm{lbmol})(18\ \mathrm{lbm/lbmol})}{(1\ \mathrm{lbmol})(16\ \mathrm{lbm/lbmol})} = \mathbf{2.25\ lbm\ H_2O/lbm\ CH_4}$$

Theoretical and Actual Combustion Processes

15-9C It represent the amount of air that contains the exact amount of oxygen needed for complete combustion.

15-10C No. The theoretical combustion is also complete, but the products of theoretical combustion does not contain any uncombined oxygen.

15-11C Case (b).

15-12C The causes of incomplete combustion are insufficient time, insufficient oxygen, insufficient mixing, and dissociation.

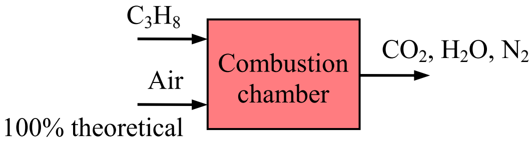
15-13C CO. Because oxygen is more strongly attracted to hydrogen than it is to carbon, and hydrogen is usually burned to completion even when there is a deficiency of oxygen.

15-14 Propane is burned with theoretical amount of air. The mass fraction of carbon dioxide and the mole and mass fractions of the water vapor in the products are to be determined.

Properties The molar masses of C_3H_8, O_2, N_2, CO_2, and H_2O are 44, 32, 28, 44, and 18 kg/kmol, respectively (Table A-1).

Analysis (*a*) The reaction in terms of undetermined coefficients is

$$C_3H_8 + x(O_2 + 3.76N_2) \longrightarrow yCO_2 + zH_2O + pN_2$$



Balancing the carbon in this reaction gives

$$y = 3$$

and the hydrogen balance gives

$$2z = 8 \longrightarrow z = 4$$

The oxygen balance produces

$$2x = 2y + z \longrightarrow x = y + z/2 = 3 + 4/2 = 5$$

A balance of the nitrogen in this reaction gives

$$2 \times 3.76x = 2p \longrightarrow p = 3.76x = 3.76 \times 5 = 18.8$$

In balanced form, the reaction is

$$C_3H_8 + 5O_2 + 18.8N_2 \longrightarrow 3CO_2 + 4H_2O + 18.8N_2$$

The mass fraction of carbon dioxide is determined from

$$\begin{aligned}\mathrm{mf}_{CO2} &= \frac{m_{CO2}}{m_{products}} = \frac{N_{CO2}M_{CO2}}{N_{CO2}M_{CO2} + N_{H2O}M_{H2O} + N_{N2}M_{N2}}\\ &= \frac{(3\ \text{kmol})(44\ \text{kg/kmol})}{(3\ \text{kmol})(44\ \text{kg/kmol}) + (4\ \text{kmol})(18\ \text{kg/kmol}) + (18.8\ \text{kmol})(28\ \text{kg/kmol})}\\ &= \frac{132\ \text{kg}}{730.4\ \text{kg}} = \mathbf{0.181}\end{aligned}$$

(*b*) The mole and mass fractions of water vapor are

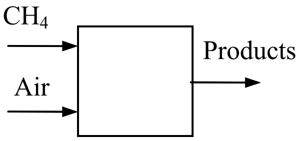
$$y_{H2O} = \frac{N_{H2O}}{N_{products}} = \frac{N_{H2O}}{N_{CO2} + N_{H2O} + N_{N2}} = \frac{4\ \text{kmol}}{3\ \text{kmol} + 4\ \text{kmol} + 18.8\ \text{kmol}} = \frac{4\ \text{kmol}}{25.8\ \text{kmol}} = \mathbf{0.155}$$

$$\begin{aligned}\mathrm{mf}_{H2O} &= \frac{m_{H2O}}{m_{products}} = \frac{N_{H2O}M_{H2O}}{N_{CO2}M_{CO2} + N_{H2O}M_{H2O} + N_{N2}M_{N2}}\\ &= \frac{(4\ \text{kmol})(18\ \text{kg/kmol})}{(3\ \text{kmol})(44\ \text{kg/kmol}) + (4\ \text{kmol})(18\ \text{kg/kmol}) + (18.8\ \text{kmol})(28\ \text{kg/kmol})}\\ &= \frac{72\ \text{kg}}{730.4\ \text{kg}} = \mathbf{0.0986}\end{aligned}$$

15-15 Methane is burned with air. The mass flow rates at the two inlets are to be determined.

Properties The molar masses of CH_4, O_2, N_2, CO_2, and H_2O are 16, 32, 28, 44, and 18 kg/kmol, respectively (Table A-1).

Analysis The stoichiometric combustion equation of CH_4 is



$$CH_4 + a_{th}[O_2 + 3.76N_2] \longrightarrow CO_2 + 2H_2O + 3.76a_{th}N_2$$

O_2 balance: $\quad a_{th} = 1 + 1 \longrightarrow a_{th} = 2$

Substituting, $\quad CH_4 + 2[O_2 + 3.76N_2] \longrightarrow CO_2 + 2H_2O + 7.52N_2$

The masses of the reactants are

$$\begin{aligned} m_{CH4} &= N_{CH4}M_{CH4} = (1\,\text{kmol})(16\,\text{kg/kmol}) = 16\,\text{kg} \\ m_{O2} &= N_{O2}M_{O2} = (2\,\text{kmol})(32\,\text{kg/kmol}) = 64\,\text{kg} \\ m_{N2} &= N_{N2}M_{N2} = (2 \times 3.76\,\text{kmol})(28\,\text{kg/kmol}) = 211\,\text{kg} \end{aligned}$$

The total mass is

$$m_{total} = m_{CH4} + m_{O2} + N_{N2} = 16 + 64 + 211 = 291\,\text{kg}$$

Then the mass fractions are

$$\begin{aligned} \text{mf}_{CH4} &= \frac{m_{CH4}}{m_{total}} = \frac{16\,\text{kg}}{291\,\text{kg}} = 0.05498 \\ \text{mf}_{O2} &= \frac{m_{O2}}{m_{total}} = \frac{64\,\text{kg}}{291\,\text{kg}} = 0.2199 \\ \text{mf}_{N2} &= \frac{m_{N2}}{m_{total}} = \frac{211\,\text{kg}}{291\,\text{kg}} = 0.7251 \end{aligned}$$

For a mixture flow of 0.5 kg/s, the mass flow rates of the reactants are

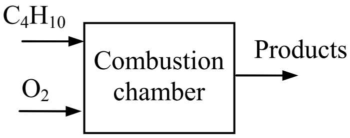
$$\begin{aligned} \dot{m}_{CH4} &= \text{mf}_{CH4}\dot{m} = (0.05498)(0.5\,\text{kg/s}) = \mathbf{0.02749\ kg/s} \\ \dot{m}_{air} &= \dot{m} - \dot{m}_{CH4} = 0.5 - 0.02749 = \mathbf{0.4725\ kg/s} \end{aligned}$$

15-16 n-Butane is burned with stoichiometric amount of oxygen. The mole fractions of CO_2 water in the products and the mole number of CO_2 in the products per mole of fuel burned are to be determined.

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

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

Analysis The combustion equation in this case is



$$C_4H_{10} + 6.5O_2 \longrightarrow 4CO_2 + 5H_2O$$

The total mole of the products are 4+5 = 9 kmol. Then the mole fractions are

$$\begin{aligned} y_{CO2} &= \frac{N_{CO2}}{N_{total}} = \frac{4\,\text{kmol}}{9\,\text{kmol}} = \mathbf{0.4444} \\ y_{CO2} &= \frac{N_{H2O}}{N_{total}} = \frac{5\,\text{kmol}}{9\,\text{kmol}} = \mathbf{0.5556} \end{aligned}$$

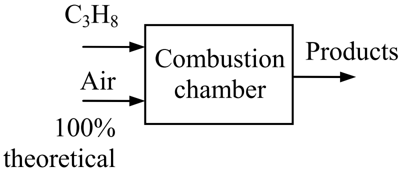
Also, $\quad N_{CO2} = \mathbf{4\ kmol\ CO_2/kmol\ C_4H_{10}}$

15-17 Propane is burned with stoichiometric amount of air. The mass fraction of each product, the mass of water and air per unit mass of fuel burned 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



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

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: $B = 3$

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

Oxygen balance: $2a_{th} = 2B + D \longrightarrow a_{th} = 0.5(2\times 3 + 4) = 5$

Nitrogen balance: $a_{th} \times 3.76 = E \longrightarrow E = 5\times 3.76 = 18.8$

Substituting, the balanced reaction equation is

$$C_3H_8 + 5\left[O_2 + 3.76N_2\right] \longrightarrow 3\,CO_2 + 4\,H_2O + 18.8\,N_2$$

The mass of each product and the total mass are

$$\begin{aligned} m_{CO2} &= N_{CO2}M_{CO2} = (3\text{ kmol})(44\text{ kg/kmol}) = 132\text{ kg} \\ m_{H2O} &= N_{H2O}M_{H2O} = (4\text{ kmol})(18\text{ kg/kmol}) = 72\text{ kg} \\ m_{N2} &= N_{N2}M_{N2} = (18.8\text{ kmol})(28\text{ kg/kmol}) = 526.4\text{ kg} \\ m_{total} &= m_{CO2} + m_{H2O} + m_{N2} = 132 + 72 + 526.4 = 730.4\text{ kg} \end{aligned}$$

Then the mass fractions are

$$\text{mf}_{CO2} = \frac{m_{CO2}}{m_{total}} = \frac{132\text{ kg}}{730.4\text{ kg}} = \mathbf{0.1807}$$

$$\text{mf}_{H2O} = \frac{m_{H2O}}{m_{total}} = \frac{72\text{ kg}}{730.4\text{ kg}} = \mathbf{0.0986}$$

$$\text{mf}_{N2} = \frac{m_{N2}}{m_{total}} = \frac{526.4\text{ kg}}{730.4\text{ kg}} = \mathbf{0.7207}$$

The mass of water per unit mass of fuel burned is

$$\frac{m_{H2O}}{m_{C3H8}} = \frac{(4\times 18)\text{ kg}}{(1\times 44)\text{ kg}} = \mathbf{1.636\ kg\ H_2O/kg\ C_3H_8}$$

The mass of air required per unit mass of fuel burned is

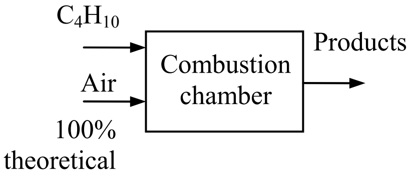
$$\frac{m_{air}}{m_{C3H8}} = \frac{(5\times 4.76\times 29)\text{ kg}}{(1\times 44)\text{ kg}} = \mathbf{15.69\ kg\ air/kg\ C_3H_8}$$

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



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

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

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

Substituting, the balanced reaction equation is

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

The mass of each product and the total mass are

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

Then the mass fractions are

$$\mathrm{mf}_{\mathrm{CO2}} = \frac{m_{\mathrm{CO2}}}{m_{\mathrm{total}}} = \frac{352\text{ kg}}{1830\text{ kg}} = \mathbf{0.1923}$$

$$\mathrm{mf}_{\mathrm{H2O}} = \frac{m_{\mathrm{H2O}}}{m_{\mathrm{total}}} = \frac{162\text{ kg}}{1830\text{ kg}} = \mathbf{0.0885}$$

$$\mathrm{mf}_{\mathrm{N2}} = \frac{m_{\mathrm{N2}}}{m_{\mathrm{total}}} = \frac{1316\text{ kg}}{1830\text{ kg}} = \mathbf{0.7191}$$

The mass of water per unit mass of fuel burned is

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

The mass of each reactant and the total mass are

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

Then the mass fractions of reactants are

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

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

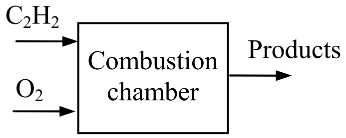
15-19 Acetylene is burned with 10 percent excess oxygen. The mass fractions of each of the products and the mass of oxygen used per unit mass of fuel burned are to be determined.

Assumptions **1** Combustion is complete. **2** The combustion products contain CO_2, H_2O, and O_2. **3** Combustion gases are ideal gases.

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

Analysis The stoichiometric combustion equation is

$$C_2H_2 + 2.5O_2 \longrightarrow 2CO_2 + H_2O$$



The combustion equation with 10% excess oxygen is

$$C_2H_2 + 2.75O_2 \longrightarrow 2CO_2 + H_2O + 0.25O_2$$

The mass of each product and the total mass are

$$\begin{aligned}
m_{CO2} &= N_{CO2}M_{CO2} = (2\text{ kmol})(44\text{ kg/kmol}) = 88\text{ kg}\\
m_{H2O} &= N_{H2O}M_{H2O} = (1\text{ kmol})(18\text{ kg/kmol}) = 18\text{ kg}\\
m_{O2} &= N_{O2}M_{O2} = (0.25\text{ kmol})(32\text{ kg/kmol}) = 8\text{ kg}\\
m_{total} &= m_{CO2} + m_{H2O} + m_{O2} = 88 + 18 + 8 = 114\text{ kg}
\end{aligned}$$

Then the mass fractions are

$$\text{mf}_{CO2} = \frac{m_{CO2}}{m_{total}} = \frac{88\text{ kg}}{114\text{ kg}} = \mathbf{0.7719}$$

$$\text{mf}_{H2O} = \frac{m_{H2O}}{m_{total}} = \frac{18\text{ kg}}{114\text{ kg}} = \mathbf{0.1579}$$

$$\text{mf}_{O2} = \frac{m_{O2}}{m_{total}} = \frac{8\text{ kg}}{114\text{ kg}} = \mathbf{0.0702}$$

The mass of oxygen per unit mass of fuel burned is determined from

$$\frac{m_{O2}}{m_{C2H2}} = \frac{(2.75\times 32)\text{ kg}}{(1\times 26)\text{ kg}} = \mathbf{3.385\text{ kg } O_2/\text{kg } C_2H_2}$$

15-20 n-Butane is burned with 100 percent excess air. The mole fractions of each of the products, the mass of carbon dioxide in the products per unit mass of the fuel, and the air-fuel ratio 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 combustion equation in this case can be written as

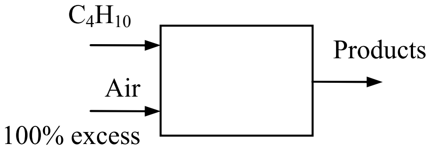
$$\mathrm{C_4H_{10}} + 2.0a_{\mathrm{th}}[\mathrm{O_2} + 3.76\mathrm{N_2}] \longrightarrow 4\mathrm{CO_2} + 5\mathrm{H_2O} + 1.0a_{\mathrm{th}}\mathrm{O_2} + (2.0\times 3.76)a_{\mathrm{th}}\mathrm{N_2}$$

where a_{th} is the stoichiometric coefficient for air. We have automatically accounted for the 100% excess air by using the factor $2.0a_{th}$ instead of a_{th} for air. The stoichiometric amount of oxygen ($a_{th}O_2$) will be used to oxidize the fuel, and the remaining excess amount ($1.0a_{th}O_2$) will appear in the products as free oxygen. The coefficient a_{th} is determined from the O_2 balance,

O_2 balance: $\quad 2.0a_{\mathrm{th}} = 4 + 2.5 + 1.0a_{\mathrm{th}} \longrightarrow a_{\mathrm{th}} = 6.5$

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

The mole fractions of the products are



$$N_m = 4 + 5 + 6.5 + 48.88 = 64.38 \text{ kmol}$$

$$y_{\mathrm{CO2}} = \frac{N_{\mathrm{CO2}}}{N_m} = \frac{4 \text{ kmol}}{64.38 \text{ kmol}} = \mathbf{0.0621}$$

$$y_{\mathrm{H2O}} = \frac{N_{\mathrm{H2O}}}{N_m} = \frac{5 \text{ kmol}}{64.38 \text{ kmol}} = \mathbf{0.0777}$$

$$y_{\mathrm{O2}} = \frac{N_{\mathrm{O2}}}{N_m} = \frac{6.5 \text{ kmol}}{64.38 \text{ kmol}} = \mathbf{0.1010}$$

$$y_{\mathrm{N2}} = \frac{N_{\mathrm{N2}}}{N_m} = \frac{48.88 \text{ kmol}}{64.38 \text{ kmol}} = \mathbf{0.7592}$$

The mass of carbon dioxide in the products per unit mass of fuel burned is

$$\frac{m_{\mathrm{CO2}}}{m_{\mathrm{C4H10}}} = \frac{(4\times 44)\text{ kg}}{(1\times 58)\text{ kg}} = \mathbf{3.034\ kg\ CO_2/kg\ C_4H_{10}}$$

The air-fuel ratio is determined by taking the ratio of the mass of the air to the mass of the fuel,

$$\mathrm{AF} = \frac{m_{\mathrm{air}}}{m_{\mathrm{fuel}}} = \frac{(13\times 4.76\text{ kmol})(29\text{ kg/kmol})}{(1\text{ kmol})(58\text{ kg/kmol})} = \mathbf{30.94\ kg\ air/kg\ fuel}$$

15-21 n-Octane is burned with 50 percent excess air. The mole fractions of each of the products, the mass of water in the products per unit mass of the fuel, 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 combustion equation in this case can be written as

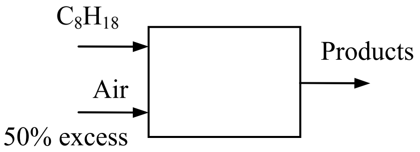
$$C_8H_{18} + 1.5a_{th}\left[O_2 + 3.76N_2\right] \longrightarrow 8CO_2 + 9H_2O + 0.5a_{th}O_2 + (1.5\times 3.76)a_{th}N_2$$

where a_{th} is the stoichiometric coefficient for air. We have automatically accounted for the 50% excess air by using the factor $1.5a_{th}$ instead of a_{th} for air. The stoichiometric amount of oxygen ($a_{th}O_2$) will be used to oxidize the fuel, and the remaining excess amount ($0.5a_{th}O_2$) will appear in the products as free oxygen. The coefficient a_{th} is determined from the O_2 balance,

O_2 balance: $\quad 1.5a_{th} = 8 + 4.5 + 0.5a_{th} \longrightarrow a_{th} = 12.5$

Substituting, $\quad C_8H_{18} + 18.75\left[O_2 + 3.76N_2\right] \longrightarrow 8CO_2 + 9H_2O + 6.25O_2 + 70.5N_2$

The mass of each product and the total mass are



$$m_{CO2} = N_{CO2}M_{CO2} = (8\text{ kmol})(44\text{ kg/kmol}) = 352\text{ kg}$$
$$m_{H2O} = N_{H2O}M_{H2O} = (9\text{ kmol})(18\text{ kg/kmol}) = 162\text{ kg}$$
$$m_{O2} = N_{O2}M_{O2} = (6.25\text{ kmol})(32\text{ kg/kmol}) = 200\text{ kg}$$
$$m_{N2} = N_{N2}M_{N2} = (70.5\text{ kmol})(28\text{ kg/kmol}) = 1974\text{ kg}$$
$$m_{total} = m_{CO2} + m_{H2O} + m_{O2} + m_{N2} = 352 + 162 + 200 + 1974 = 2688\text{ kg}$$

Then the mass fractions are

$$\text{mf}_{CO2} = \frac{m_{CO2}}{m_{total}} = \frac{352\text{ kg}}{2688\text{ kg}} = \mathbf{0.1310}$$

$$\text{mf}_{H2O} = \frac{m_{H2O}}{m_{total}} = \frac{162\text{ kg}}{2688\text{ kg}} = \mathbf{0.0603}$$

$$\text{mf}_{O2} = \frac{m_{O2}}{m_{total}} = \frac{200\text{ kg}}{2688\text{ kg}} = \mathbf{0.0744}$$

$$\text{mf}_{N2} = \frac{m_{N2}}{m_{total}} = \frac{1974\text{ kg}}{2688\text{ kg}} = \mathbf{0.7344}$$

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

$$m_{C8H18} = N_{C8H18}M_{C8H18} = (1\text{ kmol})(114\text{ kg/kmol}) = 114\text{ kg}$$
$$m_{air} = N_{air}M_{air} = (17.75\times 4.76\text{ kmol})(29\text{ kg/kmol}) = 2588\text{ kg}$$
$$m_{total} = m_{C8H18} + m_{air} = 114 + 2588 = 2702\text{ kg}$$

Then the mass fractions of reactants are

$$\text{mf}_{C8H18} = \frac{m_{C8H18}}{m_{total}} = \frac{114\text{ kg}}{2702\text{ kg}} = \mathbf{0.0422}$$

$$\text{mf}_{air} = \frac{m_{air}}{m_{total}} = \frac{2588\text{ kg}}{2702\text{ kg}} = \mathbf{0.9578}$$

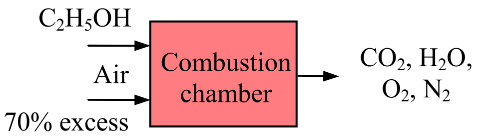
15-22 Ethyl alcohol is burned with 70% excess air. The mole fractions of the products and the reactants, the mass of water and oxygen in products per unit mass of fuel are to be determined.

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 The reaction with stoichiometric air is

$$C_2H_5OH + a_{th}[O_2 + 3.76N_2] \longrightarrow 2\,CO_2 + 3H_2O + a_{th} \times 3.76\,N_2$$



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

Substituting,

$$C_2H_5OH + 3[O_2 + 3.76N_2] \longrightarrow 2\,CO_2 + 3H_2O + 3 \times 3.76\,N_2$$

The reaction with 70% excess air can be written as

$$C_2H_5OH + 1.7 \times 3[O_2 + 3.76N_2] \longrightarrow 2\,CO_2 + 3\,H_2O + x\,O_2 + 1.7 \times 3 \times 3.76\,N_2$$

The coefficient x is determined from O_2 balance:

$$0.5 + 1.7 \times 3 = 2 + 1.5 + x \longrightarrow x = 2.1$$

Then, $\quad C_2H_5OH + 5.1[O_2 + 3.76N_2] \longrightarrow 2\,CO_2 + 3\,H_2O + 2.1\,O_2 + 19.18\,N_2$

The total moles of the products is

$$N_m = 2 + 3 + 2.1 + 19.18 = 26.28\text{ kmol}$$

The mole fractions of the products are

$$y_{CO2} = \frac{N_{CO2}}{N_m} = \frac{2\text{ kmol}}{26.28\text{ kmol}} = \mathbf{0.0761}$$

$$y_{H2O} = \frac{N_{H2O}}{N_m} = \frac{3\text{ kmol}}{26.28\text{ kmol}} = \mathbf{0.1142}$$

$$y_{O2} = \frac{N_{O2}}{N_m} = \frac{2.1\text{ kmol}}{26.28\text{ kmol}} = \mathbf{0.0799}$$

$$y_{N2} = \frac{N_{N2}}{N_m} = \frac{19.18\text{ kmol}}{26.28\text{ kmol}} = \mathbf{0.7298}$$

The total moles of the reactants is

$$N_m = 1 + 5.1 \times 4.76 = 25.28\text{ kmol}$$

The mole fractions of the reactants are

$$y_{C2H5OH} = \frac{N_{C2H5OH}}{N_m} = \frac{1\text{ kmol}}{25.28\text{ kmol}} = \mathbf{0.0396}$$

$$y_{air} = \frac{N_{air}}{N_m} = \frac{(5.1 \times 4.76)\text{ kmol}}{25.28\text{ kmol}} = \mathbf{0.9603}$$

The mass of water and oxygen in the products per unit mass of fuel burned is

$$\frac{m_{H2O}}{m_{C2H5OH}} = \frac{(3 \times 18)\text{ kg}}{(1 \times 46)\text{ kg}} = \mathbf{1.174\ kg\ H_2O/kg\ C_2H_5OH}$$

$$\frac{m_{O2}}{m_{C2H5OH}} = \frac{(2.1 \times 32)\text{ kg}}{(1 \times 46)\text{ kg}} = \mathbf{1.461 kg\ O_2/kg\ C_2H_5OH}$$

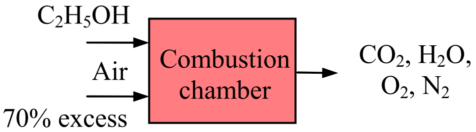
15-23 Ethyl alcohol is burned with 70% excess air. The air-fuel ratio is to be determined.

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 The reaction with stoichiometric air is

$$C_2H_5OH + a_{th}[O_2 + 3.76N_2] \longrightarrow 2\,CO_2 + 3H_2O + a_{th} \times 3.76\,N_2$$



where

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

Substituting,

$$C_2H_5OH + 3[O_2 + 3.76N_2] \longrightarrow 2\,CO_2 + 3H_2O + 3 \times 3.76\,N_2$$

The reaction with 70% excess air can be written as

$$C_2H_5OH + 1.7 \times 3[O_2 + 3.76N_2] \longrightarrow 2\,CO_2 + 3\,H_2O + x\,O_2 + 1.7 \times 3 \times 3.76\,N_2$$

The coefficient x is determined from O_2 balance:

$$0.5 + 1.7 \times 3 = 2 + 1.5 + x \longrightarrow x = 2.1$$

Then,

$$C_2H_5OH + 5.1[O_2 + 3.76N_2] \longrightarrow 2\,CO_2 + 3\,H_2O + 2.1\,O_2 + 19.18\,N_2$$

The air-fuel mass ratio is

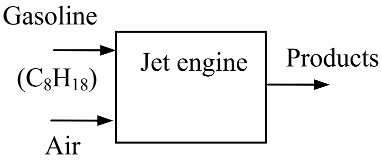
$$\mathrm{AF} = \frac{m_{\mathrm{air}}}{m_{\mathrm{fuel}}} = \frac{(5.1 \times 4.76 \times 29)\ \mathrm{kg}}{(1 \times 46)\ \mathrm{kg}} = \frac{704.0\ \mathrm{kg}}{46\ \mathrm{kg}} = \mathbf{15.30\ kg\ air/kg\ fuel}$$

15-24 Gasoline is burned steadily with air in a jet engine. The AF ratio is given. The percentage of excess air used is to be determined.

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

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

Analysis The theoretical combustion equation in this case can be written as



$$C_8H_{18} + a_{\mathrm{th}}[O_2 + 3.76N_2] \longrightarrow 8CO_2 + 9H_2O + 3.76a_{\mathrm{th}}N_2$$

where a_{th} is the stoichiometric coefficient for air. It is determined from

O_2 balance: $a_{\mathrm{th}} = 8 + 4.5 \longrightarrow a_{\mathrm{th}} = 12.5$

The air-fuel ratio for the theoretical reaction is determined by taking the ratio of the mass of the air to the mass of the fuel for,

$$\mathrm{AF_{th}} = \frac{m_{\mathrm{air,th}}}{m_{\mathrm{fuel}}} = \frac{(12.5 \times 4.76\ \mathrm{kmol})(29\ \mathrm{kg/kmol})}{(8\ \mathrm{kmol})(12\ \mathrm{kg/kmol}) + (9\ \mathrm{kmol})(2\ \mathrm{kg/kmol})} = 15.14\ \mathrm{kg\ air/kg\ fuel}$$

Then the percent theoretical air used can be determined from

$$\text{Percent theoretical air} = \frac{\mathrm{AF_{act}}}{\mathrm{AF_{th}}} = \frac{18\ \mathrm{kg\ air/kg\ fuel}}{15.14\ \mathrm{kg\ air/kg\ fuel}} = \mathbf{119\%}$$

15-25E Ethylene is burned with 175 percent theoretical air during a combustion process. The AF ratio and the dew-point temperature of the products are to be determined.

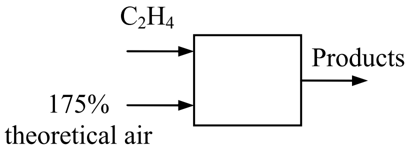
Assumptions **1** Combustion is complete. **2** The combustion products contain CO_2, H_2O, O_2, and N_2 only. **3** Combustion gases are ideal gases.

Properties The molar masses of C, H_2, and air are 12 lbm/lbmol, 2 lbm/lbmol, and 29 lbm/lbmol, respectively (Table A-1E).

Analysis (*a*) The combustion equation in this case can be written as

$$\mathrm{C_2H_4} + 1.75a_{\mathrm{th}}\left[\mathrm{O_2} + 3.76\mathrm{N_2}\right] \longrightarrow 2\mathrm{CO_2} + 2\mathrm{H_2O} + 0.75a_{\mathrm{th}}\mathrm{O_2} + (1.75\times 3.76)a_{\mathrm{th}}\mathrm{N_2}$$

where a_{th} is the stoichiometric coefficient for air. It is determined from



O_2 balance: $$1.75a_{\mathrm{th}} = 2 + 1 + 0.75a_{\mathrm{th}} \longrightarrow a_{\mathrm{th}} = 3$$

Substituting,

$$\mathrm{C_2H_4} + 5.25\left[\mathrm{O_2} + 3.76\mathrm{N_2}\right] \longrightarrow 2\mathrm{CO_2} + 2\mathrm{H_2O} + 2.25\mathrm{O_2} + 19.74\mathrm{N_2}$$

The air-fuel ratio is determined by taking the ratio of the mass of the air to the mass of the fuel,

$$\mathrm{AF} = \frac{m_{\mathrm{air}}}{m_{\mathrm{fuel}}} = \frac{(5.25\times 4.76\ \mathrm{lbmol})(29\ \mathrm{lbm/lbmol})}{(2\ \mathrm{lbmol})(12\ \mathrm{lbm/lbmol}) + (2\ \mathrm{lbmol})(2\ \mathrm{lbm/lbmol})} = \mathbf{25.9\ lbm\ air/lbm\ fuel}$$

(*b*) The dew-point temperature of a gas-vapor mixture is the saturation temperature of the water vapor in the product gases corresponding to its partial pressure. That is,

$$P_v = \left(\frac{N_v}{N_{\mathrm{prod}}}\right)P_{\mathrm{prod}} = \left(\frac{2\ \mathrm{lbmol}}{25.99\ \mathrm{lbmol}}\right)(14.5\ \mathrm{psia}) = 1.116\ \mathrm{psia}$$

Thus,

$$T_{\mathrm{dp}} = T_{\mathrm{sat@1.116\ psia}} = \mathbf{105.4^\circ F}$$

15-26 Propylene is burned with 50 percent excess air during a combustion process. The AF ratio and the temperature at which the water vapor in the products will start condensing are to be determined.

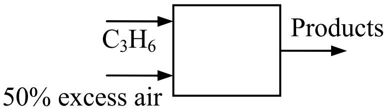
Assumptions **1** Combustion is complete. **2** The combustion products contain CO_2, H_2O, O_2, and N_2 only. **3** Combustion gases are ideal gases.

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

Analysis (*a*) The combustion equation in this case can be written as

$$C_3H_6 + 1.5a_{th}[O_2 + 3.76N_2] \longrightarrow 3CO_2 + 3H_2O + 0.5a_{th}O_2 + (1.5\times3.76)a_{th}N_2$$

where a_{th} is the stoichiometric coefficient for air. It is determined from



O_2 balance: $1.5a_{th} = 3 + 1.5 + 0.5a_{th} \longrightarrow a_{th} = 4.5$

Substituting,

$$C_3H_6 + 6.75[O_2 + 3.76N_2] \longrightarrow 3CO_2 + 3H_2O + 2.25O_2 + 25.38N_2$$

The air-fuel ratio is determined by taking the ratio of the mass of the air to the mass of the fuel,

$$\mathrm{AF} = \frac{m_{air}}{m_{fuel}} = \frac{(6.75\times4.76\text{ kmol})(29\text{ kg/kmol})}{(3\text{ kmol})(12\text{ kg/kmol}) + (3\text{ kmol})(2\text{ kg/kmol})} = \mathbf{22.2\ kg\ air/kg\ fuel}$$

(*b*) The dew-point temperature of a gas-vapor mixture is the saturation temperature of the water vapor in the product gases corresponding to its partial pressure. That is,

$$P_v = \left(\frac{N_v}{N_{prod}}\right)P_{prod} = \left(\frac{3\text{ kmol}}{33.63\text{ kmol}}\right)(105\text{ kPa}) = 9.367\text{ kPa}$$

Thus,

$$T_{dp} = T_{sat@9.367\text{ kPa}} = \mathbf{44.5°C}$$

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.

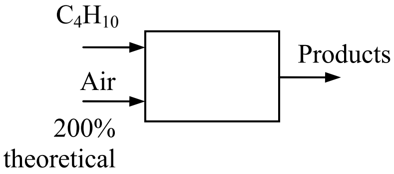
Assumptions **1** Combustion is complete. **2** The combustion products contain CO_2, H_2O, O_2, and N_2 only.

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

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

where a_{th} is the stoichiometric coefficient for air. We have automatically accounted for the 100% excess air by using the factor $2a_{\text{th}}$ 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: $2 \times 2a_{\text{th}} = 2B + D + 2E$

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

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

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

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

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

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

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

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

The vapor mole fraction is

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

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

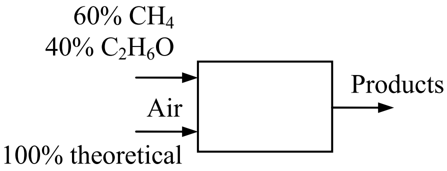
$$y_v = \frac{N_{\text{water}}}{N_{\text{total,product}}} \longrightarrow 0.1995 = \frac{5 + N_v}{4 + 5 + N_v + 6.5 + 48.88} \longrightarrow N_v = \mathbf{9.796\ kmol}$$

15-28 A fuel mixture of 60% by mass methane, CH_4, and 40% by mass ethanol, C_2H_6O, is burned completely with theoretical air. The required flow rate of air is to be determined.

Assumptions **1** Combustion is complete. **2** The combustion products contain CO_2, H_2O, 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 For 100 kg of fuel mixture, the mole numbers are



$$N_{\mathrm{CH4}} = \frac{\mathrm{mf}_{\mathrm{CH4}}}{M_{\mathrm{CH4}}} = \frac{60\ \mathrm{kg}}{16\ \mathrm{kg/kmol}} = 3.75\ \mathrm{kmol}$$

$$N_{\mathrm{C2H6O}} = \frac{\mathrm{mf}_{\mathrm{C2H6O}}}{M_{\mathrm{C2H6O}}} = \frac{40\ \mathrm{kg}}{46\ \mathrm{kg/kmol}} = 0.8696\ \mathrm{kmol}$$

The mole fraction of methane and ethanol in the fuel mixture are

$$x = \frac{N_{\mathrm{CH4}}}{N_{\mathrm{CH4}} + N_{\mathrm{C2H6O}}} = \frac{3.75\ \mathrm{kmol}}{(3.75 + 0.8696)\ \mathrm{kmol}} = 0.8118$$

$$y = \frac{N_{\mathrm{C2H6O}}}{N_{\mathrm{CH4}} + N_{\mathrm{C2H6O}}} = \frac{0.8696\ \mathrm{kmol}}{(3.75 + 0.8696)\ \mathrm{kmol}} = 0.1882$$

The combustion equation in this case can be written as

$$x\,\mathrm{CH_4} + y\,\mathrm{C_2H_6O} + a_{\mathrm{th}}[\mathrm{O_2} + 3.76\mathrm{N_2}] \longrightarrow B\,\mathrm{CO_2} + D\,\mathrm{H_2O} + F\,\mathrm{N_2}$$

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: $x + 2y = B$

Hydrogen balance: $4x + 6y = 2D$

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

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

Substituting x and y values into the equations and solving, we find the coefficients as

$$\begin{aligned} x &= 0.8118 & B &= 1.188 \\ y &= 0.1882 & D &= 2.188 \\ a_{\mathrm{th}} &= 2.188 & F &= 8.228 \end{aligned}$$

Then, we write the balanced reaction equation as

$$0.8118\,\mathrm{CH_4} + 0.1882\,\mathrm{C_2H_6O} + 2.188[\mathrm{O_2} + 3.76\mathrm{N_2}] \longrightarrow 1.188\,\mathrm{CO_2} + 2.188\,\mathrm{H_2O} + 8.228\,\mathrm{N_2}$$

The air-fuel ratio is determined by taking the ratio of the mass of the air to the mass of the fuel,

$$\begin{aligned} \mathrm{AF} = \frac{m_{\mathrm{air}}}{m_{\mathrm{fuel}}} &= \frac{(2.188 \times 4.76\ \mathrm{kmol})(29\ \mathrm{kg/kmol})}{(0.8118\ \mathrm{kmol})(12 + 4 \times 1)\mathrm{kg/kmol} + (0.1882\ \mathrm{kmol})(2 \times 12 + 6 \times 1 + 16)\mathrm{kg/kmol}} \\ &= 13.94\ \mathrm{kg\ air/kg\ fuel} \end{aligned}$$

Then, the required flow rate of air becomes

$$\dot{m}_{\mathrm{air}} = \mathrm{AF}\dot{m}_{\mathrm{fuel}} = (13.94)(10\ \mathrm{kg/s}) = \mathbf{139.4\ kg/s}$$

15-29 The volumetric fractions of the constituents of a certain natural gas are given. The AF ratio is to be determined if this gas is burned with the stoichiometric amount of dry air.

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

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

Analysis Considering 1 kmol of fuel, the combustion equation can be written as

$$(0.65CH_4 + 0.08H_2 + 0.18N_2 + 0.03O_2 + 0.06CO_2) + a_{th}(O_2 + 3.76N_2) \longrightarrow xCO_2 + yH_2O + zN_2$$

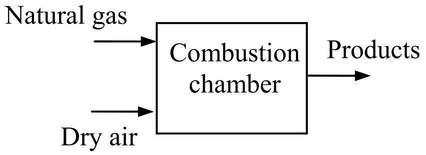
The unknown coefficients in the above equation are determined from mass balances,

$$\text{C}:\ 0.65 + 0.06 = x \longrightarrow x = 0.71$$

$$\text{H}:\ 0.65 \times 4 + 0.08 \times 2 = 2y \longrightarrow y = 1.38$$

$$\text{O}_2:\ 0.03 + 0.06 + a_{th} = x + y/2 \longrightarrow a_{th} = 1.31$$

$$\text{N}_2:\ 0.18 + 3.76a_{th} = z \longrightarrow z = 5.106$$



Thus,

$$(0.65CH_4 + 0.08H_2 + 0.18N_2 + 0.03O_2 + 0.06CO_2) + 1.31(O_2 + 3.76N_2) \longrightarrow 0.71CO_2 + 1.38H_2O + 5.106N_2$$

The air-fuel ratio for the this reaction is determined by taking the ratio of the mass of the air to the mass of the fuel,

$$m_{air} = (1.31 \times 4.76\ \text{kmol})(29\ \text{kg/kmol}) = 180.8\ \text{kg}$$

$$m_{fuel} = (0.65 \times 16 + 0.08 \times 2 + 0.18 \times 28 + 0.03 \times 32 + 0.06 \times 44)\text{kg} = 19.2\ \text{kg}$$

and

$$\text{AF}_{th} = \frac{m_{air,th}}{m_{fuel}} = \frac{180.8\ \text{kg}}{19.2\ \text{kg}} = \mathbf{9.42\ \ kg\ air/kg\ fuel}$$

15-30 The composition of a certain natural gas is given. The gas is burned with stoichiometric amount of moist air. The AF ratio is to be determined.

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

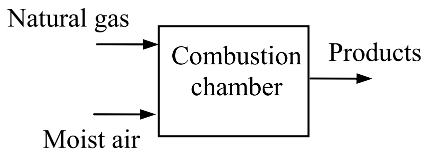
Properties The molar masses of C, H_2, N_2, O_2, and air are 12 kg/kmol, 2 kg/kmol, 28 kg/kmol, 32 kg/kmol, and 29 kg/kmol, respectively (Table A-1).

Analysis The fuel is burned completely with the stoichiometric amount of air, and thus the products will contain only H_2O, CO_2 and N_2, but no free O_2. The moisture in the air does not react with anything; it simply shows up as additional H_2O in the products. Therefore, we can simply balance the combustion equation using dry air, and then add the moisture to both sides of the equation. Considering 1 kmol of fuel, the combustion equation can be written as

$$(0.65CH_4 + 0.08H_2 + 0.18N_2 + 0.03O_2 + 0.06CO_2) + a_{th}(O_2 + 3.76N_2) \longrightarrow xCO_2 + yH_2O + zN_2$$

The unknown coefficients in the above equation are determined from mass balances,

$$\text{C}:\ 0.65 + 0.06 = x \longrightarrow x = 0.71$$
$$\text{H}:\ 0.65\times 4 + 0.08\times 2 = 2y \longrightarrow y = 1.38$$
$$O_2:\ 0.03 + 0.06 + a_{th} = x + y/2 \longrightarrow a_{th} = 1.31$$
$$N_2:\ 0.18 + 3.76a_{th} = z \longrightarrow z = 5.106$$



Thus,

$$(0.65CH_4 + 0.08H_2 + 0.18N_2 + 0.03O_2 + 0.06CO_2) + 1.31(O_2 + 3.76N_2) \longrightarrow 0.71CO_2 + 1.38H_2O + 5.106N_2$$

Next we determine the amount of moisture that accompanies $4.76a_{th} = (4.76)(1.31) = 6.24$ kmol of dry air. The partial pressure of the moisture in the air is

$$P_{v,\text{in}} = \phi_{\text{air}} P_{\text{sat@25°C}} = (0.85)(3.1698\text{ kPa}) = 2.694\text{ kPa}$$

Assuming ideal gas behavior, the number of moles of the moisture in the air ($N_{v,\text{ in}}$) is determined to be

$$N_{v,\text{in}} = \left(\frac{P_{v,\text{in}}}{P_{\text{total}}}\right) N_{\text{total}} = \left(\frac{2.694\text{ kPa}}{101.325\text{ kPa}}\right)(6.24 + N_{v,\text{in}}) \longrightarrow N_{v,\text{air}} = 0.17\text{ kmol}$$

The balanced combustion equation is obtained by substituting the coefficients determined earlier and adding 0.17 kmol of H_2O to both sides of the equation,

$$(0.65CH_4 + 0.08H_2 + 0.18N_2 + 0.03O_2 + 0.06CO_2) + 1.31(O_2 + 3.76N_2) + 0.17H_2O \longrightarrow 0.71CO_2 + 1.55H_2O + 5.106N_2$$

The air-fuel ratio for the this reaction is determined by taking the ratio of the mass of the air to the mass of the fuel,

$$m_{\text{air}} = (1.31\times 4.76\text{ kmol})(29\text{ kg/kmol}) + (0.17\text{ kmol}\times 18\text{ kg/kmol}) = 183.9\text{ kg}$$
$$m_{\text{fuel}} = (0.65\times 16 + 0.08\times 2 + 0.18\times 28 + 0.03\times 32 + 0.06\times 44)\text{kg} = 19.2\text{ kg}$$

and

$$\text{AF}_{\text{th}} = \frac{m_{\text{air,th}}}{m_{\text{fuel}}} = \frac{183.9\text{kg}}{19.2\text{kg}} = \mathbf{9.58\ kg\ air/kg\ fuel}$$

15-31 The composition of a gaseous fuel is given. It is burned with 130 percent theoretical air. The AF ratio and the fraction of water vapor that would condense if the product gases were cooled 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, N_2, and air are 12 kg/kmol, 2 kg/kmol, 28 kg/kmol, and 29 kg/kmol, respectively (Table A-1).

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

$$(0.45\mathrm{CH_4} + 0.35\mathrm{H_2} + 0.20\mathrm{N_2}) + 1.3a_{\mathrm{th}}(\mathrm{O_2} + 3.76\mathrm{N_2}) \longrightarrow x\mathrm{CO_2} + y\mathrm{H_2O} + 0.3a_{\mathrm{th}}\mathrm{O_2} + z\mathrm{N_2}$$

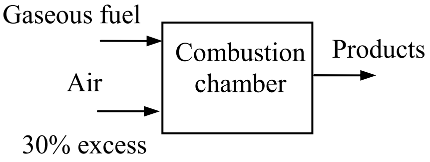
The unknown coefficients in the above equation are determined from mass balances,

$$\mathrm{C}:\ 0.45 = x \longrightarrow x = 0.45$$

$$\mathrm{H}:\ 0.45\times 4 + 0.35\times 2 = 2y \longrightarrow y = 1.2$$

$$\mathrm{O_2}:\ 1.3a_{\mathrm{th}} = x + y/2 + 0.3a_{\mathrm{th}} \longrightarrow a_{\mathrm{th}} = 1.05$$

$$\mathrm{N_2}:\ 0.20 + 3.76\times 1.3a_{\mathrm{th}} = z \longrightarrow z = 5.332$$



Thus,

$$(0.45\mathrm{CH_4} + 0.35\mathrm{H_2} + 0.20\mathrm{N_2}) + 1.365(\mathrm{O_2} + 3.76\mathrm{N_2}) \longrightarrow 0.45\mathrm{CO_2} + 1.2\mathrm{H_2O} + 0.315\mathrm{O_2} + 5.332\mathrm{N_2}$$

The air-fuel ratio for the this reaction is determined by taking the ratio of the mass of the air to the mass of the fuel,

$$m_{\mathrm{air}} = (1.365\times 4.76\text{ kmol})(29\text{ kg/kmol}) = 188.4\text{ kg}$$
$$m_{\mathrm{fuel}} = (0.45\times 16 + 0.35\times 2 + 0.2\times 28)\text{kg} = 13.5\text{ kg}$$

and

$$\mathrm{AF} = \frac{m_{\mathrm{air}}}{m_{\mathrm{fuel}}} = \frac{188.4\text{ kg}}{13.5\text{ kg}} = \mathbf{13.96\ \ kg\ air/kg\ fuel}$$

(*b*) For each kmol of fuel burned, 0.45 + 1.2 + 0.315 + 5.332 = 7.297 kmol of products are formed, including 1.2 kmol of H_2O. Assuming that the dew-point temperature of the products is above 25°C, some of the water vapor will condense as the products are cooled to 25°C. If N_w kmol of H_2O condenses, there will be 1.2 - N_w kmol of water vapor left in the products. The mole number of the products in the gas phase will also decrease to 7.297 − N_w as a result. Treating the product gases (including the remaining water vapor) as ideal gases, N_w is determined by equating the mole fraction of the water vapor to its pressure fraction,

$$\frac{N_v}{N_{\mathrm{prod,gas}}} = \frac{P_v}{P_{\mathrm{prod}}} \longrightarrow \frac{1.2 - N_w}{7.297 - N_w} = \frac{3.1698\text{ kPa}}{101.325\text{ kPa}} \longrightarrow N_w = 1.003\text{ kmol}$$

since $P_v = P_{\mathrm{sat\ @\ 25°C}} = 3.1698$ kPa. Thus the fraction of water vapor that condenses is 1.003/1.2 = 0.836 or **84%**.

15-32 Problem 15-31 is reconsidered. The effects of varying the percentages of CH4, H2 and N2 making up the fuel and the product gas temperature are to be studied.

Analysis The problem is solved using EES, and the solution is given below.

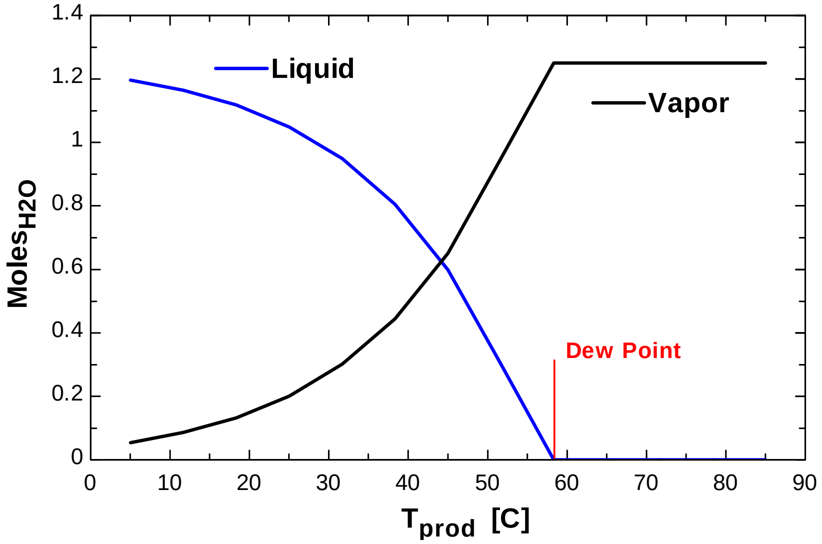
```
Let's modify this problem to include the fuels butane, ethane, methane, and propane in
pull down menu. Adiabatic Combustion of fuel CnHm entering at T_fuel with Stoichiometric Air at T_air:
Reaction:  aCxHy+bH2+cN2  +  (a*y/4 + a*x+b/2) (Theo_air/100) (O2 + 3.76 N2)
          <--> a*xCO2  +  ((a*y/2)+b) H2O  +  (c+3.76  (a*y/4 + a*x+b/2) (Theo_air/100))  N2 +  (a*y/4 + a*x+b/2)
(Theo_air/100 - 1) O2

T_prod is the product gas temperature.
Theo_air is the % theoretical air. "
Procedure
H20Cond(P_prod,T_prod,Moles_H2O,M_other:T_DewPoint,Moles_H2O_vap,Moles_H2O_liq,Result$)
P_v = Moles_H2O/(M_other+Moles_H2O)*P_prod
T_DewPoint = temperature(steam,P=P_v,x=0)
IF T_DewPoint <= T_prod then
Moles_H2O_vap = Moles_H2O
Moles_H2O_liq=0
Result$='No condensation occurred'
ELSE
Pv_new=pressure(steam,T=T_prod,x=0)
Moles_H2O_vap=Pv_new/P_prod*M_other/(1-Pv_new/P_prod)
Moles_H2O_liq = Moles_H2O - Moles_H2O_vap
Result$='There is condensation'
ENDIF
END

"Input data from the diagram window"
{P_prod = 101.325 [kPa]
Theo_air = 130 "[%]"
a=0.45
b=0.35
c=0.20
T_prod = 25 [C]}
Fuel$='CH4'
x=1
y=4

"Composition of Product gases:"
A_th = a*y/4 +a* x+b/2
AF_ratio = 4.76*A_th*Theo_air/100*molarmass(Air)/(a*16+b*2+c*28) "[kg_air/kg_fuel]"
Moles_O2=(a*y/4 +a* x+b/2) *(Theo_air/100 - 1)
Moles_N2=c+(3.76*(a*y/4 + a*x+b/2))* (Theo_air/100)
Moles_CO2=a*x
Moles_H2O=a*y/2+b
M_other=Moles_O2+Moles_N2+Moles_CO2
Call H20Cond(P_prod,T_prod,Moles_H2O,M_other:T_DewPoint,Moles_H2O_vap,Moles_H2O_liq,Result$)
Frac_cond = Moles_H2O_liq/Moles_H2O*Convert(, %) "[%]"
"Reaction:  aCxHy+bH2+cN2  +  A_th Theo_air/100 (O2 + 3.76 N2)
<--> a*xCO2  +  (a*y/2+b) H2O  +  (c+3.76  A_th Theo_air/100) N2 +  A_th (Theo_air/100 - 1) O2"
```

AF_{ratio} [kg_{air}/ kg_{fuel}]	$Frac_{cond}$ [%]	$Moles_{H2O,liq}$	$Moles_{H2O,vap}$	T_{prod} [C]
14.27	95.67	1.196	0.05409	5
14.27	93.16	1.165	0.08549	11.67
14.27	89.42	1.118	0.1323	18.33
14.27	83.92	1.049	0.201	25
14.27	75.94	0.9492	0.3008	31.67
14.27	64.44	0.8055	0.4445	38.33
14.27	47.92	0.599	0.651	45
14.27	24.06	0.3008	0.9492	51.67
14.27	0	0	1.25	58.33
14.27	0	0	1.25	65
14.27	0	0	1.25	71.67
14.27	0	0	1.25	78.33
14.27	0	0	1.25	85



15-33 Carbon is burned with dry air. The volumetric analysis of the products is given. The AF ratio and the percentage of theoretical air used are to be determined.

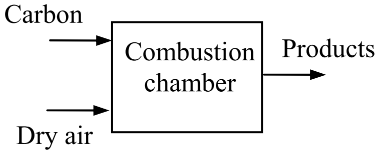
Assumptions **1** Combustion is complete. **2** The combustion products contain CO_2, CO, O_2, and N_2 only.

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

Analysis Considering 100 kmol of dry products, the combustion equation can be written as

$$\mathrm{xC} + \mathrm{a}[\mathrm{O}_2 + 3.76\mathrm{N}_2] \longrightarrow 10.06\mathrm{CO}_2 + 0.42\mathrm{CO} + 10.69\mathrm{O}_2 + 78.83\mathrm{N}_2$$

The unknown coefficients *x* and *a* are determined from mass balances,



$$\mathrm{N}_2:\ 3.76a = 78.83 \longrightarrow a = 20.965$$

$$\mathrm{C}:\ x = 10.06 + 0.42 \longrightarrow x = 10.48$$

$$(\text{Check}\,\mathrm{O}_2:\ a = 10.06 + 0.21 + 10.69 \longrightarrow 20.96 = 20.96)$$

Thus,

$$10.48\mathrm{C} + 20.96[\mathrm{O}_2 + 3.76\mathrm{N}_2] \longrightarrow 10.06\mathrm{CO}_2 + 0.42\mathrm{CO} + 10.69\mathrm{O}_2 + 78.83\mathrm{N}_2$$

The combustion equation for 1 kmol of fuel is obtained by dividing the above equation by 10.48,

$$\mathrm{C} + 2.0[\mathrm{O}_2 + 3.76\mathrm{N}_2] \longrightarrow 0.96\mathrm{CO}_2 + 0.04\mathrm{CO} + 1.02\mathrm{O}_2 + 7.52\mathrm{N}_2$$

(*a*) The air-fuel ratio is determined by taking the ratio of the mass of the air to the mass of the fuel,

$$\mathrm{AF} = \frac{m_{\mathrm{air}}}{m_{\mathrm{fuel}}} = \frac{(2.0\times 4.76\ \mathrm{kmol})(29\ \mathrm{kg/kmol})}{(1\ \mathrm{kmol})(12\ \mathrm{kg/kmol})} = \mathbf{23.0\ \ kg\ air/kg\ fuel}$$

(*b*) To find the percent theoretical air used, we need to know the theoretical amount of air, which is determined from the theoretical combustion equation of the fuel,

$$\mathrm{C} + 1[\mathrm{O}_2 + 3.76\mathrm{N}_2] \longrightarrow \mathrm{CO}_2 + 3.76\mathrm{N}_2$$

Then,

$$\text{Percent theoretical air} = \frac{m_{\mathrm{air,act}}}{m_{\mathrm{air,th}}} = \frac{N_{\mathrm{air,act}}}{N_{\mathrm{air,th}}} = \frac{(2.0)(4.76)\ \mathrm{kmol}}{(1.0)(4.76)\ \mathrm{kmol}} = \mathbf{200\%}$$

15-34 Methane is burned with dry air. The volumetric analysis of the products is given. The AF ratio and the percentage of theoretical air used 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, and air are 12 kg/kmol, 2 kg/kmol, and 29 kg/kmol, respectively (Table A-1).

Analysis Considering 100 kmol of dry products, the combustion equation can be written as

$$x\mathrm{CH_4} + a[\mathrm{O_2} + 3.76\mathrm{N_2}] \longrightarrow 5.20\mathrm{CO_2} + 0.33\mathrm{CO} + 11.24\mathrm{O_2} + 83.23\mathrm{N_2} + b\mathrm{H_2O}$$

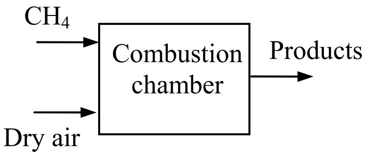
The unknown coefficients x, a, and b are determined from mass balances,

$$\mathrm{N_2}:\ 3.76a = 83.23 \longrightarrow a = 22.14$$

$$\mathrm{C}:\ x = 5.20 + 0.33 \longrightarrow x = 5.53$$

$$\mathrm{H}:\ 4x = 2b \longrightarrow b = 11.06$$

$$(\text{Check } \mathrm{O_2}:\ a = 5.20 + 0.165 + 11.24 + b/2 \longrightarrow 22.14 = 22.14)$$



Thus,

$$5.53\mathrm{CH_4} + 22.14[\mathrm{O_2} + 3.76\mathrm{N_2}] \longrightarrow 5.20\mathrm{CO_2} + 0.33\mathrm{CO} + 11.24\mathrm{O_2} + 83.23\mathrm{N_2} + 11.06\mathrm{H_2O}$$

The combustion equation for 1 kmol of fuel is obtained by dividing the above equation by 5.53,

$$\mathrm{CH_4} + 4.0[\mathrm{O_2} + 3.76\mathrm{N_2}] \longrightarrow 0.94\mathrm{CO_2} + 0.06\mathrm{CO} + 2.03\mathrm{O_2} + 15.05\mathrm{N_2} + 2\mathrm{H_2O}$$

(*a*) The air-fuel ratio is determined from its definition,

$$\mathrm{AF} = \frac{m_{\mathrm{air}}}{m_{\mathrm{fuel}}} = \frac{(4.0 \times 4.76\ \mathrm{kmol})(29\ \mathrm{kg/kmol})}{(1\ \mathrm{kmol})(12\ \mathrm{kg/kmol}) + (2\ \mathrm{kmol})(2\ \mathrm{kg/kmol})} = \mathbf{34.5\ \ kg\ air/kg\ fuel}$$

(*b*) To find the percent theoretical air used, we need to know the theoretical amount of air, which is determined from the theoretical combustion equation of the fuel,

$$\mathrm{CH_4} + a_{\mathrm{th}}[\mathrm{O_2} + 3.76\mathrm{N_2}] \longrightarrow \mathrm{CO_2} + 2\mathrm{H_2O} + 3.76a_{\mathrm{th}}\mathrm{N_2}$$

$$\mathrm{O_2}:\qquad a_{\mathrm{th}} = 1 + 1 \longrightarrow a_{\mathrm{th}} = 2.0$$

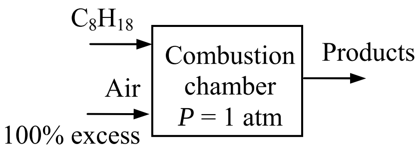
Then,

$$\text{Percent theoretical air} = \frac{m_{\mathrm{air,act}}}{m_{\mathrm{air,th}}} = \frac{N_{\mathrm{air,act}}}{N_{\mathrm{air,th}}} = \frac{(4.0)(4.76)\ \mathrm{kmol}}{(2.0)(4.76)\ \mathrm{kmol}} = \mathbf{200\%}$$

15-35 n-Octane is burned with 100% excess air. The combustion is incomplete. The mole fractions of products and the dew-point temperature of the water vapor in the products 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 The combustion reaction for stoichiometric air is

$$C_8H_{18}+12.5[O_2+3.76N_2]\longrightarrow 8CO_2+9H_2O+(12.5\times 3.76)N_2$$

The combustion equation with 100% excess air and incomplete combustion is

$$C_8H_{18}+2\times 12.5[O_2+3.76N_2]\longrightarrow (0.85\times 8)\,CO_2+(0.15\times 8)\,CO+9H_2O+x\,O_2+(2\times 12.5\times 3.76)\,N_2$$

The coefficient for CO is determined from a mass balance,

O_2 balance: $\quad 25=0.85\times 8+0.5\times 0.15\times 8+0.5\times 9+x\longrightarrow x=13.1$

Substituting,

$$C_8H_{18}+25[O_2+3.76N_2]\longrightarrow 6.8\,CO_2+1.2\,CO+9\,H_2O+13.1\,O_2+94\,N_2$$

The mole fractions of the products are

$$N_{\text{prod}}=6.8+1.2+9+13.1+94=124.1\,\text{kmol}$$

$$y_{CO2}=\frac{N_{CO2}}{N_{\text{prod}}}=\frac{6.8\text{ kmol}}{124.1\text{ kmol}}=\mathbf{0.0548}$$

$$y_{CO}=\frac{N_{CO}}{N_{\text{prod}}}=\frac{1.2\text{ kmol}}{124.1\text{ kmol}}=\mathbf{0.0097}$$

$$y_{H2O}=\frac{N_{H2O}}{N_{\text{prod}}}=\frac{9\text{ kmol}}{124.1\text{ kmol}}=\mathbf{0.0725}$$

$$y_{O2}=\frac{N_{O2}}{N_{\text{prod}}}=\frac{13.1\text{ kmol}}{124.1\text{ kmol}}=\mathbf{0.1056}$$

$$y_{N2}=\frac{N_{N2}}{N_{\text{prod}}}=\frac{94\text{ kmol}}{124.1\text{ kmol}}=\mathbf{0.7575}$$

The dew-point temperature of a gas-vapor mixture is the saturation temperature of the water vapor in the product gases corresponding to its partial pressure. That is,

$$P_v=\left(\frac{N_v}{N_{\text{prod}}}\right)P_{\text{prod}}=\left(\frac{9\text{ kmol}}{124.1\text{ kmol}}\right)(101.325\text{ kPa})=7.348\text{ kPa}$$

Thus,

$$T_{\text{dp}}=T_{\text{sat@7.348 kPa}}=\mathbf{39.9^\circ C} \qquad \text{(Table A-5 or EES)}$$

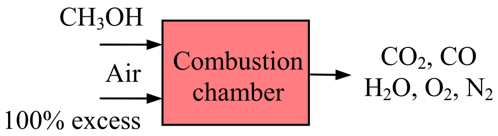
15-36 Methyl alcohol is burned with 100% excess air. The combustion is incomplete. The balanced chemical reaction is to be written and the air-fuel ratio is to be determined.

Assumptions **1** Combustion is incomplete. **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 The balanced reaction equation for stoichiometric air is

$$CH_3OH + a_{th}[O_2 + 3.76N_2] \longrightarrow CO_2 + 2\,H_2O + a_{th} \times 3.76\,N_2$$



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

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

Substituting,

$$CH_3OH + 1.5[O_2 + 3.76N_2] \longrightarrow CO_2 + 2\,H_2O + 1.5 \times 3.76\,N_2$$

The reaction with 100% excess air and incomplete combustion can be written as

$$CH_3OH + 2 \times 1.5[O_2 + 3.76N_2] \longrightarrow 0.60\,CO_2 + 0.40\,CO + 2\,H_2O + x\,O_2 + 2 \times 1.5 \times 3.76\,N_2$$

The coefficient for O_2 is determined from a mass balance,

O_2 balance: $\quad 0.5 + 2 \times 1.5 = 0.6 + 0.2 + 1 + x \longrightarrow x = 1.7$

Substituting,

$$CH_3OH + 3[O_2 + 3.76N_2] \longrightarrow 0.6\,CO_2 + 0.4\,CO + 2\,H_2O + 1.7\,O_2 + 11.28\,N_2$$

The air-fuel mass ratio is

$$\text{AF} = \frac{m_{\text{air}}}{m_{\text{fuel}}} = \frac{(3 \times 4.76 \times 29)\text{ kg}}{(1 \times 32)\text{ kg}} = \frac{414.1\text{ kg}}{32\text{ kg}} = \mathbf{12.94\ kg\ air/kg\ fuel}$$

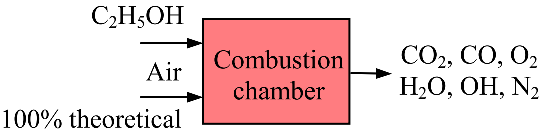
15-37 Ethyl alcohol is burned with stoichiometric amount of air. The combustion is incomplete. The apparent molecular weight of the products is to be determined.

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

Properties The molar masses of C, H_2, OH, N_2 and air are 12 kg/kmol, 2 kg/kmol, 17 kg/kmol, 28 kg/kmol, and 29 kg/kmol, respectively (Table A-1).

Analysis The reaction with stoichiometric air is

$$C_2H_5OH + a_{th}[O_2 + 3.76N_2] \longrightarrow 2\,CO_2 + 3H_2O + a_{th} \times 3.76\,N_2$$



where

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

Substituting,

$$C_2H_5OH + 3[O_2 + 3.76N_2] \longrightarrow 2\,CO_2 + 3H_2O + 3 \times 3.76\,N_2$$

The balanced reaction equation with incomplete combustion is

$$C_2H_5OH + 3[O_2 + 3.76N_2] \longrightarrow 2(0.90\,CO_2 + 0.10\,CO) + 3(0.95H_2O + 0.1\,OH) + bO_2 + 3 \times 3.76\,N_2$$

O_2 balance: $0.5+3 = 1.8+0.1+3.15/2+b \rightarrow b = 0.025$

which can be written as

$$C_2H_5OH + 3[O_2 + 3.76N_2] \longrightarrow 1.80\,CO_2 + 0.20\,CO + 2.85\,H_2O + 0.3\,OH + 0.025O_2 + 11.28\,N_2$$

The total moles of the products is

$$N_m = 1.8 + 0.2 + 2.85 + 0.3 + 0.025 + 11.28 = 16.64\text{ kmol}$$

The apparent molecular weight of the product gas is

$$M_m = \frac{m_m}{N_m} = \frac{(1.8 \times 44 + 0.20 \times 28 + 2.85 \times 18 + 0.3 \times 17 + 0.025 \times 32 + 11.28 \times 28)\text{ kg}}{16.64\text{ kmol}} = \mathbf{27.83\ kg/kmol}$$

15-38 Coal whose mass percentages are specified is burned with stoichiometric amount of air. The mass fractions of the products and the air-fuel ratio are to be determined.

Assumptions **1** Combustion is complete. **2** The combustion products contain CO_2, H_2O, SO_2, and N_2. **3** Combustion gases are ideal gases.

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

Analysis We consider 100 kg of coal for simplicity. Noting that the mass percentages in this case correspond to the masses of the constituents, the mole numbers of the constituent of the coal are determined to be

$$N_C = \frac{m_C}{M_C} = \frac{79.61\text{ kg}}{12\text{ kg/kmol}} = 6.634\text{ kmol}$$

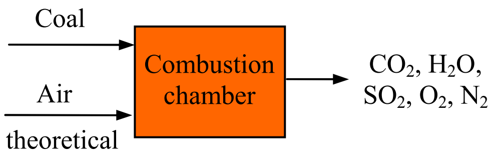
$$N_{H2} = \frac{m_{H2}}{M_{H2}} = \frac{4.66\text{ kg}}{2\text{ kg/kmol}} = 2.33\text{ kmol}$$

$$N_{O2} = \frac{m_{O2}}{M_{O2}} = \frac{4.76\text{ kg}}{32\text{ kg/kmol}} = 0.1488\text{ kmol}$$

$$N_{N2} = \frac{m_{N2}}{M_{N2}} = \frac{1.83\text{ kg}}{28\text{ kg/kmol}} = 0.06536\text{ kmol}$$

$$N_S = \frac{m_S}{M_S} = \frac{0.52\text{ kg}}{32\text{ kg/kmol}} = 0.01625\text{ kmol}$$

79.61% C
4.66% H_2
4.76% O_2
1.83% N_2
0.52% S
8.62% ash
(by mass)



Ash consists of the non-combustible matter in coal. Therefore, the mass of ash content that enters the combustion chamber is equal to the mass content that leaves. Disregarding this non-reacting component for simplicity, the combustion equation may be written as

$$6.634C + 2.33H_2 + 0.1488O_2 + 0.06536N_2 + 0.01625S + a_{th}(O_2 + 3.76N_2) \longrightarrow xCO_2 + yH_2O + zSO_2 + wN_2$$

Performing mass balances for the constituents gives

$$\text{C balance: } x = 6.634$$

$$H_2 \text{ balance: } y = 2.33$$

$$\text{S balance: } z = 0.01625$$

$$O_2 \text{ balance: } 0.1488 + a_{th} = x + 0.5y + z \longrightarrow a_{th} = 6.634 + 0.5(2.33) + 0.01625 - 0.1488 = 7.667$$

$$N_2 \text{ balance: } w = 0.06536 + 3.76a_{th} = 0.06536 + 3.76 \times 7.667 = 28.89$$

Substituting, the balanced combustion equation without the ash becomes

$$6.634C + 2.33H_2 + 0.1488O_2 + 0.06536N_2 + 0.01625S + 7.667(O_2 + 3.76N_2)$$
$$\longrightarrow 6.634CO_2 + 2.33H_2O + 0.01625SO_2 + 28.89N_2$$

The mass fractions of the products are

$$m_{total} = 6.634 \times 44 + 2.33 \times 18 + 0.01625 \times 64 + 228.89 \times 28 = 1144\text{ kg}$$

$$\text{mf}_{CO2} = \frac{m_{CO2}}{m_{total}} = \frac{(6.634 \times 44)\text{ kg}}{1144\text{ kg}} = \mathbf{0.2552}$$

$$\text{mf}_{H2O} = \frac{m_{H2O}}{m_{total}} = \frac{(2.33 \times 18)\text{ kg}}{1144\text{ kg}} = \mathbf{0.0367}$$

$$\text{mf}_{SO2} = \frac{m_{SO2}}{m_{total}} = \frac{(0.01625 \times 64)\text{ kg}}{1144\text{ kg}} = \mathbf{0.00091}$$

$$\text{mf}_{N2} = \frac{m_{N2}}{m_{total}} = \frac{(28.89 \times 28)\text{ kg}}{1144\text{ kg}} = \mathbf{0.7072}$$

The air-fuel mass ratio is then

$$\text{AF} = \frac{m_{air}}{m_{fuel}} = \frac{(7.667 \times 4.76 \times 29)\text{ kg}}{(6.634 \times 12 + 2.33 \times 2 + 0.1488 \times 32 + 0.06536 \times 28 + 0.01625 \times 32)\text{ kg}} = \frac{1058\text{ kg}}{91.38\text{ kg}} = \mathbf{11.58\text{ kg air/kg fuel}}$$

15-39 Coal whose mass percentages are specified is burned with 40% excess air. The air-fuel ratio and the apparent molecular weight of the product gas are to be determined.

Assumptions **1** Combustion is complete. **2** The combustion products contain CO_2, CO, H_2O, SO_2, and N_2. **3** Combustion gases are ideal gases.

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

Analysis We consider 100 kg of coal for simplicity. Noting that the mass percentages in this case correspond to the masses of the constituents, the mole numbers of the constituent of the coal are determined to be

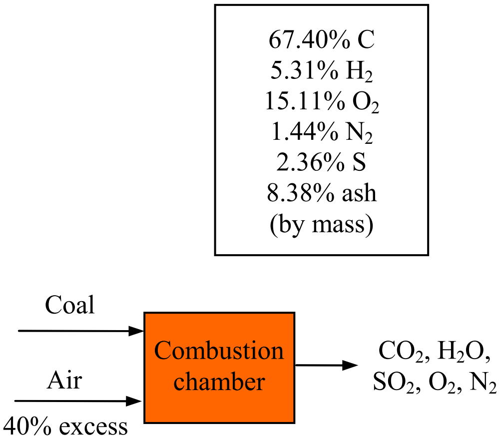
$$N_{\rm C}=\frac{m_{\rm C}}{M_{\rm C}}=\frac{67.40\text{ kg}}{12\text{ kg/kmol}}=5.617\text{ kmol}$$

$$N_{\rm H2}=\frac{m_{\rm H2}}{M_{\rm H2}}=\frac{5.31\text{ kg}}{2\text{ kg/kmol}}=2.655\text{ kmol}$$

$$N_{\rm O2}=\frac{m_{\rm O2}}{M_{\rm O2}}=\frac{15.11\text{ kg}}{32\text{ kg/kmol}}=0.4722\text{ kmol}$$

$$N_{\rm N2}=\frac{m_{\rm N2}}{M_{\rm N2}}=\frac{1.44\text{ kg}}{28\text{ kg/kmol}}=0.05143\text{ kmol}$$

$$N_{\rm S}=\frac{m_{\rm S}}{M_{\rm S}}=\frac{2.36\text{ kg}}{32\text{ kg/kmol}}=0.07375\text{ kmol}$$



The mole number of the mixture and the mole fractions are

$$N_m=5.617+2.655+0.4722+0.05143+0.07375=8.869\text{ kmol}$$

$$y_{\rm C}=\frac{N_{\rm C}}{N_m}=\frac{5.617\text{ kmol}}{8.869\text{ kmol}}=0.6333$$

$$y_{\rm H2}=\frac{N_{\rm H2}}{N_m}=\frac{2.655\text{ kmol}}{8.869\text{ kmol}}=0.2994$$

$$y_{\rm O2}=\frac{N_{\rm O2}}{N_m}=\frac{0.4722\text{ kmol}}{8.869\text{ kmol}}=0.05323$$

$$y_{\rm N2}=\frac{N_{\rm N2}}{N_m}=\frac{0.05143\text{ kmol}}{8.869\text{ kmol}}=0.00580$$

$$y_{\rm S}=\frac{N_{\rm S}}{N_m}=\frac{0.07375\text{ kmol}}{8.869\text{ kmol}}=0.00832$$

Ash consists of the non-combustible matter in coal. Therefore, the mass of ash content that enters the combustion chamber is equal to the mass content that leaves. Disregarding this non-reacting component for simplicity, the combustion equation may be written as

$$0.6333\text{C}+0.2994\text{H}_2+0.05323\text{O}_2+0.00580\text{N}_2+0.00832\text{S}+1.4a_{\rm th}(\text{O}_2+3.76\text{N}_2)$$
$$\longrightarrow x\text{CO}_2+y\text{H}_2\text{O}+z\text{SO}_2+k\text{N}_2+m\text{O}_2$$

According to the species balances,

$$\text{C balance: } x=0.6333$$
$$\text{H}_2\text{ balance: } y=0.2994$$
$$\text{S balance: } z=0.00832$$
$$\text{O}_2\text{ balance:}$$
$$0.05323+a_{\rm th}=x+0.5y+z$$
$$a_{\rm th}=0.6333+0.5\times0.2994+0.00832-0.05323=0.7381$$
$$\text{N}_2\text{ balance: } k=0.00580+1.4\times3.76a_{\rm th}=0.00580+1.4\times3.76\times0.7381=3.891$$
$$m=0.4a_{\rm th}=0.4\times0.7381=0.2952$$

Substituting,

$$0.6333\text{C} + 0.2994\text{H}_2 + 0.05323\text{O}_2 + 0.00580\text{N}_2 + 0.00832\text{S} + 1.033(\text{O}_2 + 3.76\text{N}_2)$$
$$\longrightarrow 0.6333\text{CO}_2 + 0.2994\text{H}_2\text{O} + 0.00832\text{SO}_2 + 3.891\text{N}_2 + 0.2952\text{O}_2$$

The total mass of the products is

$$m_{\text{total}} = 0.6333 \times 44 + 0.2994 \times 18 + 0.00832 \times 64 + 3.891 \times 28 + 0.2952 \times 32 = 152.2\text{ kg}$$

The total mole number of the products is

$$N_m = 0.6333 + 0.2994 + 0.00832 + 3.891 + 0.2952 = 5.127\text{ kmol}$$

The apparent molecular weight of the product gas is

$$M_m = \frac{m_m}{N_m} = \frac{152.2\text{ kg}}{5.127\text{ kmol}} = \mathbf{29.68\ kg/kmol}$$

The air-fuel mass ratio is then

$$\begin{aligned}\text{AF} = \frac{m_{\text{air}}}{m_{\text{fuel}}} &= \frac{(1.033 \times 4.76 \times 29)\text{ kg}}{(0.6333 \times 12 + 0.2994 \times 2 + 0.05323 \times 32 + 0.00580 \times 28 + 0.00832 \times 32)\text{ kg}} \\ &= \frac{142.6\text{ kg}}{10.33\text{ kg}} \\ &= \mathbf{13.80\ kg\ air/kg\ fuel}\end{aligned}$$

Enthalpy of Formation and Enthalpy of Combustion

15-40C For combustion processes the enthalpy of reaction is referred to as the enthalpy of combustion, which represents the amount of heat released during a steady-flow combustion process.

15-41C Enthalpy of formation is the enthalpy of a substance due to its chemical composition. The enthalpy of formation is related to elements or compounds whereas the enthalpy of combustion is related to a particular fuel.

15-42C The heating value is called the higher heating value when the H_2O in the products is in the liquid form, and it is called the lower heating value when the H_2O in the products is in the vapor form. The heating value of a fuel is equal to the absolute value of the enthalpy of combustion of that fuel.

15-43C If the combustion of a fuel results in a single compound, the enthalpy of formation of that compound is identical to the enthalpy of combustion of that fuel.

15-44C Yes.

15-45C No. The enthalpy of formation of N_2 is simply assigned a value of zero at the standard reference state for convenience.

15-46C 1 kmol of H_2. This is evident from the observation that when chemical bonds of H_2 are destroyed to form H_2O a large amount of energy is released.

15-47 The enthalpy of combustion of methane at a 25°C and 1 atm is to be determined using the data from Table A-26 and to be compared to the value listed in Table A-27.

Assumptions The water in the products is in the liquid phase.

Analysis The stoichiometric equation for this reaction is

$$\mathrm{CH_4} + 2[\mathrm{O_2} + 3.76\mathrm{N_2}] \longrightarrow \mathrm{CO_2} + 2\mathrm{H_2O}(\ell) + 7.52\mathrm{N_2}$$

Both the reactants and the products are at the standard reference state of 25°C and 1 atm. Also, N_2 and O_2 are stable elements, and thus their enthalpy of formation is zero. Then the enthalpy of combustion of CH_4 becomes

$$h_C = H_P - H_R = \sum N_P \bar{h}^\circ_{f,P} - \sum N_R \bar{h}^\circ_{f,R} = \left(N\bar{h}^\circ_f\right)_{\mathrm{CO_2}} + \left(N\bar{h}^\circ_f\right)_{\mathrm{H_2O}} - \left(N\bar{h}^\circ_f\right)_{\mathrm{CH_4}}$$

Using $\bar{h}^\circ_f$ values from Table A-26,

$$\begin{aligned} h_C &= (1\text{ kmol})(-393{,}520\text{ kJ/kmol}) + (2\text{ kmol})(-285{,}830\text{ kJ/kmol}) \\ &\quad -(1\text{ kmol})(-74{,}850\text{ kJ/kmol}) \\ &= \mathbf{-890{,}330\text{ kJ}}\ (\text{per kmol CH}_4) \end{aligned}$$

The listed value in Table A-27 is -890,868 kJ/kmol, which is almost identical to the calculated value. Since the water in the products is assumed to be in the liquid phase, this h_c value corresponds to the higher heating value of CH_4.

15-48 Problem 15-47 is reconsidered. The effect of temperature on the enthalpy of combustion is to be studied.

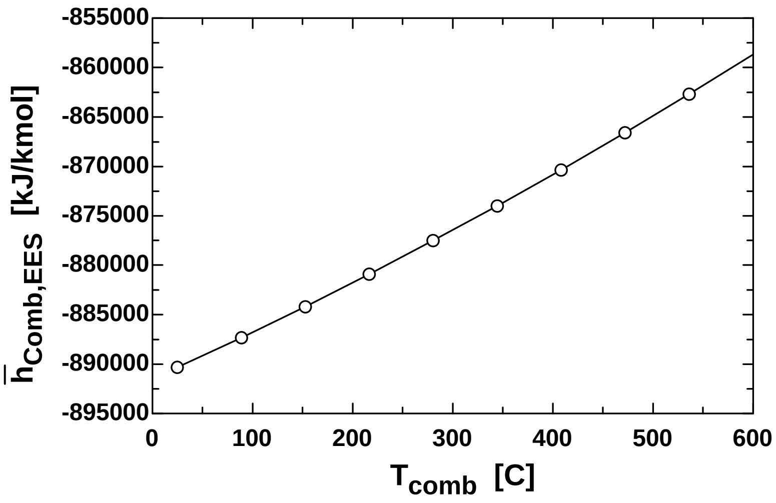
Analysis The problem is solved using EES, and the solution is given below.

```
Fuel$ = 'Methane (CH4)'
T_comb =25 [C]
T_fuel = T_comb +273 "[K]"
T_air1 = T_comb +273 "[K]"
T_prod =T_comb +273 "[K]"
h_bar_comb_TableA27 = -890360 [kJ/kmol]
"For  theoretical dry air, the complete combustion equation is"
"CH4 + A_th(O2+3.76 N2)=1 CO2+2 H2O +  A_th (3.76) N2 "

A_th*2=1*2+2*1 "theoretical O balance"

"Apply First Law SSSF"
h_fuel_EES=enthalpy(CH4,T=298) "[kJ/kmol]"
h_fuel_TableA26=-74850 "[kJ/kmol]"
h_bar_fg_H2O=enthalpy(Steam_iapws,T=298,x=1)-enthalpy(Steam_iapws,T=298,x=0) "[kJ/kmol]"
HR=h_fuel_EES+ A_th*enthalpy(O2,T=T_air1)+A_th*3.76 *enthalpy(N2,T=T_air1) "[kJ/kmol]"
HP=1*enthalpy(CO2,T=T_prod)+2*(enthalpy(H2O,T=T_prod)-h_bar_fg_H2O)+A_th*3.76*
enthalpy(N2,T=T_prod) "[kJ/kmol]"
h_bar_Comb_EES=(HP-HR) "[kJ/kmol]"
PercentError=ABS(h_bar_Comb_EES-h_bar_comb_TableA27)/ABS(h_bar_comb_TableA27)*Convert(, %) "[%]"
```

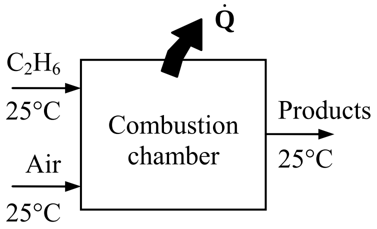
$h_{CombEES}$ [kJ/kmol]	T_{Comb} [C]
-890335	25
-887336	88.89
-884186	152.8
-880908	216.7
-877508	280.6
-873985	344.4
-870339	408.3
-866568	472.2
-862675	536.1
-858661	600



15-49 Ethane is burned with stoichiometric amount of air. The heat transfer is to be determined if both the reactants and products are at 25°C.

Assumptions The water in the products is in the vapor phase.

Analysis The stoichiometric equation for this reaction is



$$C_2H_6 + 3.5[O_2 + 3.76N_2] \longrightarrow 2CO_2 + 3H_2O + 13.16N_2$$

Since both the reactants and the products are at the standard reference state of 25°C and 1 atm, 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,

$$Q = h_C = H_P - H_R = \sum N_P \bar{h}^\circ_{f,P} - \sum N_R \bar{h}^\circ_{f,R} = \left(N\bar{h}^\circ_f\right)_{CO2} + \left(N\bar{h}^\circ_f\right)_{H2O} - \left(N\bar{h}^\circ_f\right)_{C2H6}$$

Using $\bar{h}^\circ_f$ values from Table A-26,

$$Q = h_C = (2\text{ kmol})(-393{,}520\text{ kJ/kmol}) + (3\text{ kmol})(-241{,}820\text{ kJ/kmol}) - (1\text{ kmol})(-84{,}680\text{ kJ/kmol})$$
$$= \mathbf{-1{,}427{,}820\ kJ/kmol\ C2H6}$$

15-50 Ethane is burned with stoichiometric amount of air at 1 atm and 25°C. The minimum pressure of the products which will assure that the water in the products will be in vapor form is to be determined.

Assumptions The water in the products is in the vapor phase.

Analysis The stoichiometric equation for this reaction is

$$C_2H_6 + 3.5[O_2 + 3.76N_2] \longrightarrow 2CO_2 + 3H_2O + 13.16N_2$$

At the minimum pressure, the product mixture will be saturated with water vapor and

$$P_v = P_{\text{sat@25°C}} = 3.1698\text{ kPa}$$

The mole fraction of water in the products is

$$y_v = \frac{N_{H2O}}{N_{\text{prod}}} = \frac{3\text{ kmol}}{(2+3+13.16)\text{ kmol}} = 0.1652$$

The minimum pressure of the products is then

$$P_{\min} = \frac{P_v}{y_v} = \frac{3.1698\text{ kPa}}{0.1652} = \mathbf{19.2\ kPa}$$

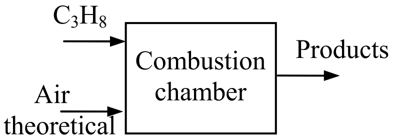
15-51 The higher and lower heating values of liquid propane are to be determined and compared to the listed values.

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

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



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,

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

The $\bar{h}^\circ_f$ of liquid propane is obtained by adding $\bar{h}_{fg}$ of propane at 25°C to $\bar{h}^\circ_f$ 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

$$\mathrm{HHV} = \frac{-h_C}{M_m} = \frac{2{,}205{,}260\text{ kJ/kmol C}_3\text{H}_8}{44.097\text{ kg/kmol C}_3\text{H}_8} = \mathbf{50{,}010\ kJ/kg\ C_3H_8}$$

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

$$\mathrm{LHV} = \frac{-h_C}{M_m} = \frac{2{,}029{,}220\text{ kJ/kmol C}_3\text{H}_8}{44.097\text{ kg/kmol C}_3\text{H}_8} = \mathbf{46{,}020\ kJ/kg\ C_3H_8}$$

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

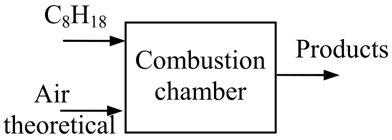
15-52 The higher and lower heating values of gaseous octane are to be determined and compared to the listed values.

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

$$C_8H_{18} + 12.5(O_2 + 3.76N_2) \longrightarrow 8CO_2 + 9H_2O + 47N_2$$



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,

$$q = h_C = H_P - H_R = \sum N_P \bar{h}^\circ_{f,P} - \sum N_R \bar{h}^\circ_{f,R} = \left(N\bar{h}^\circ_f\right)_{\text{CO2}} + \left(N\bar{h}^\circ_f\right)_{\text{H2O}} - \left(N\bar{h}^\circ_f\right)_{\text{C8H18}}$$

For the HHV, the water in the products is taken to be liquid. Then,

$$\begin{aligned} h_C &= (8\text{ kmol})(-393{,}520\text{ kJ/kmol}) + (9\text{ kmol})(-285{,}830\text{ kJ/kmol}) - (1\text{ kmol})(-208{,}450\text{ kJ/kmol}) \\ &= -5{,}512{,}180\text{ kJ/kmol octane} \end{aligned}$$

The HHV of the gaseous octane is

$$\text{HHV} = \frac{-h_C}{M_m} = \frac{5{,}512{,}180\text{ kJ/kmol }C_8H_{18}}{114.231\text{ kg/kmol }C_8H_{18}} = \mathbf{48{,}250\text{ kJ/kg }C_8H_{18}}$$

The listed value for liquid octane from Table A-27 is 47,890 kJ/kg. Adding the enthalpy of vaporization of octane to this value (47,890+363=48,253), the higher heating value of gaseous octane becomes **48,253 kJ/kg** octane. This value is practically identical to the calculated value. For the LHV, the water in the products is taken to be vapor. Then,

$$\begin{aligned} h_C &= (8\text{ kmol})(-393{,}520\text{ kJ/kmol}) + (9\text{ kmol})(-241{,}820\text{ kJ/kmol}) - (1\text{ kmol})(-208{,}450\text{ kJ/kmol}) \\ &= -5{,}116{,}090\text{ kJ/kmol octane} \end{aligned}$$

The LHV of the gaseous octane is then

$$\text{HHV} = \frac{-h_C}{M_m} = \frac{5{,}116{,}090\text{ kJ/kmol }C_8H_{18}}{114.231\text{ kg/kmol }C_8H_{18}} = \mathbf{44{,}790\text{ kJ/kg }C_8H_{18}}$$

The listed value for liquid octane from Table A-27 is 44,430 kJ/kg. Adding the enthalpy of vaporization of octane to this value (44,430+363=44,793), the lower heating value of gaseous octane becomes **44,793 kJ/kg** octane. This value is practically identical to the calculated value.

15-53 The higher and lower heating values of coal from Illinois are to be determined.

Assumptions **1** Combustion is complete. **2** The combustion products contain CO_2, CO, H_2O, SO_2, and N_2. **3** Combustion gases are ideal gases.

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

Analysis We consider 100 kg of coal for simplicity. Noting that the mass percentages in this case correspond to the masses of the constituents, the mole numbers of the constituent of the coal are determined to be

$$N_{\text{C}} = \frac{m_{\text{C}}}{M_{\text{C}}} = \frac{67.40 \text{ kg}}{12 \text{ kg/kmol}} = 5.617 \text{ kmol}$$

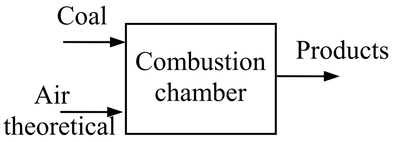
$$N_{\text{H2}} = \frac{m_{\text{H2}}}{M_{\text{H2}}} = \frac{5.31 \text{ kg}}{2 \text{ kg/kmol}} = 2.655 \text{ kmol}$$

$$N_{\text{O2}} = \frac{m_{\text{O2}}}{M_{\text{O2}}} = \frac{15.11 \text{ kg}}{32 \text{ kg/kmol}} = 0.4722 \text{ kmol}$$

$$N_{\text{N2}} = \frac{m_{\text{N2}}}{M_{\text{N2}}} = \frac{1.44 \text{ kg}}{28 \text{ kg/kmol}} = 0.05143 \text{ kmol}$$

$$N_{\text{S}} = \frac{m_{\text{S}}}{M_{\text{S}}} = \frac{2.36 \text{ kg}}{32 \text{ kg/kmol}} = 0.07375 \text{ kmol}$$

67.40% C
5.31% H_2
15.11% O_2
1.44% N_2
2.36% S
8.38% ash
(by mass)



The mole number of the mixture and the mole fractions are

$$N_m = 5.617 + 2.655 + 0.4722 + 0.05143 + 0.07375 = 8.869 \text{ kmol}$$

$$y_{\text{C}} = \frac{N_{\text{C}}}{N_m} = \frac{5.617 \text{ kmol}}{8.869 \text{ kmol}} = 0.6333$$

$$y_{\text{H2}} = \frac{N_{\text{H2}}}{N_m} = \frac{2.655 \text{ kmol}}{8.869 \text{ kmol}} = 0.2994$$

$$y_{\text{O2}} = \frac{N_{\text{O2}}}{N_m} = \frac{0.4722 \text{ kmol}}{8.869 \text{ kmol}} = 0.05323$$

$$y_{\text{N2}} = \frac{N_{\text{N2}}}{N_m} = \frac{0.05143 \text{ kmol}}{8.869 \text{ kmol}} = 0.00580$$

$$y_{\text{S}} = \frac{N_{\text{S}}}{N_m} = \frac{0.07375 \text{ kmol}}{8.869 \text{ kmol}} = 0.00832$$

Ash consists of the non-combustible matter in coal. Therefore, the mass of ash content that enters the combustion chamber is equal to the mass content that leaves. Disregarding this non-reacting component for simplicity, the combustion equation may be written as

$$0.6333\text{C} + 0.2994\text{H}_2 + 0.05323\text{O}_2 + 0.00580\text{N}_2 + 0.00832\text{S} + a_{\text{th}}(\text{O}_2 + 3.76\text{N}_2)$$
$$\longrightarrow x\text{CO}_2 + y\text{H}_2\text{O} + z\text{SO}_2 + k\text{N}_2$$

According to the species balances,

$$\text{C balance: } x = 0.6333$$
$$\text{H}_2 \text{ balance: } y = 0.2994$$
$$\text{S balance: } z = 0.00832$$
$$\text{O}_2 \text{ balance:}$$
$$0.05323 + a_{\text{th}} = x + 0.5y + z$$
$$a_{\text{th}} = 0.6333 + 0.5 \times 0.2994 + 0.00832 - 0.05323 = 0.7381$$
$$\text{N}_2 \text{ balance: } k = 0.00580 + 3.76a_{\text{th}} = 0.00580 + 3.76 \times 0.7381 = 2.781$$

Substituting,

$$0.6333\mathrm{C}+0.2994\mathrm{H}_2+0.05323\mathrm{O}_2+0.00580\mathrm{N}_2+0.00832\mathrm{S}+0.7381(\mathrm{O}_2+3.76\mathrm{N}_2)$$
$$\longrightarrow 0.6333\mathrm{CO}_2+0.2994\mathrm{H}_2\mathrm{O}+0.00832\mathrm{SO}_2+2.781\mathrm{N}_2$$

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 C, S, H_2, N_2 and O_2 are stable elements, and thus their enthalpy of formation is zero. Then,

$$q=h_C=H_P-H_R=\sum N_P\bar{h}^\circ_{f,P}-\sum N_R\bar{h}^\circ_{f,R}=\left(N\bar{h}^\circ_f\right)_{\mathrm{CO2}}+\left(N\bar{h}^\circ_f\right)_{\mathrm{H2O}}+\left(N\bar{h}^\circ_f\right)_{\mathrm{SO2}}$$

For the HHV, the water in the products is taken to be liquid. Then,

$$\begin{aligned}h_C&=(0.6333\text{ kmol})(-393{,}520\text{ kJ/kmol})+(0.2994\text{ kmol})(-285{,}830\text{ kJ/kmol})\\&\quad+(0.00832\text{ kmol})(-297{,}100\text{ kJ/kmol})\\&=-337{,}270\text{ kJ/kmol coal}\end{aligned}$$

The apparent molecular weight of the coal is

$$\begin{aligned}M_m=\frac{m_m}{N_m}&=\frac{(0.6333\times 12+0.2994\times 2+0.05323\times 32+0.00580\times 28+0.00832\times 32)\text{ kg}}{(0.6333+0.2994+0.05323+0.00580+0.00832)\text{ kmol}}\\&=\frac{10.33\text{ kg}}{1.000\text{ kmol}}=10.33\text{ kg/kmol coal}\end{aligned}$$

The HHV of the coal is then

$$\mathrm{HHV}=\frac{-h_C}{M_m}=\frac{337{,}270\text{ kJ/kmol coal}}{10.33\text{ kg/kmol coal}}=\mathbf{32{,}650\ kJ/kg\ coal}$$

For the LHV, the water in the products is taken to be vapor. Then,

$$\begin{aligned}h_C&=(0.6333\text{ kmol})(-393{,}520\text{ kJ/kmol})+(0.2994\text{ kmol})(-241{,}820\text{ kJ/kmol})\\&\quad+(0.00832\text{ kmol})(-297{,}100\text{ kJ/kmol})\\&=-324{,}090\text{ kJ/kmol coal}\end{aligned}$$

The LHV of the coal is then

$$\mathrm{LHV}=\frac{-h_C}{M_m}=\frac{324{,}090\text{ kJ/kmol coal}}{10.33\text{ kg/kmol coal}}=\mathbf{31{,}370\ kJ/kg\ coal}$$

First Law Analysis of Reacting Systems

15-54C In this case $\Delta U + W_b = \Delta H$, and the conservation of energy relation reduces to the form of the steady-flow energy relation.

15-55C The heat transfer will be the same for all cases. The excess oxygen and nitrogen enters and leaves the combustion chamber at the same state, and thus has no effect on the energy balance.

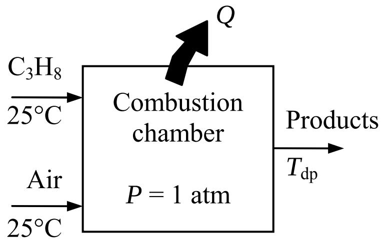
15-56C For case (*b*), which contains the maximum amount of nonreacting gases. This is because part of the chemical energy released in the combustion chamber is absorbed and transported out by the nonreacting gases.

15-57 Propane is burned with an air-fuel ratio of 25. The heat transfer per kilogram of fuel burned when the temperature of the products is such that liquid water just begins to form in the products is to be determined.

Assumptions **1** Steady operating conditions exist. **2** Air and combustion gases are ideal gases. **3** Kinetic and potential energies are negligible. **4** Combustion is complete. **5** The reactants are at 25°C and 1 atm. **6** The fuel is in vapor phase.

Properties The molar masses of propane and air are 44 kg/kmol and 29 kg/kmol, respectively (Table A-1).

Analysis The mass of air per kmol of fuel is



$$m_{\text{air}} = (\text{AF})m_{\text{fuel}}$$
$$= (25\text{ kg air/kg fuel})(1\times 44\text{ kg/kmol fuel}) = 1100\text{ kg air/kmol fuel}$$

The mole number of air per kmol of fuel is then

$$N_{\text{air}} = \frac{m_{\text{air}}}{M_{\text{air}}} = \frac{1100\text{ kg air/kmol fuel}}{29\text{ kg air/kmol air}} = 37.93\text{ kmol air/kmol fuel}$$

The combustion equation can be written as

$$C_3H_8 + (37.93/4.76)(O_2 + 3.76N_2) \longrightarrow 3CO_2 + 4H_2O + xO_2 + (37.93/4.76)\times 3.76N_2$$

The coefficient for O_2 is obtained from O_2 balance:

$$(37.93/4.76) = 3 + 2 + x \longrightarrow x = 2.968$$

Substituting, $$C_3H_8 + 7.968(O_2 + 3.76N_2) \longrightarrow 3CO_2 + 4H_2O + 2.968O_2 + 29.96N_2$$

The mole fraction of water in the products is

$$y_v = \frac{N_{\text{H2O}}}{N_{\text{prod}}} = \frac{4\text{ kmol}}{(3+4+2.968+29.96)\text{ kmol}} = \frac{4\text{ kmol}}{39.93\text{ kmol}} = 0.1002$$

The partial pressure of water vapor at 1 atm total pressure is

$$P_v = y_v P = (0.1002)(101.325\text{ kPa}) = 10.15\text{ kPa}$$

When this mixture is at the dew-point temperature, the water vapor pressure is the same as the saturation pressure. Then,

$$T_{\text{dp}} = T_{\text{sat@10.15 kPa}} = 46.1°\text{C} = 319.1\text{ K} \cong 320\text{ K}$$

We obtain properties at 320 K (instead of 319.1 K) to avoid iterations in the ideal gas tables. The heat transfer for this combustion process is determined from the energy balance $E_{\text{in}} - E_{\text{out}} = \Delta E_{\text{system}}$ applied on the combustion chamber with $W = 0$. It reduces to

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

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	$\bar{h}_{298\text{K}}$, kJ/kmol	$\bar{h}_{320\text{K}}$, kJ/kmol
C_3H_8	-103,850	---	---
O_2	0	8682	9325
N_2	0	8669	9306
H_2O (*g*)	-241,820	9904	10,639
CO_2	-393,520	9364	10,186

Substituting,

$$\begin{aligned}-\bar{Q}_{\text{out}} &= (3)(-393{,}520 + 10{,}186 - 9364) + (4)(-241{,}820 + 10{,}639 - 9904) + (2.968)(0 + 9325 - 8682)\\ &\quad + (29.96)(0 + 9306 - 8669) - (1)(-103{,}850) - 0\\ &= -2{,}017{,}590\text{ kJ/kmol } C_3H_8\end{aligned}$$

or $$\bar{Q}_{\text{out}} = 2{,}017{,}590\text{ kJ/kmol } C_3H_8$$

Then the heat transfer per kg of fuel is

$$Q_{\text{out}} = \frac{\bar{Q}_{\text{out}}}{M_{\text{fuel}}} = \frac{2{,}017{,}590\text{ kJ/kmol fuel}}{44\text{ kg/kmol}} = \mathbf{45{,}850\text{ kJ/kg } C_3H_8}$$

15-58 n-Octane is burned with 100 percent excess air. The heat transfer per kilogram of fuel burned for a product temperature of 257°C is to be determined.

Assumptions **1** Steady operating conditions exist. **2** Air and combustion gases are ideal gases. **3** Kinetic and potential energies are negligible. **4** Combustion is complete. **5** The fuel is in vapor phase.

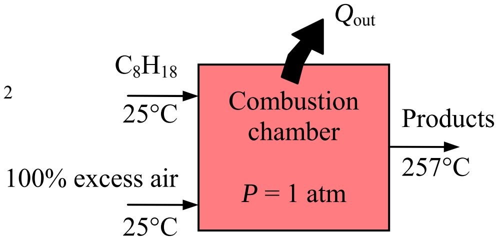
Properties The molar masses of propane and air are 44 kg/kmol and 29 kg/kmol, respectively (Table A-1).

Analysis The combustion reaction for stoichiometric air is

$$C_8H_{18} + 12.5[O_2 + 3.76N_2] \longrightarrow 8CO_2 + 9H_2O + (12.5\times 3.76)N_2$$

The combustion equation with 100% excess air is

$$C_8H_{18} + 25[O_2 + 3.76N_2] \longrightarrow 8CO_2 + 9H_2O + 12.5\,O_2 + 94N_2$$



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

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

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

Substance	$\overline{h}_f^\circ$ kJ/kmol	$\overline{h}_{298K}$ kJ/kmol	$\overline{h}_{530K}$ kJ/kmol
C_8H_{18} (*g*)	-208,450	---	---
O_2	0	8682	15,708
N_2	0	8669	15,469
H_2O (*g*)	-241,820	9904	17,889
CO_2	-393,520	9364	19,029

Substituting,

$$\begin{aligned}-\overline{Q}_{out} &= (8)(-393{,}520 + 19{,}029 - 9364) + (9)(-241{,}820 + 17{,}889 - 9904) + (12.5)(0 + 15{,}708 - 8682) \\ &\quad + (94)(0 + 15{,}469 - 8669) - (1)(-208{,}450) - 0 - 0 \\ &= -4{,}239{,}880 \text{ kJ/kmol } C_8H_{18}\end{aligned}$$

or $\overline{Q}_{out} = 4{,}239{,}880 \text{ kJ/kmol } C_8H_{18}$

Then the heat transfer per kg of fuel is

$$Q_{out} = \frac{\overline{Q}_{out}}{M_{fuel}} = \frac{4{,}239{,}880 \text{ kJ/kmol fuel}}{114 \text{ kg/kmol}} = \mathbf{37{,}200\ kJ/kg\ C_8H_{18}}$$

15-59 Propane is burned with 50 percent excess air during a steady-flow combustion process. The rate of heat transfer in the combustion chamber is to be determined.

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

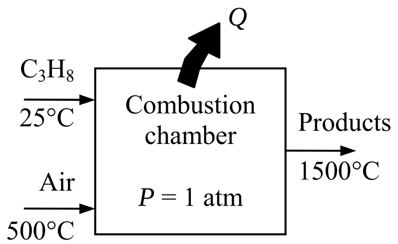
Properties The molar masses of propane and air are 44 kg/kmol and 29 kg/kmol, respectively (Table A-1).

Analysis The combustion equation can be written as

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

The stoichiometric coefficient is obtained from O_2 balance:

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



Substituting,

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

The specific volume of the air entering the system is

$$v_{air} = \frac{R_{air}T}{P} = \frac{(0.287\ \text{kJ/kmol}\cdot\text{K})(773\ \text{K})}{101.3\ \text{kPa}} = 2.190\ \text{m}^3/\text{kg}$$

and the mass flow rate of this air is

$$\dot{m}_{air} = \frac{\dot{V}_{air}}{v_{air}} = \frac{1\ \text{m}^3/\text{s}}{2.190\ \text{m}^3/\text{kg}} = 0.4566\ \text{kg/s}$$

The air-fuel ratio for this combustion process is

$$\text{AF} = \frac{m_{air}}{m_{fuel}} = \frac{(7.5 \times 4.76\ \text{kmol})(29\ \text{kg/kmol})}{(3\ \text{kmol})(12\ \text{kg/kmol}) + (4\ \text{kmol})(2\ \text{kg/kmol})} = 23.53\ \text{kg air/kg fuel}$$

The mass flow rate of fuel is

$$\dot{m}_{fuel} = \frac{\dot{m}_{air}}{\text{AF}} = \frac{0.4566\ \text{kg/s}}{23.53} = 0.01941\ \text{kg/s}$$

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

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

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

Substance	$\bar{\mathbf{h}}_\mathbf{f}^\circ$ **kJ/kmol**	$\bar{\mathbf{h}}_{\mathbf{298\ K}}$ **kJ/kmol**	$\bar{\mathbf{h}}_{\mathbf{773\ K}}$ **kJ/kmol**	$\bar{\mathbf{h}}_{\mathbf{1773\ K}}$ **kJ/kmol**
C_3H_8	-103,850	---	---	---
O_2	0	8682	23,614	59,364
N_2	0	8669	22,866	56,689
H_2O (*g*)	-241,820	9904	---	71,177
CO_2	-393,520	9364	---	87,195

Substituting,

$$\begin{aligned} -Q_{out} &= (3)(-393{,}520 + 87{,}195 - 9364) + (4)(-241{,}820 + 71{,}177 - 9904) + (2.5)(0 + 59{,}364 - 8682) \\ &\quad + (28.2)(0 + 56{,}689 - 8669) - (1)(-103{,}850 + h_{298} - h_{298}) - (7.5)(0 + 23{,}614 - 8682) \\ &\quad - (28.2)(0 + 22{,}866 - 8669) \\ &= -596{,}881\ \text{kJ/kmol}\ C_3H_8 \end{aligned}$$

or $Q_{out} = 596{,}881\ \text{kJ/kmol}\ C_3H_8$

Then the rate of heat transfer for a mass flow rate of 0.01941 kg/s for the propane becomes

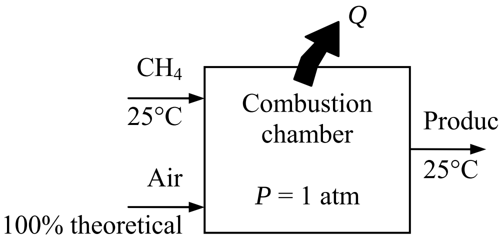
$$\dot{Q}_{out} = \dot{N}Q_{out} = \left(\frac{\dot{m}}{M}\right)Q_{out} = \left(\frac{0.01941\ \text{kg/s}}{44\ \text{kg/kmol}}\right)(596{,}881\ \text{kJ/kmol}) = \mathbf{263.3\ kW}$$

15-60 Methane is burned completely during a steady-flow combustion process. The heat transfer from the combustion chamber is to be determined for two cases.

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

Analysis The fuel is burned completely with the stoichiometric amount of air, and thus the products will contain only H_2O, CO_2 and N_2, but no free O_2. Considering 1 kmol of fuel, the theoretical combustion equation can be written as

$$\mathrm{CH_4} + a_{\mathrm{th}}(\mathrm{O_2} + 3.76\mathrm{N_2}) \longrightarrow \mathrm{CO_2} + 2\mathrm{H_2O} + 3.76a_{\mathrm{th}}\mathrm{N_2}$$



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

$$a_{\mathrm{th}} = 1 + 1 = 2$$

Substituting,

$$\mathrm{CH_4} + 2(\mathrm{O_2} + 3.76\mathrm{N_2}) \longrightarrow \mathrm{CO_2} + 2\mathrm{H_2O} + 5.64\mathrm{N_2}$$

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

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

since both the reactants and the products are at 25°C and both the air and the combustion gases can be treated as ideal gases. From the tables,

Substance	$\bar{h}_f^\circ$ **kJ/kmol**
CH_4	-74,850
O_2	0
N_2	0
H_2O (ℓ)	-285,830
CO_2	-393,520

Thus,

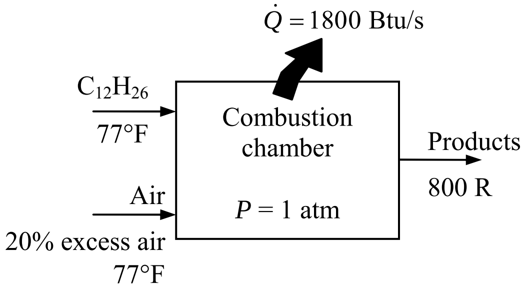
$$-Q_{\mathrm{out}} = (1)(-393{,}520) + (2)(-285{,}830) + 0 - (1)(-74{,}850) - 0 - 0 = -890{,}330 \ \mathrm{kJ/kmol\ CH_4}$$

or

$$Q_{\mathrm{out}} = \mathbf{890{,}330\ \ kJ/kmol\ CH_4}$$

If combustion is achieved with 100% excess air, the answer would still be the same since it would enter and leave at 25°C, and absorb no energy.

15-61E Diesel fuel is burned with 20 percent excess air during a steady-flow combustion process. The required mass flow rate of the diesel fuel for a specified heat transfer rate is to be determined.



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

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 of $C_{12}H_{26}$, the combustion equation can be written as

$$C_{12}H_{26} + 1.2a_{\text{th}}(O_2 + 3.76N_2) \longrightarrow 12CO_2 + 13H_2O + 0.2a_{\text{th}}O_2 + (1.2)(3.76a_{\text{th}})N_2$$

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

$$1.2a_{\text{th}} = 12 + 6.5 + 0.2a_{\text{th}} \qquad \longrightarrow \qquad a_{\text{th}} = 18.5$$

Substituting,

$$C_{12}H_{26} + 22.2(O_2 + 3.76N_2) \longrightarrow 12CO_2 + 13H_2O + 3.7O_2 + 83.47N_2$$

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

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

since all of the reactants are at 77°F. Assuming the air and the combustion products to be ideal gases, we have $h = h(T)$. From the tables,

Substance	$\bar{h}_f^\circ$ **Btu/lbmol**	$\bar{h}_{537\text{ R}}$ **Btu/lbmol**	$\bar{h}_{800\text{ R}}$ **Btu/lbmol**
$C_{12}H_{26}$	-125,190	---	---
O_2	0	3725.1	5602.0
N_2	0	3729.5	5564.4
H_2O (*g*)	-104,040	4258.0	6396.9
CO_2	-169,300	4027.5	6552.9

Thus,

$$\begin{aligned}-Q_{\text{out}} &= (12)(-169{,}300 + 6552.9 - 4027.5) + (13)(-104{,}040 + 6396.9 - 4258)\\ &\quad + (3.7)(0 + 5602.0 - 3725.1) + (83.47)(0 + 5564.4 - 3729.5) - (1)(-125{,}190) - 0 - 0\\ &= -3{,}040{,}716\text{ Btu/lbmol } C_{12}H_{26}\end{aligned}$$

or $\qquad Q_{\text{out}} = 3{,}040{,}716\text{ Btu/lbmol } C_{12}H_{26}$

Then the required mass flow rate of fuel for a heat transfer rate of 1800 Btu/s becomes

$$\dot{m} = \dot{N}M = \left(\frac{\dot{Q}}{Q}\right)M = \left(\frac{1800\text{ Btu/s}}{3{,}040{,}716\text{ Btu/lbmol}}\right)(170\text{ lbm/lbmol}) = \mathbf{0.1006\ \ lbm/s}$$

15-62 A certain coal is burned steadily with 40% excess air. The heat transfer for a given product temperature is to be determined.

Assumptions **1** Combustion is complete. **2** The combustion products contain CO_2, CO, H_2O, SO_2, and N_2. **3** Combustion gases are ideal gases.

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

Analysis We consider 100 kg of coal for simplicity. Noting that the mass percentages in this case correspond to the masses of the constituents, the mole numbers of the constituent of the coal are determined to be

39.25% C
6.93% H_2
41.11% O_2
0.72% N_2
0.79% S
11.20% ash
(by mass)

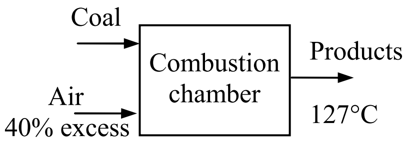
$$N_C = \frac{m_C}{M_C} = \frac{39.25\text{ kg}}{12\text{ kg/kmol}} = 3.271\text{ kmol}$$

$$N_{H2} = \frac{m_{H2}}{M_{H2}} = \frac{6.93\text{ kg}}{2\text{ kg/kmol}} = 3.465\text{ kmol}$$

$$N_{O2} = \frac{m_{O2}}{M_{O2}} = \frac{41.11\text{ kg}}{32\text{ kg/kmol}} = 1.285\text{ kmol}$$

$$N_{N2} = \frac{m_{N2}}{M_{N2}} = \frac{0.72\text{ kg}}{28\text{ kg/kmol}} = 0.0257\text{ kmol}$$

$$N_S = \frac{m_S}{M_S} = \frac{0.79\text{ kg}}{32\text{ kg/kmol}} = 0.0247\text{ kmol}$$



The mole number of the mixture and the mole fractions are

$$N_m = 3.271 + 3.465 + 1.285 + 0.0257 + 0.0247 = 8.071\text{ kmol}$$

$$y_C = \frac{N_C}{N_m} = \frac{3.271\text{ kmol}}{8.071\text{ kmol}} = 0.4052$$

$$y_{H2} = \frac{N_{H2}}{N_m} = \frac{3.465\text{ kmol}}{8.071\text{ kmol}} = 0.4293$$

$$y_{O2} = \frac{N_{O2}}{N_m} = \frac{1.285\text{ kmol}}{8.071\text{ kmol}} = 0.1592$$

$$y_{N2} = \frac{N_{N2}}{N_m} = \frac{0.0257\text{ kmol}}{8.071\text{ kmol}} = 0.00319$$

$$y_S = \frac{N_S}{N_m} = \frac{0.0247\text{ kmol}}{8.071\text{ kmol}} = 0.00306$$

Ash consists of the non-combustible matter in coal. Therefore, the mass of ash content that enters the combustion chamber is equal to the mass content that leaves. Disregarding this non-reacting component for simplicity, the combustion equation may be written as

$$0.4052\text{C} + 0.4293\text{H}_2 + 0.1592\text{O}_2 + 0.00319\text{N}_2 + 0.00306\text{S} + 1.4a_{\text{th}}(\text{O}_2 + 3.76\text{N}_2)$$
$$\longrightarrow 0.4052\text{CO}_2 + 0.4293\text{H}_2\text{O} + 0.4a_{\text{th}}\text{O}_2 + 0.00306\text{SO}_2 + 1.4a_{\text{th}} \times 3.76\text{N}_2$$

According to the O_2 mass balance,

$$0.1592 + 1.4a_{\text{th}} = 0.4052 + 0.5 \times 0.4293 + 0.4a_{\text{th}} + 0.00306 \longrightarrow a_{\text{th}} = 0.4637$$

Substituting,

$$0.4052\text{C} + 0.4293\text{H}_2 + 0.1592\text{O}_2 + 0.00319\text{N}_2 + 0.00306\text{S} + 0.6492(\text{O}_2 + 3.76\text{N}_2)$$
$$\longrightarrow 0.4052\text{CO}_2 + 0.4293\text{H}_2\text{O} + 0.1855\text{O}_2 + 0.00306\text{SO}_2 + 2.441\text{N}_2$$

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

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

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	$\bar{h}_{298\text{ K}}$ kJ/kmol	$\bar{h}_{400\text{ K}}$ kJ/kmol
O_2	0	8682	11,711
N_2	0	8669	11,640
H_2O (*g*)	-241,820	9904	13,356
CO_2	-393,520	9364	13,372
SO_2	-297,100	-	-

The enthalpy change of sulfur dioxide between the standard temperature and the product temperature using constant specific heat assumption is

$$\Delta\bar{h}_{\text{SO2}} = c_p \Delta T = (41.7\text{ kJ/kmol}\cdot\text{K})(127-25)\text{K} = 4253\text{ kJ/kmol}$$

Substituting into the energy balance relation,

$$\begin{aligned}-\bar{Q}_{\text{out}} &= (0.4052)(-393{,}520+13{,}372-9364)+(0.4293)(-241{,}820+13{,}356-9904)\\ &\quad+(0.1855)(0+11{,}711-8682)+(2.441)(0+11{,}640-8669)+(0.00306)(-297{,}100+4253)-0\\ &= -253{,}244\text{ kJ/kmol C}_8\text{H}_{18}\end{aligned}$$

or $$\bar{Q}_{\text{out}} = 253{,}244\text{ kJ/kmol fuel}$$

Then the heat transfer per kg of fuel is

$$\begin{aligned}Q_{\text{out}} &= \frac{\bar{Q}_{\text{out}}}{M_{\text{fuel}}} = \frac{253{,}244\text{ kJ/kmol fuel}}{(0.4052\times 12+0.4293\times 2+0.1592\times 32+0.00319\times 28+0.00306\times 32)\text{ kg/kmol}}\\ &= \frac{253{,}244\text{ kJ/kmol fuel}}{11.00\text{ kg/kmol}}\\ &= \mathbf{23{,}020\text{ kJ/kg coal}}\end{aligned}$$

15-63 Octane gas is burned with 30 percent excess air during a steady-flow combustion process. The heat transfer per unit mass of octane is to be determined.

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

Properties The molar mass of C_8H_{18} is 114 kg/kmol (Table A-1).

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. The moisture in the air does not react with anything; it simply shows up as additional H_2O in the products. Therefore, for simplicity, we will balance the combustion equation using dry air, and then add the moisture to both sides of the equation. Considering 1 kmol of C_8H_{18}, the combustion equation can be written as

$$\mathrm{C_8H_{18}(g)} + 1.8a_{\mathrm{th}}(\mathrm{O_2} + 3.76\mathrm{N_2}) \longrightarrow 8\mathrm{CO_2} + 9\mathrm{H_2O} + 0.8a_{\mathrm{th}}\mathrm{O_2} + (1.8)(3.76a_{\mathrm{th}})\mathrm{N_2}$$

where a_{th} is the stoichiometric coefficient for air. It is determined from

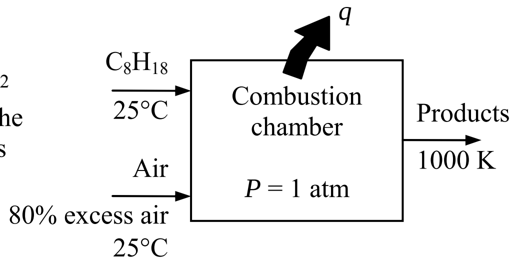
$$\mathrm{O_2}\ \text{balance}: 1.8a_{\mathrm{th}} = 8 + 4.5 + 0.8a_{\mathrm{th}} \longrightarrow a_{\mathrm{th}} = 12.5$$

Thus,

$$\mathrm{C_8H_{18}(g)} + 22.5(\mathrm{O_2} + 3.76\mathrm{N_2}) \longrightarrow 8\mathrm{CO_2} + 9\mathrm{H_2O} + 10\mathrm{O_2} + 84.6\mathrm{N_2}$$

Therefore, 22.5 × 4.76 = 107.1 kmol of dry air will be used per kmol of the fuel. The partial pressure of the water vapor present in the incoming air is

$$P_{v,\mathrm{in}} = \phi_{\mathrm{air}} P_{\mathrm{sat@25^\circ C}} = (0.40)(3.1698\ \mathrm{kPa}) = 1.268\ \mathrm{kPa}$$



Assuming ideal gas behavior, the number of moles of the moisture that accompanies 107.1 kmol of incoming dry air is determined to be

$$N_{v,\mathrm{in}} = \left(\frac{P_{v,\mathrm{in}}}{P_{\mathrm{total}}}\right) N_{\mathrm{total}} = \left(\frac{1.268\ \mathrm{kPa}}{101.325\ \mathrm{kPa}}\right)(107.1 + N_{v,\mathrm{in}}) \longrightarrow N_{v,\mathrm{in}} = 1.36\ \mathrm{kmol}$$

The balanced combustion equation is obtained by adding 1.36 kmol of H_2O to both sides of the equation,

$$\mathrm{C_8H_{18}(g)} + 22.5(\mathrm{O_2} + 3.76\mathrm{N_2}) + 1.36\mathrm{H_2O} \longrightarrow 8\mathrm{CO_2} + 10.36\mathrm{H_2O} + 10\mathrm{O_2} + 84.6\mathrm{N_2}$$

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

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

since all of the reactants are at 25°C. 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	$\bar{h}_{1000\ \mathrm{K}}$, kJ/kmol
C_8H_{18} (*g*)	-208,450	---	---
O_2	0	8682	31,389
N_2	0	8669	30,129
H_2O (*g*)	-241,820	9904	35,882
CO_2	-393,520	9364	42,769

Substituting,

$$\begin{aligned} -Q_{\mathrm{out}} &= (8)(-393{,}520 + 42{,}769 - 9364) + (10.36)(-241{,}820 + 35{,}882 - 9904) \\ &\quad + (10)(0 + 31{,}389 - 8682) + (84.6)(0 + 30{,}129 - 8669) \\ &\quad - (1)(-208{,}450) - (1.36)(-241{,}820) - 0 - 0 \\ &= -2{,}537{,}130\ \mathrm{kJ/kmol\ C_8H_{18}} \end{aligned}$$

Thus 2,537,130 kJ of heat is transferred from the combustion chamber for each kmol (114 kg) of C_8H_{18}. Then the heat transfer per kg of C_8H_{18} becomes

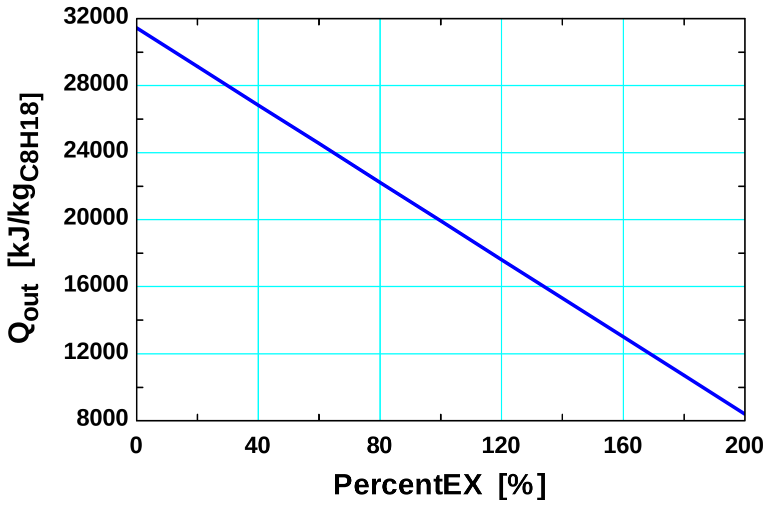
$$q = \frac{Q_{\mathrm{out}}}{M} = \frac{2{,}537{,}130\ \mathrm{kJ}}{114\mathrm{kg}} = \mathbf{22{,}260\ kJ/kg\ C_8H_{18}}$$

15-64 Problem 15-63 is reconsidered. The effect of the amount of excess air on the heat transfer for the combustion process is to be investigated.

Analysis The problem is solved using EES, and the solution is given below.

```
Fuel$ = 'Octane (C8H18)'
T_fuel = (25+273) "[K]"
{PercentEX = 80 "[%]"}
Ex = PercentEX/100 "[%Excess air/100]"
P_air1 = 101.3 [kPa]
T_air1 = 25+273 "[K]"
RH_1 = 40/100  "[%]"
T_prod = 1000 [K]
M_air = 28.97 [kg/kmol]
M_water = 18 [kg/kmol]
M_C8H18=(8*12+18*1) "[kg/kmol]"
"For  theoretical dry air, the complete combustion equation is"
"C8H18 + A_th(O2+3.76 N2)=8 CO2+9 H2O +  A_th (3.76) N2 "
A_th*2=8*2+9*1 "theoretical O balance"
"now to find the amount of water vapor associated with the dry air"
w_1=HUMRAT(AirH2O,T=T_air1,P=P_air1,R=RH_1) "Humidity ratio, kgv/kga"
N_w=w_1*(A_th*4.76*M_air)/M_water "Moles of water in the atmoshperic air, kmol/kmol_fuel"
"The balanced combustion equation with Ex% excess moist air is"
"C8H18 + (1+EX)[A_th(O2+3.76 N2)+N_w H2O]=8 CO2+(9+(1+Ex)*N_w) H2O + (1+Ex) A_th (3.76) N2+ Ex(
A_th) O2 "
"Apply First Law SSSF"
H_fuel = -208450 [kJ/kmol] "from Table A-26"
HR=H_fuel+ (1+Ex)*A_th*enthalpy(O2,T=T_air1)+(1+Ex)*A_th*3.76
*enthalpy(N2,T=T_air1)+(1+Ex)*N_w*enthalpy(H2O,T=T_air1)
HP=8*enthalpy(CO2,T=T_prod)+(9+(1+Ex)*N_w)*enthalpy(H2O,T=T_prod)+(1+Ex)*A_th*3.76*
enthalpy(N2,T=T_prod)+Ex*A_th*enthalpy(O2,T=T_prod)
Q_net=(HP-HR)"kJ/kmol"/(M_C8H18 "kg/kmol")   "[kJ/kg_C8H18]"
Q_out = -Q_net "[kJ/kg_C8H18]"
"This solution used the humidity ratio form psychrometric data to determine the moles of water vapor in
atomspheric air.  One should calculate the moles of water contained in the atmospheric air by the method shown
in Chapter 14 which uses the relative humidity to find the partial pressure of the water vapor and, thus, the moles
of water vapor. Explore what happens to the results as you vary the percent excess air, relative humidity, and
product temperature."
```

PercentEX [%]	Q_{out} [kJ/kgC8H18]
0	31444
20	29139
40	26834
60	24529
80	22224
100	19919
120	17614
140	15309
160	13003
180	10698
200	8393



15-65 Propane gas is burned with 100% excess air. The combustion is incomplete. The balanced chemical reaction is to be written, and the dew-point temperature of the products and the heat transfer from the combustion chamber are to be determined.

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

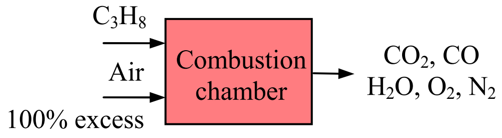
Properties The molar masses of C, H_2, O_2, N_2 and air are 12 kg/kmol, 2 kg/kmol, 32 kg/kmol, 28 kg/kmol, and 29 kg/kmol, respectively (Table A-1).

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

$$\mathrm{C_3H_8} + a_{\mathrm{th}}[\mathrm{O_2} + 3.76\mathrm{N_2}] \longrightarrow 3\,\mathrm{CO_2} + 4\,\mathrm{H_2O} + a_{\mathrm{th}} \times 3.76\,\mathrm{N_2}$$

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

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



Substituting,

$$\mathrm{C_3H_8} + 5[\mathrm{O_2} + 3.76\mathrm{N_2}] \longrightarrow 3\,\mathrm{CO_2} + 4\,\mathrm{H_2O} + 5 \times 3.76\,\mathrm{N_2}$$

The reaction with 100% excess air and incomplete combustion can be written as

$$\mathrm{C_3H_8} + 2 \times 5[\mathrm{O_2} + 3.76\mathrm{N_2}] \longrightarrow 0.90 \times 3\,\mathrm{CO_2} + 0.10 \times 3\,\mathrm{CO} + 4\,\mathrm{H_2O} + x\,\mathrm{O_2} + 2 \times 5 \times 3.76\,\mathrm{N_2}$$

The coefficient for O_2 is determined from a mass balance,

O_2 balance: $\quad 10 = 0.9 \times 3 + 0.05 \times 3 + 2 + x \longrightarrow x = 5.15$

Substituting,

$$\mathrm{C_3H_8} + 10[\mathrm{O_2} + 3.76\mathrm{N_2}] \longrightarrow 2.7\,\mathrm{CO_2} + 0.3\,\mathrm{CO} + 4\,\mathrm{H_2O} + 5.15\,\mathrm{O_2} + 37.6\,\mathrm{N_2}$$

(b) The partial pressure of water vapor is

$$P_v = \frac{N_{\mathrm{H2O}}}{N_{\mathrm{total}}} P_{\mathrm{total}} = \frac{4}{2.7 + 0.3 + 4 + 5.15 + 37.6}(100\text{ kPa}) = \frac{4\text{ kmol}}{49.75\text{ kmol}}(100\text{ kPa}) = 8.040\text{ kPa}$$

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

$$T_{dp} = T_{\mathrm{sat@8.04\,kPa}} = \mathbf{41.5°C} \quad \text{(Table A-5)}$$

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

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

Both the reactants and products are at 25 °C. Assuming the air and the combustion products to be ideal gases, we have $h = h(T)$. Also, since the temperature of products (25 °C) is lower than the dew-po,nt temperature, some water vapor will condense. Noting from Table A-5 that $P_{\mathrm{sat@25°C}} = 3.17\text{ kPa}$, the molar amount of water that remain as vapor is determined from

$$P_v = \frac{N_{\mathrm{H2O}}}{N_{\mathrm{total}}} P_{\mathrm{total}} \longrightarrow 3.17\text{ kPa} = \frac{N_{\mathrm{H2O,vapor}}}{N_{\mathrm{H2O,vapor}} + 2.7 + 0.3 + 5.15 + 37.6}(100\text{ kPa}) \longrightarrow N_{\mathrm{H2O,vapor}} = 1.5\text{ kmol}$$

Thus, $N_{\mathrm{H2O,liquid}} = 4 - 1.5 = 2.5\text{ kmol}$

Then, using the values given in the table,

$$\begin{aligned} -Q_{\mathrm{out}} &= (2.7)(-393{,}520) + (0.3)(-110{,}530) + (1.5)(-241{,}820) + (2.5)(-285{,}830) - (1)(-103{,}850) \\ &= -\,2{,}069{,}120\text{ kJ/kmol C}_3\text{H}_8 \end{aligned}$$

or $\quad Q_{\mathrm{out}} = 2{,}069{,}120\text{ kJ/kmol C}_3\text{H}_8$

Then the heat transfer for a 100 kmol fuel becomes

$$Q_{\mathrm{out}} = NQ_{\mathrm{out}} = \left(\frac{\dot{m}}{N}\right)Q_{\mathrm{out}} = (100\text{ kmol fuel})(2{,}069{,}120\text{ kJ/kmol fuel}) = \mathbf{2.069 \times 10^8\ kJ}$$

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.

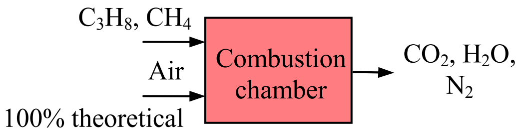
Properties The molar masses of C, H_2, O_2, N_2 and air are 12 kg/kmol, 2 kg/kmol, 32 kg/kmol, 28 kg/kmol, and 29 kg/kmol, respectively (Table A-1).

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:

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

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

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

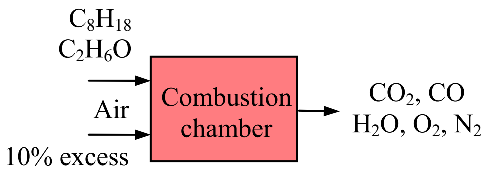
The air-fuel ratio is

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

The mass flow rate of air is then

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

15-67 A mixture of ethanol and octane is burned with 10% excess air. The combustion is incomplete. The balanced chemical reaction is to be written, and the dew-point temperature of the products, the heat transfer for the process, and the relative humidity of atmospheric air for specified conditions are to be determined.



Assumptions **1** Combustion is incomplete. **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

$$0.1\,C_2H_6O + 0.9\,C_8H_{18} + a_{th}[O_2 + 3.76N_2] \longrightarrow 7.4\,CO_2 + 8.4\,H_2O + a_{th} \times 3.76\,N_2$$

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

$$0.1/2 + a_{th} = 7.4 + 8.4/2 \longrightarrow a_{th} = 11.55$$

Substituting,

$$0.1\,C_2H_6O + 0.9\,C_8H_{18} + 11.55[O_2 + 3.76N_2] \longrightarrow 7.4\,CO_2 + 8.4\,H_2O + 11.55 \times 3.76\,N_2$$

The reaction with 10% excess air and incomplete combustion can be written as

$$0.1\,C_2H_6O + 0.9\,C_8H_{18} + 1.1 \times 11.55[O_2 + 3.76N_2]$$
$$\longrightarrow 0.9 \times 7.4\,CO_2 + 0.1 \times 7.4\,CO + 8.4\,H_2O + x\,O_2 + 1.1 \times 11.55 \times 3.76\,N_2$$

The coefficient for O_2 is determined from a mass balance,

O_2 balance: $0.5 \times 0.1 + 1.1 \times 11.55 = 0.9 \times 7.4 + 0.5 \times (0.1 \times 7.4) + 0.5 \times 8.4 + x \longrightarrow x = 1.525$

Substituting,

$$0.1\,C_2H_6O + 0.9\,C_8H_{18} + 12.705[O_2 + 3.76N_2]$$
$$\longrightarrow 6.66\,CO_2 + 0.74\,CO + 8.4\,H_2O + 1.525\,O_2 + 47.77\,N_2$$

(b) The partial pressure of water vapor is

$$P_v = \frac{N_{\text{H2O}}}{N_{\text{total}}} P_{\text{total}} = \frac{8.4}{6.66 + 0.74 + 8.4 + 1.525 + 47.77}(100\text{ kPa}) = \frac{8.4\text{ kmol}}{65.10\text{ kmol}}(100\text{ kPa}) = 12.9\text{ kPa}$$

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

$$T_{dp} = T_{\text{sat@12.9 kPa}} = \mathbf{50.5°C} \quad \text{(Table A-5)}$$

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

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

Both the reactants and products are at 25 °C. Assuming the air and the combustion products to be ideal gases, we have $h = h(T)$. Then, using the values given in the table,

$$-Q_{\text{out}} = (6.66)(-393{,}520) + (0.74)(-110{,}530) + (8.4)(-241{,}820) - (0.1)(-235{,}310) - (0.9)(-208{,}450)$$
$$= -4{,}522{,}790\text{ kJ/kmol fuel}$$

or $\quad Q_{\text{out}} = 4{,}522{,}790$ kJ/kmol fuel

The molar mass of the fuel is

$$M = 0.1 \times 46 + 0.9 \times 114 = 107.2\text{ kg/kmol}$$

Then the heat transfer for a 2.5 kg of fuel becomes

$$Q_{\text{out}} = NQ_{\text{out}} = \left(\frac{m}{M}\right)Q_{\text{out}} = \frac{2.5\text{ kg}}{107.2\text{ kg/kmol}}(4{,}522{,}790\text{ kJmol}) = \mathbf{105{,}480\ kJ}$$

(d) For 9.57 kmol of water vapor in the products, the air must carry 9.57 − 8.4 = 1.17 kmol of water vapor in the atmospheric air. The partial pressure of this water vapor in the stmospheric air is

$$P_v = \frac{N_{\text{H2O}}}{N_{\text{total}}}P_{\text{total}} = \frac{1.17}{12.705 \times 4.76 + 1.17}(100\text{ kPa}) = \frac{1.17\text{ kmol}}{61.65\text{ kmol}}(100\text{ kPa}) = 1.8979\text{ kPa}$$

The saturation pressure of water at 25°C is 3.17 kPa (Table A-4). The relative humidity of water vapor in the atmospheric air is then

$$\phi = \frac{P_v}{P_{\text{total}}} = \frac{1.8979\text{ kPa}}{3.17\text{ kPa}} = 0.599 = \mathbf{59.9\%}$$

15-68 A mixture of methane and oxygen contained in a tank is burned at constant volume. The final pressure in the tank and the heat transfer during this process are to be determined.

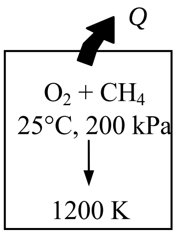
Assumptions **1** Air and combustion gases are ideal gases. **2** Combustion is complete.

Properties The molar masses of CH_4 and O_2 are 16 kg/kmol and 32 kg/kmol, respectively (Table A-1).

Analysis (*a*) The combustion is assumed to be complete, and thus all the carbon in the methane burns to CO_2 and all of the hydrogen to H_2O. The number of moles of CH_4 and O_2 in the tank are

$$N_{CH_4} = \frac{m_{CH_4}}{M_{CH_4}} = \frac{0.12\text{ kg}}{16\text{ kg/kmol}} = 7.5\times10^{-3}\text{ kmol} = 7.5\text{ mol}$$

$$N_{O_2} = \frac{m_{O_2}}{M_{O_2}} = \frac{0.6\text{ kg}}{32\text{ kg/kmol}} = 18.75\times10^{-3}\text{ kmol} = 18.75\text{ mol}$$



Then the combustion equation can be written as

$$7.5CH_4 + 18.75O_2 \longrightarrow 7.5CO_2 + 15H_2O + 3.75O_2$$

At 1200 K, water exists in the gas phase. Assuming both the reactants and the products to be ideal gases, the final pressure in the tank is determined to be

$$\left.\begin{aligned} P_R \boldsymbol{V} &= N_R R_u T_R \\ P_P \boldsymbol{V} &= N_P R_u T_P \end{aligned}\right\} P_P = P_R \left(\frac{N_P}{N_R}\right)\left(\frac{T_P}{T_R}\right)$$

Substituting,

$$P_P = (200\text{ kPa})\left(\frac{26.25\text{ mol}}{26.25\text{ mol}}\right)\left(\frac{1200\text{ K}}{298\text{ K}}\right) = \mathbf{805\text{ kPa}}$$

which is relatively low. Therefore, the ideal gas assumption utilized earlier is appropriate.

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

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

Since both the reactants and products are assumed to be ideal gases, all the internal energy and enthalpies depend on temperature only, and the $P\bar{\boldsymbol{v}}$ terms in this equation can be replaced by R_uT. It yields

$$-Q_{\text{out}} = \sum N_P\left(\bar{h}_f^\circ + \bar{h}_{1200\text{ K}} - \bar{h}_{298\text{ K}} - R_uT\right)_P - \sum N_R\left(\bar{h}_f^\circ - R_uT\right)_R$$

since the reactants are at the standard reference temperature of 25°C. From the tables,

Substance	$\bar{h}_f^\circ$ **kJ/kmol**	$\bar{h}_{298\text{ K}}$ **kJ/kmol**	$\bar{h}_{1200\text{ K}}$ **kJ/kmol**
CH_4	-74,850	---	---
O_2	0	8682	38,447
H_2O (*g*)	-241,820	9904	44,380
CO_2	-393,520	9364	53,848

Thus,

$$\begin{aligned} -Q_{\text{out}} = &(7.5)(-393{,}520 + 53{,}848 - 9364 - 8.314\times1200) \\ &+ (15)(-241{,}820 + 44{,}380 - 9904 - 8.314\times1200) \\ &+ (3.75)(0 + 38{,}447 - 8682 - 8.314\times1200) \\ &- (7.5)(-74{,}850 - 8.314\times298) - (18.75)(-8.314\times298) \\ = &-5{,}251{,}791\text{ J} = -5252\text{ kJ} \end{aligned}$$

Thus $Q_{\text{out}} = \mathbf{5252\text{ kJ}}$ of heat is transferred from the combustion chamber as 120 g of CH_4 burned in this combustion chamber.

15-69 Problem 15-68 is reconsidered. The effect of the final temperature on the final pressure and the heat transfer for the combustion process is to be investigated.

Analysis The problem is solved using EES, and the solution is given below.

```
"Input Data"
T_reac = (25+273) "[K]"              "reactant mixture temperature"
P_reac = 200 [kPa]                   "reactant mixture pressure"
{T_prod = 1200 [K]}                  "product mixture temperature"
m_O2=0.600 [kg]                      "initial mass of O2"
Mw_O2 = 32 [kg/kmol]
m_CH4 = 0.120 [kg]                   "initial mass of CH4"
Mw_CH4=(1*12+4*1) "[kg/kmol]"
R_u = 8.314 [kJ/kmol-K]              "universal gas constant"

"For  theoretical oxygen, the complete combustion equation is"
"CH4 + A_th O2=1 CO2+2 H2O  "
2*A_th=1*2+2*1"theoretical O balance"

"now to find the actual moles of O2 supplied  per mole of fuel"
N_O2 = m_O2/Mw_O2/N_CH4
N_CH4= m_CH4/Mw_CH4

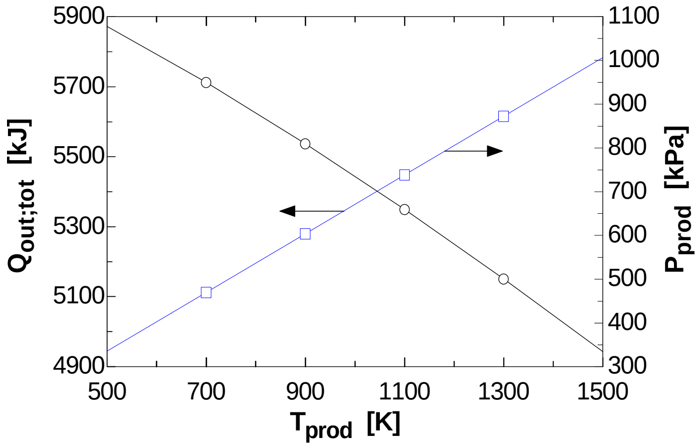
"The balanced complete combustion equation with Ex% excess O2 is"
"CH4 + (1+EX) A_th O2=1 CO2+ 2 H2O + Ex( A_th) O2 "
N_O2 = (1+Ex)*A_th

"Apply First Law to the closed system combustion chamber and assume ideal gas
behavior. (At 1200 K, water exists in the gas phase.)"
E_in - E_out = DELTAE_sys
E_in = 0
E_out = Q_out "kJ/kmol_CH4"  "No work is done because volume is constant"
DELTAE_sys = U_prod - U_reac "neglect KE and PE and note: U = H - PV = N(h - R_u T)"
U_reac = 1*(enthalpy(CH4, T=T_reac) - R_u*T_reac) +(1+EX)*A_th*(enthalpy(O2,T=T_reac) - R_u*T_reac)
U_prod = 1*(enthalpy(CO2, T=T_prod) - R_u*T_prod) +2*(enthalpy(H2O, T=T_prod) -
R_u*T_prod)+EX*A_th*(enthalpy(O2,T=T_prod) - R_u*T_prod)

"The total heat transfer out, in kJ, is:"
Q_out_tot=Q_out"kJ/kmol_CH4"/(Mw_CH4 "kg/kmol_CH4") *m_CH4"kg"  "kJ"

"The final pressure in the tank is the pressure of the product gases.  Assuming
ideal gas behavior for the gases in the constant volume tank, the ideal gas law gives:"
P_reac*V =N_reac * R_u *T_reac
P_prod*V = N_prod * R_u * T_prod
N_reac = N_CH4*(1 + N_O2)
N_prod = N_CH4*(1 + 2 + Ex*A_th)
```

T_{prod} [K]	$Q_{out,tot}$ [kJ]	P_{prod} [kPa]
500	5872	335.6
700	5712	469.8
900	5537	604
1100	5349	738.3
1300	5151	872.5
1500	4943	1007



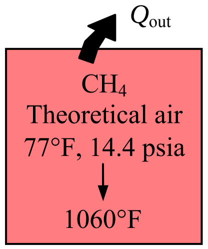
15-70E Methane is burned with stoichiometric amount of air in a rigid container. The heat rejected from the container is to be determined.

Assumptions **1** Air and combustion gases are ideal gases. **2** Combustion is complete.

Properties The molar masses of CH_4 and air are 16 lbm/lbmol and 29 lbm/lbmol, respectively (Table A-1E).

Analysis The combustion equation for 1 lbmol of fuel is

$$\mathrm{CH_4} + 2(\mathrm{O_2} + 3.76N_2) \longrightarrow \mathrm{CO_2} + 2\mathrm{H_2O} + 7.52\mathrm{N_2}$$



The heat transfer for this constant volume 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

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

Since both the reactants and products are assumed to be ideal gases, all the internal energy and enthalpies depend on temperature only, and the $P\bar{\boldsymbol{v}}$ terms in this equation can be replaced by R_uT. It yields

$$-Q_{out} = \sum N_P \left(\bar{h}_f^\circ + \bar{h}_{1520\,\mathrm{R}} - \bar{h}_{537\,\mathrm{R}} - R_uT\right)_P - \sum N_R \left(\bar{h}_f^\circ - R_uT\right)_R$$

since the reactants are at the standard reference temperature of 77°F. From the tables,

Substance	$\bar{\mathbf{h}}_\mathbf{f}^\circ$ **Btu/lbmol**	$\bar{\mathbf{h}}_{537\mathrm{R}}$ **Btu/lbmol**	$\bar{\mathbf{h}}_{1520\mathrm{R}}$ **Btu/lbmol**
CH_4	-32,210	---	---
O_2	0	3725.1	11,179.6
N_2	0	3729.5	10,800.4
H_2O (*g*)	-104,040	4258.0	12,738.8
CO_2	-169,300	4027.5	14,824.9

Thus,

$$\begin{aligned}
-Q_{out} &= (1)(-169{,}300 + 14{,}824.9 - 4027.5 - 1.9858 \times 1520) \\
&\quad + (2)(-104{,}040 + 12{,}738.8 - 4258.0 - 1.9858 \times 1520) \\
&\quad + (7.52)(0 + 10{,}800.4 - 3729.5 - 1.9858 \times 1520) \\
&\quad - (1)(-32{,}210 - 1.9858 \times 537) - (2)(-1.9858 \times 537) - (7.52)(-1.9858 \times 537) \\
&= -284{,}800 \text{ Btu/lbmol } \mathrm{CH_4}
\end{aligned}$$

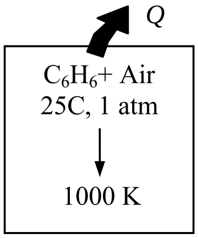
Thus

$$Q_{out} = \mathbf{284{,}800\ Btu/lbmol\ CH_4}$$

15-71 A mixture of benzene gas and 30 percent excess air contained in a constant-volume tank is ignited. The heat transfer from the combustion chamber is to be determined.

Assumptions **1** Both the reactants and products are ideal gases. **2** Combustion is complete.

Analysis The theoretical combustion equation of C_6H_6 with stoichiometric amount of air is



$$C_6H_6(g) + a_{th}(O_2 + 3.76N_2) \longrightarrow 6CO_2 + 3H_2O + 3.76a_{th}N_2$$

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

$$a_{th} = 6 + 1.5 = 7.5$$

Then the actual combustion equation with 30% excess air becomes

$$C_6H_6(g) + 9.75(O_2 + 3.76N_2) \longrightarrow 5.52CO_2 + 0.48CO + 3H_2O + 2.49O_2 + 36.66N_2$$

The heat transfer for this constant volume 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

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

Since both the reactants and the products behave as ideal gases, all the internal energy and enthalpies depend on temperature only, and the $P\bar{v}$ terms in this equation can be replaced by R_uT.

It yields

$$-Q_{out} = \sum N_P\left(\bar{h}_f^\circ + \bar{h}_{1000\text{ K}} - \bar{h}_{298\text{ K}} - R_uT\right)_P - \sum N_R\left(\bar{h}_f^\circ - R_uT\right)_R$$

since the reactants are at the standard reference temperature of 25°C. From the tables,

Substance	$\bar{\mathbf{h}}_\mathbf{f}^\circ$ **kJ/kmol**	$\bar{\mathbf{h}}_{\mathbf{298\ K}}$ **kJ/kmol**	$\bar{\mathbf{h}}_{\mathbf{1000\ K}}$ **kJ/kmol**
C_6H_6 (*g*)	82,930	---	---
O_2	0	8682	31,389
N_2	0	8669	30,129
H_2O (*g*)	-241,820	9904	35,882
CO	-110,530	8669	30,355
CO_2	-393,520	9364	42,769

Thus,

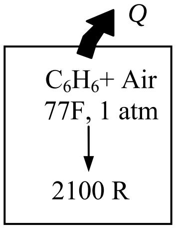
$$\begin{aligned}-Q_{out} &= (5.52)(-393{,}520 + 42{,}769 - 9364 - 8.314\times1000)\\&\quad + (0.48)(-110{,}530 + 30{,}355 - 8669 - 8.314\times1000)\\&\quad + (3)(-241{,}820 + 35{,}882 - 9904 - 8.314\times1000)\\&\quad + (2.49)(0 + 31{,}389 - 8682 - 8.314\times1000)\\&\quad + (36.66)(0 + 30{,}129 - 8669 - 8.314\times1000)\\&\quad - (1)(82{,}930 - 8.314\times298) - (9.75)(4.76)(-8.314\times298)\\&= -2{,}200{,}433\text{ kJ}\end{aligned}$$

or $Q_{out} = \mathbf{2{,}200{,}433\ kJ}$

15-72E A mixture of benzene gas and 60 percent excess air contained in a constant-volume tank is ignited. The heat transfer from the combustion chamber is to be determined.

Assumptions **1** Both the reactants and products are ideal gases. **2** Combustion is complete.

Analysis The theoretical combustion equation of C_6H_6 with stoichiometric amount of air is



$$C_6H_6(g) + a_{th}(O_2 + 3.76N_2) \longrightarrow 6CO_2 + 3H_2O + 3.76a_{th}N_2$$

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

$$a_{th} = 6 + 1.5 = 7.5$$

Then the actual combustion equation with 60% excess air becomes

$$C_6H_6(g) + 12(O_2 + 3.76N_2) \longrightarrow 5.52CO_2 + 0.48CO + 3H_2O + 4.74O_2 + 45.12N_2$$

The heat transfer for this constant volume 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

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

Since both the reactants and the products behave as ideal gases, all the internal energy and enthalpies depend on temperature only, and the $P\bar{v}$ terms in this equation can be replaced by R_uT.

It yields

$$-Q_{out} = \sum N_P\left(\bar{h}_f^\circ + \bar{h}_{1800R} - \bar{h}_{537R} - R_uT\right)_P - \sum N_R\left(\bar{h}_f^\circ - R_uT\right)_R$$

since the reactants are at the standard reference temperature of 77°F. From the tables,

Substance	$\bar{h}_f^\circ$ **Btu/lbmol**	$\bar{h}_{537\ R}$ **Btu/lbmol**	$\bar{h}_{2100R}$ **Btu/lbmol**
C_6H_6 (*g*)	35,680	---	---
O_2	0	3725.1	16,011
N_2	0	3729.5	15,334
H_2O (*g*)	-104,040	4258.0	18,467
CO	-47,540	3725.1	15,463
CO_2	-169,300	4027.5	22,353

Thus,

$$\begin{aligned}
-Q_{out} &= (5.52)(-169{,}300 + 22{,}353 - 4027.5 - 1.986 \times 2100) \\
&\quad + (0.48)(-47{,}540 + 15{,}463 - 3725.1 - 1.986 \times 2100) \\
&\quad + (3)(-104{,}040 + 18{,}467 - 4258.0 - 1.986 \times 2100) \\
&\quad + (4.74)(0 + 16{,}011 - 3725.1 - 1.986 \times 2100) \\
&\quad + (45.12)(0 + 15{,}334 - 3729.5 - 1.986 \times 2100) \\
&\quad - (1)(35{,}680 - 1.986 \times 537) - (12)(4.76)(-1.986 \times 537) \\
&= -757{,}400 \text{ Btu}
\end{aligned}$$

or $Q_{out} =$ **757,400 Btu**

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.

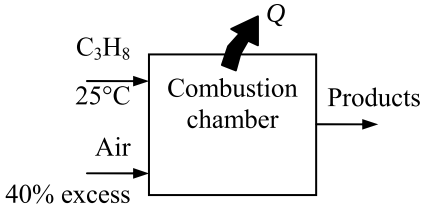
Assumptions **1** Combustion is complete. **2** The combustion products contain CO_2, H_2O, O_2, and N_2 only.

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

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

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

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

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

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

15-74 Wheat straw that is being considered as an alternative fuel is tested in a bomb calorimeter. The heating value of this straw is to be determined and compared to the higher heating value of propane.

Assumptions **1** Combustion is complete.

Analysis The heat released by the combustion is

$$Q = mc_v \Delta T = (100\ \text{kJ/K})(1.8\ \text{K}) = 180\ \text{kJ}$$

The heating value is then

$$\text{HV} = \frac{Q}{m} = \frac{180\ \text{kJ}}{0.010\ \text{kg}} = \mathbf{18{,}000\ kJ/kg}$$

From Table A-27, the higher heating value of propane is

$$\text{HHV} = \mathbf{50{,}330\ kJ/kg}$$

Adiabatic Flame Temperature

15-75C For the case of stoichiometric amount of pure oxygen since we have the same amount of chemical energy released but a smaller amount of mass to absorb it.

15-76C Under the conditions of complete combustion with stoichiometric amount of air.

15-77 Hydrogen is burned with 50 percent excess air during a steady-flow combustion process. The exit temperature of product gases is to be determined.

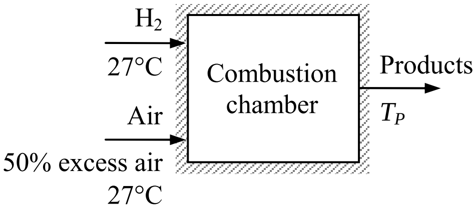
Assumptions **1** Steady operating conditions exist. **2** Air and combustion gases are ideal gases. **3** Kinetic and potential energies are negligible. **4** There are no work interactions. **5** The combustion chamber is adiabatic.

Analysis Adiabatic flame temperature is the temperature at which the products leave the combustion chamber under adiabatic conditions ($Q = 0$) with no work interactions ($W = 0$). Under steady-flow conditions the energy balance $E_{\text{in}} - E_{\text{out}} = \Delta E_{\text{system}}$ applied on the combustion chamber reduces to

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

The combustion equation of H_2 with 50% excess air is

$$H_2 + 0.75(O_2 + 3.76N_2) \longrightarrow H_2O + 0.25O_2 + 2.82N_2$$



From the tables,

Substance	$\bar{h}_f^\circ$ kJ/kmol	$\bar{h}_{300\text{K}}$ kJ/kmol	$\bar{h}_{298\text{ K}}$ kJ/kmol
H_2	0	8522	8468
O_2	0	8736	8682
N_2	0	8723	8669
H_2O (*g*)	-241,820	9966	9904

Thus,

$$\begin{aligned}&(1)\left(-241{,}820 + \bar{h}_{H_2O} - 9904\right) + (0.25)\left(0 + \bar{h}_{O_2} - 8682\right) + (2.82)\left(0 + \bar{h}_{N_2} - 8669\right)\\ &\quad = (1)(0 + 8522 - 8468) + (0.75)(0 + 8736 - 8682) + (2.82)(0 + 8723 - 8669)\end{aligned}$$

It yields

$$\bar{h}_{H_2O} + 0.25\bar{h}_{O_2} + 2.82\bar{h}_{N_2} = 278{,}590 \text{ kJ}$$

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 278,590/(1 + 0.25 + 2.82) = 68,450 kJ/kmol. This enthalpy value corresponds to about 2100 K for N_2. Noting that the majority of the moles are N_2, T_P will be close to 2100 K, but somewhat under it because of the higher specific heat of H_2O.

At 2000 K:
$$\begin{aligned}\bar{h}_{H_2O} + 0.25\bar{h}_{O_2} + 2.82\bar{h}_{N_2} &= (1)(82{,}593) + (0.25)(67{,}881) + (2.82)(64{,}810)\\ &= 282{,}330 \text{ kJ (Higher than } 278{,}590 \text{ kJ)}\end{aligned}$$

At 1960 K:
$$\begin{aligned}\bar{h}_{H_2O} + 0.25\bar{h}_{O_2} + 2.826\bar{h}_{N_2} &= (1)(80{,}555) + (0.25)(66{,}374) + (2.82)(63{,}381)\\ &= 275{,}880 \text{ kJ (Lower than } 278{,}590 \text{ kJ)}\end{aligned}$$

By interpolation, T_P = **1977 K**

Discussion The adiabatic flame temperature cam be obtained by using EES without a trial and error approach. We found the temperature to be **1978 K** by EES. The results are practically identical.

15-78 Problem 15-77 is reconsidered. This problem is to be modified to include the fuels butane, ethane, methane, and propane as well as H2; to include the effects of inlet air and fuel temperatures; and the percent theoretical air supplied.

Analysis The problem is solved using EES, and the solution is given below.

```
Adiabatic Combustion of fuel CnHm entering at T_fuel with Stoichiometric Air at T_air:
Reaction:  CxHy  +  (y/4 + x) (Theo_air/100) (O2 + 3.76 N2)
 <--> xCO2  +  (y/2) H2O  +  3.76  (y/4 + x) (Theo_air/100)  N2 +  (y/4 + x) (Theo_air/100 - 1) O2
T_prod is the adiabatic combustion temperature, assuming no dissociation.
Theo_air is the % theoretical air. " "The initial guess value of T_prod = 450K ."
Procedure Fuel(Fuel$:x,y,Name$)
"This procedure takes the fuel name and returns the moles of C and moles of H"
If fuel$='C2H6' then
          x=2;y=6
             Name$='ethane'
else
If fuel$='C3H8' then
          x=3; y=8
             Name$='propane'
else
If fuel$='C4H10' then
          x=4; y=10
            Name$='butane'
else
if fuel$='CH4' then
          x=1; y=4
            Name$='methane'
else
if fuel$='H2' then
          x=0; y=2
            Name$='hydrogen'
endif; endif; endif; endif; endif
end

{"Input data from the diagram window"
T_fuel = 300 [K]
T_air = 300  [K]
Theo_air = 150  "%"
Fuel$='H2'}
Call Fuel(fuel$:x,y,Name$)
HR=enthalpy(Fuel$,T=T_fuel)+ (y/4 + x) *(Theo_air/100) *enthalpy(O2,T=T_air)+3.76*(y/4 + x) *(Theo_air/100)
*enthalpy(N2,T=T_air)
HP=HR  "Adiabatic"
HP=x*enthalpy(CO2,T=T_prod)+(y/2)*enthalpy(H2O,T=T_prod)+3.76*(y/4 + x)*
(Theo_air/100)*enthalpy(N2,T=T_prod)+(y/4 + x) *(Theo_air/100 - 1)*enthalpy(O2,T=T_prod)
Moles_O2=(y/4 + x) *(Theo_air/100 - 1)
Moles_N2=3.76*(y/4 + x)* (Theo_air/100)
Moles_CO2=x;   Moles_H2O=y/2
T[1]=T_prod;  xa[1]=Theo_air  "array variable are plotted in Plot Window 1"
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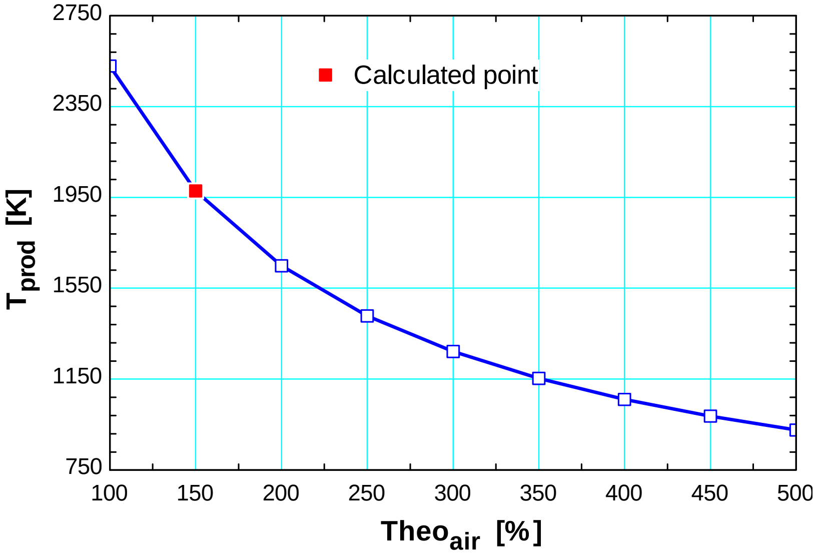
$Theo_{air}$ [%]	T_{prod} [K]
100	2528
150	1978
200	1648
250	1428
300	1271
350	1153
400	1060
450	986.3
500	925.5



15-79 Acetylene is burned with stoichiometric amount of oxygen. The adiabatic flame temperature is to be determined.

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.

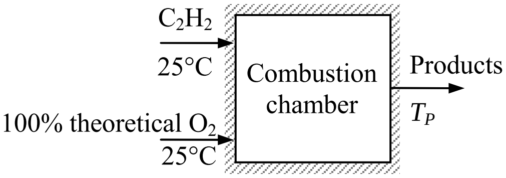
Analysis Under steady-flow conditions the energy balance $E_{\text{in}} - E_{\text{out}} = \Delta E_{\text{system}}$ applied on the combustion chamber with $Q = W = 0$ reduces to

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

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

$$C_2H_2 + 2.5O_2 \longrightarrow 2\,CO_2 + 1\,H_2O$$

From the tables,



Substance	$\bar{\mathbf{h}}_{\mathbf{f}}^\circ$ **kJ/kmol**	$\bar{\mathbf{h}}_{\mathbf{298K}}$ **kJ/kmol**
C_2H_2 (*g*)	226,730	---
O_2	0	8682
N_2	0	8669
H_2O (*g*)	-241,820	9904
CO_2	-393,520	9364

Thus,

$$(2)\left(-393{,}520 + \bar{h}_{\text{CO2}} - 9364\right) + (1)\left(-241{,}820 + \bar{h}_{\text{H2O}} - 9904\right) = (1)(226{,}730) + 0 + 0$$

It yields $$2\bar{h}_{\text{CO2}} + 1\bar{h}_{\text{H2O}} = 1{,}284{,}220 \text{ kJ}$$

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 2,284,220/(2+1) = 428,074 kJ/kmol. The ideal gas tables do not list enthalpy values this high. Therefore, we cannot use the tables to estimate the adiabatic flame temperature. In Table A-2b, the highest available value of specific heat is $c_p = 1.234$ kJ/kg·K for CO_2 at 1000 K. The specific heat of water vapor is $c_p = 1.8723$ kJ/kg·K (Table A-2a). Using these specific heat values,

$$(2)\left(-393{,}520 + \bar{c}_p\Delta T\right) + (1)\left(-241{,}820 + \bar{c}_p\Delta T\right) = (1)(226{,}730) + 0 + 0$$

where $\Delta T = (T_{\text{af}} - 25)°\text{C}$. The specific heats on a molar base are

$$\bar{c}_{p,\text{CO2}} = \bar{c}_p M = (1.234 \text{ kJ/kg}\cdot\text{K})(44 \text{ kg/kmol}) = 54.3 \text{ kJ/kmol}\cdot\text{K}$$
$$\bar{c}_{p,\text{H2O}} = \bar{c}_p M = (1.8723 \text{ kJ/kg}\cdot\text{K})(18 \text{ kg/kmol}) = 33.7 \text{ kJ/kmol}\cdot\text{K}$$

Substituting,

$$(2)\left(-393{,}520 + 54.3\Delta T\right) + (1)\left(-241{,}820 + 33.7\Delta T\right) = 226{,}730$$
$$(2 \times 54.3)\Delta T + 33.7\Delta T = 1{,}255{,}590$$
$$\Delta T = \frac{1{,}255{,}590 \text{ kJ/kmol}}{(2 \times 54.3 + 33.7) \text{ kJ/kmol}\cdot\text{K}} = 8824 \text{ K}$$

Then the adiabatic flame temperature is estimated as

$$T_{\text{af}} = \Delta T + 25 = 8824 + 25 = \mathbf{8849°C}$$

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

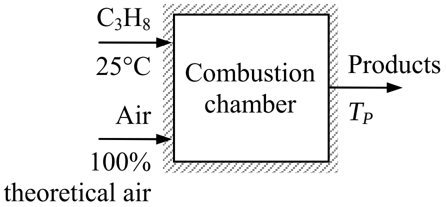
$$\sum N_P\left(\bar{h}_f^\circ + \bar{h} - \bar{h}^\circ\right)_P = \sum N_R\left(\bar{h}_f^\circ + \bar{h} - \bar{h}^\circ\right)_R \longrightarrow \sum N_P\left(\bar{h}_f^\circ + \bar{h} - \bar{h}^\circ\right)_P = \sum N_R \bar{h}_{f,R}^\circ$$

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	$\bar{\mathbf{h}}_\mathbf{f}^\circ$ **kJ/kmol**	$\bar{\mathbf{h}}_{\mathbf{298K}}$ **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,

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

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

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:

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

At 2450 K:

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

At 2400 K:

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

At 2350 K:

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

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

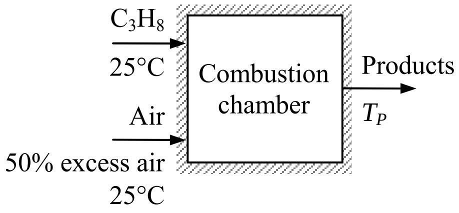
$$C_3H_8 + 1.5\times a_{th}(O_2 + 3.76N_2) \longrightarrow 3CO_2 + 4H_2O + 0.5\times a_{th}O_2 + 1.5\times a_{th}\times 3.76N_2$$

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,

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

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:

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

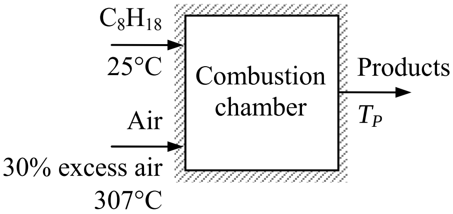
At 1840 K:

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

By interpolation,

T_P = **1830 K = 1557°C**

15-81 Octane is burned with 40 percent excess air adiabatically during a steady-flow combustion process. The exit temperature of product gases is to be determined.



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

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

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

$$C_8H_{18}(l) + 1.4a_{\text{th}}(O_2 + 3.76N_2) \longrightarrow 8CO_2 + 9H_2O + 0.4a_{\text{th}}O_2 + (1.4)(3.76)a_{\text{th}}N_2$$

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

$$1.4a_{\text{th}} = 8 + 4.5 + 0.4a_{\text{th}} \longrightarrow a_{\text{th}} = 12.5$$

Thus, $C_8H_{18}(g) + 17.5(O_2 + 3.76N_2) \longrightarrow 8CO_2 + 9H_2O + 5O_2 + 65.8N_2$

From the tables,

Substance	$\bar{h}_f^\circ$ kJ/kmol	$\bar{h}_{298K}$ kJ/kmol	$\bar{h}_{580K}$ kJ/kmol
C_8H_{18} (*l*)	-249,950	---	---
O_2	0	8682	17,290
N_2	0	8669	16,962
H_2O (*g*)	-241,820	9904	---
CO_2	-393,520	9364	---

Thus,

$$\begin{aligned}&(8)(-393{,}520 + \bar{h}_{CO2} - 9364) + (9)(-241{,}820 + \bar{h}_{H2O} - 9904) + (5)(0 + \bar{h}_{O2} - 8682)\\ &+ (65.8)(0 + \bar{h}_{N2} - 8669) = (1)(-249{,}950) + (17.5)(0 + 17{,}290 - 8682) + (65.8)(16{,}962 - 8669)\end{aligned}$$

It yields $\quad 8\bar{h}_{CO2} + 9\bar{h}_{H2O} + 5\bar{h}_{O2} + 65.8\bar{h}_{N2} = 6{,}548{,}788 \text{ kJ}$

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 6,548,788/(8 + 9 + 5 + 65.8) = 74,588 kJ/kmol. This enthalpy value corresponds to about 2250 K for N_2. Noting that the majority of the moles are N_2, T_P will be close to 2250 K, but somewhat under it because of the higher specific heat of H_2O.

At 2100 K:

$$\begin{aligned}8\bar{h}_{CO2} + 9\bar{h}_{H2O} + 5\bar{h}_{O2} + 65.8\bar{h}_{N2} &= (8)(106{,}864) + (9)(87{,}735) + (5)(71{,}668) + (65.8)(68{,}417)\\ &= 6{,}504{,}706 \text{ kJ (Lower than } 6{,}548{,}788 \text{ kJ)}\end{aligned}$$

At 2150 K:

$$\begin{aligned}8\bar{h}_{CO2} + 9\bar{h}_{H2O} + 5\bar{h}_{O2} + 65.8\bar{h}_{N2} &= (8)(109{,}898) + (9)(90{,}330) + (5)(73{,}573) + (65.8)(70{,}226)\\ &= 6{,}680{,}890 \text{ kJ (Higher than } 6{,}548{,}788 \text{ kJ)}\end{aligned}$$

By interpolation,

T_P = **2113 K = 1840°C**

15-82 A certain coal is burned with 100 percent excess air adiabatically during a steady-flow combustion process. The temperature of product gases is to be determined for complete combustion and incomplete combustion 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.

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

Analysis We consider 100 kg of coal for simplicity. Noting that the mass percentages in this case correspond to the masses of the constituents, the mole numbers of the constituent of the coal are determined to be

$$N_C = \frac{m_C}{M_C} = \frac{84.36\text{ kg}}{12\text{ kg/kmol}} = 7.03\text{ kmol}$$

$$N_{H2} = \frac{m_{H2}}{M_{H2}} = \frac{1.89\text{ kg}}{2\text{ kg/kmol}} = 0.945\text{ kmol}$$

$$N_{O2} = \frac{m_{O2}}{M_{O2}} = \frac{4.40\text{ kg}}{32\text{ kg/kmol}} = 0.1375\text{ kmol}$$

$$N_{N2} = \frac{m_{N2}}{M_{N2}} = \frac{0.63\text{ kg}}{28\text{ kg/kmol}} = 0.0225\text{ kmol}$$

$$N_S = \frac{m_S}{M_S} = \frac{0.89\text{ kg}}{32\text{ kg/kmol}} = 0.0278\text{ kmol}$$

84.36% C
1.89% H_2
4.40% O_2
0.63% N_2
0.89% S
7.83% ash
(by mass)

The mole number of the mixture and the mole fractions are

$$N_m = 7.03 + 0.945 + 0.1375 + 0.0225 + 0.0278 = 8.163\text{ kmol}$$

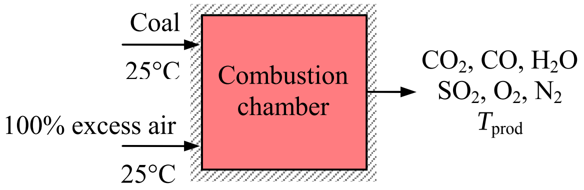
$$y_C = \frac{N_C}{N_m} = \frac{7.03\text{ kmol}}{8.163\text{ kmol}} = 0.8611$$

$$y_{H2} = \frac{N_{H2}}{N_m} = \frac{0.945\text{ kmol}}{8.163\text{ kmol}} = 0.1158$$

$$y_{O2} = \frac{N_{O2}}{N_m} = \frac{0.1375\text{ kmol}}{8.163\text{ kmol}} = 0.01684$$

$$y_{N2} = \frac{N_{N2}}{N_m} = \frac{0.0225\text{ kmol}}{8.163\text{ kmol}} = 0.00276$$

$$y_S = \frac{N_S}{N_m} = \frac{0.0278\text{ kmol}}{8.163\text{ kmol}} = 0.003407$$



Ash consists of the non-combustible matter in coal. Therefore, the mass of ash content that enters the combustion chamber is equal to the mass content that leaves. Disregarding this non-reacting component for simplicity, the combustion equation may be written as

$$0.8611\text{C} + 0.1158\text{H}_2 + 0.01684\text{O}_2 + 0.00276\text{N}_2 + 0.00341\text{S} + 2a_{\text{th}}(\text{O}_2 + 3.76\text{N}_2)$$
$$\longrightarrow x(0.97\text{CO}_2 + 0.03\text{CO}) + y\text{H}_2\text{O} + z\text{SO}_2 + k\text{N}_2 + a_{\text{th}}\text{O}_2$$

According to the species balances,

$$\text{C balance:}\quad x = 0.8611$$
$$\text{H}_2\text{ balance:}\quad y = 0.1158$$
$$\text{S balance:}\quad z = 0.00341$$
$$\text{O}_2\text{ balance:}$$
$$0.01684 + 2a_{\text{th}} = 0.97x + 0.015x + 0.5y + z + a_{\text{th}}$$
$$a_{\text{th}} = (0.97)(0.8611) + (0.015)(0.8611) + (0.5)(0.1158) + 0.00341 - 0.01684 = 0.8927$$
$$\text{N}_2\text{ balance:}\quad 0.00276 + 2\times 3.76a_{\text{th}} = k \longrightarrow k = 0.00276 + 2\times 3.76\times 0.8927 = 6.72$$

Substituting,

$$0.8611\text{C} + 0.1158\text{H}_2 + 0.01684\text{O}_2 + 0.00276\text{N}_2 + 0.003407\text{S} + 1.785(\text{O}_2 + 3.76\text{N}_2)$$
$$\longrightarrow 0.8353\text{CO}_2 + 0.0258\text{CO} + 0.1158\text{H}_2\text{O} + 0.00341\text{SO}_2 + 6.72\text{N}_2 + 0.8927\text{O}_2$$

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

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

From the tables,

Substance	$\bar{\mathbf{h}}_{\mathbf{f}}^\circ$ **kJ/kmol**	$\bar{\mathbf{h}}_{\mathbf{298K}}$ **kJ/kmol**
O_2	0	8682
N_2	0	8669
H_2O (*g*)	-241,820	9904
CO	-110,530	8669
CO_2	-110,530	8669

Thus,

$$(0.8353)\left(-393{,}520 + \bar{h}_{\text{CO2}} - 9364\right) + (0.0258)\left(-110{,}530 + \bar{h}_{\text{CO}} - 8669\right) + (0.1158)\left(-241{,}820 + \bar{h}_{\text{H2O}} - 9904\right)$$
$$+ (0.8927)\left(0 + \bar{h}_{\text{O2}} - 8682\right) + (6.72)\left(0 + \bar{h}_{\text{N2}} - 8669\right) = 0$$

It yields $\quad 0.8353\bar{h}_{\text{CO2}} + 0.0258\bar{h}_{\text{CO}} + 0.1158\bar{h}_{\text{H2O}} + 0.8927\bar{h}_{\text{O2}} + 6.72\bar{h}_{\text{N2}} = 434{,}760 \text{ kJ}$

The product 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

434,760/(0.8353+0.0258+0.1158+0.00341+6.72+0.8927) = 50,595 kJ/kmol.

This enthalpy value corresponds to about 1600 K for N_2. Noting that the majority of the moles are N_2, T_P will be close to 1600 K, but somewhat under it because of the higher specific heat of H_2O.

At 1500 K:

$$0.8353\bar{h}_{\text{CO2}} + 0.0258\bar{h}_{\text{CO}} + 0.1158\bar{h}_{\text{H2O}} + 0.9095\bar{h}_{\text{O2}} + 6.842\bar{h}_{\text{N2}}$$
$$= (0.8353)(71{,}078) + (0.0258)(47{,}517) + (0.1158)(57{,}999) + (0.8927)(49{,}292) + (6.72)(47{,}073)$$
$$= 427{,}647 \text{ kJ} \quad \text{(Lower than 434,760 kJ)}$$

At 1520 K:

$$0.8353\bar{h}_{\text{CO2}} + 0.0258\bar{h}_{\text{CO}} + 0.1158\bar{h}_{\text{H2O}} + 0.9095\bar{h}_{\text{O2}} + 6.842\bar{h}_{\text{N2}}$$
$$= (0.8353)(72{,}246) + (0.0258)(48{,}222) + (0.1158)(58{,}942) + (0.8927)(50{,}024) + (6.72)(47{,}771)$$
$$= 434{,}094 \text{ kJ} \quad \text{(Lower than 434,760 kJ)}$$

By extrapolation, $\quad T_P =$ **1522 K = 1249°C**

We repeat the calculations for the complete combustion now:

The combustion equation in this case may be written as

$$0.8611\text{C} + 0.1158\text{H}_2 + 0.01684\text{O}_2 + 0.00276\text{N}_2 + 0.00341\text{S} + 2a_{\text{th}}(\text{O}_2 + 3.76\text{N}_2)$$
$$\longrightarrow x\text{CO}_2 + y\text{H}_2\text{O} + z\text{SO}_2 + k\text{N}_2 + a_{\text{th}}\text{O}_2$$

According to the species balances,

$$\text{C balance: } x = 0.8611$$

$$\text{H}_2 \text{ balance: } y = 0.1158$$

$$\text{S balance: } z = 0.00341$$

$$\text{O}_2 \text{ balance:}$$

$$0.01684 + 2a_{\text{th}} = x + 0.5y + z + a_{\text{th}} \longrightarrow a_{\text{th}} = 0.8611 + (0.5)(0.1158) + 0.00341 - 0.01684 = 0.9056$$

$$\text{N}_2 \text{ balance: } 0.00276 + 2 \times 3.76 a_{\text{th}} = k \longrightarrow k = 0.00276 + 2 \times 3.76 \times 0.9056 = 6.81$$

Substituting,

$$0.8611\text{C} + 0.1158\text{H}_2 + 0.01684\text{O}_2 + 0.00276\text{N}_2 + 0.003407\text{S} + 1.819(\text{O}_2 + 3.76\text{N}_2)$$
$$\longrightarrow 0.8611\text{CO}_2 + 0.1158\text{H}_2\text{O} + 0.00341\text{SO}_2 + 6.81\text{N}_2 + 0.9056\text{O}_2$$

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

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

From the tables,

Substance	$\bar{\mathbf{h}}_\mathbf{f}^\circ$ **kJ/kmol**	$\bar{\mathbf{h}}_{\mathbf{298K}}$ **kJ/kmol**
O_2	0	8682
N_2	0	8669
H_2O (*g*)	-241,820	9904
CO	-110,530	8669
CO_2	0	8682

Thus,

$$(0.8611)\left(-393{,}520 + \bar{h}_{\text{CO2}} - 9364\right) + (0.1158)\left(-241{,}820 + \bar{h}_{\text{H2O}} - 9904\right)$$
$$+ (0.9056)\left(0 + \bar{h}_{\text{O2}} - 8682\right) + (6.81)\left(0 + \bar{h}_{\text{N2}} - 8669\right) = 0$$

It yields $$0.8611\bar{h}_{\text{CO2}} + 0.1158\bar{h}_{\text{H2O}} + 0.9056\bar{h}_{\text{O2}} + 6.81\bar{h}_{\text{N2}} = 442{,}971 \text{ kJ}$$

The product 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

442,971/(0.8611+0.1158+0.00341+6.81+0.9056) = 50,940 kJ/kmol.

This enthalpy value corresponds to about 1600 K for N_2. Noting that the majority of the moles are N_2, T_P will be close to 1600 K, but somewhat under it because of the higher specific heat of H_2O.

At 1500 K:

$$0.8611\bar{h}_{\text{CO2}} + 0.1158\bar{h}_{\text{H2O}} + 0.9056\bar{h}_{\text{O2}} + 6.81\bar{h}_{\text{N2}}$$
$$= (0.8611)(71{,}078) + (0.1158)(57{,}999) + (0.9056)(49{,}292) + (6.81)(47{,}073)$$
$$= 433{,}128 \text{ kJ} \quad (\text{Lower than } 442{,}971 \text{ kJ})$$

At 1520 K:

$$0.8611\bar{h}_{\text{CO2}} + 0.1158\bar{h}_{\text{H2O}} + 0.9056\bar{h}_{\text{O2}} + 6.81\bar{h}_{\text{N2}}$$
$$= (0.8611)(72{,}246) + (0.1158)(58{,}942) + (0.9056)(50{,}024) + (6.81)(47{,}771)$$
$$= 439{,}658 \text{ kJ} \quad (\text{Lower than } 442{,}971 \text{ kJ})$$

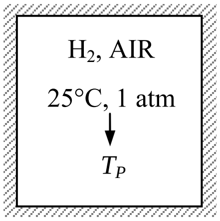
By extrapolation, T_P = **1530 K = 1257°C**

15-83 A mixture of hydrogen and the stoichiometric amount of air contained in a constant-volume tank is ignited. The final temperature in the tank is to be determined.

Assumptions **1** The tank is adiabatic. **2** Both the reactants and products are ideal gases. **3** There are no work interactions. **4** Combustion is complete.

Analysis The combustion equation of H_2 with stoichiometric amount of air is

$$H_2 + 0.5(O_2 + 3.76N_2) \longrightarrow H_2O + 1.88N_2$$



The final temperature in the tank is determined from the energy balance relation $E_{in} - E_{out} = \Delta E_{system}$ for reacting closed systems under adiabatic conditions ($Q = 0$) with no work interactions ($W = 0$),

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

Since both the reactants and the products behave as ideal gases, all the internal energy and enthalpies depend on temperature only, and the $P\bar{v}$ terms in this equation can be replaced by R_uT.

It yields

$$\sum N_P\left(\bar{h}_f^\circ + \bar{h}_{T_P} - \bar{h}_{298\text{ K}} - R_uT\right)_P = \sum N_R\left(\bar{h}_f^\circ R_uT\right)_R$$

since the reactants are at the standard reference temperature of 25°C. From the tables,

Substance	$\bar{\mathbf{h}}_\mathbf{f}^\circ$ **kJ/kmol**	$\bar{\mathbf{h}}_{\mathbf{298\ K}}$ **kJ/kmol**
H_2	0	8468
O_2	0	8682
N_2	0	8669
H_2O (*g*)	-241,820	9904

Thus,

$$\begin{aligned}(1)(-241{,}820 + \bar{h}_{H_2O} - 9904 - 8.314 \times T_P) + (1.88)(0 + \bar{h}_{N_2} - 8669 - 8.314 \times T_P) \\ = (1)(0 - 8.314 \times 298) + (0.5)(0 - 8.314 \times 298) + (1.88)(0 - 8.314 \times 298)\end{aligned}$$

It yields $\bar{h}_{H_2O} + 1.88\bar{h}_{N_2} - 23.94 \times T_P = 259{,}648 \text{ kJ}$

The temperature of the product gases is obtained from a trial and error solution,

At 3050 K: $$\begin{aligned}\bar{h}_{H_2O} + 1.88\bar{h}_{N_2} - 23.94 \times T_P &= (1)(139{,}051) + (1.88)(103{,}260) - (23.94)(3050) \\ &= 260{,}163 \text{ kJ (Higher than } 259{,}648 \text{ kJ)}\end{aligned}$$

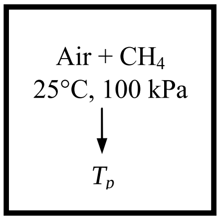
At 3000 K: $$\begin{aligned}\bar{h}_{H_2O} + 1.88\bar{h}_{N_2} - 23.94 \times T_P &= (1)(136{,}264) + (1.88)(101{,}407) - (23.94)(3000) \\ &= 255{,}089 \text{ kJ (Lower than } 259{,}648 \text{ kJ)}\end{aligned}$$

By interpolation, T_P = **3045 K**

15-84 Methane is burned with 300 percent excess air adiabatically in a constant volume container. The final pressure and temperature of product gases are to be determined.

Assumptions **1** 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 The combustion equation is



$$CH_4 + 4a_{th}[O_2 + 3.76N_2] \longrightarrow CO_2 + 2\,H_2O + 4a_{th} \times 3.76\,N_2 + 3a_{th}O_2$$

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

$$4a_{th} = 1 + 1 + 3a_{th} \longrightarrow a_{th} = 2$$

Substituting, $CH_4 + 8[O_2 + 3.76N_2] \longrightarrow CO_2 + 2\,H_2O + 30.08\,N_2 + 6\,O_2$

For this constant-volume process, the energy balance $E_{in} - E_{out} = \Delta E_{system}$ applied on the combustion chamber with $Q = W = 0$ reduces to

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

Since both the reactants and products are assumed to be ideal gases, all the internal energy and enthalpies depend on temperature only, and the $P\bar{\boldsymbol{v}}$ terms in this equation can be replaced by R_uT. It yields

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

From the tables,

Substance	$\bar{\mathbf{h}}_\mathbf{f}^\circ$ **kJ/kmol**	$\bar{\mathbf{h}}_\mathbf{298K}$ **kJ/kmol**
CH_4 (*g*)	-74,850	---
O_2	0	8682
N_2	0	8669
H_2O (*g*)	-241,820	9904
CO_2	-393,520	9364

Thus,

$$\begin{aligned}&(1)\left(-393{,}520 + \bar{h}_{CO2} - 9364 - 8.314 \times T_p\right) + (2)\left(-241{,}820 + \bar{h}_{H2O} - 9904 - 8.314 \times T_p\right)\\ &+ (6)\left(0 + \bar{h}_{O2} - 8682 - 8.314 \times T_p\right) + (30.08)\left(0 + \bar{h}_{N2} - 8669 - 8.314 \times T_p\right)\\ &= (1)(-74{,}850 - 8.314 \times 298) + (8)(0 - 8.314 \times 298) + (30.08)(-8.314 \times 298)\end{aligned}$$

It yields $\bar{h}_{CO2} + 2\bar{h}_{H2O} + 6\bar{h}_{O2} + 30.08\bar{h}_{N2} - 324.9T_p = -171{,}674 + 1{,}219{,}188 = 1{,}047{,}514 \text{ kJ}$

The adiabatic flame temperature is obtained from a trial and error solution. A first guess may be obtained by assuming all the products are nitrogen and using nitrogen enthalpy in the above equation. That is,

$$39.08\bar{h}_{N2} - 324.9T_p = 1{,}047{,}514 \text{ kJ}$$

An investigation of Table A-18 shows that this equation is satisfied at a temperature close to 1200 K but it will be somewhat under it because of the higher specific heat of H_2O.

At 1100 K: $(48{,}258) + (2)(40{,}071) + (6)(34{,}899) + (30.08)(33{,}426) - (324.9)(1100) = 985{,}858$

(Lower than 1,047,514 kJ)

At 1200 K: $(53{,}848) + (2)(44{,}380) + (6)(38{,}447) + (30.08)(36{,}777) - (324.9)(1200) = 1{,}089{,}662$

(Higher than 1,047,514 kJ)

By interpolation, T_P = **1159 K**

The volume of reactants when 1 kmol of fuel is burned is

$$\boldsymbol{V} = \boldsymbol{V}_{fuel} + \boldsymbol{V}_{air} = (N_{fuel} + N_{air})\frac{R_uT}{P} = (1 + 38.08)\text{ kmol})\frac{(8.314\text{ kJ/kmol}\cdot\text{K})(298\text{ K})}{101.3\text{ kPa}} = 955.8\text{ m}^3$$

The final pressure is then

$$P = N_{prod}\frac{R_uT}{\boldsymbol{V}} = (39.08\text{ kmol})\frac{(8.314\text{ kJ/kmol}\cdot\text{K})(1159\text{ K})}{955.8\text{ m}^3} = \mathbf{394\ kPa}$$

Entropy Change and Second Law Analysis of Reacting Systems

15-85C Assuming the system exchanges heat with the surroundings at T_0, the increase-in-entropy principle can be expressed as

$$S_{\text{gen}} = \sum N_P \bar{s}_P - \sum N_R \bar{s}_R + \frac{Q_{\text{out}}}{T_0}$$

15-86C By subtracting $R\ln(P/P_0)$ from the tabulated value at 1 atm. Here P is the actual pressure of the substance and P_0 is the atmospheric pressure.

15-87C It represents the reversible work associated with the formation of that compound.

15-88 Hydrogen is burned steadily with oxygen. The reversible work and 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 The combustion equation is

$$2H_2 + 1O_2 \longrightarrow 2H_2O.$$

The H_2, the O_2, and the H_2O are at 25°C and 1 atm, which is the standard reference state and also the state of the surroundings. Therefore, the reversible work in this case is simply the difference between the Gibbs function of formation of the reactants and that of the products,

$$\begin{aligned} W_{\text{rev}} &= \sum N_R \bar{g}_{f,R}^\circ - \sum N_P \bar{g}_{f,P}^\circ = N_{H_2}\overset{\nearrow 0}{\bar{g}_{f,H_2}^\circ} + N_{O_2}\overset{\nearrow 0}{\bar{g}_{f,O_2}^\circ} - N_{H_2O}\bar{g}_{f,H_2O}^\circ = -N_{H_2O}\bar{g}_{f,H_2O}^\circ \\ &= -(2\text{ kmol})(-237{,}180\text{ kJ/kmol}) = \mathbf{474{,}360\ kJ} \quad (\text{for 2 kmol of } H_2) \end{aligned}$$

since the g_f° of stable elements at 25°C and 1 atm is zero. Therefore, 474,360 kJ of work could be done as 2 kmol of H_2 is burned with 1 kmol of O_2 at 25°C and 1 atm in an environment at the same state. The reversible work in this case represents the exergy of the reactants since the product (the H_2O) is at the state of the surroundings.

This process involves no actual work. Therefore, the reversible work and exergy destruction are identical,

$X_{\text{destruction}} =$ **474,360 kJ** (for 2 kmol of H_2)

We could also determine the reversible work without involving the Gibbs function,

$$\begin{aligned} W_{\text{rev}} &= \sum N_R\left(\bar{h}_f^\circ + \bar{h} - \bar{h}^\circ - T_0\bar{s}\right)_R - \sum N_P\left(\bar{h}_f^\circ + \bar{h} - \bar{h}^\circ - T_0\bar{s}\right)_P \\ &= \sum N_R\left(\bar{h}_f^\circ - T_0\bar{s}\right)_R - \sum N_P\left(\bar{h}_f^\circ - T_0\bar{s}\right)_P \\ &= N_{H_2}\left(\bar{h}_f^\circ - T_0\bar{s}^\circ\right)_{H_2} + N_{O_2}\left(\bar{h}_f^\circ - T_0\bar{s}^\circ\right)_{O_2} - N_{H_2O}\left(\bar{h}_f^\circ - T_0\bar{s}^\circ\right)_{H_2O} \end{aligned}$$

Substituting,

$$W_{\text{rev}} = (2)(0 - 298\times130.58) + (1)(0 - 298\times205.03) - (2)(-285{,}830 - 298\times69.92) = 474{,}400\text{ kJ}$$

which is almost identical to the result obtained before.

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.

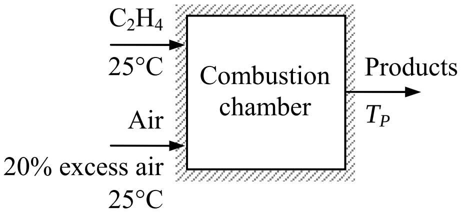
Assumptions **1** Combustion is complete. **2** Steady operating conditions exist. **3** Air and the combustion gases are ideal gases. **4** Changes in kinetic and potential energies are negligible.

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

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

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

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



Thus,

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

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

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

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	$\bar{h}_{298\text{ K}}$ 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,

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

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

By trial and error,

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

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

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

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 $P_i = y_i\, P_{total}$, where y_i is the mole fraction of component *i* . Also,

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

The entropy calculations can be presented in tabular form as

	N_i	y_i	$\bar{s}_i^\circ(T,1\text{atm})$	$R_u\ln(y_i P_m)$	$N_i\bar{s}_i$
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*) $$X_{\text{destroyed}} = T_0 S_{\text{gen}} = (298\text{ K})(1311.28\text{ kJ/kmol}\cdot\text{K C}_2\text{H}_4) = \mathbf{390{,}760\ \ kJ}\ \ (\text{per kmol C}_2\text{H}_4)$$

15-90 Liquid octane is burned steadily with 50 percent excess air. The heat transfer rate from the combustion chamber, the entropy generation rate, and the reversible work and exergy destruction rate 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 C_8H_{18}, the combustion equation can be written as

$$C_8H_{18}(\ell)+1.5a_{\text{th}}(O_2+3.76N_2)\longrightarrow 8CO_2+9H_2O+0.5a_{\text{th}}O_2+(1.5)(3.76)a_{\text{th}}N_2$$

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

$$1.5a_{\text{th}}=8+4.5+0.5a_{\text{th}} \qquad \longrightarrow \qquad a_{\text{th}}=12.5$$

Thus,

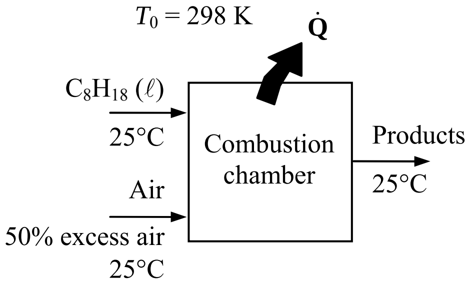
$$C_8H_{18}(\ell)+18.75(O_2+3.76N_2)\longrightarrow 8CO_2+9H_2O+6.25O_2+70.5N_2$$

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

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

since all of the reactants are at 25°C. Assuming the air and the combustion products to be ideal gases, we have $h=h(T)$. From the tables,

Substance	$\bar{\mathbf{h}}_\mathbf{f}^\circ$ **kJ/kmol**
C_8H_{18} (ℓ)	-249,950
O_2	0
N_2	0
H_2O (l)	-285,830
CO_2	-393,520



Substituting,

$$-Q_{\text{out}}=(8)(-393{,}520)+(9)(-285{,}830)+0+0-(1)(-249{,}950)-0-0=-5{,}470{,}680\text{ kJ/kmol of }C_8H_{18}\text{ or}$$

$$Q_{\text{out}}=5{,}470{,}680\text{ kJ/kmol of }C_8H_{18}$$

The C_8H_{18} is burned at a rate of 0.25 kg/min or

$$\dot{N}=\frac{\dot{m}}{M}=\frac{0.25\text{ kg/min}}{[(8)(12)+(18)(1)]\text{ kg/kmol}}=2.193\times10^{-3}\text{ kmol/min}$$

Thus,

$$\dot{Q}_{\text{out}}=\dot{N}Q_{\text{out}}=(2.193\times10^{-3}\text{ kmol/min})(5{,}470{,}680\text{ kJ/kmol})=\mathbf{11{,}997\text{ kJ/min}}$$

The heat transfer for this process is also equivalent to the enthalpy of combustion of liquid C_8H_{18}, which could easily be de determined from Table A-27 to be $\bar{h}_C$ = 5,470,740 kJ/kmol C_8H_{18}.

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

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

The C_8H_{18} is at 25°C and 1 atm, and thus its absolute entropy is $\bar{s}_{C_8H_{18}} = 360.79$ 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 $P_i = y_i P_{total}$, 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

	N_i	y_i	$\bar{s}_i^\circ(T,1atm)$	$R_u \ln(y_i P_m)$	$N_i \bar{s}_i$
C_8H_{18}	1	1.00	360.79	---	360.79
O_2	18.75	0.21	205.14	-12.98	4089.75
N_2	70.50	0.79	191.61	-1.96	13646.69
				S_R =	18,097.23 kJ/K
CO_2	8	0.0944	213.80	-19.62	1867.3
H_2O (ℓ)	9	---	69.92	---	629.3
O_2	6.25	0.0737	205.04	-21.68	1417.6
N_2	70.50	0.8319	191.61	-1.53	13,616.3
				S_P =	17,531 kJ/K

Thus,

$$S_{gen} = S_P - S_R + \frac{Q_{surr}}{T_{surr}} = 17{,}531 - 18{,}097 + \frac{5{,}470{,}523 \text{ kJ}}{298 \text{ K}} = 17{,}798 \text{ kJ/kmol}\cdot\text{K}$$

and

$$\dot{S}_{gen} = \dot{N} S_{gen} = (2.193\times10^{-3} \text{ kmol/min})(17{,}798 \text{ kJ/kmol}\cdot\text{K}) = \mathbf{39.03 \text{ kJ/min}\cdot\text{K}}$$

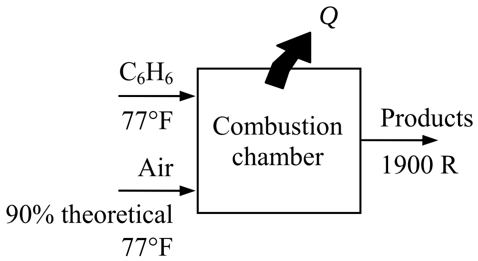
(*c*) The exergy destruction rate associated with this process is determined from

$$\dot{X}_{destroyed} = T_0 \dot{S}_{gen} = (298 \text{ K})(39.03 \text{ kJ/min}\cdot\text{K}) = 11{,}632 \text{ kJ/min} = \mathbf{193.9 \text{ kW}}$$

15-91E Benzene gas is burned steadily with 90 percent theoretical air. The heat transfer rate from the combustion chamber and the exergy destruction are to be determined.

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

Analysis (*a*) The fuel is burned with insufficient amount of air, and thus the products will contain some CO as well as CO_2, H_2O, and N_2. The theoretical combustion equation of C_6H_6 is



$$C_6H_6 + a_{th}(O_2 + 3.76N_2) \longrightarrow 6CO_2 + 3H_2O + 3.76a_{th}N_2$$

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

$$a_{th} = 6 + 1.5 = 7.5$$

Then the actual combustion equation can be written as

$$C_6H_6 + (0.90)(7.5)(O_2 + 3.76N_2) \longrightarrow xCO_2 + (6-x)CO + 3H_2O + 25.38N_2$$

The value of x is determined from an O_2 balance,

$$(0.90)(7.5) = x + (6-x)/2 + 1.5 \longrightarrow x = 4.5$$

Thus,

$$C_6H_6 + 6.75(O_2 + 3.76N_2) \longrightarrow 4.5CO_2 + 1.5CO + 3H_2O + 25.38N_2$$

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

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

since all of the reactants are at 77°F. Assuming the air and the combustion products to be ideal gases, we have $h = h(T)$. From the tables,

Substance	$\bar{h}_f^\circ$ Btu/lbmol	$\bar{h}_{537\ R}$ Btu/lbmol	$\bar{h}_{1900R}$ Btu/lbmol
C_6H_6 (*g*)	35,680	---	---
O_2	0	3725.1	14,322
N_2	0	3729.5	13,742
H_2O (*g*)	-104,040	4258.0	16,428
CO	-47,540	3725.1	13,850
CO_2	-169,300	4027.5	19,698

Thus,

$$\begin{aligned}-Q_{out} &= (4.5)(-169{,}300 + 19{,}698 - 4027.5) + (1.5)(-47{,}540 + 13{,}850 - 3725.1)\\ &\quad + (3)(-104{,}040 + 16{,}428 - 4258) + (25.38)(0 + 13{,}742 - 3729.5) - (1)(35{,}680) - 0 - 0\\ &= \mathbf{-804{,}630\ Btu/lbmol\ of\ C_6H_6}\end{aligned}$$

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

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

The C_6H_6 is at 77°F and 1 atm, and thus its absolute entropy is $\bar{s}_{C_6H_6} = 64.34$ Btu/lbmol·R (Table A-26E). 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 $P_i = y_i P_{\text{total}}$, 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

	N_i	y_i	$\bar{s}_i^\circ(T,1\text{atm})$	$R_u \ln(y_i P_m)$	$N_i \bar{s}_i$
C_6H_6	1	1.00	64.34	---	64.34
O_2	6.75	0.21	49.00	-3.10	351.68
N_2	25.38	0.79	45.77	-0.47	1173.57
					S_R = 1589.59 Btu/R
CO_2	4.5	0.1309	64.999	-4.038	310.67
CO	1.5	0.0436	56.509	-6.222	94.10
H_2O (*g*)	3	0.0873	56.097	-4.843	182.82
N_2	25.38	0.7382	54.896	-0.603	1408.56
					S_P = 1996.15 Btu/R

Thus,

$$S_{\text{gen}} = S_P - S_R + \frac{Q_{\text{out}}}{T_{\text{surr}}} = 1996.15 - 1589.59 + \frac{804{,}630}{537} = 1904.9 \text{ Btu/R}$$

Then the exergy destroyed is determined from

$$X_{\text{destroyed}} = T_0 S_{\text{gen}} = (537 \text{ R})(1904.9 \text{ Btu/lbmol}\cdot\text{R}) = \mathbf{1{,}022{,}950 \text{ Btu/R}} \ (\text{per lbmol } C_6H_6)$$

15-92 Liquid propane is burned steadily with 150 percent excess air. The mass flow rate of air, the heat transfer rate from the combustion chamber, and the rate of entropy generation 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.

Properties The molar masses of C_3H_8 and air are 44 kg/kmol and 29 kg/kmol, respectively (Table A-1).

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_3H_8, the combustion equation can be written as

$$\mathrm{C_3H_8}(\ell)+2.5a_{\mathrm{th}}(\mathrm{O_2}+3.76\mathrm{N_2})\longrightarrow 3\mathrm{CO_2}+4\mathrm{H_2O}+1.5a_{\mathrm{th}}\mathrm{O_2}+(2.5)(3.76)a_{\mathrm{th}}\mathrm{N_2}$$

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

$$2.5a_{\mathrm{th}}=3+2+1.5a_{\mathrm{th}} \quad\longrightarrow\quad a_{\mathrm{th}}=5$$

Substituting,

$$\mathrm{C_3H_8}(\ell)+12.5(\mathrm{O_2}+3.76\mathrm{N_2})\longrightarrow 3\mathrm{CO_2}+4\mathrm{H_2O}+7.5\mathrm{O_2}+47\mathrm{N_2}$$

The air-fuel ratio for this combustion process is

$$\mathrm{AF}=\frac{m_{\mathrm{air}}}{m_{\mathrm{fuel}}}=\frac{(12.5\times 4.76\ \mathrm{kmol})(29\ \mathrm{kg/kmol})}{(3\ \mathrm{kmol})(12\ \mathrm{kg/kmol})+(4\ \mathrm{kmol})(2\ \mathrm{kg/kmol})}=39.2\ \mathrm{kg\ air/kg\ fuel}$$

Thus,

$$\dot m_{\mathrm{air}}=(\mathrm{AF})(\dot m_{\mathrm{fuel}})=(39.2\ \mathrm{kg\ air/kg\ fuel})(0.4\ \mathrm{kg\ fuel/min})=\mathbf{15.7\ kg\ air/min}$$

(*b*) 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, (The $\bar h_f^\circ$ of liquid propane is obtained by adding the h_{fg} at 25°C to $\bar h_f^\circ$ of gaseous propane).

Substance	$\bar{\mathbf{h}}_\mathbf{f}^\circ$ kJ/kmol	$\bar{\mathbf{h}}_{\mathbf{285\ K}}$ kJ/kmol	$\bar{\mathbf{h}}_{\mathbf{298\ K}}$ kJ/kmol	$\bar{\mathbf{h}}_{\mathbf{1200\ K}}$ kJ/kmol
C_3H_8 (ℓ)	-118,910	---	---	---
O_2	0	8296.5	8682	38,447
N_2	0	8286.5	8669	36,777
H_2O (*g*)	-241,820	---	9904	44,380
CO_2	-393,520	---	9364	53,848

Thus,

$$\begin{aligned}-Q_{\mathrm{out}}&=(3)(-393{,}520+53{,}848-9364)+(4)(-241{,}820+44{,}380-9904)\\&\quad+(7.5)(0+38{,}447-8682)+(47)(0+36{,}777-8669)-(1)(-118{,}910+h_{298}-h_{298})\\&\quad-(12.5)(0+8296.5-8682)-(47)(0+8286.5-8669)\\&=-190{,}464\ \mathrm{kJ/kmol\ of\ C_3H_8}\end{aligned}$$

Thus 190,464 kJ of heat is transferred from the combustion chamber for each kmol (44 kg) of propane. This corresponds to 190,464/44 = 4328.7 kJ of heat transfer per kg of propane. Then the rate of heat transfer for a mass flow rate of 0.4 kg/min for the propane becomes

$$\dot{Q}_{\text{out}} = \dot{m}q_{\text{out}} = (0.4\ \text{kg/min})(4328.7\ \text{kJ/kg}) = \mathbf{1732\ kJ/min}$$

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

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

The C_3H_8 is at 25°C and 1 atm, and thus its absolute entropy for the gas phase is $\bar{s}_{C_3H_8} = 269.91$ kJ/kmol·K (Table A-26). Then the entropy of $C_3H_8(\ell)$ is obtained from

$$s_{C_3H_8}(\ell) \cong s_{C_3H_8}(\text{g}) - s_{fg} = s_{C_3H_8}(\text{g}) - \frac{\bar{h}_{fg}}{T} = 269.91 - \frac{15{,}060}{298.15} = 219.4\ \text{kJ/kmol}\cdot\text{K}$$

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 $P_i = y_i P_{\text{total}}$, 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	$\bar{s}_i^\circ(T,1\text{atm})$	$R_u\ln(y_iP_m)$	$N_i\bar{s}_i$
C_3H_8	1	---	219.40	---	219.40
O_2	12.5	0.21	203.70	-12.98	2708.50
N_2	47	0.79	190.18	-1.96	9030.58
				S_R =	11,958.48 kJ/K
CO_2	3	0.0488	279.307	-25.112	913.26
H_2O (*g*)	4	0.0650	240.333	-22.720	1052.21
O_2	7.5	0.1220	249.906	-17.494	2005.50
N_2	47	0.7642	234.115	-2.236	11108.50
				S_P =	15,079.47 kJ/K

Thus,

$$S_{\text{gen}} = S_P - S_R + \frac{Q_{\text{out}}}{T_{\text{surr}}} = 15{,}079.47 - 11{,}958.48 + \frac{190{,}464}{298} = 3760.1\ \text{kJ/K}\ (\text{per kmol}\ C_3H_8)$$

Then the rate of entropy generation becomes

$$\dot{S}_{\text{gen}} = (\dot{N})(S_{\text{gen}}) = \left(\frac{0.4}{44}\ \text{kmol/min}\right)(3760.1\ \text{kJ/kmol}\cdot\text{K}) = \mathbf{34.2\ kJ/min\cdot K}$$

15-93 Problem 15-92 is reconsidered. The effect of the surroundings temperature on the rate of exergy destruction is to be studied.

Analysis The problem is solved using EES, and the solution is given below.

```
Fuel$ = 'Propane (C3H8)_liq'
T_fuel = (25 + 273.15) "[K]"
P_fuel = 101.3 [kPa]
m_dot_fuel = 0.4 [kg/min]*Convert(kg/min, kg/s)
Ex = 1.5 "Excess air"
P_air = 101.3 [kPa]
T_air = (12+273.15) "[K]"
T_prod = 1200 [K]
P_prod = 101.3 [kPa]
Mw_air = 28.97 "lbm/lbmol_air"
Mw_C3H8=(3*12+8*1) "kg/kmol_C3H8"
{TsurrC = 25 [C]}
T_surr = TsurrC+273.15 "[K]"

"For  theoretical dry air, the complete combustion equation is"
"C3H8 + A_th(O2+3.76 N2)=3 CO2+4 H2O +  A_th (3.76) N2 "

2*A_th=3*2+4*1"theoretical O balance"

"The balanced combustion equation with Ex%/100 excess moist air is"
"C3H8 + (1+EX)A_th(O2+3.76 N2)=3 CO2+ 4 H2O + (1+Ex) A_th (3.76) N2+ Ex( A_th) O2 "

"The air-fuel ratio on a mass basis is:"
AF = (1+Ex)*A_th*4.76*Mw_air/(1*Mw_C3H8) "kg_air/kg_fuel"

"The air mass flow rate is:"
m_dot_air = m_dot_fuel * AF

"Apply First Law SSSF to the combustion process per kilomole of fuel:"
E_in - E_out = DELTAE_cv
E_in =HR

"Since EES gives the enthalpy of gasesous components, we adjust the
EES calculated enthalpy to get the liquid enthalpy.  Subtracting the enthalpy
of vaporization from the gaseous enthalpy gives  the enthalpy of the liquid fuel.
h_fuel(liq) = h_fuel(gas) - h_fg_fuel"
h_fg_fuel = 15060 "kJ/kmol from Table A-27"
HR = 1*(enthalpy(C3H8, T=T_fuel) - h_fg_fuel)+ (1+Ex)*A_th*enthalpy(O2,T=T_air)+(1+Ex)*A_th*3.76
*enthalpy(N2,T=T_air)
E_out = HP + Q_out
HP=3*enthalpy(CO2,T=T_prod)+4*enthalpy(H2O,T=T_prod)+(1+Ex)*A_th*3.76*
enthalpy(N2,T=T_prod)+Ex*A_th*enthalpy(O2,T=T_prod)
DELTAE_cv = 0 "Steady-flow requirement"

"The heat transfer rate from the combustion chamber is:"
Q_dot_out=Q_out"kJ/kmol_fuel"/(Mw_C3H8 "kg/kmol_fuel")*m_dot_fuel"kg/s"  "kW"

"Entopy Generation due to the combustion process and heat rejection to the surroundings:"

"Entopy of the reactants per kilomole of fuel:"
P_O2_reac= 1/4.76*P_air  "Dalton's law of partial pressures for O2 in air"
s_O2_reac=entropy(O2,T=T_air,P=P_O2_reac)
P_N2_reac= 3.76/4.76*P_air "Dalton's law of partial pressures for N2 in air"
s_N2_reac=entropy(N2,T=T_air,P=P_N2_reac)
```

```
s_C3H8_reac=entropy(C3H8, T=T_fuel,P=P_fuel) - s_fg_fuel "Adjust the EES gaseous value by s_fg"
"For phase change, s_fg is given by:"
s_fg_fuel = h_fg_fuel/T_fuel
SR = 1*s_C3H8_reac + (1+Ex)*A_th*s_O2_reac + (1+Ex)*A_th*3.76*s_N2_reac

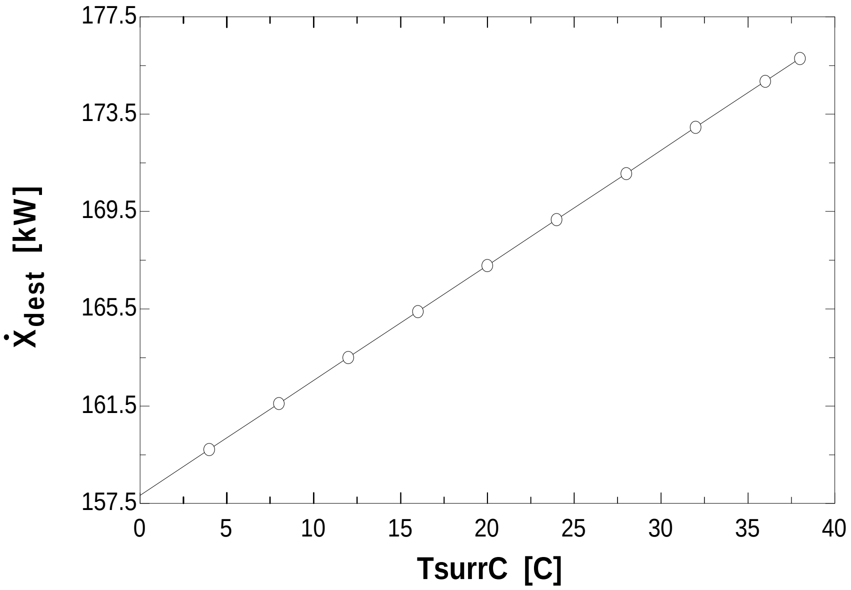
"Entopy of the products per kilomle of fuel:"
"By Dalton's law the partial pressures of the product gases is the product
of the mole fraction and P_prod"
N_prod = 3 + 4 + (1+Ex)*A_th*3.76 + Ex*A_th  "total kmol of products"
P_O2_prod = Ex*A_th/N_prod*P_prod "Patrial pressure O2 in products"
s_O2_prod=entropy(O2,T=T_prod,P=P_O2_prod)
P_N2_prod = (1+Ex)*A_th*3.76/N_prod*P_prod "Patrial pressure N2 in products"
s_N2_prod=entropy(N2,T=T_prod,P=P_N2_prod)
P_CO2_prod = 3/N_prod*P_prod       "Patrial pressure CO2 in products"
s_CO2_prod=entropy(CO2, T=T_prod,P=P_CO2_prod)
P_H2O_prod = 4/N_prod*P_prod       "Patrial pressure H2O in products"
s_H2O_prod=entropy(H2O, T=T_prod,P=P_H2O_prod)

SP = 3*s_CO2_prod + 4*s_H2O_prod + (1+Ex)*A_th*3.76*s_N2_prod + Ex*A_th*s_O2_prod

"Since Q_out is the heat rejected to the surroundings per kilomole fuel,
the entropy of the surroundings is:"
S_surr = Q_out/T_surr

"Rate of entropy generation:"
S_dot_gen = (SP - SR +S_surr)"kJ/kmol_fuel"/(Mw_C3H8 "kg/kmol_fuel")*m_dot_fuel"kg/s"  "kW/K"
X_dot_dest = T_surr*S_dot_gen"[kW]"
```

TsurrC [C]	X_{dest} [kW]
0	157.8
4	159.7
8	161.6
12	163.5
16	165.4
20	167.3
24	169.2
28	171.1
32	173
36	174.9
38	175.8



15-94 Liquid octane is burned steadily with 70 percent excess air. The entropy generation and exergy destruction per unit mass of the fuel are to be determined.

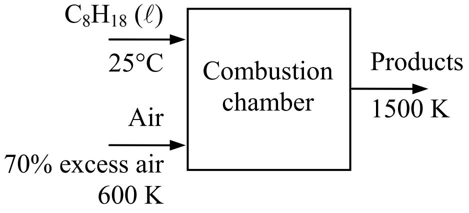
Assumptions **1** Combustion is complete. **2** Steady operating conditions exist. **3** Air and the combustion gases are ideal gases. **4** Changes in kinetic and potential energies are negligible.

Properties The molar masses of C_8H_{18} and air are 114 kg/kmol and 29 kg/kmol, respectively (Table A-1).

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 C_8H_{18}, the combustion equation can be written as

$$\mathrm{C_8H_{18}}(\ell)+1.7a_{\mathrm{th}}(\mathrm{O_2}+3.76\mathrm{N_2})\longrightarrow 8\mathrm{CO_2}+9\mathrm{H_2O}+0.7a_{\mathrm{th}}\mathrm{O_2}+(1.7)(3.76)a_{\mathrm{th}}\mathrm{N_2}$$

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



$$1.7a_{\mathrm{th}}=8+4.5+0.7a_{\mathrm{th}}\longrightarrow a_{\mathrm{th}}=12.5$$

Thus,

$$\mathrm{C_8H_{18}}(\ell)+21.25(\mathrm{O_2}+3.76\mathrm{N_2})\longrightarrow 8\mathrm{CO_2}+9\mathrm{H_2O}+8.75\mathrm{O_2}+79.9\mathrm{N_2}$$

(*b*) 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	$\bar{\mathbf{h}}_\mathbf{f}^\circ$ **kJ/kmol**	$\bar{\mathbf{h}}_{\mathbf{298K}}$ **kJ/kmol**	$\bar{\mathbf{h}}_{\mathbf{600K}}$ **kJ/kmol**	$\bar{\mathbf{h}}_{\mathbf{1500K}}$ **kJ/kmol**
C_8H_{18} (ℓ)	-249,950	---	---	---
O_2	0	8682	17,929	49,292
N_2	0	8669	17,563	47,073
H_2O (*g*)	-241,820	9904	---	57,999
CO_2	-393,520	9364	---	71,078

Thus,

$$\begin{aligned}-Q_{\mathrm{out}}&=(8)(-393{,}520+71{,}078-9364)+(9)(-241{,}820+57{,}999-9904)\\&\quad+(8.75)(0+49{,}292-8682)+(79.9)(0+47{,}073-8669)-(1)(-249{,}950)\\&\quad-(21.25)(0+17{,}929-8682)-(79.9)(0+17{,}563-8669)\\&=-1{,}631{,}335\ \text{kJ/kmol of }\mathrm{C_8H_{18}}\end{aligned}$$

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 P_m = 600 kPa (=600/101.325=5.92 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*. Then,

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

The entropy calculations can be presented in tabular form as

	$\mathbf{N_i}$	$\mathbf{y_i}$	$\bar{s}_i^\circ(\mathbf{T,1atm})$	$\mathbf{R_u \ln(y_i P_m)}$	$\mathbf{N_i \bar{s}_i}$
C_8H_{18}	1	---	466.73	14.79	451.94
O_2	21.25	0.21	226.35	1.81	4771.48
N_2	79.9	0.79	212.07	12.83	15,919.28
					S_R = 21,142.70 kJ/K
CO_2	8	0.0757	292.11	-6.673	2390.26
H_2O (*g*)	9	0.0852	250.45	-5.690	2305.26
O_2	8.75	0.0828	257.97	-5.928	2309.11
N_2	79.9	0.7563	241.77	12.46	18,321.87
					S_P = 25,326.50 kJ/K

Thus,

$$S_{\text{gen}} = S_P - S_R + \frac{Q_{\text{out}}}{T_{\text{surr}}} = 25{,}326.50 - 21{,}142.70 + \frac{1{,}631{,}335}{298} = 9658.1 \text{ kJ/K } \left(\text{per kmol } C_8H_{18}\right)$$

The exergy destruction is

$$X_{\text{dest}} = T_0 S_{\text{gen}} = (298)(9658.1 \text{ kJ/K}) = 2{,}878{,}114 \text{ kJ/K } \left(\text{per kmol } C_8H_{18}\right)$$

The entropy generation and exergy destruction per unit mass of the fuel are

$$S_{\text{gen}} = \frac{\bar{S}_{\text{gen}}}{M_{\text{fuel}}} = \frac{9658.1 \text{ kJ/K}\cdot\text{kmol}}{114 \text{ kg/kmol}} = \mathbf{84.72\ kJ/K\cdot kg\ C_8H_{18}}$$

$$X_{\text{dest}} = \frac{\bar{X}_{\text{dest}}}{M_{\text{fuel}}} = \frac{2{,}878{,}114 \text{ kJ/K}\cdot\text{kmol}}{114 \text{ kg/kmol}} = \mathbf{25{,}250\ kJ/kg\ C_8H_{18}}$$

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.

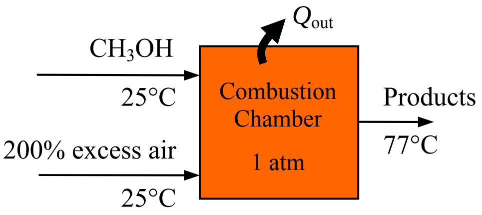
Assumptions **1** Combustion is complete. **2** Steady operating conditions exist. **3** Air and the combustion gases are ideal gases. **4** Changes in kinetic and potential energies are negligible.

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 $P_i = y_i P_{\mathrm{total}}$, 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,

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

Review Problems

15-96 A sample of a certain fluid is burned in a bomb calorimeter. The heating value of the fuel is to be determined.

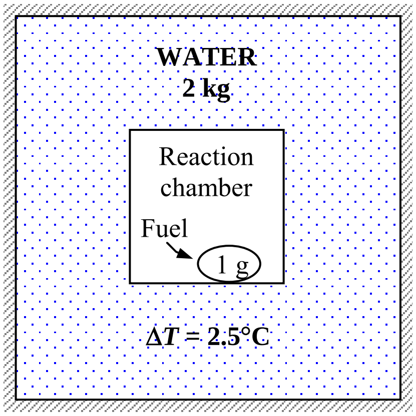
Properties The specific heat of water is 4.18 kJ/kg.°C (Table A-3).

Analysis We take the water as the system, which is a closed system, for which the energy balance on the system $E_{\text{in}} - E_{\text{out}} = \Delta E_{\text{system}}$ with $W = 0$ can be written as

$$Q_{\text{in}} = \Delta U$$

or

$$\begin{aligned} Q_{\text{in}} &= mc\Delta T \\ &= (2\text{ kg})(4.18\text{ kJ/kg}\cdot{}^\circ\text{C})(2.5^\circ\text{C}) \\ &= 20.90\text{ kJ} \ (\text{per gram of fuel}) \end{aligned}$$



Therefore, heat transfer per kg of the fuel would be **20,900 kJ/kg fuel**. Disregarding the slight energy stored in the gases of the combustion chamber, this value corresponds to the heating value of the fuel.

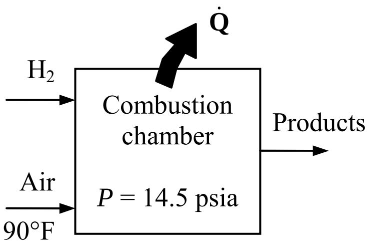
15-97E Hydrogen is burned with 100 percent excess air. The AF ratio and the volume flow rate of air are to be determined.

Assumptions **1** Combustion is complete. **2** Air and the combustion gases are ideal gases.

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

Analysis (*a*) The combustion is complete, and thus products will contain only H_2O, O_2 and N_2. The moisture in the air does not react with anything; it simply shows up as additional H_2O in the products. Therefore, for simplicity, we will balance the combustion equation using dry air, and then add the moisture to both sides of the equation. The combustion equation in this case can be written as

$$H_2 + 2a_{th}(O_2 + 3.76N_2) \longrightarrow H_2O + a_{th}O_2 + (2)(3.76)a_{th}N_2$$



where a_{th} is the stoichiometric coefficient for air. It is determined from

O_2 balance: $\quad 2a_{th} = 0.5 + a_{th} \longrightarrow a_{th} = 0.5$

Substituting, $\quad H_2 + (O_2 + 3.76N_2) \longrightarrow H_2O + 0.5O_2 + 3.76N_2$

Therefore, 4.76 lbmol of dry air will be used per kmol of the fuel. The partial pressure of the water vapor present in the incoming air is

$$P_{v,\text{in}} = \phi_{\text{air}} P_{\text{sat@90°F}} = (0.60)(0.69904\text{ psi}) = 0.419\text{ psia}$$

The number of moles of the moisture that accompanies 4.76 lbmol of incoming dry air ($N_{v,\text{ in}}$) is determined to be

$$N_{v,\text{in}} = \left(\frac{P_{v,\text{in}}}{P_{\text{total}}}\right)N_{\text{total}} = \left(\frac{0.419\text{ psia}}{14.5\text{ psia}}\right)(4.76 + N_{v,\text{in}}) \longrightarrow N_{v,\text{in}} = 0.142\text{ lbmol}$$

The balanced combustion equation is obtained by substituting the coefficients determined earlier and adding 0.142 lbmol of H_2O to both sides of the equation,

$$H_2 + (O_2 + 3.76N_2) + 0.142H_2O \longrightarrow 1.142H_2O + 0.5O_2 + 3.76N_2$$

The air-fuel ratio is determined by taking the ratio of the mass of the air to the mass of the fuel,

$$\text{AF} = \frac{m_{\text{air}}}{m_{\text{fuel}}} = \frac{(4.76\text{ lbmol})(29\text{ lbm/lbmol}) + (0.142\text{ lbmol})(18\text{ lbm/lbmol})}{(1\text{ lbmol})(2\text{ lbm/lbmol})} = \mathbf{70.3\text{ lbm air/lbmfuel}}$$

(*b*) The mass flow rate of H_2 is given to be 10 lbm/h. Since we need 70.3 lbm air per lbm of H_2, the required mass flow rate of air is

$$\dot{m}_{\text{air}} = (\text{AF})(\dot{m}_{\text{fuel}}) = (70.3)(25\text{ lbm/h}) = 1758\text{ lbm/h}$$

The mole fractions of water vapor and the dry air in the incoming air are

$$y_{H_2O} = \frac{N_{H_2O}}{N_{\text{total}}} = \frac{0.142}{4.76 + 0.142} = 0.029 \text{ and } y_{\text{dryair}} = 1 - 0.029 = 0.971$$

Thus,

$$M = (yM)_{H_2O} + (yM)_{\text{dryair}} = (0.029)(18) + (0.971)(29) = 28.7\text{ lbm/lbmol}$$

$$v = \frac{RT}{P} = \frac{(10.73/28.7\text{ psia}\cdot\text{ft}^3/\text{lbm}\cdot\text{R})(550\text{ R})}{14.5\text{ psia}} = 14.18\text{ ft}^3/\text{lbm}$$

$$\dot{V} = \dot{m}v = (1758\text{ lbm/h})(14.18\text{ ft}^3/\text{lbm}) = \mathbf{24{,}928\text{ ft}^3/\text{h}}$$

15-98 A gaseous fuel with a known composition is burned with dry air, and the volumetric analysis of products gases is determined. The AF ratio, the percent theoretical air used, and the volume flow rate of air are to be determined.

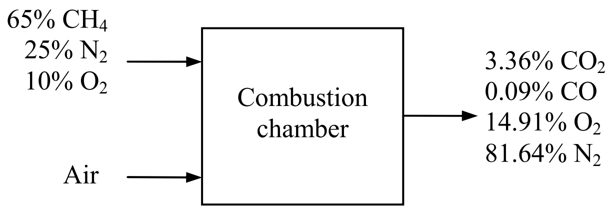
Assumptions **1** Combustion is complete. **2** Air and the combustion gases are ideal gases.

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

Analysis Considering 100 kmol of dry products, the combustion equation can be written as

$$x(0.65\mathrm{CH_4}+0.25\mathrm{N_2}+0.10\mathrm{O_2})+a(\mathrm{O_2}+3.76\mathrm{N_2})$$
$$\longrightarrow 3.36\mathrm{CO_2}+0.09\mathrm{CO}+14.91\mathrm{O_2}+81.64\mathrm{N_2}+b\mathrm{H_2O}$$

The unknown coefficients *x*, *a*, and *b* are determined from mass balances,



$$\mathrm{C}:\ 0.65x=3.36+0.09 \longrightarrow x=5.31$$
$$\mathrm{H}:\ (4)(0.65)x=2b \longrightarrow b=6.90$$
$$\mathrm{N_2}:\ 0.25x+3.76a=81.64 \longrightarrow a=21.36$$

$$\left[\text{Check}\ \ \mathrm{O_2}:\ 0.10x+a=3.36+0.045+14.91+b/2 \longrightarrow a=21.23\right]$$

The N_2 balance and O_2 balance gives two different *a* values. There must be a small error in the volumetric analysis of the products and the mass balance is not completely satisfied. Yet we solve the problem with an *a* value of 21.36 being aware of this situation. Then,

$$5.31(0.65\mathrm{CH_4}+0.25\mathrm{N_2}+0.10\mathrm{O_2})+21.36(\mathrm{O_2}+3.76\mathrm{N_2})$$
$$\longrightarrow 3.36\mathrm{CO_2}+0.09\mathrm{CO}+14.91\mathrm{O_2}+81.64\mathrm{N_2}+6.9\mathrm{H_2O}$$

The combustion equation for 1 kmol of fuel is obtained by dividing the above equation by 5.31,

$$(0.65\mathrm{CH_4}+0.25\mathrm{N_2}+0.10\mathrm{O_2})+4.02(\mathrm{O_2}+3.76\mathrm{N_2})$$
$$\longrightarrow 0.633\mathrm{CO_2}+0.017\mathrm{CO}+2.81\mathrm{O_2}+15.37\mathrm{N_2}+1.3\mathrm{H_2O}$$

(*a*) The air-fuel ratio is determined from its definition,

$$\mathrm{AF}=\frac{m_{\text{air}}}{m_{\text{fuel}}}=\frac{(4.02\times 4.76\ \text{kmol})(29\ \text{kg/kmol})}{0.65\times 16+0.25\times 28+0.10\times 32}=\mathbf{26.9\ \ kg\ air/kg\ fuel}$$

(*b*) To find the percent theoretical air used, we need to know the theoretical amount of air, which is determined from the theoretical combustion equation of the fuel,

$$(0.65\mathrm{CH_4}+0.25\mathrm{N_2}+0.10\mathrm{O_2})+a_{\text{th}}(\mathrm{O_2}+3.76\mathrm{N_2})\longrightarrow 0.65\mathrm{CO_2}+1.3\mathrm{H_2O}+(0.25+3.76a_{\text{th}})\mathrm{N_2}$$

$$\mathrm{O_2}:\ 0.10+a_{\text{th}}=0.65+0.65 \longrightarrow a_{\text{th}}=1.2$$

Then, $$\text{Percent theoretical air}=\frac{m_{\text{air,act}}}{m_{\text{air,th}}}=\frac{N_{\text{air,act}}}{N_{\text{air,th}}}=\frac{(4.02)(4.76)\ \text{kmol}}{(1.2)(4.76)\ \text{kmol}}=3.35=\mathbf{335\%}$$

(*c*) The specific volume, mass flow rate, and the volume flow rate of air at the inlet conditions are

$$v=\frac{RT}{P}=\frac{(0.287\ \text{kPa}\cdot\text{m}^3/\text{kg}\cdot\text{K})(298\ \text{K})}{100\ \text{kPa}}=0.855\ \text{m}^3/\text{kg}$$

$$\dot m_{\text{air}}=(\text{AF})\dot m_{\text{fuel}}=(26.9\ \text{kg air/kg fuel})(3.5\ \text{kg fuel/min})=94.15\ \text{m}^3/\text{min}$$

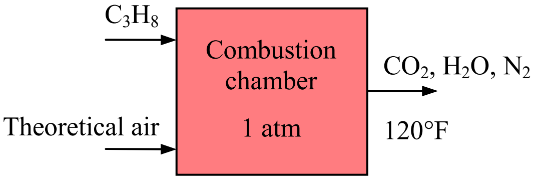
$$\dot V_{\text{air}}=(\dot m v)_{\text{air}}=(94.15\ \text{kg/min})(0.855\ \text{m}^3/\text{kg})=\mathbf{80.5\ m^3/min}$$

15-99E Propane is burned with stoichiometric amount of air. The fraction of the water in the products that is vapor is to be determined.

Assumptions **1** Combustion is complete. **2** Steady operating conditions exist. **3** Air and the combustion gases are ideal gases.

Analysis The fuel is burned completely with the air, and thus the products will contain only CO_2, H_2O, and N_2. Considering 1 kmol C_3H_8, the combustion equation can be written as

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



The mole fraction of water in the products is

$$y = \frac{N_{\mathrm{H2O}}}{N_{\mathrm{prod}}} = \frac{4\ \mathrm{kmol}}{(3+4+18.8)\mathrm{kmol}} = 0.1550$$

The saturation pressure for the water vapor is

$$P_v = P_{\mathrm{sat@120°F}} = 1.6951\ \mathrm{psia}$$

When the combustion gases are saturated, the mole fraction of the water vapor will be

$$y_g = \frac{P_v}{P} = \frac{1.6951\ \mathrm{kPa}}{14.696\ \mathrm{kPa}} = 0.1153$$

Thus, the fraction of water vapor in the combustion products is

$$f_{\mathrm{vapor}} = \frac{y_g}{y} = \frac{0.1153}{0.1550} = \mathbf{0.744}$$

15-100 Coal whose mass percentages are specified is burned with 20% excess air. The dew-point temperature of the products is to be determined.

Assumptions **1** Combustion is complete. **2** The combustion products contain CO_2, CO, H_2O, SO_2, and N_2. **3** Combustion gases are ideal gases.

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

Analysis We consider 100 kg of coal for simplicity. Noting that the mass percentages in this case correspond to the masses of the constituents, the mole numbers of the constituent of the coal are determined to be

61.40% C
5.79% H_2
25.31% O_2
1.09% N_2
1.41% S
5.00% ash
(by mass)

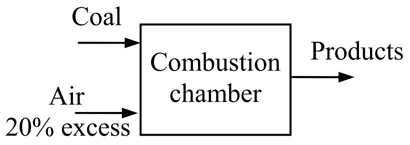
$$N_C = \frac{m_C}{M_C} = \frac{61.40\text{ kg}}{12\text{ kg/kmol}} = 5.117\text{ kmol}$$

$$N_{H2} = \frac{m_{H2}}{M_{H2}} = \frac{5.79\text{ kg}}{2\text{ kg/kmol}} = 2.895\text{ kmol}$$

$$N_{O2} = \frac{m_{O2}}{M_{O2}} = \frac{25.31\text{ kg}}{32\text{ kg/kmol}} = 0.7909\text{ kmol}$$

$$N_{N2} = \frac{m_{N2}}{M_{N2}} = \frac{1.09\text{ kg}}{28\text{ kg/kmol}} = 0.03893\text{ kmol}$$

$$N_S = \frac{m_S}{M_S} = \frac{1.41\text{ kg}}{32\text{ kg/kmol}} = 0.04406\text{ kmol}$$



The mole number of the mixture and the mole fractions are

$$N_m = 5.117 + 2.895 + 0.7909 + 0.03893 + 0.04406 = 8.886\text{ kmol}$$

$$y_C = \frac{N_C}{N_m} = \frac{5.117\text{ kmol}}{8.886\text{ kmol}} = 0.5758$$

$$y_{H2} = \frac{N_{H2}}{N_m} = \frac{2.895\text{ kmol}}{8.886\text{ kmol}} = 0.3258$$

$$y_{O2} = \frac{N_{O2}}{N_m} = \frac{0.7909\text{ kmol}}{8.886\text{ kmol}} = 0.0890$$

$$y_{N2} = \frac{N_{N2}}{N_m} = \frac{0.03893\text{ kmol}}{8.886\text{ kmol}} = 0.00438$$

$$y_S = \frac{N_S}{N_m} = \frac{0.04406\text{ kmol}}{8.886\text{ kmol}} = 0.00496$$

Ash consists of the non-combustible matter in coal. Therefore, the mass of ash content that enters the combustion chamber is equal to the mass content that leaves. Disregarding this non-reacting component for simplicity, the combustion equation may be written as

$$0.5758\text{C} + 0.3258\text{H}_2 + 0.0890\text{O}_2 + 0.00438\text{N}_2 + 0.00496\text{S} + 1.25a_{\text{th}}(\text{O}_2 + 3.76\text{N}_2)$$
$$\longrightarrow 0.5758\text{CO}_2 + 0.3258\text{H}_2\text{O} + 0.00496\text{SO}_2 + 0.25a_{\text{th}}\text{O}_2 + 1.25a_{\text{th}} \times 3.76\text{N}_2$$

According to the oxygen balance,

$$\text{O}_2\text{ balance: }\ 0.0890 + 1.25a_{\text{th}} = 0.5758 + 0.5 \times 0.3258 + 0.00496 + 0.25a_{\text{th}} \longrightarrow a_{\text{th}} = 0.6547$$

Substituting,

$$0.5758\text{C} + 0.3258\text{H}_2 + 0.0890\text{O}_2 + 0.00438\text{N}_2 + 0.00496\text{S} + 0.8184(\text{O}_2 + 3.76\text{N}_2)$$
$$\longrightarrow 0.5758\text{CO}_2 + 0.3258\text{H}_2\text{O} + 0.00496\text{SO}_2 + 0.1637\text{O}_2 + 3.077\text{N}_2$$

The dew-point temperature of a gas-vapor mixture is the saturation temperature of the water vapor in the product gases corresponding to its partial pressure. That is,

$$P_v = \left(\frac{N_v}{N_{\text{prod}}}\right)P_{\text{prod}} = \left(\frac{0.3258\text{ kmol}}{(0.5758 + 0.3258 + 0.00496 + 0.1637 + 3.077)\text{ kmol}}\right)(101.3\text{ kPa}) = 7.96\text{ kPa}$$

Thus, $T_{\text{dp}} = T_{\text{sat@7.96 kPa}} = \mathbf{41.3°C}$ (Table A-5)

15-101 Methane is burned steadily with 50 percent excess air. The dew-point temperature of the water vapor in the products is to be determined.

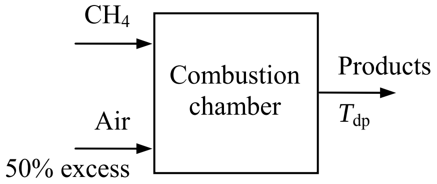
Assumptions **1** Combustion is complete. **2** Steady operating conditions exist. **3** Air and the combustion gases are ideal gases.

Properties The molar masses of CH_4 and air are 16 kg/kmol and 29 kg/kmol, respectively (Table A-1).

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_4, the combustion equation can be written as

$$\mathrm{CH_4} + 1.5a_{\mathrm{th}}(\mathrm{O_2} + 3.76\mathrm{N_2}) \longrightarrow \mathrm{CO_2} + 2\mathrm{H_2O} + 0.5a_{\mathrm{th}}\mathrm{O_2} + (1.5)(3.76)a_{\mathrm{th}}\mathrm{N_2}$$

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



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

Thus,

$$\mathrm{CH_4} + 3(\mathrm{O_2} + 3.76\mathrm{N_2}) \longrightarrow \mathrm{CO_2} + 2\mathrm{H_2O} + \mathrm{O_2} + 11.28\mathrm{N_2}$$

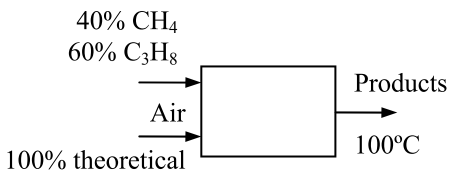
The dew-point temperature of a gas-vapor mixture is the saturation temperature of the water vapor in the product gases corresponding to its partial pressure. That is,

$$P_v = \left(\frac{N_v}{N_{\mathrm{prod}}}\right)P_{\mathrm{prod}} = \left(\frac{2\ \mathrm{kmol}}{(1+2+1+11.28)\ \mathrm{kmol}}\right)(101.325\ \mathrm{kPa}) = 13.26\ \mathrm{kPa}$$

Thus,

$$T_{\mathrm{dp}} = T_{\mathrm{sat@13.26\,kPa}} = \mathbf{51.4^\circ C} \qquad \text{(from EES)}$$

15-102 A mixture of 40% by volume methane, CH4, and 60% by volume propane, C3H8, is burned completely with theoretical air. The amount of water formed during combustion process that will be condensed is to be determined.



Assumptions **1** Combustion is complete. **2** The combustion products contain CO_2, H_2O, 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 combustion equation in this case can be written as

$$0.4\,CH_4 + 0.6\,C_3H_8 + a_{th}\left[O_2 + 3.76N_2\right] \longrightarrow B\,CO_2 + D\,H_2O + F\,N_2$$

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: $B = 0.4 + 3 \times 0.6 = 2.2$

Hydrogen balance: $2D = 4 \times 0.4 + 8 \times 0.6 = 2D \longrightarrow D = 3.2$

Oxygen balance: $2a_{th} = 2B + D \longrightarrow 2a_{th} = 2(2.2) + 3.2 \longrightarrow a_{th} = 3.8$

Nitrogen balance: $3.76a_{th} = F \longrightarrow 3.76(3.8) = F \longrightarrow F = 14.29$

Then, we write the balanced reaction equation as

$$0.4\,CH_4 + 0.6\,C_3H_8 + 3.8\left[O_2 + 3.76N_2\right] \longrightarrow 2.2\,CO_2 + 3.2\,H_2O + 14.29\,N_2$$

The vapor mole fraction in the products is

$$y_v = \frac{3.2}{2.2 + 3.2 + 14.29} = 0.1625$$

The partial pressure of water in the products is

$$P_{\text{v,prod}} = y_v P_{\text{prod}} = (0.1625)(100\text{ kPa}) = 16.25\text{ kPa}$$

The dew point temperature of the products is

$$T_{\text{dp}} = T_{\text{sat@16.25 kPa}} = 55.64°\text{C}$$

The partial pressure of the water vapor remaining in the products at the product temperature is

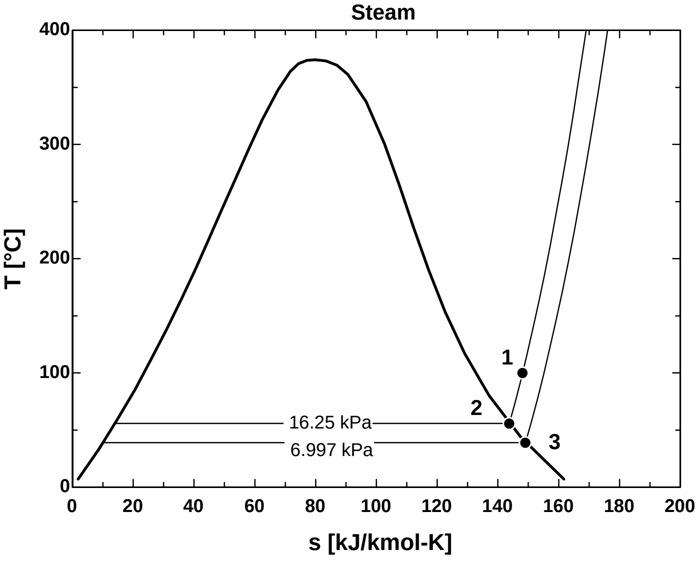
$$P_{\text{v}} = P_{\text{sat @ 39°C}} = 7.0\text{ kPa}$$

The kmol of water vapor in the products at the product temperature is

$$P_v = \frac{N_v}{N_{\text{total,product}}} P_{\text{prod}}$$

$$7.0\text{ kPa} = \frac{N_v}{2.2 + N_v + 14.29}$$

$$N_v = 1.241\text{ kmol}$$



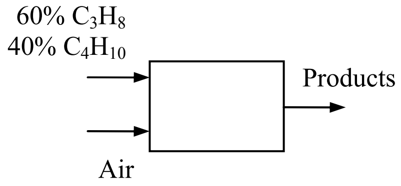
The kmol of water condensed is

$$N_w = 3.2 - 1.241 = \mathbf{1.96\ kmol\ water/kmol\ fuel}$$

15-103 A gaseous fuel mixture of 60% propane, C_3H_8, and 40% butane, C_4H_{10}, on a volume basis is burned with an air-fuel ratio of 25. The moles of nitrogen in the air supplied to the combustion process, the moles of water formed in the combustion process, and the moles of oxygen in the product gases are to be determined.

Assumptions **1** Combustion is complete. **2** The combustion products contain CO_2, H_2O, 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 theoretical combustion equation in this case can be written as

$$0.6\,C_3H_8 + 0.4\,C_4H_{10} + a_{th}\left[O_2 + 3.76N_2\right] \longrightarrow B\,CO_2 + D\,H_2O + F\,N_2$$

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: $B = 3 \times 0.6 + 4 \times 0.4 = 3.4$

Hydrogen balance: $8 \times 0.6 + 10 \times 0.4 = 2D \longrightarrow D = 4.4$

Oxygen balance: $2a_{th} = 2B + D \longrightarrow 2a_{th} = 2 \times 3.4 + 4.4 \longrightarrow a_{th} = 5.6$

Nitrogen balance: $3.76a_{th} = F \longrightarrow 3.76 \times 5.6 = F \longrightarrow F = 21.06$

Then, we write the balanced theoretical reaction equation as

$$0.6\,C_3H_8 + 0.4\,C_4H_{10} + 5.6\left[O_2 + 3.76N_2\right] \longrightarrow 3.4\,CO_2 + 4.4\,H_2O + 21.06\,N_2$$

The air-fuel ratio for the theoretical reaction is determined from

$$\text{AF}_{\text{th}} = \frac{m_{\text{air}}}{m_{\text{fuel}}} = \frac{(5.6 \times 4.76\text{ kmol})(29\text{ kg/kmol})}{(0.6 \times 44 + 0.4 \times 58)\text{ kg}} = 15.59\text{ kg air/kg fuel}$$

The percent theoretical air is

$$\text{PercentTH}_{\text{air}} = \frac{\text{AF}_{\text{actual}}}{\text{AF}_{\text{th}}} = \frac{25}{15.59} \times 100 = 160.4\%$$

The moles of nitrogen supplied is

$$N_{\text{N2}} = \frac{\text{PercentTH}_{\text{air}}}{100} \times a_{th} \times 3.76 = \frac{160.4}{100}(5.6)(3.76) = \mathbf{33.8\text{ kmol}}\text{ per kmol fuel}$$

The moles of water formed in the combustion process is

$$N_{\text{H2O}} = D = \mathbf{4.4\text{ kmol}}\text{ per kmol fuel}$$

The moles of oxygen in the product gases is

$$N_{\text{O2}} = \left(\frac{\text{PercentTH}_{\text{air}}}{100} - 1\right)a_{th} = \left(\frac{160.4}{100} - 1\right)(5.6) = \mathbf{3.38\text{ kmol}}\text{ per kmol fuel}$$

15-104 Ethane is completely burned with air. Various parameters are to be determined for the given reaction.

Assumptions The water in the products is in the vapor phase.

Analysis (a) The reaction equation is given as

$$C_2H_6 + 4.788[O_2 + 3.76N_2] \longrightarrow 2CO_2 + 3H_2O + 1.288O_2 + 18N_2$$

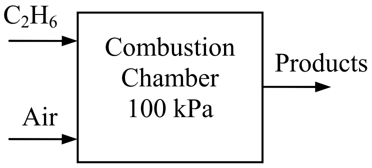
The partial pressure of water vapor is

$$P_v = \frac{N_{\text{H2O}}}{N_{\text{total}}} P_{\text{total}} = \frac{3}{2+3+1.288+18}(100\text{ kPa}) = \frac{3\text{ kmol}}{24.288\text{ kmol}}(100\text{ kPa}) = 12.35\text{ kPa}$$

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

$$T_{dp} = T_{\text{sat@12.35 kPa}} = \mathbf{49.6°C} \quad \text{(Table A-5)}$$

(b) The partial pressure of oxygen is



$$P_{\text{O2}} = \frac{N_{\text{O2}}}{N_{\text{total}}} P_{\text{total}} = \frac{1.288\text{ kmol}}{24.288\text{ kmol}}(100\text{ kPa}) = 5.303\text{ kPa}$$

The specific volume of oxygen is then,

$$\boldsymbol{v}_{\text{O2}} = \frac{R_{\text{O2}}T}{P_{\text{O2}}} = \frac{(0.2598\text{ kJ/kmol}\cdot\text{K})(373\text{ K})}{5.303\text{ kPa}} = \mathbf{18.3\ m^3/kg}$$

(c) The combustion reaction with stoichiometric air is

$$C_2H_6 + 3.5(O_2 + 3.76N_2) \longrightarrow 2CO_2 + 3H_2O + 3.5\times 3.76N_2$$

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,

$$q = h_C = H_P - H_R = \sum N_P \bar{h}^\circ_{f,P} - \sum N_R \bar{h}^\circ_{f,R} = \left(N\bar{h}^\circ_f\right)_{\text{CO2}} + \left(N\bar{h}^\circ_f\right)_{\text{H2O}} - \left(N\bar{h}^\circ_f\right)_{\text{C2H6}}$$

For the LHV, the water in the products is taken to be vapor. Then,

$$\begin{aligned} h_C &= (2\text{ kmol})(-393{,}520\text{ kJ/kmol}) + (3\text{ kmol})(-241{,}820\text{ kJ/kmol}) - (1\text{ kmol})(-84{,}680\text{ kJ/kmol}) \\ &= -1{,}427{,}820\text{ kJ/kmol ethane} \end{aligned}$$

The LHV per unit kmol of the fuel is the negative of the enthalpy of combustion:

$$\text{LHV} = -h_C = \mathbf{1{,}427{,}820\ kJ/kmol\ C_2H_6}$$

(d) The average molar mass of the product gas is

$$\begin{aligned} M &= \frac{N_{\text{CO2}}M_{\text{CO2}} + N_{\text{H2O}}M_{\text{H2O}} + N_{\text{O2}}M_{\text{O2}} + N_{\text{N2}}M_{\text{N2}}}{N_{\text{total}}} \\ &= \frac{(2\text{ kmol})(44\text{ kg/kmol}) + (3\text{ kmol})(18\text{ kg/kmol}) + (1.288\text{ kmol})(32\text{ kg/kmol}) + (18\text{ kmol})(28\text{ kg/kmol})}{24.288\text{ kmol}} \\ &= \frac{687.2\text{ kmol}}{24.288\text{ kmol}} = \mathbf{28.29\ kg/kmol} \end{aligned}$$

(e) The average molar constant pressure specific heat of the product gas is

$$\begin{aligned} \bar{c}_p &= \frac{N_{\text{CO2}}\bar{c}_{p,\text{CO2}} + N_{\text{H2O}}\bar{c}_{p,\text{H2O}} + N_{\text{O2}}\bar{c}_{p,\text{O2}} + N_{\text{N2}}\bar{c}_{p,\text{N2}}}{N_{\text{total}}} \\ &= \frac{(2\times 41.16 + 3\times 34.28 + 1.288\times 30.14 + 18\times 29.27)\text{ kJ/K}}{24.288\text{ kmol}} = \frac{750.8\text{ kJ/K}}{24.288\text{ kmol}} = \mathbf{30.91\ kJ/kmol\cdot K} \end{aligned}$$

(f) The air-fuel mass ratio is

$$\text{AF} = \frac{m_{\text{air}}}{m_{\text{fuel}}} = \frac{(4.788\times 4.76\times 29)\text{ kg}}{(1\times 30)\text{ kg}} = \frac{660.9\text{ kg}}{30\text{ kg}} = \mathbf{22.03\ kg\ air/kg\ fuel}$$

(g) For a molar fuel flow rate is 0.1 kmol/min, the mass flow rate of water in the product gases is

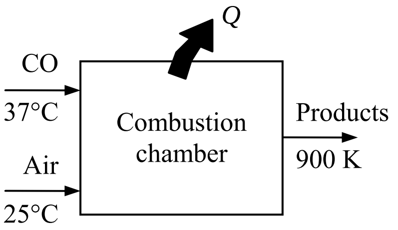
$$\dot{m}_{\text{H2O}} = \dot{N}_{\text{C2H6}} \frac{N_{\text{H2O}}}{N_{\text{C2H6}}} M_{\text{H2O}} = (0.1\text{ kmol/min})\frac{3\text{ kmol}}{1\text{ kmol}}(18\text{ kg/kmol}) = \mathbf{5.4\ kg/min}$$

15-105 CO gas is burned with air during a steady-flow combustion process. The rate of heat transfer from the combustion chamber is to be determined.

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** Combustion is complete.

Properties The molar masses of CO and air are 28 kg/kmol and 29 kg/kmol, respectively (Table A-1).

Analysis We first need to calculate the amount of air used per kmol of CO before we can write the combustion equation,



$$v_{\mathrm{CO}} = \frac{RT}{P} = \frac{(0.2968\ \mathrm{kPa\cdot m^3/kg\cdot K})(310\ \mathrm{K})}{(110\ \mathrm{kPa})} = 0.836\ \mathrm{m^3/kg}$$

$$\dot{m}_{\mathrm{CO}} = \frac{\dot{V}_{\mathrm{CO}}}{v_{\mathrm{CO}}} = \frac{0.4\ \mathrm{m^3/min}}{0.836\ \mathrm{m^3/kg}} = 0.478\ \mathrm{kg/min}$$

Then the molar air-fuel ratio becomes

$$\overline{\mathrm{AF}} = \frac{N_{\mathrm{air}}}{N_{\mathrm{fuel}}} = \frac{\dot{m}_{\mathrm{air}} / M_{\mathrm{air}}}{\dot{m}_{\mathrm{fuel}} / M_{\mathrm{fuel}}} = \frac{(1.5\ \mathrm{kg/min})/(29\ \mathrm{kg/kmol})}{(0.478\ \mathrm{kg/min})/(28\ \mathrm{kg/kmol})} = 3.03\ \mathrm{kmol\ air/kmol\ fuel}$$

Thus the number of moles of O_2 used per mole of CO is 3.03/4.76 = 0.637. Then the combustion equation in this case can be written as

$$\mathrm{CO} + 0.637(\mathrm{O_2} + 3.76\mathrm{N_2}) \longrightarrow \mathrm{CO_2} + 0.137\mathrm{O_2} + 2.40\mathrm{N_2}$$

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**	$\bar{h}_{310\ \mathrm{K}}$ **kJ/kmol**	$\bar{h}_{900\ \mathrm{K}}$ **kJ/kmol**
CO	-110,530	8669	9014	27,066
O_2	0	8682	---	27,928
N_2	0	8669	---	26,890
CO_2	-393,520	9364	---	37,405

Thus,

$$\begin{aligned} -Q_{\mathrm{out}} &= (1)(-393{,}520 + 37{,}405 - 9364) + (0.137)(0 + 27{,}928 - 8682) \\ &\quad + (2.4)(0 + 26{,}890 - 8669) - (1)(-110{,}530 + 9014 - 8669) - 0 - 0 \\ &= -208{,}927\ \mathrm{kJ/kmol\ of\ CO} \end{aligned}$$

Then the rate of heat transfer for a mass flow rate of 0.956 kg/min for CO becomes

$$\dot{Q}_{\mathrm{out}} = \dot{N}Q_{\mathrm{out}} = \left(\frac{\dot{m}}{N}\right)Q_{\mathrm{out}} = \left(\frac{0.478\ \mathrm{kg/min}}{28\ \mathrm{kg/kmol}}\right)(208{,}927\ \mathrm{kJ/kmol}) = \mathbf{3567\ kJ/min}$$

15-106 Ethanol gas is burned with 10% excess air. The combustion is incomplete. The theoretical kmols of oxygen in the reactants, the balanced chemical reaction, and the rate of heat transfer are to be determined.

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

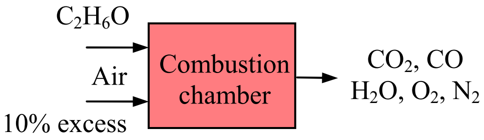
Properties The molar masses of C, H_2, O_2, N_2 and air are 12 kg/kmol, 2 kg/kmol, 32 kg/kmol, 28 kg/kmol, and 29 kg/kmol, respectively (Table A-1).

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

$$C_2H_6O + a_{th}[O_2 + 3.76N_2] \longrightarrow 2\,CO_2 + 3\,H_2O + a_{th} \times 3.76\,N_2$$

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

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



Substituting,

$$C_2H_6O + 3[O_2 + 3.76N_2] \longrightarrow 2\,CO_2 + 3\,H_2O + 11.28\,N_2$$

Therefore, **3 kmol of oxygen** is required to burn 1 kmol of ethanol.

(b) The reaction with 10% excess air and incomplete combustion can be written as

$$C_2H_6O + 1.1 \times 3[O_2 + 3.76N_2] \longrightarrow 2(0.9\,CO_2 + 0.1\,CO) + 3\,H_2O + x\,O_2 + 1.1 \times 3 \times 3.76\,N_2$$

The coefficient for O_2 is determined from a mass balance,

O_2 balance: $\quad 0.5 + 1.1 \times 3 = 0.9 \times 2 + 0.5 \times (0.1 \times 2) + 0.5 \times 3 + x \longrightarrow x = 0.4$

Substituting,

$$C_2H_6O + 3.3[O_2 + 3.76N_2] \longrightarrow 1.8\,CO_2 + 0.2\,CO + 3\,H_2O + 0.4\,O_2 + 12.408\,N_2$$

(b) 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

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

Both the reactants and products are at 25 °C. Assuming the air and the combustion products to be ideal gases, we have $h = h(T)$. Then, using the values given in the table,

$$\begin{aligned} -Q_{out} &= (1.8)(-393{,}520) + (0.2)(-110{,}530) + (3)(-241{,}820) - (1)(-235{,}310) \\ &= -1{,}220{,}590 \text{ kJ/kmol fuel} \end{aligned}$$

or

$$Q_{out} = 1{,}220{,}590 \text{ kJ/kmol fuel}$$

For a 3.5 kg/h of fuel burned, the rate of heat transfer is

$$\dot{Q}_{out} = \dot{N}Q_{out} = \left(\frac{\dot{m}}{M}\right)Q_{out} = \frac{3.5 \text{ kg/h}}{46 \text{ kg/kmol}}(1{,}220{,}590 \text{ kJ/kmol}) = 92{,}870 \text{ kJ/h} = \mathbf{25.80\ kW}$$

15-107 Propane gas is burned with air during a steady-flow combustion process. The adiabatic flame temperature is to be determined for different 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.

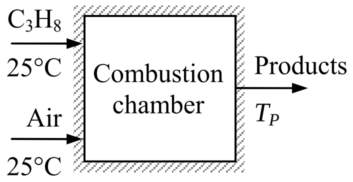
Analysis Adiabatic flame temperature is the temperature at which the products leave the combustion chamber under adiabatic conditions ($Q = 0$) with no work interactions ($W = 0$). Under steady-flow conditions the energy balance $E_{\text{in}} - E_{\text{out}} = \Delta E_{\text{system}}$ applied on the combustion chamber reduces to

$$\sum N_P\left(\bar{h}_f^\circ + \bar{h} - \bar{h}^\circ\right)_P = \sum N_R\left(\bar{h}_f^\circ + \bar{h} - \bar{h}^\circ\right)_R \longrightarrow \sum N_P\left(\bar{h}_f^\circ + \bar{h}_T - \bar{h}^\circ\right)_P = \left(N\bar{h}_f^\circ\right)_{\text{C}_3\text{H}_8}$$

since all the reactants are at the standard reference temperature of 25°C, and $\bar{h}_f^\circ = 0$ for O_2 and N_2.

(*a*) The theoretical combustion equation of C_3H_8 with stoichiometric amount of air is

$$\text{C}_3\text{H}_8(g) + 5(\text{O}_2 + 3.76\text{N}_2) \longrightarrow 3\text{CO}_2 + 4\text{H}_2\text{O} + 18.8\text{N}_2$$



From the tables,

Substance	$\bar{\mathbf{h}}_\mathbf{f}^\circ$ **kJ/kmol**	$\bar{\mathbf{h}}_{\mathbf{298\ K}}$ **kJ/kmol**
C_3H_8 (*g*)	-103,850	---
O_2	0	8682
N_2	0	8669
H_2O (*g*)	-241,820	9904
CO	-110,530	8669
CO_2	-393,520	9364

Thus,

$$(3)\left(-393{,}520 + \bar{h}_{\text{CO}_2} - 9364\right) + (4)\left(-241{,}820 + \bar{h}_{\text{H}_2\text{O}} - 9904\right) + (18.8)\left(0 + \bar{h}_{\text{N}_2} - 8669\right) = (1)(-103{,}850)$$

It yields

$$3\bar{h}_{\text{CO}_2} + 4\bar{h}_{\text{H}_2\text{O}} + 18.8\bar{h}_{\text{N}_2} = 2{,}274{,}675 \text{ kJ}$$

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 2,274,675 / (3 + 4 + 18.8) = 88,165 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 heats of CO_2 and H_2O.

At 2400 K: $$3\bar{h}_{\text{CO}_2} + 4\bar{h}_{\text{H}_2\text{O}} + 18.8\bar{h}_{\text{N}_2} = (3)(125{,}152) + (4)(103{,}508) + (18.8)(79{,}320)$$
$$= 2{,}280{,}704 \text{ kJ (Higher than 2,274,675 kJ)}$$

At 2350 K: $$3\bar{h}_{\text{CO}_2} + 4\bar{h}_{\text{H}_2\text{O}} + 18.8\bar{h}_{\text{N}_2} = (3)(122{,}091) + (4)(100{,}846) + (18.8)(77{,}496)$$
$$= 2{,}226{,}582 \text{ kJ (Lower than 2,274,675 kJ)}$$

By interpolation, T_P = **2394 K**

(*b*) The balanced combustion equation for complete combustion with 200% theoretical air is

$$\text{C}_3\text{H}_8(g) + 10(\text{O}_2 + 3.76\text{N}_2) \longrightarrow 3\text{CO}_2 + 4\text{H}_2\text{O} + 5\text{O}_2 + 37.6\text{N}_2$$

Substituting known numerical values,

$$(3)(-393{,}520+\bar{h}_{CO_2}-9364)+(4)(-241{,}820+\bar{h}_{H_2O}-9904)$$
$$+(5)(0+\bar{h}_{O_2}-8682)+(37.6)(0+\bar{h}_{N_2}-8669)=(1)(-103{,}850)$$

which yields

$$3\bar{h}_{CO_2}+4\bar{h}_{H_2O}+5\bar{h}_{O_2}+37.6\bar{h}_{N_2}=2{,}481{,}060\text{ kJ}$$

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 2,481,060 / (3 + 4 + 5 + 37.6) = 50,021 kJ/kmol. This enthalpy value corresponds to about 1580 K for N_2. Noting that the majority of the moles are N_2, T_P will be close to 1580 K, but somewhat under it because of the higher specific heats of CO_2 and H_2O.

At 1540 K: $$3\bar{h}_{CO_2}+4\bar{h}_{H_2O}+5\bar{h}_{O_2}+37.6\bar{h}_{N_2}=(3)(73{,}417)+(4)(59{,}888)+(5)(50{,}756)+(37.6)(48{,}470)$$
$$=2{,}536{,}055\text{ kJ (Higher than 2,481,060 kJ)}$$

At 1500 K: $$3\bar{h}_{CO_2}+4\bar{h}_{H_2O}+5\bar{h}_{O_2}+37.6\bar{h}_{N_2}=(3)(71{,}078)+(4)(57{,}999)+(5)(49{,}292)+(37.6)(47{,}073)$$
$$=2{,}461{,}630\text{ kJ (Lower than 2,481,060 kJ)}$$

By interpolation, T_P = **1510 K**

(*c*) The balanced combustion equation for incomplete combustion with 95% theoretical air is

$$C_3H_8(g)+4.75(O_2+3.76N_2)\longrightarrow 2.5CO_2+0.5CO+4H_2O+17.86N_2$$

Substituting known numerical values,

$$(2.5)(-393{,}520+\bar{h}_{CO_2}-9364)+(0.5)(-110{,}530+\bar{h}_{CO}-8669)$$
$$+(4)(-241{,}820+\bar{h}_{H_2O}-9904)+(17.86)(0+\bar{h}_{N_2}-8669)=(1)(-103{,}850)$$

which yields

$$2.5\bar{h}_{CO_2}+0.5\bar{h}_{CO}+4\bar{h}_{H_2O}+17.86\bar{h}_{N_2}=2{,}124{,}684\text{ kJ}$$

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 2,124,684 / (2.5 + 4 + 0.5 + 17.86) = 85,466 kJ/kmol. This enthalpy value corresponds to about 2550 K for N_2. Noting that the majority of the moles are N_2, T_P will be close to 2550 K, but somewhat under it because of the higher specific heats of CO_2 and H_2O.

At 2350 K:

$$2.5\bar{h}_{CO_2}+0.5\bar{h}_{CO}+4\bar{h}_{H_2O}+17.86\bar{h}_{N_2}=(2.5)(122{,}091)+(0.5)(78{,}178)+(4)(100{,}846)+(17.86)(77{,}496)$$
$$=2{,}131{,}779\text{ kJ (Higher than 2,124,684 kJ)}$$

At 2300 K:

$$2.5\bar{h}_{CO_2}+0.5\bar{h}_{CO}+4\bar{h}_{H_2O}+17.86\bar{h}_{N_2}=(2.5)(119{,}035)+(0.5)(76{,}345)+(4)(98{,}199)+(17.86)(75{,}676)$$
$$=2{,}080{,}129\text{ kJ (Lower than 2,124,684 kJ)}$$

By interpolation, T_P = **2343 K**

15-108 The highest possible temperatures that can be obtained when liquid gasoline is burned steadily with air and with pure oxygen are to be determined.

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 The highest possible temperature that can be achieved during a combustion process is the temperature which occurs when a fuel is burned completely with stoichiometric amount of air in an adiabatic combustion chamber. It is determined from

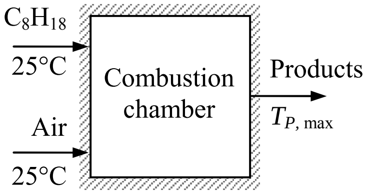
$$\sum N_P\left(\bar{h}_f^\circ+\bar{h}-\bar{h}^\circ\right)_P=\sum N_R\left(\bar{h}_f^\circ+\bar{h}-\bar{h}^\circ\right)_R \longrightarrow \sum N_P\left(\bar{h}_f^\circ+\bar{h}_T-\bar{h}^\circ\right)_P=\left(N\bar{h}_f^\circ\right)_{C_8H_{18}}$$

since all the reactants are at the standard reference temperature of 25°C, and for O_2 and N_2. The theoretical combustion equation of C_8H_{18} air is

$$C_8H_{18}+12.5\left(O_2+3.76N_2\right)\longrightarrow 8CO_2+9H_2O+47N_2$$

From the tables,

Substance	$\bar{h}_f^\circ$ kJ/kmol	$\bar{h}_{298\ K}$ kJ/kmol
C_8H_{18} (ℓ)	-249,950	---
O_2	0	8682
N_2	0	8669
H_2O (*g*)	-241,820	9904
CO_2	-393,520	9364



Thus,

$$(8)\left(-393{,}520+\bar{h}_{CO_2}-9364\right)+(9)\left(-241{,}820+\bar{h}_{H_2O}-9904\right)+(47)\left(0+\bar{h}_{N_2}-8669\right)=(1)(-249{,}950)$$

It yields $\quad 8\bar{h}_{CO_2}+9\bar{h}_{H_2O}+47\bar{h}_{N_2}=5{,}646{,}081\text{ kJ}$

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 5,646,081/(8 + 9 + 47) = 88,220 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 2400 K: $\quad 8\bar{h}_{CO_2}+9\bar{h}_{H_2O}+47\bar{h}_{N_2}=(8)(125{,}152)+(9)(103{,}508)+(47)(79{,}320)$
$\qquad =5{,}660{,}828\text{ kJ (Higher than 5,646,081 kJ)}$

At 2350 K: $\quad 8\bar{h}_{CO_2}+9\bar{h}_{H_2O}+47\bar{h}_{N_2}=(8)(122{,}091)+(9)(100{,}846)+(47)(77{,}496)$
$\qquad =5{,}526{,}654\text{ kJ (Lower than 5,646,081 kJ)}$

By interpolation, $\quad T_P$ = **2395 K**

If the fuel is burned with stoichiometric amount of pure O_2, the combustion equation would be

$$C_8H_{18}+12.5O_2\longrightarrow 8CO_2+9H_2O$$

Thus, $\quad (8)\left(-393{,}520+\bar{h}_{CO_2}-9364\right)+(9)\left(-241{,}820+\bar{h}_{H_2O}-9904\right)=(1)(-249{,}950)$

It yields $\quad 8\bar{h}_{CO_2}+9\bar{h}_{H_2O}=5{,}238{,}638\text{ kJ}$

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 5,238,638/(8 + 9) = 308,155 kJ/kmol. This enthalpy value is higher than the highest enthalpy value listed for H_2O and CO_2. Thus an estimate of the adiabatic flame temperature can be obtained by extrapolation.

At 3200 K: $\quad 8\bar{h}_{CO_2}+9\bar{h}_{H_2O}=(8)(174{,}695)+(9)(147{,}457)=2{,}724{,}673\text{ kJ}$

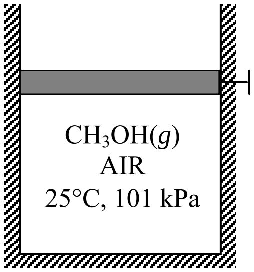
At 3250 K: $\quad 8\bar{h}_{CO_2}+9\bar{h}_{H_2O}=(8)(177{,}822)+(9)(150{,}272)=2{,}775{,}024\text{ kJ}$

By extrapolation, we get T_P = **3597 K**. However, the solution of this problem using EES gives **5645 K**. The large difference between these two values is due to extrapolation.

15-109 Methyl alcohol vapor is burned with the stoichiometric amount of air in a combustion chamber. The maximum pressure that can occur in the combustion chamber if the combustion takes place at constant volume and the maximum volume of the combustion chamber if the combustion occurs at constant pressure are to be determined.

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

Analysis (*a*) The combustion equation of $CH_3OH(g)$ with stoichiometric amount of air is



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

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

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

Thus,

$$\mathrm{CH_3OH} + 1.5(\mathrm{O_2} + 3.76\mathrm{N_2}) \longrightarrow \mathrm{CO_2} + 2\mathrm{H_2O} + 5.64\mathrm{N_2}$$

The final temperature in the tank is determined from the energy balance relation $E_{\mathrm{in}} - E_{\mathrm{out}} = \Delta E_{\mathrm{system}}$ for reacting closed systems under adiabatic conditions ($Q = 0$) with no work interactions ($W = 0$),

$$0 = \sum N_P\left(\bar{h}_f^\circ + \bar{h} - \bar{h}^\circ - P\bar{v}\right)_P - \sum N_R\left(\bar{h}_f^\circ + \bar{h} - \bar{h}^\circ - P\bar{v}\right)_R$$

Assuming both the reactants and the products to behave as ideal gases, all the internal energy and enthalpies depend on temperature only, and the $P\bar{v}$ terms in this equation can be replaced by R_uT. It yields

$$\sum N_P\left(\bar{h}_f^\circ + \bar{h}_{T_P} - \bar{h}_{298K} - R_uT\right)_P = \sum N_R\left(\bar{h}_f^\circ - R_uT\right)_R$$

since the reactants are at the standard reference temperature of 25°C. From the tables,

Substance	$\bar{\mathbf{h}}_{\mathbf{f}}^\circ$ **kJ/kmol**	$\bar{\mathbf{h}}_{\mathbf{298\ K}}$ **kJ/kmol**
CH_3OH	-200,670	---
O_2	0	8682
N_2	0	8669
H_2O (*g*)	-241,820	9904
CO_2	-393,520	9364

Thus,

$$\begin{aligned}&(1)\left(-393{,}520 + \bar{h}_{\mathrm{CO_2}} - 9364 - 8.314\times T_P\right) + (2)\left(-241{,}820 + \bar{h}_{\mathrm{H_2O}} - 9904 - 8.314\times T_P\right)\\ &+ (5.64)\left(0 + \bar{h}_{\mathrm{N_2}} - 8669 - 8.314\times T_P\right) = (1)(-200{,}670 - 8.314\times 298) + (1.5)(0 - 8.314\times 298)\\ &\qquad + (5.64)(0 - 8.314\times 298)\end{aligned}$$

It yields

$$\bar{h}_{\mathrm{CO_2}} + 2\bar{h}_{\mathrm{H_2O}} + 5.64\bar{h}_{\mathrm{N_2}} - 71.833\times T_P = 734{,}388\ \mathrm{kJ}$$

The temperature of the product gases is obtained from a trial and error solution,

At 2850 K:

$$\begin{aligned}\bar{h}_{\mathrm{CO_2}} + 2\bar{h}_{\mathrm{H_2O}} + 5.64\bar{h}_{\mathrm{N_2}} - 71.833\times T_P &= (1)(152{,}908) + (2)(127{,}952) + (5.64)(95{,}859) - (71.833)(2850)\\ &= 744{,}733\ \mathrm{kJ}\ (\text{Higher than } 734{,}388\ \mathrm{kJ})\end{aligned}$$

At 2800 K:

$$\begin{aligned}\bar{h}_{\mathrm{CO_2}} + 2\bar{h}_{\mathrm{H_2O}} + 5.64\bar{h}_{\mathrm{N_2}} - 71.833\times T_P &= (1)(149{,}808) + (2)(125{,}198) + (5.64)(94{,}014) - (71.833)(2800)\\ &= 729{,}311\ \mathrm{kJ}\ (\text{Lower than } 734{,}388\ \mathrm{kJ})\end{aligned}$$

By interpolation $T_P = 2816\text{ K}$

Since both the reactants and the products behave as ideal gases, the final (maximum) pressure that can occur in the combustion chamber is determined to be

$$\frac{P_1\mathbf{V}}{P_2\mathbf{V}} = \frac{N_1 R_u T_1}{N_2 R_u T_2} \longrightarrow P_2 = \frac{N_2 T_2}{N_1 T_1} P_1 = \frac{(8.64\text{ kmol})(2816\text{ K})}{(8.14\text{ kmol})(298\text{ K})}(101\text{ kPa}) = \mathbf{1013\ kPa}$$

(*b*) The combustion equation of $CH_3OH(g)$ remains the same in the case of constant pressure. Further, the boundary work in this case can be combined with the *u* terms so that the first law relation can be expressed in terms of enthalpies just like the steady-flow process,

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

Since both the reactants and the products behave as ideal gases, we have $h = h(T)$. Also noting that $Q = 0$ for an adiabatic combustion process, the 1st law relation reduces to

$$\sum N_P\left(\bar{h}_f^\circ + \bar{h}_{T_P} - \bar{h}_{298\text{ K}}\right)_P = \sum N_R\left(\bar{h}_f^\circ\right)_R$$

since the reactants are at the standard reference temperature of 25°C. Then using data from the mini table above, we get

$$(1)(-393{,}520 + \bar{h}_{CO_2} - 9364) + (2)(-241{,}820 + \bar{h}_{H_2O} - 9904) + (5.64)(0 + \bar{h}_{N_2} - 8669)$$
$$= (1)(-200{,}670) + (1.5)(0) + (5.64)(0)$$

It yields

$$\bar{h}_{CO_2} + 2\bar{h}_{H_2O} + 5.64\bar{h}_{N_2} = 754{,}555\text{ kJ}$$

The temperature of the product gases is obtained from a trial and error solution,

At 2350 K: $\bar{h}_{CO_2} + 2\bar{h}_{H_2O} + 5.64\bar{h}_{N_2} = (1)(122{,}091) + (2)(100{,}846) + (5.64)(77{,}496)$
$= 760{,}860\text{ kJ}\ (\text{Higher than } 754{,}555\text{ kJ})$

At 2300 K: $\bar{h}_{CO_2} + 2\bar{h}_{H_2O} + 5.64\bar{h}_{N_2} = (1)(119{,}035) + (2)(98{,}199) + (5.64)(75{,}676)$
$= 742{,}246\text{ kJ}\ (\text{Lower than } 754{,}555\text{ kJ})$

By interpolation, $T_P = 2333\text{ K}$

Treating both the reactants and the products as ideal gases, the final (maximum) volume that the combustion chamber can have is determined to be

$$\frac{P\mathbf{V}_1}{P\mathbf{V}_2} = \frac{N_1 R_u T_1}{N_2 R_u T_2} \longrightarrow \mathbf{V}_2 = \frac{N_2 T_2}{N_1 T_1}\mathbf{V}_1 = \frac{(8.64\text{ kmol})(2333\text{ K})}{(8.14\text{ kmol})(298\text{ K})}(1.5\text{ L}) = \mathbf{12.5\ L}$$

15-110 Problem 15–109 is reconsidered. The effect of the initial volume of the combustion chamber on the maximum pressure of the chamber for constant volume combustion or the maximum volume of the chamber for constant pressure combustion is to be investigated.

Analysis The problem is solved using EES, and the solution is given below.

```
"Given"
V1=1.5 [L]
T1=(25+273) [K]
P1=101 [kPa]
T0=25+273 [K]

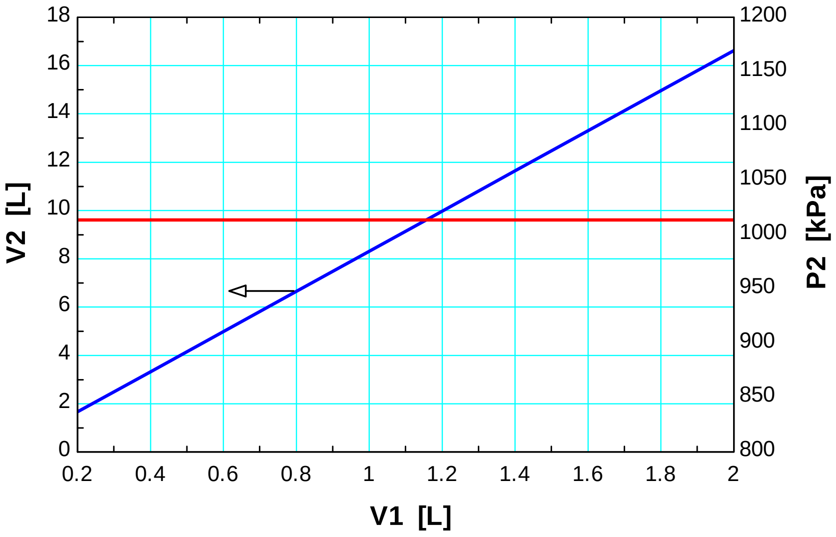
"Properties"
R_u=8.314 "[kJ/kmol-K]"

"Analysis"
"The stoichiometric combustion equation is:  CH3OH + a_th (O2+3.76N2) = CO2 + 2 H2O + 3.76*a_th N2"
1+2*a_th=2+2 "O balance"
"Mol numbers of reactants and products in kmol"
N_CH3OH=1
N_O2=a_th
N_N2=a_th*3.76
N_CO2=1
N_H2O=2
"Enthalpy of formation data from Table A-26 in kJ/kmol"
h_f_CH3OH=-200670
"Enthalpies of reactants in kJ/kmol"
h_O2=enthalpy(O2, T=T1)
h_N2_R=enthalpy(N2, T=T1)
"Enthalpies of products in kJ/kmol"
h_N2_P_a=enthalpy(N2, T=T2_a)
h_CO2_a=enthalpy(CO2, T=T2_a)
h_H2O_a=enthalpy(H2O, T=T2_a)

H_P_a=N_CO2*(h_CO2_a-R_u*T2_a)+N_H2O*(h_H2O_a-R_u*T2_a)+N_N2*(h_N2_P_a-R_u*T2_a)
H_R_a=N_CH3OH*(h_f_CH3OH-R_u*T1)+N_O2*(h_O2-R_u*T1)+N_N2*(h_N2_R-R_u*T1)
H_P_a=H_R_a
P2=(N_P/N_R)*(T2_a/T1)*P1 "Final pressure"
N_R=1+4.76*a_th
N_P=1+2+3.76*a_th
"(b)"
"Now ideal gas enthalpies of products are, in kJ/kmol"
h_N2_P_b=enthalpy(N2, T=T2_b)
h_CO2_b=enthalpy(CO2, T=T2_b)
h_H2O_b=enthalpy(H2O, T=T2_b)

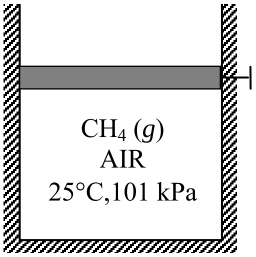
H_P_b=N_CO2*h_CO2_b+N_H2O*h_H2O_b+N_N2*h_N2_P_b
H_R_b=N_CH3OH*h_f_CH3OH+N_O2*h_O2+N_N2*h_N2_R
H_P_b=H_R_b
V2=(N_P/N_R)*(T2_b/T1)*V1 "Final pressure"
```

V1 [L]	V2 [L]	P2 [kPa]
0.2	1.663	1013
0.4	3.325	1013
0.6	4.988	1013
0.8	6.651	1013
1	8.313	1013
1.2	9.976	1013
1.4	11.64	1013
1.6	13.3	1013
1.8	14.96	1013
2	16.63	1013



15-111 Methane is burned with the stoichiometric amount of air in a combustion chamber. The maximum pressure that can occur in the combustion chamber if the combustion takes place at constant volume and the maximum volume of the combustion chamber if the combustion occurs at constant pressure are to be determined.

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



Analysis (*a*) The combustion equation of $CH_4(g)$ with stoichiometric amount of air is

$$CH_4 + a_{th}(O_2 + 3.76N_2) \longrightarrow CO_2 + 2H_2O + 3.76a_{th}N_2$$

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

$$a_{th} = 1 + 1 \quad \longrightarrow \quad a_{th} = 2$$

Thus,

$$CH_4 + 2(O_2 + 3.76N_2) \longrightarrow CO_2 + 2H_2O + 7.52N_2$$

The final temperature in the tank is determined from the energy balance relation $E_{in} - E_{out} = \Delta E_{system}$ for reacting closed systems under adiabatic conditions (Q = 0) with no work interactions (W = 0),

$$0 = \sum N_P\left(\bar{h}_f^\circ + \bar{h} - \bar{h}^\circ - P\bar{v}\right)_P - \sum N_R\left(\bar{h}_f^\circ + \bar{h} - \bar{h}^\circ - P\bar{v}\right)_R$$

Since both the reactants and the products behave as ideal gases, all the internal energy and enthalpies depend on temperature only, and the $P\bar{v}$ terms in this equation can be replaced by R_uT. It yields

$$\sum N_P\left(\bar{h}_f^\circ + \bar{h}_{T_P} - \bar{h}_{298\text{K}} - R_uT\right)_P = \sum N_R\left(\bar{h}_f^\circ - R_uT\right)_R$$

since the reactants are at the standard reference temperature of 25°C. From the tables,

Substance	$\bar{h}_f^\circ$ **kJ/kmol**	$\bar{h}_{298\text{ K}}$ **kJ/kmol**
CH_4	-74,850	---
O_2	0	8682
N_2	0	8669
H_2O (*g*)	-241,820	9904
CO_2	-393,520	9364

Thus,

$$\begin{aligned}&(1)(-393{,}520 + \bar{h}_{CO_2} - 9364 - 8.314 \times T_P) + (2)(-241{,}820 + \bar{h}_{H_2O} - 9904 - 8.314 \times T_P)\\ &+ (7.52)(0 + \bar{h}_{N_2} - 8669 - 8.314 \times T_P) = (1)(-74{,}850 - 8.314 \times 298) + (2)(0 - 8.314 \times 298)\\ &\qquad + (7.52)(0 - 8.314 \times 298)\end{aligned}$$

It yields

$$\bar{h}_{CO_2} + 2\bar{h}_{H_2O} + 7.52\bar{h}_{N_2} - 87.463 \times T_P = 870{,}609 \text{ kJ}$$

The temperature of the product gases is obtained from a trial and error solution,

At 2850 K:

$$\begin{aligned}\bar{h}_{CO_2} + 2\bar{h}_{H_2O} + 7.52\bar{h}_{N_2} - 87.463 \times T_P &= (1)(152{,}908) + (2)(127{,}952) + (7.52)(95{,}859) - (87.463)(2850)\\ &= 880{,}402 \text{ kJ (Higher than 870,609 kJ)}\end{aligned}$$

At 2800 K:

$$\begin{aligned}\bar{h}_{CO_2} + 2\bar{h}_{H_2O} + 7.52\bar{h}_{N_2} - 87.463 \times T_P &= (1)(149{,}808) + (2)(125{,}198) + (7.52)(94{,}014) - (87.463)(2800)\\ &= 862{,}293 \text{ kJ (Lower than 870,609 kJ)}\end{aligned}$$

By interpolation, $T_P = 2823\text{ K}$

Treating both the reactants and the products as ideal gases, the final (maximum) pressure that can occur in the combustion chamber is determined to be

$$\frac{P_1 V}{P_2 V} = \frac{N_1 R_u T_1}{N_2 R_u T_2} \longrightarrow P_2 = \frac{N_2 T_2}{N_1 T_1} P_1 = \frac{(10.52\text{ kmol})(2823\text{ K})}{(10.52\text{ kmol})(298\text{ K})}(101\text{ kPa}) = \mathbf{957\ \ kPa}$$

(*b*) The combustion equation of $CH_4(g)$ remains the same in the case of constant pressure. Further, the boundary work in this case can be combined with the *u* terms so that the first law relation can be expressed in terms of enthalpies just like the steady-flow process,

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

Again since both the reactants and the products behave as ideal gases, we have $h = h(T)$. Also noting that $Q = 0$ for an adiabatic combustion process, the energy balance relation reduces to

$$\sum N_P \left(\bar{h}_f^\circ + \bar{h}_{T_P} - \bar{h}_{298\text{ K}}\right)_P = \sum N_R \left(\bar{h}_f^\circ\right)_R$$

since the reactants are at the standard reference temperature of 25°C. Then using data from the mini table above, we get

$$\begin{aligned}(1)(-393{,}520 + \bar{h}_{CO_2} - 9364) + (2)(-241{,}820 + \bar{h}_{H_2O} - 9904) + (7.52)(0 + \bar{h}_{N_2} - 8669)\\ = (1)(-74{,}850) + (2)(0) + (7.52)(0)\end{aligned}$$

It yields

$$\bar{h}_{CO_2} + 2\bar{h}_{H_2O} + 7.52\bar{h}_{N_2} = 896{,}673\text{ kJ}$$

The temperature of the product gases is obtained from a trial and error solution,

At 2350 K:
$$\begin{aligned}\bar{h}_{CO_2} + 2\bar{h}_{H_2O} + 7.52\bar{h}_{N_2} &= (1)(122{,}091) + (2)(100{,}846) + (7.52)(77{,}496)\\ &= 906{,}553\text{ kJ (Higher than 896,673 kJ)}\end{aligned}$$

At 2300 K:
$$\begin{aligned}\bar{h}_{CO_2} + 2\bar{h}_{H_2O} + 7.52\bar{h}_{N_2} &= (1)(119{,}035) + (2)(98{,}199) + (7.52)(75{,}676)\\ &= 884{,}517\text{ kJ (Lower than 896,673 kJ)}\end{aligned}$$

By interpolation, $T_P = 2328\text{ K}$

Treating both the reactants and the products as ideal gases, the final (maximum) volume that the combustion chamber can have is determined to be

$$\frac{P V_1}{P V_2} = \frac{N_1 R_u T_1}{N_2 R_u T_2} \longrightarrow V_2 = \frac{N_2 T_2}{N_1 T_1} V_1 = \frac{(10.52\text{ kmol})(2328\text{ K})}{(10.52\text{ kmol})(298\text{ K})}(1.5\text{ L}) = \mathbf{11.7\ \ L}$$

15-112 n-Octane is burned with 100 percent excess air. The combustion is incomplete. The maximum work that can be produced is to be determined.

Assumptions **1** Combustion is incomplete. **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 combustion equation with 100% excess air and 10% CO is

$$C_8H_{18} + 2\times 12.5(O_2 + 3.76N_2) \longrightarrow 8(0.90\ CO_2 + 0.10\,CO) + 9\,H_2O + xO_2 + 2\times 12.5\times 3.76\,N_2$$

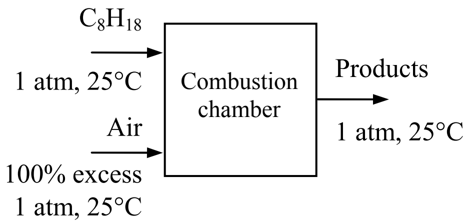
The coefficient for O_2 is determined from its mass balance as

$$25 = 7.2 + 0.4 + 4.5 + x \longrightarrow x = 12.9$$

Substituting,

$$C_8H_{18} + 25(O_2 + 3.76N_2) \longrightarrow 7.2\ CO_2 + 0.8\,CO + 9\,H_2O + 12.9\,O_2 + 94\,N_2$$

The reactants and products are at 25°C and 1 atm, which is the standard reference state and also the state of the surroundings. Therefore, the reversible work in this case is simply the difference between the Gibbs function of formation of the reactants and that of the products,



$$\begin{aligned} W_{\text{rev}} &= \sum N_R \bar{g}^{\circ}_{f,R} - \sum N_P \bar{g}^{\circ}_{f,P} \\ &= (1)(16{,}530) - (7.2)(-394{,}360) - (0.8)(-137{,}150) - (9)(-228{,}590) \\ &= 5{,}022{,}952\ \text{kJ}\quad \text{(per kmol of fuel)} \end{aligned}$$

since the g°_f of stable elements at 25°C and 1 atm is zero. Per unit mass basis,

$$W_{\text{rev}} = \frac{5{,}022{,}952\ \text{kJ/kmol}}{114\ \text{kg/kmol}} = \mathbf{44{,}060\ kJ/kg\ fuel}$$

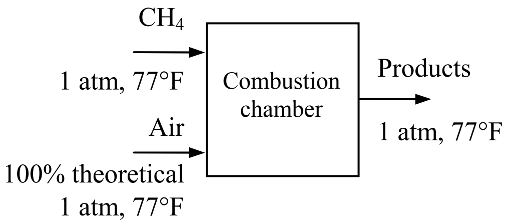
15-113E Methane is burned with stoichiometric air. The maximum work that can be produced is to be determined.

Assumptions **1** Combustion is incomplete. **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 combustion equation is

$$CH_4 + 2(O_2 + 3.76N_2) \longrightarrow CO_2 + 2H_2O + 7.52N_2$$

The reactants and products are at 77°F and 1 atm, which is the standard reference state and also the state of the surroundings. Therefore, the reversible work in this case is simply the difference between the Gibbs function of formation of the reactants and that of the products,



$$\begin{aligned} W_{\text{rev}} &= \sum N_R \bar{g}^{\circ}_{f,R} - \sum N_P \bar{g}^{\circ}_{f,P} \\ &= (1)(-21{,}860) - (1)(-169{,}680) - (2)(-98{,}350) \\ &= 344{,}520\ \text{Btu}\quad \text{(per lbmol of fuel)} \end{aligned}$$

since the g°_f of stable elements at 77°F and 1 atm is zero. Per unit mass basis,

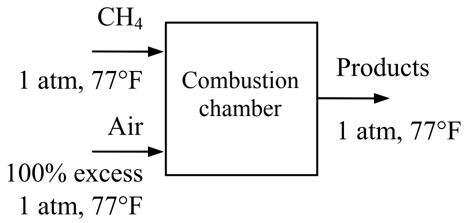
$$W_{\text{rev}} = \frac{344{,}520\ \text{Btu/lbmol}}{16\ \text{lbm/lbmol}} = \mathbf{21{,}530\ Btu/lbm\ fuel}$$

15-114E Methane is burned with 100% excess air. The maximum work that can be produced is to be determined and compared to when methane is burned with stoichiometric air.

Assumptions **1** Combustion is incomplete. **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 combustion equation with 100% excess air is

$$\mathrm{CH_4} + 4(\mathrm{O_2} + 3.76\mathrm{N_2}) \longrightarrow \mathrm{CO_2} + 2\mathrm{H_2O} + 2\mathrm{O_2} + 15.04\mathrm{N_2}$$



The reactants and products are at 77°F and 1 atm, which is the standard reference state and also the state of the surroundings. Therefore, the reversible work in this case is simply the difference between the Gibbs function of formation of the reactants and that of the products,

$$\begin{aligned} W_{\text{rev}} &= \sum N_R \bar{g}^{\circ}_{f,R} - \sum N_P \bar{g}^{\circ}_{f,P} \\ &= (1)(-21{,}860) - (1)(-169{,}680) - (2)(-98{,}350) \\ &= 344{,}520 \text{ Btu} \quad \text{(per lbmol of fuel)} \end{aligned}$$

since the g°_f of stable elements at 77°F and 1 atm is zero. Per unit mass basis,

$$W_{\text{rev}} = \frac{344{,}52 \text{ Btu/lbmol}}{16 \text{ lbm/lbmol}} = \mathbf{21{,}530\ Btu/lbm\ fuel}$$

The excess air only adds oxygen and nitrogen to the reactants and products. The excess air then does not change the maximum work.

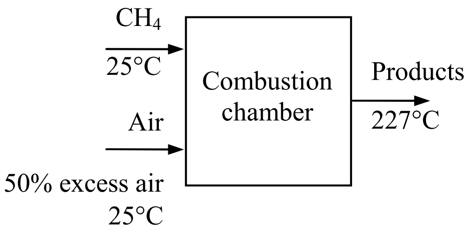
15-115 Methane is burned steadily with 50 percent excess air in a steam boiler. The amount of steam generated per unit of fuel mass burned, the change in the exergy of the combustion streams, the change in the exergy of the steam stream, and the lost work potential are to be determined.

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.

Properties The molar masses of CH_4 and air are 16 kg/kmol and 29 kg/kmol, respectively (Table A-1).

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 CH_4, the combustion equation can be written as

$$CH_4 + 3(O_2 + 3.76N_2) \longrightarrow CO_2 + 2H_2O + O_2 + 11.28N_2$$



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

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

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	$\bar{h}_{298K}$ kJ/kmol	$\bar{h}_{500K}$ kJ/kmol
CH_4	-74,850	---	---
O_2	0	8682	14,770
N_2	0	8669	14,581
H_2O (*g*)	-241,820	9904	16,828
CO_2	-393,520	9364	17,678

Thus,

$$\begin{aligned} -Q_{out} &= (1)(-393{,}520 + 17{,}678 - 9364) + (2)(-241{,}820 + 16{,}828 - 9904) \\ &\quad + (1)(0 + 14{,}770 - 8682) + (11.28)(0 + 14{,}581 - 8669) - (1)(-74{,}850) \\ &= -707{,}373 \text{ kJ/kmol of fuel} \end{aligned}$$

The heat loss per unit mass of the fuel is

$$Q_{out} = \frac{707{,}373 \text{ kJ/kmol of fuel}}{16 \text{ kg/kmol of fuel}} = 44{,}211 \text{ kJ/kg fuel}$$

The amount of steam generated per unit mass of fuel burned is determined from an energy balance to be (Enthalpies of steam are from tables A-4 and A-6)

$$\frac{m_s}{m_f} = \frac{Q_{out}}{\Delta h_s} = \frac{44{,}211 \text{ kJ/kg fuel}}{(3214.5 - 852.26) \text{ kJ/kg steam}} = \mathbf{18.72 \text{ kg steam/kg fuel}}$$

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

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

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 $P_i = y_i P_{total}$, 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	$\bar{s}_i^\circ(T,1\text{atm})$	$R_u\ln(y_iP_m)$	$N_i\bar{s}_i$
CH_4	1	---	186.16	0	186.16
O_2	3	0.21	205.04	-12.98	654.06
N_2	11.28	0.79	191.61	-1.960	2183.47
					S_R = 3023.69 kJ/K
CO_2	1	0.0654	234.814	-22.67	257.48
H_2O (*g*)	2	0.1309	206.413	-16.91	446.65
O_2	1	0.0654	220.589	-22.67	243.26
N_2	11.28	0.7382	206.630	-2.524	2359.26
					S_P = 3306.65 kJ/K

Thus,

$$S_{\text{gen}} = S_P - S_R + \frac{Q_{\text{out}}}{T_{\text{surr}}} = 3306.65 - 3023.69 + \frac{707{,}373}{298} = 2657\text{ kJ/K }(\text{per kmol fuel})$$

The exergy change of the combustion streams is equal to the exergy destruction since there is no actual work output. That is,

$$\Delta X_{\text{gases}} = -X_{\text{dest}} = -T_0 S_{\text{gen}} = -(298\text{ K})(2657\text{ kJ/K}) = -791{,}786\text{ kJ/kmol fuel}$$

Per unit mass basis,

$$\Delta X_{\text{gases}} = \frac{-791{,}786\text{ kJ/kmol fuel}}{16\text{ kg/kmol}} = \mathbf{-49{,}490\text{ kJ/kg fuel}}$$

Note that the exergy change is negative since the exergy of combustion gases decreases.

(*c*) The exergy change of the steam stream is

$$\Delta X_{\text{steam}} = \Delta h - T_0\Delta s = (3214.5 - 852.26) - (298)(6.7714 - 2.3305) = \mathbf{1039\text{ kJ/kg steam}}$$

(*d*) The lost work potential is the negative of the net exergy change of both streams:

$$\begin{aligned} X_{\text{dest}} &= -\left(\frac{m_s}{m_f}\Delta X_{\text{steam}} + \Delta X_{\text{gases}}\right) \\ &= -[(18.72\text{ kg steam/kg fuel})(1039\text{ kJ/kg steam}) + (-49{,}490\text{ kJ/kg fuel})] \\ &= \mathbf{30{,}040\text{ kJ/kg fuel}} \end{aligned}$$

15-116 A coal from Utah is burned steadily with 50 percent excess air in a steam boiler. The amount of steam generated per unit of fuel mass burned, the change in the exergy of the combustion streams, the change in the exergy of the steam stream, and the lost work potential 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. **5** The effect of sulfur on the energy and entropy balances is negligible.

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

Analysis (*a*) We consider 100 kg of coal for simplicity. Noting that the mass percentages in this case correspond to the masses of the constituents, the mole numbers of the constituent of the coal are determined to be

$$N_C = \frac{m_C}{M_C} = \frac{61.40\text{ kg}}{12\text{ kg/kmol}} = 5.117\text{ kmol}$$

$$N_{H2} = \frac{m_{H2}}{M_{H2}} = \frac{5.79\text{ kg}}{2\text{ kg/kmol}} = 2.895\text{ kmol}$$

$$N_{O2} = \frac{m_{O2}}{M_{O2}} = \frac{25.31\text{ kg}}{32\text{ kg/kmol}} = 0.7909\text{ kmol}$$

$$N_{N2} = \frac{m_{N2}}{M_{N2}} = \frac{1.09\text{ kg}}{28\text{ kg/kmol}} = 0.03893\text{ kmol}$$

$$N_S = \frac{m_S}{M_S} = \frac{1.41\text{ kg}}{32\text{ kg/kmol}} = 0.04406\text{ kmol}$$

61.40% C
5.79% H_2
25.31% O_2
1.09% N_2
1.41% S
5.00% ash
(by mass)

The mole number of the mixture and the mole fractions are

$$N_m = 5.117 + 2.895 + 0.7909 + 0.03893 + 0.04406 = 8.886\text{ kmol}$$

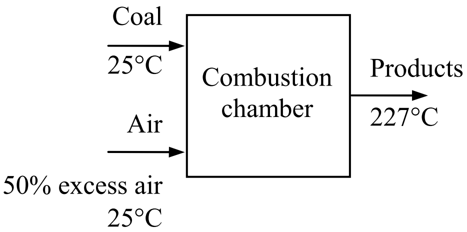
$$y_C = \frac{N_C}{N_m} = \frac{5.117\text{ kmol}}{8.886\text{ kmol}} = 0.5758$$

$$y_{H2} = \frac{N_{H2}}{N_m} = \frac{2.895\text{ kmol}}{8.886\text{ kmol}} = 0.3258$$

$$y_{O2} = \frac{N_{O2}}{N_m} = \frac{0.7909\text{ kmol}}{8.886\text{ kmol}} = 0.0890$$

$$y_{N2} = \frac{N_{N2}}{N_m} = \frac{0.03893\text{ kmol}}{8.886\text{ kmol}} = 0.00438$$

$$y_S = \frac{N_S}{N_m} = \frac{0.04406\text{ kmol}}{8.886\text{ kmol}} = 0.00496$$



Ash consists of the non-combustible matter in coal. Therefore, the mass of ash content that enters the combustion chamber is equal to the mass content that leaves. Disregarding this non-reacting component for simplicity, the combustion equation may be written as

$$0.5758\text{C} + 0.3258\text{H}_2 + 0.0890\text{O}_2 + 0.00438\text{N}_2 + 0.00496\text{S} + 1.5a_{\text{th}}(\text{O}_2 + 3.76\text{N}_2)$$
$$\longrightarrow 0.5758\text{CO}_2 + 0.3258\text{H}_2\text{O} + 0.00496\text{SO}_2 + 0.5a_{\text{th}}\text{O}_2 + 1.5a_{\text{th}} \times 3.76\text{N}_2$$

According to the oxygen balance,

$$\text{O}_2\text{ balance: } 0.0890 + 1.5a_{\text{th}} = 0.5758 + 0.5\times 0.3258 + 0.00496 + 0.5a_{\text{th}} \longrightarrow a_{\text{th}} = 0.6547$$

Substituting,

$$0.5758\text{C} + 0.3258\text{H}_2 + 0.0890\text{O}_2 + 0.00438\text{N}_2 + 0.00496\text{S} + 0.9821(\text{O}_2 + 3.76\text{N}_2)$$
$$\longrightarrow 0.5758\text{CO}_2 + 0.3258\text{H}_2\text{O} + 0.00496\text{SO}_2 + 0.3274\text{O}_2 + 3.693\text{N}_2$$

The apparent molecular weight of the coal is

$$M_m = \frac{m_m}{N_m} = \frac{(0.5758\times 12 + 0.3258\times 2 + 0.0890\times 32 + 0.00438\times 28 + 0.00496\times 32)\text{ kg}}{(0.5758 + 0.3258 + 0.0890 + 0.00438 + 0.00496)\text{ kmol}}$$
$$= \frac{10.69\text{ kg}}{1.0\text{ kmol}} = 10.69\text{ kg/kmol coal}$$

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

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

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

Substance	$\bar{\mathbf{h}}_{\mathbf{f}}^\circ$ **kJ/kmol**	$\bar{\mathbf{h}}_{\mathbf{298K}}$ **kJ/kmol**	$\bar{\mathbf{h}}_{\mathbf{500K}}$ **kJ/kmol**
O_2	0	8682	14,770
N_2	0	8669	14,581
H_2O (*g*)	-241,820	9904	16,828
CO_2	-393,520	9364	17,678

Thus,

$$\begin{aligned}-Q_{\text{out}} &= (0.5758)(-393{,}520 + 17{,}678 - 9364) + (0.3258)(-241{,}820 + 16{,}828 - 9904)\\ &\quad + (0.3274)(0 + 14{,}770 - 8682) + (3.693)(0 + 14{,}581 - 8669) - 0\\ &= -274{,}505\text{ kJ/kmol of fuel}\end{aligned}$$

The heat loss per unit mass of the fuel is

$$Q_{\text{out}} = \frac{274{,}505\text{ kJ/kmol of fuel}}{10.69\text{ kg/kmol of fuel}} = 25{,}679\text{ kJ/kg fuel}$$

The amount of steam generated per unit mass of fuel burned is determined from an energy balance to be (Enthalpies of steam are from tables A-4 and A-6)

$$\frac{m_s}{m_f} = \frac{Q_{\text{out}}}{\Delta h_s} = \frac{25{,}679\text{ kJ/kg fuel}}{(3214.5 - 852.26)\text{ kJ/kg steam}} = \mathbf{10.87\ kg\ steam/kg\ fuel}$$

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

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

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 $P_i = y_i P_{\text{total}}$, 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

	$\mathbf{N_i}$	$\mathbf{y_i}$	$\bar{\mathbf{s}}_\mathbf{i}^\circ(\mathbf{T,1atm})$	$\mathbf{R_u ln(y_i P_m)}$	$\mathbf{N_i \bar{s}_i}$
C	0.5758	0.5758	5.74	-4.589	5.95
H_2	0.3258	0.3258	130.68	-9.324	45.61
O_2	0.0890	0.0890	205.04	-20.11	20.04
N_2	0.00438	0.00438	191.61	-45.15	1.04
O_2	0.9821	0.21	205.04	-12.98	214.12
N_2	3.693	0.79	191.61	-1.960	714.85
					S_R = 1001.61 kJ/K
CO_2	0.5758	0.1170	234.814	-17.84	145.48
H_2O (*g*)	0.3258	0.0662	206.413	-22.57	74.60
O_2	0.3274	0.0665	220.589	-22.54	79.60
N_2	3.693	0.7503	206.630	-2.388	771.90
					S_P = 1071.58 kJ/K

Thus,

$$S_{\text{gen}} = S_P - S_R + \frac{Q_{\text{out}}}{T_{\text{surr}}} = 1071.58 - 1001.61 + \frac{274{,}505}{298} = 991.1 \text{ kJ/K } (\text{per kmol fuel})$$

The exergy change of the combustion streams is equal to the exergy destruction since there is no actual work output. That is,

$$\Delta X_{\text{gases}} = -X_{\text{dest}} = -T_0 S_{\text{gen}} = -(298\text{ K})(991.1\text{ kJ/K}) = -295{,}348\text{ kJ/kmol fuel}$$

Per unit mass basis,

$$\Delta X_{\text{gases}} = \frac{-295{,}348\text{ kJ/K}}{10.69\text{ kg/kmol}} = \mathbf{-27{,}630\text{ kJ/kg fuel}}$$

Note that the exergy change is negative since the exergy of combustion gases decreases.

(*c*) The exergy change of the steam stream is

$$\Delta X_{\text{steam}} = \Delta h - T_0 \Delta s = (3214.5 - 852.26) - (298)(6.7714 - 2.3305) = \mathbf{1039\text{ kJ/kg steam}}$$

(*d*) The lost work potential is the negative of the net exergy change of both streams:

$$\begin{aligned} X_{\text{dest}} &= -\left(\frac{m_s}{m_f}\Delta X_{\text{steam}} + \Delta X_{\text{gases}}\right) \\ &= -[(10.87\text{ kg steam/kg fuel})(1039\text{ kJ/kg steam}) + (-27{,}630\text{ kJ/kg fuel})] \\ &= \mathbf{16{,}340\text{ kJ/kg fuel}} \end{aligned}$$

15-117 An expression for the HHV of a gaseous alkane C_nH_{2n+2} in terms of n is to be developed.

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

Analysis The complete reaction balance for 1 kmol of fuel is

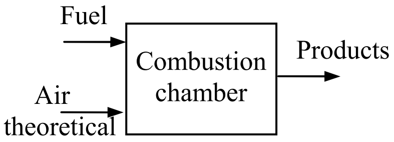
$$\mathrm{C}_n\mathrm{H}_{2n+2}+\frac{3n+1}{2}(\mathrm{O}_2+3.76\mathrm{N}_2)\longrightarrow n\mathrm{CO}_2+(n+1)\mathrm{H}_2\mathrm{O}+\frac{3n+1}{2}(3.76)\mathrm{N}_2$$

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,

$$q=h_C=H_P-H_R=\sum N_P\bar{h}^\circ_{f,P}-\sum N_R\bar{h}^\circ_{f,R}=\left(N\bar{h}^\circ_f\right)_{\mathrm{CO2}}+\left(N\bar{h}^\circ_f\right)_{\mathrm{H2O}}-\left(N\bar{h}^\circ_f\right)_{\mathrm{fuel}}$$

For the HHV, the water in the products is taken to be liquid. Then,

$$h_C=n(-393{,}520)+(n+1)(-285{,}830)-\left(\bar{h}^\circ_f\right)_{\mathrm{fuel}}$$



The HHV of the fuel is

$$\mathrm{HHV}=\frac{-h_C}{M_{\mathrm{fuel}}}=\frac{n(-393{,}520)+(n+1)(-285{,}830)-\left(\bar{h}^\circ_f\right)_{\mathrm{fuel}}}{M_{\mathrm{fuel}}}$$

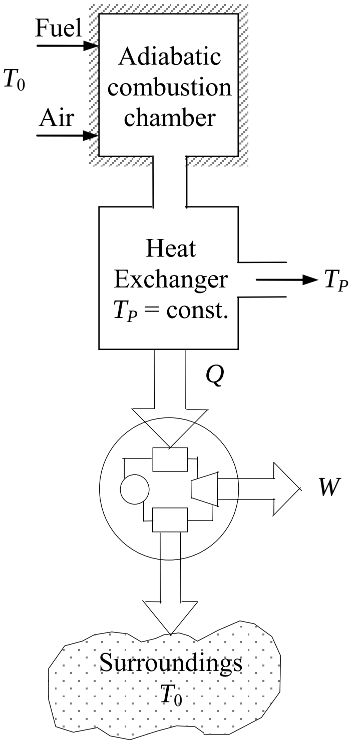
For the LHV, the water in the products is taken to be vapor. Then,

$$\mathrm{LHV}=\frac{n(-393{,}520)+(n+1)(-241{,}820)-\left(\bar{h}^\circ_f\right)_{\mathrm{fuel}}}{M_{\mathrm{fuel}}}$$

15-118 It is to be shown that the work output of the Carnot engine will be maximum when $T_p = \sqrt{T_0 T_{af}}$. It is also to be shown that the maximum work output of the Carnot engine in this case becomes $w = CT_{\text{af}}\left(1-\dfrac{\sqrt{T_0}}{\sqrt{T_{\text{af}}}}\right)^2$.

Analysis The combustion gases will leave the combustion chamber and enter the heat exchanger at the adiabatic flame temperature T_{af} since the chamber is adiabatic and the fuel is burned completely. The combustion gases experience no change in their chemical composition as they flow through the heat exchanger. Therefore, we can treat the combustion gases as a gas stream with a constant specific heat c_p. Noting that the heat exchanger involves no work interactions, the energy balance equation for this single-stream steady-flow device can be written as

$$\dot{Q} = \dot{m}(h_e - h_i) = \dot{m}C(T_p - T_{\text{af}})$$



where $\dot{Q}$ is the negative of the heat supplied to the heat engine. That is,

$$\dot{Q}_H = -\dot{Q} = \dot{m}C(T_{\text{af}} - T_p)$$

Then the work output of the Carnot heat engine can be expressed as

$$\dot{W} = \dot{Q}_H\left(1-\frac{T_0}{T_p}\right) = \dot{m}C(T_{\text{af}} - T_p)\left(1-\frac{T_0}{T_p}\right) \qquad (1)$$

Taking the partial derivative of $\dot{W}$ with respect to T_p while holding T_{af} and T_0 constant gives

$$\frac{\partial W}{\partial T_p} = 0 \longrightarrow -\dot{m}C\left(1-\frac{T_0}{T_p}\right) + \dot{m}C(T_p - T_{\text{af}})\frac{T_0}{T_p^2} = 0$$

Solving for T_p we obtain

$$T_p = \sqrt{T_0 T_{\text{af}}}$$

which the temperature at which the work output of the Carnot engine will be a maximum. The maximum work output is determined by substituting the relation above into Eq. (1),

$$\dot{W} = \dot{m}C(T_{\text{af}} - T_p)\left(1-\frac{T_0}{T_p}\right) = \dot{m}C\left(T_{\text{af}} - \sqrt{T_0 T_{\text{af}}}\right)\left(1-\frac{T_0}{\sqrt{T_0 T_{\text{af}}}}\right)$$

It simplifies to

$$\dot{W} = \dot{m}CT_{\text{af}}\left(1-\frac{\sqrt{T_0}}{\sqrt{T_{\text{af}}}}\right)^2$$

or

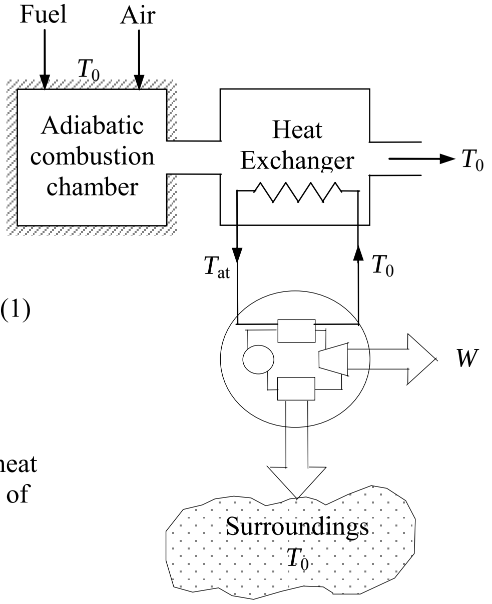
$$w = CT_{\text{af}}\left(1-\frac{\sqrt{T_0}}{\sqrt{T_{\text{af}}}}\right)^2$$

which is the desired relation.

15-119 It is to be shown that the work output of the reversible heat engine operating at the specified conditions is $\dot{W}_{\text{rev}} = \dot{m}CT_0\left(\frac{T_{\text{af}}}{T_0} - 1 - \ln\frac{T_{\text{af}}}{T_0}\right)$. It is also to be shown that the effective flame temperature T_e of the furnace considered is $T_e = \frac{T_{\text{af}} - T_0}{\ln(T_{\text{af}}/T_0)}$.

Analysis The combustion gases will leave the combustion chamber and enter the heat exchanger at the adiabatic flame temperature T_{af} since the chamber is adiabatic and the fuel is burned completely. The combustion gases experience no change in their chemical composition as they flow through the heat exchanger. Therefore, we can treat the combustion gases as a gas stream with a constant specific heat c_p. Also, the work output of the reversible heat engine is equal to the reversible work W_{rev} of the heat exchanger as the combustion gases are cooled from T_{af} to T_0. That is,

$$\begin{aligned}\dot{W}_{\text{rev}} &= \dot{m}\left(h_i - h_e - T_0(s_i - s_e)\right)\\ &= \dot{m}C\left(T_{\text{af}} - T_0 - T_0\left(C\ln\frac{T_{\text{af}}}{T_0} - R\ln\frac{P_{\text{af}}}{P_0}^{\nearrow 0}\right)\right)\\ &= \dot{m}C\left(T_{\text{af}} - T_0 - T_0 C\ln\frac{T_{\text{af}}}{T_0}\right)\end{aligned}$$



which can be rearranged as

$$\dot{W}_{\text{rev}} = \dot{m}CT_0\left(\frac{T_{\text{af}}}{T_0} - 1 - \ln\frac{T_{\text{af}}}{T_0}\right) \quad \text{or} \quad w_{\text{rev}} = CT_0\left(\frac{T_{\text{af}}}{T_0} - 1 - \ln\frac{T_{\text{af}}}{T_0}\right) \quad (1)$$

which is the desired result.

The effective flame temperature T_e can be determined from the requirement that a Carnot heat engine which receives the same amount of heat from a heat reservoir at constant temperature T_e produces the same amount of work. The amount of heat delivered to the heat engine above is

$$\dot{Q}_H = \dot{m}(h_i - h_e) = \dot{m}C(T_{\text{af}} - T_0)$$

A Carnot heat engine which receives this much heat at a constant temperature T_e will produce work in the amount of

$$\dot{W} = \dot{Q}_H \eta_{\text{th,Carnot}} = \dot{m}C(T_{\text{af}} - T_0)\left(1 - \frac{T_0}{T_e}\right) \quad (2)$$

Setting equations (1) and (2) equal to each other yields

$$\dot{m}CT_0\left(\frac{T_{\text{af}}}{T_0} - 1 - \ln\frac{T_{\text{af}}}{T_0}\right) = \dot{m}C(T_{\text{af}} - T_0)\left(1 - \frac{T_0}{T_e}\right)$$

$$T_{\text{af}} - T_0 - T_0\ln\frac{T_{\text{af}}}{T_0} = T_{\text{af}} - T_{\text{af}}\frac{T_0}{T_e} - T_0 + T_0\frac{T_0}{T_e}$$

Simplifying and solving for T_e, we obtain

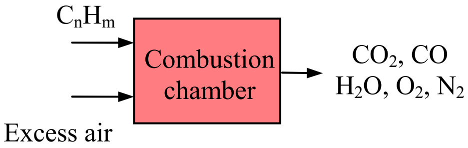
$$T_e = \frac{T_{\text{af}} - T_0}{\ln(T_{\text{af}}/T_0)}$$

which is the desired relation.

15-120 The combustion of a hydrocarbon fuel C_nH_m with excess air and incomplete combustion is considered. The coefficients of the reactants and products are to be written in terms of other parameters.

Analysis The balanced reaction equation for stoichiometric air is

$$C_nH_m + A_{th}[O_2 + 3.76N_2] \longrightarrow n\,CO_2 + (m/2)\,H_2O + 3.76A_{th}\;N_2$$



The stoichiometric coefficient A_{th} is determined from an O_2 balance:

$$A_{th} = n + m/4$$

The reaction with excess air and incomplete combustion is

$$C_nH_m + (1+B)A_{th}[O_2 + 3.76N_2] \longrightarrow an\,CO_2 + bn\,CO + (m/2)\,H_2O + G\,O_2 + 3.76(1+B)A_{th}\;N_2$$

The given reaction is

$$C_nH_m + (1+B)A_{th}[O_2 + 3.76N_2] \longrightarrow D\,CO_2 + E\,CO + F\,H_2O + G\,O_2 + J\,N_2$$

Thus,

$$D = an$$
$$E = bn$$
$$F = m/2$$
$$J = 3.76(1+B)A_{th}$$

The coefficient G for O_2 is determined from a mass balance,

O_2 balance:

$$(1+B)A_{th} = an + \frac{bn}{2} + \frac{m}{4} + G$$

$$(1+B)\left(n + \frac{m}{4}\right) = an + \frac{bn}{2} + \frac{m}{4} + G$$

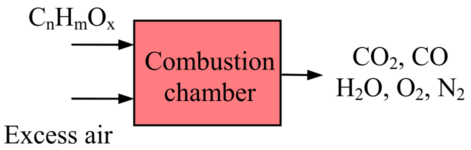
$$\left(n + \frac{m}{4}\right) + BA_{th} = an + \frac{bn}{2} + \frac{m}{4} + G$$

$$\begin{aligned} G &= n + BA_{th} - an - \frac{bn}{2} \\ &= n(1-a) + BA_{th} - \frac{bn}{2} \\ &= nb - \frac{bn}{2} + BA_{th} \\ &= \frac{bn}{2} + BA_{th} \end{aligned}$$

15-121 The combustion of an alcohol fuel ($C_nH_mO_x$) with excess air and incomplete combustion is considered. The coefficients of the reactants and products are to be written in terms of other parameters.

Analysis The balanced reaction equation for stoichiometric air is

$$C_nH_mO_x + A_{th}[O_2 + 3.76N_2] \longrightarrow n\,CO_2 + (m/2)\,H_2O + 3.76A_{th}\ N_2$$



The stoichiometric coefficient A_{th} is determined from an O_2 balance:

$$x/2 + A_{th} = n + m/4 \longrightarrow A_{th} = n + m/4 - x/2$$

The reaction with excess air and incomplete combustion is

$$C_nH_mO_x + (1+B)A_{th}[O_2 + 3.76N_2] \longrightarrow an\,CO_2 + bn\,CO + (m/2)\,H_2O + G\,O_2 + 3.76(1+B)A_{th}\ N_2$$

The given reaction is

$$C_nH_mO_x + (1+B)A_{th}[O_2 + 3.76N_2] \longrightarrow D\,CO_2 + E\,CO + F\,H_2O + G\,O_2 + J\,N_2$$

Thus,

$$D = an$$

$$E = bn$$

$$F = m/2$$

$$J = 3.76(1+B)A_{th}$$

The coefficient G for O_2 is determined from a mass balance,

O_2 balance:

$$\frac{x}{2} + (1+B)A_{th} = an + \frac{bn}{2} + \frac{m}{4} + G$$

$$\frac{x}{2} + (1+B)\left(n + \frac{m}{4} - \frac{x}{2}\right) = an + \frac{bn}{2} + \frac{m}{4} + G$$

$$\frac{x}{2} + \left(n + \frac{m}{4} - \frac{x}{2}\right) + B\left(n + \frac{m}{4} - \frac{x}{2}\right) = an + \frac{bn}{2} + \frac{m}{4} + G$$

$$\frac{x}{2} + \left(n + \frac{m}{4} - \frac{x}{2}\right) + BA_{th} = an + \frac{bn}{2} + \frac{m}{4} + G$$

$$G = n - an + BA_{th} - \frac{bn}{2}$$

$$= n(1-a) - \frac{bn}{2} + BA_{th}$$

$$= nb - \frac{bn}{2} + BA_{th}$$

$$= \frac{bn}{2} + BA_{th}$$

15-122 The combustion of a mixture of an alcohol fuel ($C_nH_mO_x$) and a hydrocarbon fuel (C_wH_z) with excess air and incomplete combustion is considered. The coefficients of the reactants and products are to be written in terms of other parameters.

Analysis The balanced reaction equation for stoichiometric air is

$$y_1\,C_nH_mO_x + y_2\,C_wH_z + A_{th}\left[O_2 + 3.76N_2\right] \longrightarrow (y_1 n + y_2 w)\,CO_2 + 0.5(y_1 m + y_2 z)\,H_2O + 3.76A_{th}\,N_2$$

The stoichiometric coefficient A_{th} is determined from an O_2 balance:

$$y_1 x/2 + A_{th} = (y_1 n + y_2 w) + (y_1 m + y_2 z)/4 \longrightarrow A_{th} = (y_1 n + y_2 w) + (y_1 m + y_2 z)/4 - y_1 x/2$$

The reaction with excess air and incomplete combustion is

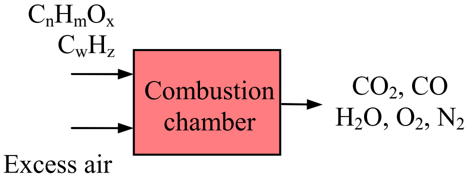
$$\begin{aligned} & y_1\,C_nH_mO_x + y_2\,C_wH_z + (1+B)A_{th}\left[O_2 + 3.76N_2\right] \\ & \longrightarrow a(y_1 n + y_2 w)\,CO_2 + b(y_1 n + y_2 w)\,CO + 0.5(y_1 m + y_2 z)\,H_2O + G\,O_2 + 3.76(1+B)A_{th}\,N_2 \end{aligned}$$

The given reaction is

$$y_1\,C_nH_mO_x + y_2\,C_wH_z + (1+B)A_{th}\left[O_2 + 3.76N_2\right] \longrightarrow D\,CO_2 + E\,CO + F\,H_2O + G\,O_2 + J\,N_2$$

Thus,

$$\begin{aligned} D &= a(y_1 n + y_2 w) \\ E &= b(y_1 n + y_2 w) \\ F &= 0.5(y_1 m + y_2 z) \\ J &= 3.76(1+B)A_{th} \end{aligned}$$



The coefficient G for O_2 is determined from a mass balance,

O_2 balance:

$$\begin{aligned} 0.5y_1 x + (1+B)A_{th} &= a(y_1 n + y_2 w) + 0.5b(y_1 n + y_2 w) + 0.25(y_1 m + y_2 z) + G \\ 0.5y_1 x + (y_1 n + y_2 w) + 0.25(y_1 m + y_2 z) - 0.5y_1 x + BA_{th} &= a(y_1 n + y_2 w) + 0.5b(y_1 n + y_2 w) + 0.25(y_1 m + y_2 z) + G \\ G &= (y_1 n + y_2 w) - a(y_1 n + y_2 w) - 0.5b(y_1 n + y_2 w) + BA_{th} \\ &= (y_1 n + y_2 w)(1-a) - 0.5b(y_1 n + y_2 w) + BA_{th} \\ &= b(y_1 n + y_2 w) - 0.5b(y_1 n + y_2 w) + BA_{th} \\ &= 0.5b(y_1 n + y_2 w) + BA_{th} \end{aligned}$$

15-123 The effect of the amount of air on the adiabatic flame temperature of liquid octane (C_8H_{18}) is to be investigated.

Analysis The problem is solved using EES, and the solution is given below.

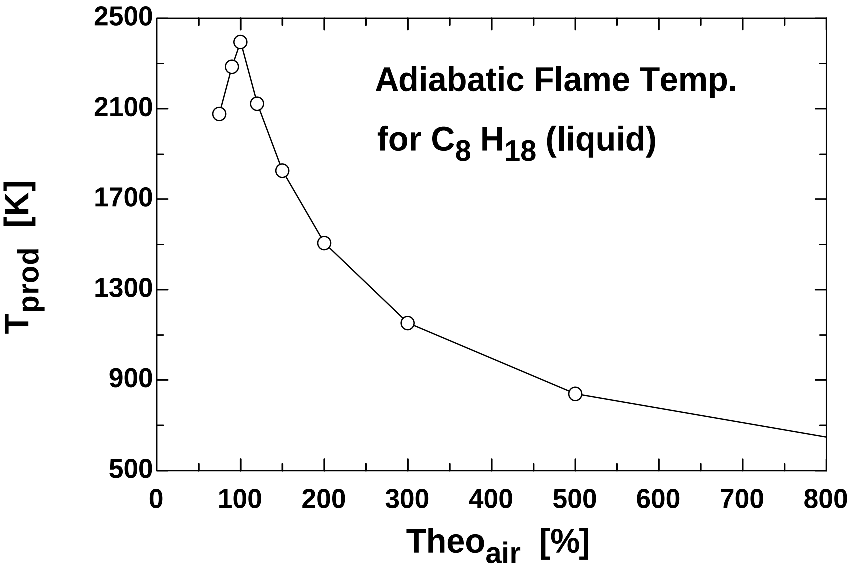
```
Adiabatic Combustion of fuel CnHm entering at T_fuel with Stoichiometric Air at T_air:
Reaction:  CxHyOz  +  (y/4 + x-z/2) (Theo_air/100) (O2 + 3.76 N2)
           <--> xCO2  +  (y/2) H2O  +  3.76  (y/4 + x-z/2) (Theo_air/100)  N2 +  (y/4 + x-z/2) (Theo_air/100 - 1) O2"
"For  theoretical oxygen, the complete combustion equation for CH3OH is"
"CH3OH + A_th O2=1 CO2+2 H2O  "
"1+ 2*A_th=1*2+2*1""theoretical O balance"
"Adiabatic, Incomplete Combustion of fuel CnHm entering at T_fuel with Stoichiometric Air at T_air:
Reaction:  CxHyOz  +  (y/4 + x-z/2) (Theo_air/100) (O2 + 3.76 N2)
           <--> (x-w)CO2 +wCO +  (y/2) H2O  +  3.76  (y/4 + x-z/2) (Theo_air/100)  N2 +  ((y/4 + x-z/2)
(Theo_air/100 - 1) +w/2)O2"

"T_prod is the adiabatic combustion temperature, assuming no dissociation.
Theo_air is the % theoretical air. "
"The initial guess value of T_prod = 450K ."

Procedure Fuel(Fuel$,T_fuel:x,y,z,h_fuel,Name$)
"This procedure takes the fuel name and returns the moles of C and moles of H"
If fuel$='C2H2(g)' then
         x=2;y=2; z=0
            Name$='Acetylene'
                                              h_fuel = 226730
else
If fuel$='C3H8(l)' then
         x=3; y=8; z=0
            Name$='Propane(liq)'
                                              h_fuel = -103850-15060
else
If fuel$='C8H18(l)' then
         x=8; y=18; z=0
           Name$='Octane(liq)'
                                              h_fuel = -249950
else
if fuel$='CH4(g)' then
         x=1; y=4; z=0
           Name$='Methane'
                                              h_fuel = enthalpy(CH4,T=T_fuel)
else
if fuel$='CH3OH(g)' then
         x=1; y=4; z=1
           Name$='Methyl alcohol'
                                              h_fuel = -200670
endif; endif; endif; endif; endif
end
Procedure Moles(x,y,z,Th_air,A_th:w,MolO2,SolMeth$)
ErrTh =(2*x + y/2 - z - x)/(2*A_th)*100
IF Th_air >= 1 then
SolMeth$ = '>= 100%, the solution assumes complete combustion.'
{MolCO = 0
MolCO2 = x}
w=0
MolO2 = A_th*(Th_air - 1)
GOTO 10
ELSE
 w = 2*x + y/2 - z - 2*A_th*Th_air
```

```
IF w > x  then
Call ERROR('The moles of CO2 are negative, the percent theoretical air must be >= xxxF3 %',ErrTh)
Else
SolMeth$ = '< 100%, the solution assumes incomplete combustion with no O_2 in products.'
MolO2 = 0
endif; endif
10:
END
{"Input data from the diagram window"
T_air = 298    [K]
Theo_air = 200  "%"
Fuel$='CH4(g)'}
T_fuel = 298   [K]
Call Fuel(Fuel$,T_fuel:x,y,z,h_fuel,Name$)
A_th =x +  y/4 - z/2
Th_air = Theo_air/100
Call Moles(x,y,z,Th_air,A_th:w,MolO2,SolMeth$)
HR=h_fuel+ (x+y/4-z/2) *(Theo_air/100) *enthalpy(O2,T=T_air)+3.76*(x+y/4-z/2) *(Theo_air/100)
*enthalpy(N2,T=T_air)
HP=HR  "Adiabatic"
HP=(x-w)*enthalpy(CO2,T=T_prod)+w*enthalpy(CO,T=T_prod)+(y/2)*enthalpy(H2O,T=T_prod)+3.76*(x+y/4-
z/2)* (Theo_air/100)*enthalpy(N2,T=T_prod)+MolO2*enthalpy(O2,T=T_prod)
Moles_O2=MolO2
Moles_N2=3.76*(x+y/4-z/2)* (Theo_air/100)
Moles_CO2=x-w
Moles_CO=w
Moles_H2O=y/2
```

$Theo_{air}$ [%]	T_{prod} [K]
75	2077
90	2287
100	2396
120	2122
150	1827
200	1506
300	1153
500	840.1
800	648.4



15-124 A general program is to be written to determine the adiabatic flame temperature during the complete combustion of a hydrocarbon fuel C_nH_m at 25°C in a steady-flow combustion chamber when the percent of excess air and its temperature are specified.

Analysis The problem is solved using EES, and the solution is given below.

```
Adiabatic Combustion of fuel CnHm entering at T_fuel with Stoichiometric Air at T_air:
Reaction:  CxHyOz  +  (y/4 + x-z/2) (Theo_air/100) (O2 + 3.76 N2)
            <--> xCO2  +  (y/2) H2O  +  3.76  (y/4 + x-z/2) (Theo_air/100)  N2 +  (y/4 + x-z/2) (Theo_air/100 - 1) O2"
"For  theoretical oxygen, the complete combustion equation for CH3OH is"
"CH3OH + A_th O2=1 CO2+2 H2O  "
"1+ 2*A_th=1*2+2*1""theoretical O balance"
"Adiabatic, Incomplete Combustion of fuel CnHm entering at T_fuel with Stoichiometric Air at T_air:
Reaction:  CxHyOz  +  (y/4 + x-z/2) (Theo_air/100) (O2 + 3.76 N2)
            <--> (x-w)CO2 +wCO +  (y/2) H2O  +  3.76  (y/4 + x-z/2) (Theo_air/100)  N2 +  ((y/4 + x-z/2)
(Theo_air/100 - 1) +w/2)O2"

"T_prod is the adiabatic combustion temperature, assuming no dissociation.
Theo_air is the % theoretical air. "
"The initial guess value of T_prod = 450K ."

Procedure Fuel(Fuel$,T_fuel:x,y,z,h_fuel,Name$)
"This procedure takes the fuel name and returns the moles of C and moles of H"
If fuel$='C2H2(g)' then
          x=2;y=2; z=0
             Name$='acetylene'
                                                  h_fuel = 226730
else
If fuel$='C3H8(l)' then
          x=3; y=8; z=0
             Name$='propane(liq)'
                                                  h_fuel = -103850-15060
else
If fuel$='C8H18(l)' then
          x=8; y=18; z=0
            Name$='octane(liq)'
                                                  h_fuel = -249950
else
if fuel$='CH4(g)' then
          x=1; y=4; z=0
            Name$='methane'
                                                  h_fuel = enthalpy(CH4,T=T_fuel)
else
if fuel$='CH3OH(g)' then
          x=1; y=4; z=1
            Name$='methyl alcohol'
                                                  h_fuel = -200670

endif; endif; endif; endif; endif
end

Procedure Moles(x,y,z,Th_air,A_th:w,MolO2,SolMeth$)

ErrTh =(2*x + y/2 - z - x)/(2*A_th)*100

IF Th_air >= 1 then
SolMeth$ = '>= 100%, the solution assumes complete combustion.'
{MolCO = 0
MolCO2 = x}
```

```
w=0
MolO2 = A_th*(Th_air - 1)
GOTO 10
ELSE
 w = 2*x + y/2 - z - 2*A_th*Th_air
IF w > x  then
Call ERROR('The moles of CO2 are negative, the percent theoretical air must be >= xxxF3 %',ErrTh)
Else
SolMeth$ = '< 100%, the solution assumes incomplete combustion with no O_2 in products.'
MolO2 = 0
endif; endif
10:
END

{"Input data from the diagram window"
T_air = 298 [K]
Theo_air = 120 [%]
Fuel$='CH4(g)'}
T_fuel = 298 [K]

Call Fuel(Fuel$,T_fuel:x,y,z,h_fuel,Name$)
A_th =x +  y/4 - z/2
Th_air = Theo_air/100
Call Moles(x,y,z,Th_air,A_th:w,MolO2,SolMeth$)
HR=h_fuel+ (x+y/4-z/2) *(Theo_air/100) *enthalpy(O2,T=T_air)+3.76*(x+y/4-z/2) *(Theo_air/100)
*enthalpy(N2,T=T_air)
HP=HR  "Adiabatic"
HP=(x-w)*enthalpy(CO2,T=T_prod)+w*enthalpy(CO,T=T_prod)+(y/2)*enthalpy(H2O,T=T_prod)+3.76*(x+y/4-
z/2)* (Theo_air/100)*enthalpy(N2,T=T_prod)+MolO2*enthalpy(O2,T=T_prod)
Moles_O2=MolO2
Moles_N2=3.76*(x+y/4-z/2)* (Theo_air/100)
Moles_CO2=x-w
Moles_CO=w
Moles_H2O=y/2
```

SOLUTION for the sample calculation

```
A_th=5                        fuel$='C3H8(l)'
HP=-119035 [kJ/kg]            HR=-119035 [kJ/kg]
h_fuel=-118910                Moles_CO=0.000
Moles_CO2=3.000               Moles_H2O=4
Moles_N2=22.560               Moles_O2=1.000
MolO2=1                       Name$='propane(liq)'
SolMeth$='>= 100%, the solution assumes complete combustion.'
Theo_air=120 [%]              Th_air=1.200
T_air=298 [K]                 T_fuel=298 [K]
T_prod=2112 [K]               w=0
x=3                           y=8
z=0
```

15-125 The minimum percent of excess air that needs to be used for the fuels CH_4(g), C_2H_2(g), CH_3OH(g), C_3H_8(g), and C_8H_{18}(*l*) if the adiabatic flame temperature is not to exceed 1500 K is to be determined.

Analysis The problem is solved using EES, and the solution is given below.

```
Adiabatic Combustion of fuel CnHm entering at T_fuel with Stoichiometric Air at T_air:
Reaction:  CxHyOz  +  (y/4 + x-z/2) (Theo_air/100) (O2 + 3.76 N2)
             <--> xCO2  +  (y/2) H2O  +  3.76  (y/4 + x-z/2) (Theo_air/100)  N2 +  (y/4 + x-z/2) (Theo_air/100 - 1) O2"
{"For  theoretical oxygen, the complete combustion equation for CH3OH is"
"CH3OH + A_th O2=1 CO2+2 H2O  "
1+ 2*A_th=1*2+2*1"theoretical O balance"}

"T_prod is the adiabatic combustion temperature, assuming no dissociation.
Theo_air is the % theoretical air. "
"The initial guess value of T_prod = 450K ."

Procedure Fuel(Fuel$,T_fuel:x,y,z,h_fuel,Name$)
"This procedure takes the fuel name and returns the moles of C and moles of H"
If fuel$='C2H2(g)' then
          x=2;y=2; z=0
             Name$='acetylene'
                                                  h_fuel = 226730
else
If fuel$='C3H8(g)' then
          x=3; y=8; z=0
             Name$='propane'
                                                  h_fuel = enthalpy(C3H8,T=T_fuel)
else
If fuel$='C8H18(l)' then
          x=8; y=18; z=0
            Name$='octane'
                                                  h_fuel = -249950
else
if fuel$='CH4(g)' then
          x=1; y=4; z=0
            Name$='methane'
                                                  h_fuel = enthalpy(CH4,T=T_fuel)
else
if fuel$='CH3OH(g)' then
          x=1; y=4; z=1
            Name$='methyl alcohol'
                                                  h_fuel = -200670

endif; endif; endif; endif; endif
end

{"Input data from the diagram window"
T_air = 298 [K]
Fuel$='CH4(g)'}
T_fuel = 298 [K]

Excess_air=Theo_air - 100 "[%]"
Call Fuel(Fuel$,T_fuel:x,y,z,h_fuel,Name$)
A_th = y/4 + x-z/2
Th_air = Theo_air/100
HR=h_fuel+ (y/4 + x-z/2) *(Theo_air/100) *enthalpy(O2,T=T_air)+3.76*(y/4 + x-z/2) *(Theo_air/100)
*enthalpy(N2,T=T_air)
HP=HR  "Adiabatic"
```

```
HP=x*enthalpy(CO2,T=T_prod)+(y/2)*enthalpy(H2O,T=T_prod)+3.76*(y/4 + x-z/2)*
(Theo_air/100)*enthalpy(N2,T=T_prod)+(y/4 + x-z/2) *(Theo_air/100 - 1)*enthalpy(O2,T=T_prod)

Moles_O2=(y/4 + x-z/2) *(Theo_air/100 - 1)
Moles_N2=3.76*(y/4 + x-z/2)* (Theo_air/100)
Moles_CO2=x
Moles_H2O=y/2
T[1]=T_prod;  xa[1]=Theo_air
```

SOLUTION for a sample calculation

```
A_th=2.5                      Excess_air=156.251 [%]
fuel$='C2H2(g)'               HP=226596 [kJ/kg]
HR=226596 [kJ/kg]             h_fuel=226730
Moles_CO2=2                   Moles_H2O=1
Moles_N2=24.09                Moles_O2=3.906
Name$='acetylene'             Theo_air=256.3 [%]
Th_air=2.563                  T[1]=1500 [K]
T_air=298 [K]                 T_fuel=298 [K]
T_prod=1500 [K]               x=2
xa[1]=256.3                   y=2
z=0
```

15-126 The minimum percentages of excess air that need to be used for the fuels CH_4(g), C_2H_2(g), CH_3OH(g), C_3H_8(g), and C8H18(l) AFOR adiabatic flame temperatures of 1200 K, 1750 K, and 2000 K are to be determined.

Analysis The problem is solved using EES, and the solution is given below.

```
Adiabatic Combustion of fuel CnHm entering at T_fuel with Stoichiometric Air at T_air:
Reaction:  CxHyOz  +  (y/4 + x-z/2) (Theo_air/100) (O2 + 3.76 N2)
           <--> xCO2  +  (y/2) H2O  +  3.76  (y/4 + x-z/2) (Theo_air/100)  N2 +  (y/4 + x-z/2) (Theo_air/100 - 1) O2"
{"For  theoretical oxygen, the complete combustion equation for CH3OH is"
"CH3OH + A_th O2=1 CO2+2 H2O  "
1+ 2*A_th=1*2+2*1"theoretical O balance"}

"T_prod is the adiabatic combustion temperature, assuming no dissociation.
Theo_air is the % theoretical air. "
"The initial guess value of T_prod = 450K ."

Procedure Fuel(Fuel$,T_fuel:x,y,z,h_fuel,Name$)
"This procedure takes the fuel name and returns the moles of C and moles of H"
If fuel$='C2H2(g)' then
         x=2;y=2; z=0
            Name$='acetylene'
                                             h_fuel = 226730
else
If fuel$='C3H8(g)' then
         x=3; y=8; z=0
            Name$='propane'
                                             h_fuel = enthalpy(C3H8,T=T_fuel)
else
If fuel$='C8H18(l)' then
         x=8; y=18; z=0
           Name$='octane'
                                             h_fuel = -249950
else
if fuel$='CH4(g)' then
         x=1; y=4; z=0
           Name$='methane'
                                             h_fuel = enthalpy(CH4,T=T_fuel)
else
if fuel$='CH3OH(g)' then
         x=1; y=4; z=1
           Name$='methyl alcohol'
                                             h_fuel = -200670

endif; endif; endif; endif; endif
end

{"Input data from the diagram window"
T_air = 298 [K]
Fuel$='CH4(g)'}
T_fuel = 298 [K]

Excess_air=Theo_air - 100 "[%]"
Call Fuel(Fuel$,T_fuel:x,y,z,h_fuel,Name$)
A_th = y/4 + x-z/2
Th_air = Theo_air/100
HR=h_fuel+ (y/4 + x-z/2) *(Theo_air/100) *enthalpy(O2,T=T_air)+3.76*(y/4 + x-z/2) *(Theo_air/100)
*enthalpy(N2,T=T_air)
HP=HR  "Adiabatic"
```

```
HP=x*enthalpy(CO2,T=T_prod)+(y/2)*enthalpy(H2O,T=T_prod)+3.76*(y/4 + x-z/2)*
(Theo_air/100)*enthalpy(N2,T=T_prod)+(y/4 + x-z/2) *(Theo_air/100 - 1)*enthalpy(O2,T=T_prod)

Moles_O2=(y/4 + x-z/2) *(Theo_air/100 - 1)
Moles_N2=3.76*(y/4 + x-z/2)* (Theo_air/100)
Moles_CO2=x
Moles_H2O=y/2
T[1]=T_prod;  xa[1]=Theo_air
```

SOLUTION for a sample calculation

```
A_th=5                          Excess_air=31.395 [%]
fuel$='C3H8(g)'                 HP=-103995 [kJ/kg]
HR=-103995 [kJ/kg]              h_fuel=-103858
Moles_CO2=3                     Moles_H2O=4
Moles_N2=24.7                   Moles_O2=1.570
Name$='propane'                 Theo_air=131.4 [%]
Th_air=1.314                    T[1]=2000 [K]
T_air=298 [K]                   T_fuel=298 [K]
T_prod=2000 [K]                 x=3
xa[1]=131.4                     y=8
z=0
```

15-127 The adiabatic flame temperature of $CH_4(g)$ is to be determined when both the fuel and the air enter the combustion chamber at 25°C for the cases of 0, 20, 40, 60, 80, 100, 200, 500, and 1000 percent excess air.

Analysis The problem is solved using EES, and the solution is given below.

```
Adiabatic Combustion of fuel CnHm entering at T_fuel with Stoichiometric Air at T_air:
Reaction:  CxHyOz  +  (y/4 + x-z/2) (Theo_air/100) (O2 + 3.76 N2)
              <--> xCO2  +  (y/2) H2O  +  3.76  (y/4 + x-z/2) (Theo_air/100)  N2 +  (y/4 + x-z/2) (Theo_air/100 - 1) O2"
"For  theoretical oxygen, the complete combustion equation for CH3OH is"
"CH3OH + A_th O2=1 CO2+2 H2O  "
"1+ 2*A_th=1*2+2*1""theoretical O balance"
"Adiabatic, Incomplete Combustion of fuel CnHm entering at T_fuel with Stoichiometric Air at T_air:
Reaction:  CxHyOz  +  (y/4 + x-z/2) (Theo_air/100) (O2 + 3.76 N2)
              <--> (x-w)CO2 +wCO +  (y/2) H2O  +  3.76  (y/4 + x-z/2) (Theo_air/100)  N2 +  ((y/4 + x-z/2)
(Theo_air/100 - 1) +w/2)O2"
"T_prod is the adiabatic combustion temperature, assuming no dissociation.
Theo_air is the % theoretical air. "
"The initial guess value of T_prod = 450K ."

Procedure Fuel(Fuel$,T_fuel:x,y,z,h_fuel,Name$)
"This procedure takes the fuel name and returns the moles of C and moles of H"
If fuel$='C2H2(g)' then
            x=2;y=2; z=0
                Name$='acetylene'
                                                            h_fuel = 226730
else
If fuel$='C3H8(g)' then
            x=3; y=8; z=0
                Name$='propane'
                                                            h_fuel = enthalpy(C3H8,T=T_fuel)
else
If fuel$='C8H18(l)' then
            x=8; y=18; z=0
               Name$='octane'
                                                            h_fuel = -249950
else
if fuel$='CH4(g)' then
            x=1; y=4; z=0
               Name$='methane'
                                                            h_fuel = enthalpy(CH4,T=T_fuel)
else
if fuel$='CH3OH(g)' then
            x=1; y=4; z=1
               Name$='methyl alcohol'
                                                            h_fuel = -200670

endif; endif; endif; endif; endif
end

Procedure Moles(x,y,z,Th_air,A_th:w,MolO2,SolMeth$)

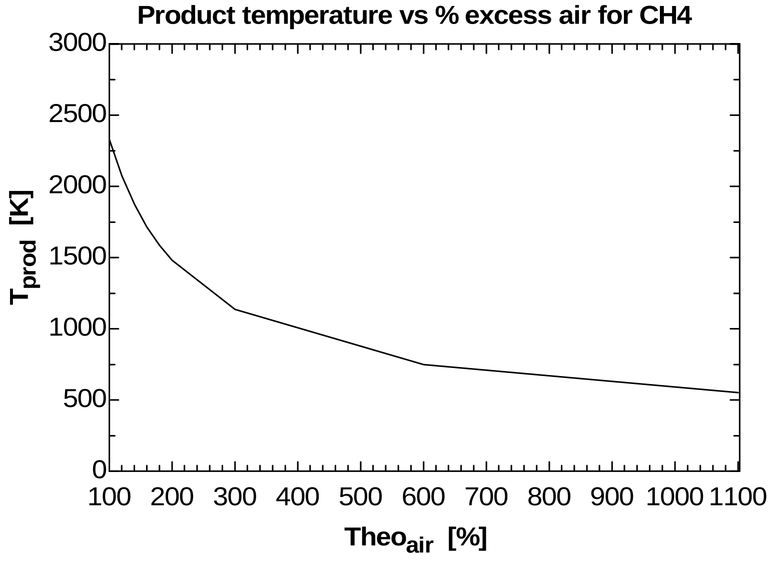
ErrTh =(2*x + y/2 - z - x)/(2*A_th)*100

IF Th_air >= 1 then
SolMeth$ = '>= 100%, the solution assumes complete combustion.'
{MolCO = 0
MolCO2 = x}
w=0
MolO2 = A_th*(Th_air - 1)
```

```
GOTO 10
ELSE
  w = 2*x + y/2 - z - 2*A_th*Th_air
IF w > x  then
Call ERROR('The moles of CO2 are negative, the percent theoretical air must be >= xxxF3 %',ErrTh)
Else
SolMeth$ = '< 100%, the solution assumes incomplete combustion with no O_2 in products.'
MolO2 = 0
endif; endif
10:
END

{"Input data from the diagram window"
T_air = 298 [K]
Theo_air = 200 [%]
Fuel$='CH4(g)'}
T_fuel = 298 [K]
Call Fuel(Fuel$,T_fuel:x,y,z,h_fuel,Name$)
A_th =x +  y/4 - z/2
Th_air = Theo_air/100
Call Moles(x,y,z,Th_air,A_th:w,MolO2,SolMeth$)
HR=h_fuel+ (x+y/4-z/2) *(Theo_air/100) *enthalpy(O2,T=T_air)+3.76*(x+y/4-z/2) *(Theo_air/100)
*enthalpy(N2,T=T_air)
HP=HR  "Adiabatic"
HP=(x-w)*enthalpy(CO2,T=T_prod)+w*enthalpy(CO,T=T_prod)+(y/2)*enthalpy(H2O,T=T_prod)+3.76*(x+y/4-
z/2)* (Theo_air/100)*enthalpy(N2,T=T_prod)+MolO2*enthalpy(O2,T=T_prod)
Moles_O2=MolO2
Moles_N2=3.76*(x+y/4-z/2)* (Theo_air/100)
Moles_CO2=x-w
Moles_CO=w
Moles_H2O=y/2
```

$Theo_{air}$ [%]	T_{prod} [K]
100	2329
120	2071
140	1872
160	1715
180	1587
200	1480
300	1137
600	749.5
1100	553



15-128 The fuel among CH_4(g), C_2H_2(g), C_2H_6(g), C_3H_8(g), and C_8H_{18}(l) that gives the highest temperature when burned completely in an adiabatic constant-volume chamber with the theoretical amount of air is to be determined.

Analysis The problem is solved using EES, and the solution is given below.

```
Adiabatic Combustion of fuel CnHm  with Stoichiometric Air at T_fuel =T_air=T_reac in a constant volume,
closed system:
Reaction:  CxHyOz  +  (x+y/4-z/2) (Theo_air/100) (O2 + 3.76 N2)
          --> xCO2  +  (y/2) H2O  +  3.76  (x+y/4-z/2) (Theo_air/100)  N2 +  (x+y/4-z/2) (Theo_air/100 - 1) O2"
"For  theoretical oxygen, the complete combustion equation for CH3OH is"
"CH3OH + A_th O2=1 CO2+2 H2O  "
"1+ 2*A_th=1*2+2*1""theoretical O balance"

"Adiabatic, Incomplete Combustion of fuel CnHm with Stoichiometric Air at T_fuel =T_air=T_reac in a constant
volume, closed system:
Reaction:  CxHyOz  +  (x+y/4-z/2) (Theo_air/100) (O2 + 3.76 N2)
    --> (x-w)CO2 +wCO +  (y/2) H2O  +  3.76  (x+y/4-z/2) (Theo_air/100)  N2 +  ((x+y/4-z/2) (Theo_air/100 - 1)
+w/2)O2"

"T_prod is the adiabatic combustion temperature, assuming no dissociation.
Theo_air is the % theoretical air. "

"The initial guess value of T_prod = 450K ."

Procedure Fuel(Fuel$,T_fuel:x,y,z,h_fuel,Name$)
"This procedure takes the fuel name and returns the moles of C and moles of H"
If fuel$='C2H2(g)' then
          x=2;y=2; z=0
              Name$='acetylene'
                                                       h_fuel = 226730"Table A.26"
else
If fuel$='C3H8(g)' then
          x=3; y=8; z=0
              Name$='propane'
                                                       h_fuel = enthalpy(C3H8,T=T_fuel)
else
If fuel$='C8H18(l)' then
          x=8; y=18; z=0
             Name$='octane'
                                                       h_fuel = -249950"Table A.26"
else
if fuel$='CH4(g)' then
          x=1; y=4; z=0
             Name$='methane'
                                                       h_fuel = enthalpy(CH4,T=T_fuel)
else
if fuel$='CH3OH(g)' then
          x=1; y=4; z=1
             Name$='methyl alcohol'
                                                       h_fuel = -200670"Table A.26"
endif; endif; endif; endif; endif
end

Procedure Moles(x,y,z,Th_air,A_th:w,MolO2,SolMeth$)
ErrTh =(2*x + y/2 - z - x)/(2*A_th)*100
IF Th_air >= 1 then
SolMeth$ = '>= 100%, the solution assumes complete combustion.'
w=0
MolO2 = A_th*(Th_air - 1)
```

```
GOTO 10
ELSE
  w = 2*x + y/2 - z - 2*A_th*Th_air
IF w > x  then
Call ERROR('The moles of CO2 are negative, the percent theoretical air must be >= xxxF3 %',ErrTh)
Else
SolMeth$ = '< 100%, the solution assumes incomplete combustion with no O_2 in products.'
MolO2 = 0
endif; endif
10:
END

{"Input data from the diagram window"
Theo_air = 200 [%]
Fuel$='CH4(g)'}

T_reac = 298  [K]
T_air = T_reac
T_fuel = T_reac
R_u = 8.314 [kJ/kmol-K]
Call Fuel(Fuel$,T_fuel:x,y,z,h_fuel,Name$)
A_th =x +  y/4 - z/2
Th_air = Theo_air/100
Call Moles(x,y,z,Th_air,A_th:w,MolO2,SolMeth$)
UR=(h_fuel-R_u*T_fuel)+ (x+y/4-z/2) *(Theo_air/100) *(enthalpy(O2,T=T_air)-R_u*T_air)+3.76*(x+y/4-z/2)
*(Theo_air/100) *(enthalpy(N2,T=T_air)-R_u*T_air)
UP=(x-w)*(enthalpy(CO2,T=T_prod)-R_u*T_prod)+w*(enthalpy(CO,T=T_prod)-
R_u*T_prod)+(y/2)*(enthalpy(H2O,T=T_prod)-R_u*T_prod)+3.76*(x+y/4-z/2)*
(Theo_air/100)*(enthalpy(N2,T=T_prod)-R_u*T_prod)+MolO2*(enthalpy(O2,T=T_prod)-R_u*T_prod)

UR =UP "Adiabatic, constant volume conservation of energy"
Moles_O2=MolO2
Moles_N2=3.76*(x+y/4-z/2)* (Theo_air/100)
Moles_CO2=x-w
Moles_CO=w
Moles_H2O=y/2
```

SOLUTION for CH4

```
A_th=2                      fuel$='CH4(g)'              h_fuel=-74875
Moles_CO=0.000              Moles_CO2=1.000             Moles_H2O=2
Moles_N2=7.520              Moles_O2=0.000              MolO2=0
Name$='methane'             R_u=8.314 [kJ/kmol-K]
SolMeth$='>= 100%, the solution assumes complete combustion.'
Theo_air=100 [%]            Th_air=1.000                T_air=298 [K]
T_fuel=298 [K]              T_prod=2824 [K]             T_reac=298 [K]
UP=-100981                  UR=-100981                  w=0
x=1                         y=4                         z=0
```

SOLUTION for C2H2

```
A_th=2.5                    fuel$='C2H2(g)'             h_fuel=226730
Moles_CO=0.000              Moles_CO2=2.000             Moles_H2O=1
Moles_N2=9.400              Moles_O2=0.000              MolO2=0
Name$='acetylene'           R_u=8.314 [kJ/kmol-K]
SolMeth$='>= 100%, the solution assumes complete combustion.'
Theo_air=100 [%]            Th_air=1.000                T_air=298 [K]
T_fuel=298 [K]              T_prod=3535 [K]             T_reac=298 [K]
UP=194717                   UR=194717                   w=0
x=2                         y=2                         z=0
```

SOLUTION for CH3OH

```
A_th=1.5                          fuel$='CH3OH(g)'                  h_fuel=-200670
Moles_CO=0.000                    Moles_CO2=1.000                   Moles_H2O=2
Moles_N2=5.640                    Moles_O2=0.000                    MolO2=0
Name$='methyl alcohol'            R_u=8.314 [kJ/kmol-K]
SolMeth$='>= 100%, the solution assumes complete combustion.'
Theo_air=100 [%]                  Th_air=1.000                      T_air=298 [K]
T_fuel=298 [K]                    T_prod=2817 [K]                   T_reac=298 [K]
UP=-220869                        UR=-220869                        w=0
x=1                               y=4                               z=1
```

SOLUTION for C3H8

```
A_th=5                            fuel$='C3H8(g)'                   h_fuel=-103858
Moles_CO=0.000                    Moles_CO2=3.000                   Moles_H2O=4
Moles_N2=18.800                   Moles_O2=0.000                    MolO2=0
Name$='propane'                   R_u=8.314 [kJ/kmol-K]
SolMeth$='>= 100%, the solution assumes complete combustion.'
Theo_air=100 [%]                  Th_air=1.000                      T_air=298 [K]
T_fuel=298 [K]                    T_prod=2909 [K]                   T_reac=298 [K]
UP=-165406                        UR=-165406                        w=0
x=3                               y=8                               z=0
```

SOLUTION for C8H18

```
A_th=12.5                         fuel$='C8H18(l)'                  h_fuel=-249950
Moles_CO=0.000                    Moles_CO2=8.000                   Moles_H2O=9
Moles_N2=47.000                   Moles_O2=0.000                    MolO2=0
Name$='octane'                    R_u=8.314 [kJ/kmol-K]
SolMeth$='>= 100%, the solution assumes complete combustion.'
Theo_air=100 [%]                  Th_air=1.000                      T_air=298 [K]
T_fuel=298 [K]                    T_prod=2911 [K]                   T_reac=298 [K]
UP=-400104                        UR=-400104                        w=0
x=8                               y=18                              z=0
```

Fundamentals of Engineering (FE) Exam Problems

15-129 A fuel is burned with 70 percent theoretical air. This is equivalent to

(a) 30% excess air (b) 70% excess air (c) 30% deficiency of air
(d) 70% deficiency of air (e) stoichiometric amount of air

Answer (c) 30% deficiency ofair

Solution Solved by EES Software. Solutions can be verified by copying-and-pasting the following lines on a blank EES screen. (Similar problems and their solutions can be obtained easily by modifying numerical values).

```
air_th=0.7
"air_th=air_access+1"
air_th=1-air_deficiency
```

15-130 Propane C_3H_8 is burned with 150 percent theoretical air. The air-fuel mass ratio for this combustion process is

(a) 5.3 (b) 10.5 (c) 15.7 (d) 23.4 (e) 39.3

Answer (d) 23.4

Solution Solved by EES Software. Solutions can be verified by copying-and-pasting the following lines on a blank EES screen. (Similar problems and their solutions can be obtained easily by modifying numerical values).

```
n_C=3
n_H=8
m_fuel=n_H*1+n_C*12
a_th=n_C+n_H/4
coeff=1.5  "coeff=1 for theoretical combustion, 1.5 for 50% excess air"
n_O2=coeff*a_th
n_N2=3.76*n_O2
m_air=n_O2*32+n_N2*28
AF=m_air/m_fuel
```

15-131 One kmol of methane (CH_4) is burned with an unknown amount of air during a combustion process. If the combustion is complete and there are 1 kmol of free O_2 in the products, the air-fuel mass ratio is

(a) 34.6 (b) 25.7 (c) 17.2 (d) 14.3 (e) 11.9

Answer (b) 25.7

Solution Solved by EES Software. Solutions can be verified by copying-and-pasting the following lines on a blank EES screen. (Similar problems and their solutions can be obtained easily by modifying numerical values).

```
n_C=1
n_H=4
m_fuel=n_H*1+n_C*12
a_th=n_C+n_H/4
(coeff-1)*a_th=1  "O2 balance: Coeff=1 for theoretical combustion, 1.5 for 50% excess air"
n_O2=coeff*a_th
n_N2=3.76*n_O2
m_air=n_O2*32+n_N2*28
AF=m_air/m_fuel

"Some Wrong Solutions with Common Mistakes:"
W1_AF=1/AF "Taking the inverse of AF"
W2_AF=n_O2+n_N2 "Finding air-fuel mole ratio"
W3_AF=AF/coeff "Ignoring excess air"
```

15-132 A fuel is burned steadily in a combustion chamber. The combustion temperature will be the highest except when

(a) the fuel is preheated.
(b) the fuel is burned with a deficiency of air.
(c) the air is dry.
(d) the combustion chamber is well insulated.
(e) the combustion is complete.

Answer (b) the fuel is burned with a deficiency of air.

15-133 An equimolar mixture of carbon dioxide and water vapor at 1 atm and 60°C enter a dehumidifying section where the entire water vapor is condensed and removed from the mixture, and the carbon dioxide leaves at 1 atm and 60°C. The entropy change of carbon dioxide in the dehumidifying section is

(a) –2.8 kJ/kg·K (b) –0.13 kJ/kg·K (c) 0 (d) 0.13 kJ/kg·K (e) 2.8 kJ/kg·K

Answer (b) –0.13 kJ/kg·K

Solution Solved by EES Software. Solutions can be verified by copying-and-pasting the following lines on a blank EES screen. (Similar problems and their solutions can be obtained easily by modifying numerical values).

```
Cp_CO2=0.846
R_CO2=0.1889
T1=60+273 "K"
T2=T1
P1= 1 "atm"
P2=1 "atm"
y1_CO2=0.5; P1_CO2=y1_CO2*P1
y2_CO2=1; P2_CO2=y2_CO2*P2
Ds_CO2=Cp_CO2*ln(T2/T1)-R_CO2*ln(P2_CO2/P1_CO2)

"Some Wrong Solutions with Common Mistakes:"
W1_Ds=0 "Assuming no entropy change"
W2_Ds=Cp_CO2*ln(T2/T1)-R_CO2*ln(P1_CO2/P2_CO2) "Using pressure fractions backwards"
```

15-134 Methane (CH_4) is burned completely with 80% excess air during a steady-flow combustion process. If both the reactants and the products are maintained at 25°C and 1 atm and the water in the products exists in the liquid form, the heat transfer from the combustion chamber per unit mass of methane is

(a) 890 MJ/kg (b) 802 MJ/kg (c) 75 MJ/kg (d) 56 MJ/kg (e) 50 MJ/kg

Answer (d) 56 MJ/kg

Solution Solved by EES Software. Solutions can be verified by copying-and-pasting the following lines on a blank EES screen. (Similar problems and their solutions can be obtained easily by modifying numerical values).

```
T= 25 "C"
P=1 "atm"
EXCESS=0.8
"Heat transfer in this case is the HHV at room temperature,"
HHV_CH4 =55.53 "MJ/kg"
LHV_CH4 =50.05 "MJ/kg"

"Some Wrong Solutions with Common Mistakes:"
W1_Q=LHV_CH4 "Assuming lower heating value"
W2_Q=EXCESS*hHV_CH4 "Assuming Q to be proportional to excess air"
```

15-135 The higher heating value of a hydrocarbon fuel C_nH_m with $m = 8$ is given to be 1560 MJ/kmol of fuel. Then its lower heating value is

(a) 1384 MJ/kmol (b) 1208 MJ/kmol (c) 1402 MJ/kmol (d) 1540 MJ/kmol (e) 1550 MJ/kmol

Answer (a) 1384 MJ/kmol

Solution Solved by EES Software. Solutions can be verified by copying-and-pasting the following lines on a blank EES screen. (Similar problems and their solutions can be obtained easily by modifying numerical values).

```
HHV=1560 "MJ/kmol fuel"
h_fg=2.4423 "MJ/kg, Enthalpy of vaporization of water at 25C"
n_H=8
n_water=n_H/2
m_water=n_water*18
LHV=HHV-h_fg*m_water

"Some Wrong Solutions with Common Mistakes:"
W1_LHV=HHV - h_fg*n_water  "Using mole numbers instead of mass"
W2_LHV= HHV - h_fg*m_water*2 "Taking mole numbers of H2O to be m instead of m/2"
W3_LHV= HHV - h_fg*n_water*2 "Taking mole numbers of H2O to be m instead of m/2, and using mole
numbers"
```

15-136 Acetylene gas (C_2H_2) is burned completely during a steady-flow combustion process. The fuel and the air enter the combustion chamber at 25°C, and the products leave at 1500 K. If the enthalpy of the products relative to the standard reference state is –404 MJ/kmol of fuel, the heat transfer from the combustion chamber is

(a) 177 MJ/kmol (b) 227 MJ/kmol (c) 404 MJ/kmol (d) 631 MJ/kmol (e) 751 MJ/kmol

Answer (d) 631 MJ/kmol

Solution Solved by EES Software. Solutions can be verified by copying-and-pasting the following lines on a blank EES screen. (Similar problems and their solutions can be obtained easily by modifying numerical values).

```
hf_fuel=226730/1000 "MJ/kmol fuel"
H_prod=-404 "MJ/kmol fuel"
H_react=hf_fuel
Q_out=H_react-H_prod

"Some Wrong Solutions with Common Mistakes:"
W1_Qout= -H_prod "Taking Qout to be H_prod"
W2_Qout= H_react+H_prod "Adding enthalpies instead of subtracting them"
```

15-137 Benzene gas (C_6H_6) is burned with 95 percent theoretical air during a steady-flow combustion process. The mole fraction of the CO in the products is

(a) 8.3% (b) 4.7% (c) 2.1% (d) 1.9% (e) 14.3%

Answer (c) 2.1%

Solution Solved by EES Software. Solutions can be verified by copying-and-pasting the following lines on a blank EES screen. (Similar problems and their solutions can be obtained easily by modifying numerical values).

```
n_C=6
n_H=6
a_th=n_C+n_H/4
coeff=0.95  "coeff=1 for theoretical combustion, 1.5 for 50% excess air"
"Assuming all the H burns to H2O, the combustion equation is
   C6H6+coeff*a_th(O2+3.76N2)----- (n_CO2) CO2+(n_CO)CO+(n_H2O) H2O+(n_N2) N2"
n_O2=coeff*a_th
n_N2=3.76*n_O2
n_H2O=n_H/2
n_CO2+n_CO=n_C
2*n_CO2+n_CO+n_H2O=2*n_O2  "Oxygen balance"
n_prod=n_CO2+n_CO+n_H2O+n_N2  "Total mole numbers of product gases"
y_CO=n_CO/n_prod "mole fraction of CO in product gases"

"Some Wrong Solutions with Common Mistakes:"
W1_yCO=n_CO/n1_prod; n1_prod=n_CO2+n_CO+n_H2O "Not including N2 in n_prod"
W2_yCO=(n_CO2+n_CO)/n_prod "Using both CO and CO2 in calculations"
```

15-138 A fuel is burned during a steady-flow combustion process. Heat is lost to the surroundings at 300 K at a rate of 1120 kW. The entropy of the reactants entering per unit time is 17 kW/K and that of the products is 15 kW/K. The total rate of exergy destruction during this combustion process is

(a) 520 kW (b) 600 kW (c) 1120 kW (d) 340 kW (e) 739 kW

Answer (a) 520 kW

Solution Solved by EES Software. Solutions can be verified by copying-and-pasting the following lines on a blank EES screen. (Similar problems and their solutions can be obtained easily by modifying numerical values).

```
To=300 "K"
Q_out=1120 "kW"
S_react=17 "kW'K"
S_prod= 15 "kW/K"
S_react-S_prod-Q_out/To+S_gen=0 "Entropy balance for steady state operation, Sin-Sout+Sgen=0"
X_dest=To*S_gen

"Some Wrong Solutions with Common Mistakes:"
W1_Xdest=S_gen "Taking Sgen as exergy destruction"
W2_Xdest=To*S_gen1; S_react-S_prod-S_gen1=0 "Ignoring Q_out/To"
```

15-139 ··· 15-144 Design and Essay Problems

15-139 A certain industrial process generates a liquid solution of ethanol and water as the waste product. The solution is to be burned using methane. A combustion process is to be developed to accomplish this incineration process with minimum amount of methane.

Analysis The mass flow rate of the liquid ethanol-water solution is given to be 10 kg/s. Considering that the mass fraction of ethanol in the solution is 0.2,

$$\dot{m}_{\text{ethanol}} = (0.2)(10\text{ kg/s}) = 2\text{ kg/s}$$
$$\dot{m}_{\text{water}} = (0.8)(10\text{ kg/s}) = 8\text{ kg/s}$$

Noting that the molar masses $M_{\text{ethanol}} = 46$ and $M_{\text{water}} = 18$ kg/kmol and that mole numbers $N = m/M$, the mole flow rates become

$$\dot{N}_{\text{ethanol}} = \frac{\dot{m}_{\text{ethanol}}}{M_{\text{ethanol}}} = \frac{2\text{ kg/s}}{46\text{ kg/kmol}} = 0.04348\text{ kmol/s}$$
$$\dot{N}_{\text{water}} = \frac{\dot{m}_{\text{water}}}{M_{\text{water}}} = \frac{8\text{ kg/s}}{18\text{ kg/kmol}} = 0.44444\text{ kmol/s}$$

Note that

$$\frac{\dot{N}_{\text{water}}}{\dot{N}_{\text{ethanol}}} = \frac{0.44444}{0.04348} = 10.222\text{ kmol H}_2\text{O/kmol C}_2\text{H}_5\text{OH}$$

That is, 10.222 moles of liquid water is present in the solution for each mole of ethanol.

Assuming complete combustion, the combustion equation of C_2H_5OH (ℓ) with stoichiometric amount of air is

$$C_2H_5OH(\ell) + a_{\text{th}}(O_2 + 3.76N_2) \longrightarrow 2CO_2 + 3H_2O + 3.76a_{\text{th}}N_2$$

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

$$1 + 2a_{\text{th}} = 4 + 3 \longrightarrow a_{\text{th}} = 3$$

Thus,

$$C_2H_5OH(\ell) + 3(O_2 + 3.76N_2) \longrightarrow 2CO_2 + 3H_2O + 11.28N_2$$

Noting that 10.222 kmol of liquid water accompanies each kmol of ethanol, the actual combustion equation can be written as

$$C_2H_5OH(\ell) + 3(O_2 + 3.76N_2) + 10.222H_2O(\ell) \longrightarrow 2CO_2 + 3H_2O(g) + 11.28N_2 + 10.222H_2O(l)$$

The heat transfer for this combustion process is determined from the steady-flow energy balance equation with $W = 0$,

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

Assuming the air and the combustion products to be ideal gases, we have $h = h(T)$. We assume all the reactants to enter the combustion chamber at the standard reference temperature of 25°C. Furthermore, we assume the products to leave the combustion chamber at 1400 K which is a little over the required temperature of 1100°C. From the tables,

Substance	$\bar{h}_f^\circ$ kJ/kmol	$\bar{h}_{298\ K}$ kJ/kmol	$\bar{h}_{1400\ K}$ kJ/kmol
C_2H_5OH (ℓ)	-277,690	---	---
CH_4	-74,850	---	---
O_2	0	8682	45,648
N_2	0	8669	43,605
H_2O (*g*)	-241,820	9904	53,351
H_2O (ℓ)	-285,830	---	---
CO_2	-393,520	9364	65,271

Thus,

$$\begin{aligned} Q &= (2)(-393{,}520+65{,}271-9364)+(3)(-241{,}820+53{,}351-9904) \\ &\quad +(11.28)(0+43{,}605-8669)-(1)(-277{,}690)-0-0 \\ &\quad +(10.222)(-241{,}820+53{,}351-9904)-(10.222)(-285{,}830) \\ &= 295{,}409 \text{ kJ/kmol of } C_2H_5OH \end{aligned}$$

The positive sign indicates that 295,409 kJ of heat must be supplied to the combustion chamber from another source (such as burning methane) to ensure that the combustion products will leave at the desired temperature of 1400 K. Then the rate of heat transfer required for a mole flow rate of 0.04348 kmol C_2H_5OH/s CO becomes

$$\dot{Q} = \dot{N}Q = (0.04348 \text{ kmol/s})(295{,}409 \text{ kJ/kmol}) = 12{,}844 \text{ kJ/s}$$

Assuming complete combustion, the combustion equation of $CH_4(g)$ with stoichiometric amount of air is

$$CH_4 + a_{th}(O_2 + 3.76N_2) \longrightarrow CO_2 + 2H_2O + 3.76a_{th}N_2$$

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

Thus,

$$a_{th} = 1+1 \longrightarrow a_{th} = 2$$

$$CH_4 + 2(O_2 + 3.76N_2) \longrightarrow CO_2 + 2H_2O + 7.52N_2$$

The heat transfer for this combustion process is determined from the steady-flow energy balance $E_{in} - E_{out} = \Delta E_{system}$ equation as shown above under the same assumptions and using the same mini table:

$$\begin{aligned} Q &= (1)(-393{,}520+65{,}271-9364)+(2)(-241{,}820+53{,}351-9904) \\ &\quad +(7.52)(0+43{,}605-8669)-(1)(-74{,}850)-0-0 \\ &= -396{,}790 \text{ kJ/kmol of } CH_4 \end{aligned}$$

That is, 396,790 kJ of heat is supplied to the combustion chamber for each kmol of methane burned. To supply heat at the required rate of 12,844 kJ/s, we must burn methane at a rate of

$$\dot{N}_{CH_4} = \frac{\dot{Q}}{Q} = \frac{12{,}844 \text{ kJ/s}}{396{,}790 \text{ kJ/kmol}} = 0.03237 \text{ kmolCH}_4\text{/s}$$

or,

$$\dot{m}_{CH_4} = M_{CH_4}\dot{N}_{CH_4} = (16 \text{ kg/kmol})(0.03237 \text{ kmolCH}_4\text{/s}) = \mathbf{0.5179\ \ kg/s}$$

Therefore, we must supply methane to the combustion chamber at a minimum rate 0.5179 kg/s in order to maintain the temperature of the combustion chamber above 1400 K.

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