Chapter 15

CHEMICAL REACTIONS

Fuels and Combustion

15-1C Gasoline is C_8H_{18}, diesel fuel is $C_{12}H_{26}$, and natural gas is CH_4.

15-2C 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.

15-3C 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.

15-4C 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.

15-5C The number of atoms are preserved during a chemical reaction, but the total mole numbers are not.

15-6C 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.

15-7C No. Because the molar mass of the fuel and the molar mass of the air, in general, are different.

15-8 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

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

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

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

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

Then,

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

and

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

15-9E 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

$$\mathrm{CH}_4 + 2\mathrm{O}_2 \longrightarrow \mathrm{CO}_2 + 2\mathrm{H}_2\mathrm{O}$$

$$\boxed{\begin{array}{l}\mathrm{CH}_4 + 2\mathrm{O}_2 \\ \to \mathrm{CO}_2 + 2\mathrm{H}_2\mathrm{O}\end{array}}$$

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

15-10C The causes of incomplete combustion are insufficient time, insufficient oxygen, insufficient mixing, and dissociation.

15-11C 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.

15-12C It represent the amount of air that contains the exact amount of oxygen needed for complete combustion.

15-13C No. The theoretical combustion is also complete, but the products of theoretical combustion does not contain any uncombined oxygen.

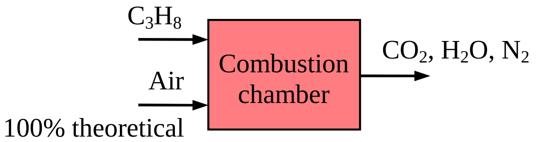
15-14C Case (b).

15-15 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

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

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

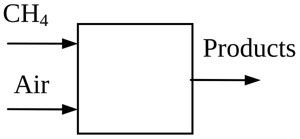
$$y_{\mathrm{H2O}} = \frac{N_{\mathrm{H2O}}}{N_{\mathrm{products}}} = \frac{N_{\mathrm{H2O}}}{N_{\mathrm{CO2}} + N_{\mathrm{H2O}} + N_{\mathrm{N2}}} = \frac{4\ \mathrm{kmol}}{3\ \mathrm{kmol} + 4\ \mathrm{kmol} + 18.8\ \mathrm{kmol}} = \frac{4\ \mathrm{kmol}}{25.8\ \mathrm{kmol}} = \mathbf{0.155}$$

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

15-16 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



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

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

Substituting, $\quad \mathrm{CH_4} + 2\left[\mathrm{O_2} + 3.76\mathrm{N_2}\right] \longrightarrow \mathrm{CO_2} + 2\mathrm{H_2O} + 7.52\mathrm{N_2}$

The masses of the reactants are

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

The total mass is

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

Then the mass fractions are

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

For a mixture flow of 0.01 kg/s, the mass flow rates of the reactants are

$$\begin{aligned} \dot{m}_{\mathrm{CH4}} &= \mathrm{mf}_{\mathrm{CH4}}\dot{m} = (0.05498)(0.01\mathrm{kg/s}) = \mathbf{0.0005498\ kg/s} \\ \dot{m}_{\mathrm{air}} &= \dot{m} - \dot{m}_{\mathrm{CH4}} = 0.01 - 0.0005498 = \mathbf{0.009450\ kg/s} \end{aligned}$$

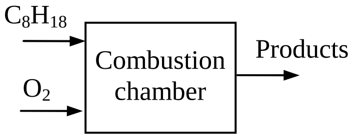
15-17 n-Octane is burned with stoichiometric amount of oxygen. The mass fractions of each of the products and the mass of water in the products per unit mass of fuel burned are to be determined.

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

$$C_8H_{18} + 12.5O_2 \longrightarrow 8CO_2 + 9H_2O$$



The mass of each product and the total mass are

$$\begin{aligned} m_{CO2} &= N_{CO2}M_{CO2} = (8\text{ kmol})(44\text{ kg/kmol}) = 352\text{ kg} \\ m_{H2O} &= N_{H2O}M_{H2O} = (9\text{ kmol})(18\text{ kg/kmol}) = 162\text{ kg} \\ m_{total} &= m_{CO2} + m_{H2O} = 352 + 162 = 514\text{ kg} \end{aligned}$$

Then the mass fractions are

$$\text{mf}_{CO2} = \frac{m_{CO2}}{m_{total}} = \frac{352\text{ kg}}{514\text{ kg}} = \mathbf{0.6848}$$

$$\text{mf}_{H2O} = \frac{m_{H2O}}{m_{total}} = \frac{162\text{ kg}}{514\text{ kg}} = \mathbf{0.3152}$$

The mass of water in the products per unit mass of fuel burned is determined from

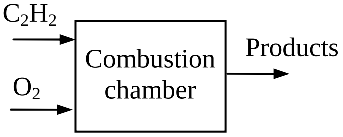
$$\frac{m_{H2O}}{m_{C8H18}} = \frac{(9\times 18)\text{ kg}}{(1\times 114)\text{ kg}} = \mathbf{1.42\ kg\ H_2O/kg\ C_8H_{18}}$$

15-18 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

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

15-19 Coal whose mass percentages are specified is burned with 50 percent excess air. The fuel-air ratio is to be determined.

Assumptions **1** Combustion is complete. **2** The combustion products contain CO_2, H_2O, SO_2, N_2, and O_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 The mass fractions of the constituent of the coal when the ash is substituted are

$$\mathrm{mf}_{\mathrm{C}} = \frac{m_{\mathrm{C}}}{m_{\mathrm{total}}} = \frac{79.61\ \mathrm{kg}}{(100-8.62)\ \mathrm{kg}} = \frac{79.61\ \mathrm{kg}}{91.38\ \mathrm{kg}} = 0.8712$$

$$\mathrm{mf}_{\mathrm{H2}} = \frac{m_{\mathrm{H2}}}{m_{\mathrm{total}}} = \frac{4.66\ \mathrm{kg}}{91.38\ \mathrm{kg}} = 0.05100$$

$$\mathrm{mf}_{\mathrm{O2}} = \frac{m_{\mathrm{O2}}}{m_{\mathrm{total}}} = \frac{4.76\ \mathrm{kg}}{91.38\ \mathrm{kg}} = 0.05209$$

$$\mathrm{mf}_{\mathrm{N2}} = \frac{m_{\mathrm{N2}}}{m_{\mathrm{total}}} = \frac{1.83\ \mathrm{kg}}{91.38\ \mathrm{kg}} = 0.02003$$

$$\mathrm{mf}_{\mathrm{S}} = \frac{m_{\mathrm{S}}}{m_{\mathrm{total}}} = \frac{0.52\ \mathrm{kg}}{91.38\ \mathrm{kg}} = 0.00569$$

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

We now consider 100 kg of this mixture. Then the mole numbers of each component are

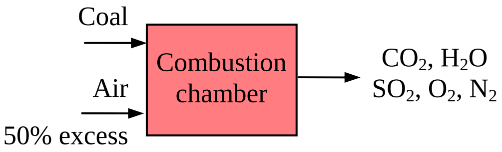
$$N_{\mathrm{C}} = \frac{m_{\mathrm{C}}}{M_{\mathrm{C}}} = \frac{87.12\ \mathrm{kg}}{12\ \mathrm{kg/kmol}} = 7.26\ \mathrm{kmol}$$

$$N_{\mathrm{H2}} = \frac{m_{\mathrm{H2}}}{M_{\mathrm{H2}}} = \frac{5.10\ \mathrm{kg}}{2\ \mathrm{kg/kmol}} = 2.55\ \mathrm{kmol}$$

$$N_{\mathrm{O2}} = \frac{m_{\mathrm{O2}}}{M_{\mathrm{O2}}} = \frac{5.209\ \mathrm{kg}}{32\ \mathrm{kg/kmol}} = 0.1628\ \mathrm{kmol}$$

$$N_{\mathrm{N2}} = \frac{m_{\mathrm{N2}}}{M_{\mathrm{N2}}} = \frac{2.003\ \mathrm{kg}}{28\ \mathrm{kg/kmol}} = 0.07154\ \mathrm{kmol}$$

$$N_{\mathrm{S}} = \frac{m_{\mathrm{S}}}{M_{\mathrm{S}}} = \frac{0.569\ \mathrm{kg}}{32\ \mathrm{kg/kmol}} = 0.01778\ \mathrm{kmol}$$



The mole number of the mixture and the mole fractions are

$$N_m = 7.26 + 2.55 + 0.1628 + 0.07154 + 0.01778 = 10.06\ \mathrm{kmol}$$

$$y_{\mathrm{C}} = \frac{N_{\mathrm{C}}}{N_m} = \frac{7.26\ \mathrm{kmol}}{10.06\ \mathrm{kmol}} = 0.7215$$

$$y_{\mathrm{H2}} = \frac{N_{\mathrm{H2}}}{N_m} = \frac{2.55\ \mathrm{kmol}}{10.06\ \mathrm{kmol}} = 0.2535$$

$$y_{\mathrm{O2}} = \frac{N_{\mathrm{O2}}}{N_m} = \frac{0.1628\ \mathrm{kmol}}{10.06\ \mathrm{kmol}} = 0.01618$$

$$y_{\mathrm{N2}} = \frac{N_{\mathrm{N2}}}{N_m} = \frac{0.07154\ \mathrm{kmol}}{10.06\ \mathrm{kmol}} = 0.00711$$

$$y_{\mathrm{S}} = \frac{N_{\mathrm{S}}}{N_m} = \frac{0.01778\ \mathrm{kmol}}{10.06\ \mathrm{kmol}} = 0.00177$$

Then, the combustion equation in this case may be written as

$$0.7215\text{C} + 0.2535\text{H}_2 + 0.01618\text{O}_2 + 0.00711\text{N}_2 + 0.00177\text{S} + 1.5a_{\text{th}}(\text{O}_2 + 3.76\text{N}_2)$$
$$\longrightarrow x\text{CO}_2 + y\text{H}_2\text{O} + z\text{SO}_2 + k\text{N}_2 + 0.5a_{th}\text{O}_2$$

According to the species balances,

$$\text{C balance:} \quad x = 0.7215$$
$$\text{H}_2 \text{ balance:} \quad y = 0.2535$$
$$\text{S balance:} \quad z = 0.00177$$
$$\text{O}_2 \text{ balance:}$$
$$0.01618 + 1.5a_{\text{th}} = x + 0.5y + z + 0.5a_{\text{th}}$$
$$1.5a_{\text{th}} - 0.5a_{\text{th}} = 0.7215 + 0.5(0.2535) + 0.00177 - 0.01617 \longrightarrow a_{\text{th}} = 0.8339$$
$$\text{N}_2 \text{ balance:} \quad 0.00711 + 1.5 \times 3.76a_{\text{th}} = k \longrightarrow k = 0.00711 + 1.5 \times 3.76 \times 0.8339 = 4.710$$

Substituting,

$$0.7215\text{C} + 0.2535\text{H}_2 + 0.01617\text{O}_2 + 0.00711\text{N}_2 + 0.00177\text{S} + 1.2509(\text{O}_2 + 3.76\text{N}_2)$$
$$\longrightarrow 0.7215\text{CO}_2 + 0.2535\text{H}_2\text{O} + 0.0018\text{SO}_2 + 4.71\text{N}_2 + 0.4170\text{O}_2$$

The fuel-air mass ratio is then

$$\text{FA} = \frac{m_{\text{fuel}}}{m_{\text{air}}} = \frac{(0.7215 \times 12 + 0.2535 \times 2 + 0.01617 \times 32 + 0.00711 \times 28 + 0.00177 \times 32)\text{ kg}}{(1.2509 \times 4.76 \times 29)\text{ kg}}$$
$$= \frac{9.938\text{ kg}}{172.5\text{ kg}} = \mathbf{0.0576\ kg\ fuel/kg\ air}$$

15-20 Propane is burned with 75 percent excess air during a combustion process. The AF ratio is to be determined.

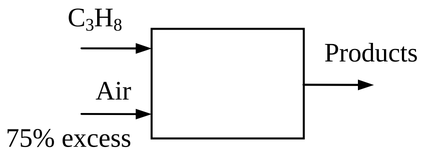
Assumptions **1** Combustion is complete. **2** The combustion products contain CO_2, H_2O, O_2, and N_2 only.

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

Analysis The combustion equation in this case can be written as

$$C_3H_8 + 1.75a_{th}\left[O_2 + 3.76N_2\right] \longrightarrow 3CO_2 + 4H_2O + 0.75a_{th}O_2 + (1.75\times 3.76)a_{th}N_2$$

where a_{th} is the stoichiometric coefficient for air. We have automatically accounted for the 75% excess air by using the factor $1.75a_{th}$ 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 ($0.75a_{th}O_2$) will appear in the products as free oxygen. The coefficient a_{th} is determined from the O_2 balance,



O_2 balance: $$1.75a_{th} = 3 + 2 + 0.75a_{th} \longrightarrow a_{th} = 5$$

Substituting,

$$C_3H_8 + 8.75\left[O_2 + 3.76N_2\right] \longrightarrow 3CO_2 + 4H_2O + 3.75O_2 + 32.9N_2$$

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

$$\text{AF} = \frac{m_{\text{air}}}{m_{\text{fuel}}} = \frac{(8.75\times 4.76\text{ kmol})(29\text{ kg/kmol})}{(3\text{ kmol})(12\text{ kg/kmol}) + (4\text{ kmol})(2\text{ kg/kmol})} = \mathbf{27.5\ kg\ air/kg\ fuel}$$

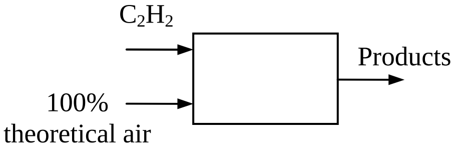
15-21 Acetylene is burned with the stoichiometric amount of air during a combustion process. The AF ratio is to be determined on a mass and on a mole basis.

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 This is a theoretical combustion process since C_2H_2 is burned completely with stoichiometric amount of air. The stoichiometric combustion equation of C_2H_2 is

$$C_2H_2 + a_{th}\left[O_2 + 3.76N_2\right] \longrightarrow 2CO_2 + H_2O + 3.76a_{th}N_2$$



O_2 balance: $a_{th} = 2 + 0.5 \longrightarrow a_{th} = 2.5$

Substituting,

$$C_2H_2 + 2.5\left[O_2 + 3.76N_2\right] \longrightarrow 2CO_2 + H_2O + 9.4N_2$$

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

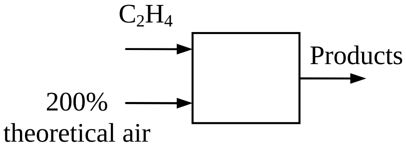
$$\text{AF} = \frac{m_{\text{air}}}{m_{\text{fuel}}} = \frac{(2.5\times 4.76\text{ kmol})(29\text{ kg/kmol})}{(2\text{ kmol})(12\text{ kg/kmol}) + (1\text{ kmol})(2\text{ kg/kmol})} = \mathbf{13.3\ \ kg\ air/kg\ fuel}$$

On a mole basis, the air-fuel ratio is expressed as the ratio of the mole numbers of the air to the mole numbers of the fuel,

$$\text{AF}_{\text{mole basis}} = \frac{N_{\text{air}}}{N_{\text{fuel}}} = \frac{(2.5\times 4.76)\text{ kmol}}{1\text{ kmol fuel}} = \mathbf{11.9\ kmol\ air/kmol\ fuel}$$

15-22E Ethylene is burned with 200 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

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

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

O_2 balance: $\quad 2a_{th} = 2 + 1 + a_{th} \longrightarrow a_{th} = 3$

Substituting,

$$C_2H_4 + 6[O_2 + 3.76N_2] \longrightarrow 2CO_2 + 2H_2O + 3O_2 + 22.56N_2$$

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

$$\mathrm{AF} = \frac{m_{air}}{m_{fuel}} = \frac{(6\times 4.76 \text{ lbmol})(29 \text{ lbm/lbmol})}{(2 \text{ lbmol})(12 \text{ lbm/lbmol}) + (2 \text{ lbmol})(2 \text{ lbm/lbmol})} = \mathbf{29.6 \text{ lbm air/lbm fuel}}$$

(*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,

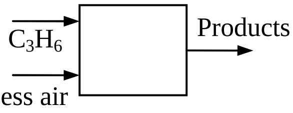
$$P_v = \left(\frac{N_v}{N_{prod}}\right)P_{prod} = \left(\frac{2 \text{ lbmol}}{29.56 \text{ lbmol}}\right)(14.5 \text{ psia}) = 0.981 \text{ psia}$$

Thus,

$$T_{dp} = T_{sat@0.981 \text{ psia}} = \mathbf{101°F}$$

15-23 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

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

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

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

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,

$$\mathrm{AF} = \frac{m_{\mathrm{air}}}{m_{\mathrm{fuel}}} = \frac{(6.75\times 4.76\text{ kmol})(29\text{ kg/kmol})}{(3\text{ kmol})(12\text{ kg/kmol}) + (3\text{ kmol})(2\text{ kg/kmol})} = \mathbf{22.2\ kg\ air/kg\ fuel}$$

(*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,

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

Thus,

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

15-24 Butane C_4H_{10} is burned with 200 percent theoretical air. The kmol of water that needs to be sprayed into the combustion chamber per kmol of fuel is to be determined.

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

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

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

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

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

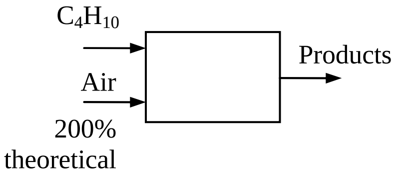
Carbon balance: $B = 4$

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

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

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

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



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

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

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

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

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

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

The vapor mole fraction is

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

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

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

15-25 A fuel mixture of 20% by mass methane, CH_4, and 80% 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 The combustion equation in this case can be written as

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

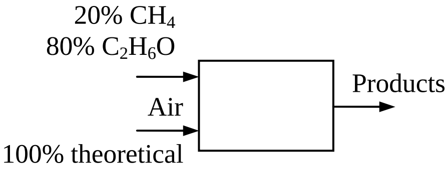
where a_{th} is the stoichiometric coefficient for air. The coefficient a_{th} and other coefficients are to be determined from the mass balances

Carbon balance: $x + 2y = B$

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

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

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



Solving the above equations, we find the coefficients as

$$x = 0.4182 \qquad B = 1.582$$
$$y = 0.5818 \qquad D = 2.582$$
$$a_{\mathrm{th}} = 2.582 \qquad F = 9.708$$

Then, we write the balanced reaction equation as

$$0.4182\,\mathrm{CH_4} + 0.5818\,\mathrm{C_2H_6O} + 2.582\left[\mathrm{O_2} + 3.76\mathrm{N_2}\right] \longrightarrow 1.582\,\mathrm{CO_2} + 2.582\,\mathrm{H_2O} + 9.708\,\mathrm{N_2}$$

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

$$\begin{aligned}\mathrm{AF} &= \frac{m_{\mathrm{air}}}{m_{\mathrm{fuel}}}\\ &= \frac{(2.582\times 4.76\ \mathrm{kmol})(29\ \mathrm{kg/kmol})}{(0.4182\ \mathrm{kmol})(12+4\times 1)\mathrm{kg/kmol} + (0.5818\ \mathrm{kmol})(2\times 12 + 6\times 1 + 16)\mathrm{kg/kmol}}\\ &= 10.64\ \mathrm{kg\ air/kg\ fuel}\end{aligned}$$

Then, the required flow rate of air becomes

$$\dot{m}_{\mathrm{air}} = \mathrm{AF}\dot{m}_{\mathrm{fuel}} = (10.64)(31\ \mathrm{kg/s}) = \mathbf{330\ kg/s}$$

15-26 Octane is burned with 250 percent theoretical air during a combustion process. The AF ratio and the dew-pint temperature of the products are to be determined.

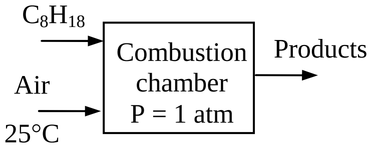
Assumptions **1** Combustion is complete. **2** The combustion products contain CO_2, H_2O, O_2, and N_2 only. **3** Combustion gases are ideal gases.

Properties The molar masses of C, H_2, and air are 12 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

$$C_8H_{18} + 2.5a_{th}\left[O_2 + 3.76N_2\right] \longrightarrow 8CO_2 + 9H_2O + 1.5a_{th}O_2 + (2.5\times 3.76)a_{th}N_2$$

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



O_2 balance: $2.5a_{th} = 8 + 4.5 + 1.5a_{th} \longrightarrow a_{th} = 12.5$

Substituting,

$$C_8H_{18} + 31.25\left[O_2 + 3.76N_2\right] \rightarrow 8CO_2 + 9H_2O + 18.75O_2 + 117.5N_2$$

Thus,

$$\text{AF} = \frac{m_{\text{air}}}{m_{\text{fuel}}} = \frac{(31.25\times 4.76\text{ kmol})(29\text{ kg/kmol})}{(8\text{ kmol})(12\text{ kg/kmol}) + (9\text{ kmol})(2\text{ kg/kmol})} = \mathbf{37.8\ kg\ air/kg\ fuel}$$

(*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,

$$P_v = \left(\frac{N_v}{N_{\text{prod}}}\right)P_{\text{prod}} = \left(\frac{9\text{ kmol}}{153.25\text{ kmol}}\right)(101.325\text{ kPa}) = 5.951\text{ kPa}$$

Thus,

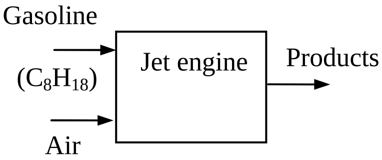
$$T_{\text{dp}} = T_{\text{sat@5.951 kPa}} = \mathbf{36.0°C}$$

15-27 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



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

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

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

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,

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

Then the percent theoretical air used can be determined from

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

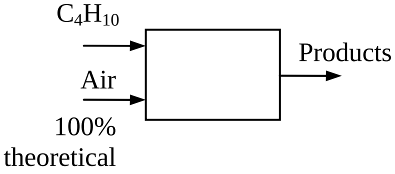
15-28 n-butane is burned with stoichiometric amount of air. The mass fraction of each product, the mass of CO_2 and air per unit mass of fuel burned are to be determined.

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

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



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: $B = 4$

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

Oxygen balance: $2a_{\mathrm{th}} = 2B + D \longrightarrow a_{\mathrm{th}} = 0.5(2\times4+5) = 6.5$

Nitrogen balance: $a_{\mathrm{th}} \times 3.76 = E \longrightarrow E = 6.5\times3.76 = 24.44$

Substituting, the balanced reaction equation is

$$\mathrm{C_4H_{10}} + 6.5[\mathrm{O_2} + 3.76\mathrm{N_2}] \longrightarrow 4\,\mathrm{CO_2} + 5\,\mathrm{H_2O} + 24.44\,\mathrm{N_2}$$

The mass of each product and the total mass are

$$\begin{aligned} m_{\mathrm{CO2}} &= N_{\mathrm{CO2}}M_{\mathrm{CO2}} = (4\text{ kmol})(44\text{ kg/kmol}) = 176\text{ kg} \\ m_{\mathrm{H2O}} &= N_{\mathrm{H2O}}M_{\mathrm{H2O}} = (5\text{ kmol})(18\text{ kg/kmol}) = 90\text{ kg} \\ m_{\mathrm{N2}} &= N_{\mathrm{N2}}M_{\mathrm{N2}} = (24.44\text{ kmol})(28\text{ kg/kmol}) = 684\text{ kg} \\ m_{\mathrm{total}} &= m_{\mathrm{CO2}} + m_{\mathrm{H2O}} = 176 + 90 + 684 = 950\text{ kg} \end{aligned}$$

Then the mass fractions are

$$\mathrm{mf}_{\mathrm{CO2}} = \frac{m_{\mathrm{CO2}}}{m_{\mathrm{total}}} = \frac{176\text{ kg}}{950\text{ kg}} = \mathbf{0.1853}$$

$$\mathrm{mf}_{\mathrm{H2O}} = \frac{m_{\mathrm{H2O}}}{m_{\mathrm{total}}} = \frac{90\text{ kg}}{950\text{ kg}} = \mathbf{0.0947}$$

$$\mathrm{mf}_{\mathrm{O2}} = \frac{m_{\mathrm{O2}}}{m_{\mathrm{total}}} = \frac{684\text{ kg}}{950\text{ kg}} = \mathbf{0.720}$$

The mass of carbon dioxide per unit mass of fuel burned is

$$\frac{m_{\mathrm{CO2}}}{m_{\mathrm{C4H10}}} = \frac{(4\times44)\text{ kg}}{(1\times58)\text{ kg}} = \mathbf{3.034\ kg\ CO_2/kg\ C_4H_{10}}$$

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

$$\frac{m_{\mathrm{air}}}{m_{\mathrm{C4H10}}} = \frac{(6.5\times4.76\times29)\text{ kg}}{(1\times58)\text{ kg}} = \mathbf{15.47\ kg\ air/kg\ C_4H_{10}}$$

15-29 Coal whose mass percentages are specified is burned with stoichiometric amount of air. The mole fractions of the products, the apparent molecular weight of the product gas, 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 The mass fractions of the constituent of the coal when the ash is substituted are

$$\mathrm{mf}_{\mathrm{C}} = \frac{m_{\mathrm{C}}}{m_{\mathrm{total}}} = \frac{84.36\ \mathrm{kg}}{(100-7.83)\ \mathrm{kg}} = \frac{84.36\ \mathrm{kg}}{92.17\ \mathrm{kg}} = 0.9153$$

$$\mathrm{mf}_{\mathrm{H2}} = \frac{m_{\mathrm{H2}}}{m_{\mathrm{total}}} = \frac{1.89\ \mathrm{kg}}{92.17\ \mathrm{kg}} = 0.02051$$

$$\mathrm{mf}_{\mathrm{O2}} = \frac{m_{\mathrm{O2}}}{m_{\mathrm{total}}} = \frac{4.40\ \mathrm{kg}}{92.17\ \mathrm{kg}} = 0.04774$$

$$\mathrm{mf}_{\mathrm{N2}} = \frac{m_{\mathrm{N2}}}{m_{\mathrm{total}}} = \frac{0.63\ \mathrm{kg}}{92.17\ \mathrm{kg}} = 0.006835$$

$$\mathrm{mf}_{\mathrm{S}} = \frac{m_{\mathrm{S}}}{m_{\mathrm{total}}} = \frac{0.89\ \mathrm{kg}}{92.17\ \mathrm{kg}} = 0.009656$$

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

We now consider 100 kg of this mixture. Then the mole numbers of each component are

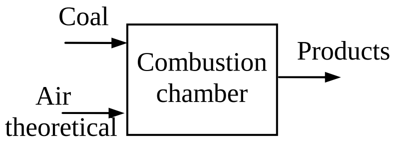
$$N_{\mathrm{C}} = \frac{m_{\mathrm{C}}}{M_{\mathrm{C}}} = \frac{91.53\ \mathrm{kg}}{12\ \mathrm{kg/kmol}} = 7.628\ \mathrm{kmol}$$

$$N_{\mathrm{H2}} = \frac{m_{\mathrm{H2}}}{M_{\mathrm{H2}}} = \frac{2.051\ \mathrm{kg}}{2\ \mathrm{kg/kmol}} = 1.026\ \mathrm{kmol}$$

$$N_{\mathrm{O2}} = \frac{m_{\mathrm{O2}}}{M_{\mathrm{O2}}} = \frac{4.774\ \mathrm{kg}}{32\ \mathrm{kg/kmol}} = 0.1492\ \mathrm{kmol}$$

$$N_{\mathrm{N2}} = \frac{m_{\mathrm{N2}}}{M_{\mathrm{N2}}} = \frac{0.6835\ \mathrm{kg}}{28\ \mathrm{kg/kmol}} = 0.02441\ \mathrm{kmol}$$

$$N_{\mathrm{S}} = \frac{m_{\mathrm{S}}}{M_{\mathrm{S}}} = \frac{0.9656\ \mathrm{kg}}{32\ \mathrm{kg/kmol}} = 0.03018\ \mathrm{kmol}$$



The mole number of the mixture and the mole fractions are

$$N_m = 7.628 + 1.026 + 0.1492 + 0.02441 + 0.03018 = 8.858\ \mathrm{kmol}$$

$$y_{\mathrm{C}} = \frac{N_{\mathrm{C}}}{N_m} = \frac{7.628\ \mathrm{kmol}}{8.858\ \mathrm{kmol}} = 0.8611$$

$$y_{\mathrm{H2}} = \frac{N_{\mathrm{H2}}}{N_m} = \frac{1.026\ \mathrm{kmol}}{8.858\ \mathrm{kmol}} = 0.1158$$

$$y_{\mathrm{O2}} = \frac{N_{\mathrm{O2}}}{N_m} = \frac{0.1492\ \mathrm{kmol}}{8.858\ \mathrm{kmol}} = 0.01684$$

$$y_{\mathrm{N2}} = \frac{N_{\mathrm{N2}}}{N_m} = \frac{0.02441\ \mathrm{kmol}}{8.858\ \mathrm{kmol}} = 0.002756$$

$$y_{\mathrm{S}} = \frac{N_{\mathrm{S}}}{N_m} = \frac{0.03018\ \mathrm{kmol}}{8.858\ \mathrm{kmol}} = 0.003407$$

Then, the combustion equation in this case may be written as

$$0.8611\mathrm{C}+0.1158\mathrm{H}_2+0.01684\mathrm{O}_2+0.002756\mathrm{N}_2+0.003407\mathrm{S}+a_{\mathrm{th}}(\mathrm{O}_2+3.76\mathrm{N}_2)$$
$$\longrightarrow x\mathrm{CO}_2+y\mathrm{H}_2\mathrm{O}+z\mathrm{SO}_2+k\mathrm{N}_2$$

According to the species balances,

$$\begin{aligned}
&\text{C balance:}\quad x=0.8611\\
&\mathrm{H}_2\text{ balance:}\quad y=0.1158\\
&\text{S balance:}\quad z=0.003407\\
&\mathrm{O}_2\text{ balance:}\\
&0.01684+a_{\mathrm{th}}=x+0.5y+z\\
&a_{\mathrm{th}}=0.8611+0.5(0.1158)+0.003407-0.01684=0.9056\\
&\mathrm{N}_2\text{ balance:}\quad k=0.002756+3.76a_{\mathrm{th}}=0.002756+3.76\times0.9056=3.408
\end{aligned}$$

Substituting,

$$0.8611\mathrm{C}+0.1158\mathrm{H}_2+0.01684\mathrm{O}_2+0.002756\mathrm{N}_2+0.003407\mathrm{S}+0.9056(\mathrm{O}_2+3.76\mathrm{N}_2)$$
$$\longrightarrow 0.8611\mathrm{CO}_2+0.1158\mathrm{H}_2\mathrm{O}+0.00341\mathrm{SO}_2+3.408\mathrm{N}_2$$

The mole fractions of the products are

$$N_m=0.8611+0.1158+0.00341+3.408=4.388\text{ kmol}$$

$$y_{\mathrm{CO2}}=\frac{N_{\mathrm{CO2}}}{N_m}=\frac{0.8611\text{ kmol}}{4.388\text{ kmol}}=\mathbf{0.1962}$$

$$y_{\mathrm{H2O}}=\frac{N_{\mathrm{H2O}}}{N_m}=\frac{0.1158\text{ kmol}}{4.388\text{ kmol}}=\mathbf{0.02639}$$

$$y_{\mathrm{SO2}}=\frac{N_{\mathrm{SO2}}}{N_m}=\frac{0.00341\text{ kmol}}{4.388\text{ kmol}}=\mathbf{0.00078}$$

$$y_{\mathrm{N2}}=\frac{N_{\mathrm{N2}}}{N_m}=\frac{3.408\text{ kmol}}{4.388\text{ kmol}}=\mathbf{0.7767}$$

The apparent molecular weight of the product gas is

$$M_m=\frac{m_m}{N_m}=\frac{(0.8611\times44+0.1158\times18+0.00341\times64+3.408\times28)\text{ kg}}{4.388\text{ kmol}}=\mathbf{30.91\ kg/kmol}$$

The air-fuel mass ratio is then

$$\mathrm{AF}=\frac{m_{\mathrm{air}}}{m_{\mathrm{fuel}}}=\frac{(0.9056\times4.76\times29)\text{ kg}}{(0.8611\times12+0.1158\times2+0.01684\times32+0.002756\times28+0.003408\times32)\text{ kg}}$$
$$=\frac{125.0\text{ kg}}{11.29\text{ kg}}=\mathbf{11.07\ kg\ air/kg\ fuel}$$

15-30 Methyl alcohol is burned with stoichiometric amount of air. The mole fraction of each product, the apparent molar mass of the product gas, and the mass of water in the products per unit mass of fuel burned are to be determined.

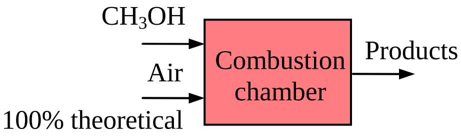
Assumptions **1** Combustion is complete. **2** The combustion products contain CO_2, H_2O, and N_2 only.

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 balanced reaction equation for stoichiometric air is

$$\mathrm{CH_3OH} + 1.5[\mathrm{O_2} + 3.76\mathrm{N_2}] \longrightarrow \mathrm{CO_2} + 2\,\mathrm{H_2O} + 5.64\,\mathrm{N_2}$$

The mole fractions of the products are



$$N_m = 1 + 2 + 5.64 = 8.64\ \text{kmol}$$

$$y_{\mathrm{CO2}} = \frac{N_{\mathrm{CO2}}}{N_m} = \frac{1\ \text{kmol}}{8.64\ \text{kmol}} = \mathbf{0.1157}$$

$$y_{\mathrm{H2O}} = \frac{N_{\mathrm{H2O}}}{N_m} = \frac{2\ \text{kmol}}{8.64\ \text{kmol}} = \mathbf{0.2315}$$

$$y_{\mathrm{N2}} = \frac{N_{\mathrm{N2}}}{N_m} = \frac{5.64\ \text{kmol}}{8.64\ \text{kmol}} = \mathbf{0.6528}$$

The apparent molecular weight of the product gas is

$$M_m = \frac{m_m}{N_m} = \frac{(1 \times 44 + 2 \times 18 + 5.64 \times 28)\ \text{kg}}{8.64\ \text{kmol}} = \mathbf{27.54\ kg/kmol}$$

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

$$\frac{m_{\mathrm{H2O}}}{m_{\mathrm{CH3OH}}} = \frac{(2 \times 18)\ \text{kg}}{(1 \times 32)\ \text{kg}} = \mathbf{1.125\ kg\ H_2O/kg\ CH_3OH}$$

15-31 Coal whose mass percentages are specified is burned with stoichiometric amount of air. The combustion is incomplete. The mass fractions of the products, the apparent molecular weight of the product gas, and the air-fuel ratio 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 The mass fractions of the constituent of the coal when the ash is substituted are

$$\mathrm{mf}_\mathrm{C} = \frac{m_\mathrm{C}}{m_\mathrm{total}} = \frac{61.40\ \mathrm{kg}}{(100-5.00)\ \mathrm{kg}} = \frac{61.40\ \mathrm{kg}}{95.00\ \mathrm{kg}} = 0.6463$$

$$\mathrm{mf}_\mathrm{H2} = \frac{m_\mathrm{H2}}{m_\mathrm{total}} = \frac{5.79\ \mathrm{kg}}{95.00\ \mathrm{kg}} = 0.06095$$

$$\mathrm{mf}_\mathrm{O2} = \frac{m_\mathrm{O2}}{m_\mathrm{total}} = \frac{25.31\ \mathrm{kg}}{95.00\ \mathrm{kg}} = 0.2664$$

$$\mathrm{mf}_\mathrm{N2} = \frac{m_\mathrm{N2}}{m_\mathrm{total}} = \frac{1.09\ \mathrm{kg}}{95.00\ \mathrm{kg}} = 0.01147$$

$$\mathrm{mf}_\mathrm{S} = \frac{m_\mathrm{S}}{m_\mathrm{total}} = \frac{1.41\ \mathrm{kg}}{95.00\ \mathrm{kg}} = 0.01484$$

61.40% C
5.79% H_2
25.31% O_2
1.09% N_2
1.41% S
5.00% ash
(by mass)

We now consider 100 kg of this mixture. Then the mole numbers of each component are

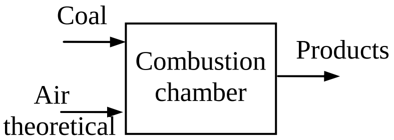
$$N_\mathrm{C} = \frac{m_\mathrm{C}}{M_\mathrm{C}} = \frac{64.63\ \mathrm{kg}}{12\ \mathrm{kg/kmol}} = 5.386\ \mathrm{kmol}$$

$$N_\mathrm{H2} = \frac{m_\mathrm{H2}}{M_\mathrm{H2}} = \frac{6.095\ \mathrm{kg}}{2\ \mathrm{kg/kmol}} = 3.048\ \mathrm{kmol}$$

$$N_\mathrm{O2} = \frac{m_\mathrm{O2}}{M_\mathrm{O2}} = \frac{26.64\ \mathrm{kg}}{32\ \mathrm{kg/kmol}} = 0.8325\ \mathrm{kmol}$$

$$N_\mathrm{N2} = \frac{m_\mathrm{N2}}{M_\mathrm{N2}} = \frac{1.147\ \mathrm{kg}}{28\ \mathrm{kg/kmol}} = 0.04096\ \mathrm{kmol}$$

$$N_\mathrm{S} = \frac{m_\mathrm{S}}{M_\mathrm{S}} = \frac{1.484\ \mathrm{kg}}{32\ \mathrm{kg/kmol}} = 0.04638\ \mathrm{kmol}$$



The mole number of the mixture and the mole fractions are

$$N_m = 5.386 + 3.048 + 0.8325 + 0.04096 + 0.04638 = 9.354\ \mathrm{kmol}$$

$$y_\mathrm{C} = \frac{N_\mathrm{C}}{N_m} = \frac{5.386\ \mathrm{kmol}}{9.354\ \mathrm{kmol}} = 0.5758$$

$$y_\mathrm{H2} = \frac{N_\mathrm{H2}}{N_m} = \frac{3.048\ \mathrm{kmol}}{9.354\ \mathrm{kmol}} = 0.3258$$

$$y_\mathrm{O2} = \frac{N_\mathrm{O2}}{N_m} = \frac{0.8325\ \mathrm{kmol}}{9.354\ \mathrm{kmol}} = 0.0890$$

$$y_\mathrm{N2} = \frac{N_\mathrm{N2}}{N_m} = \frac{0.04096\ \mathrm{kmol}}{9.354\ \mathrm{kmol}} = 0.00438$$

$$y_\mathrm{S} = \frac{N_\mathrm{S}}{N_m} = \frac{0.04638\ \mathrm{kmol}}{9.354\ \mathrm{kmol}} = 0.00496$$

Then, the combustion equation in this case may be written as

$$0.5758\text{C} + 0.3258\text{H}_2 + 0.0890\text{O}_2 + 0.00438\text{N}_2 + 0.00496\text{S} + a_{\text{th}}(\text{O}_2 + 3.76\text{N}_2)$$
$$\longrightarrow x(0.95\text{CO}_2 + 0.05\text{CO}) + y\text{H}_2\text{O} + z\text{SO}_2 + k\text{N}_2$$

According to the species balances,

$$\text{C balance:} \quad x = 0.5758$$
$$\text{H}_2 \text{ balance:} \quad y = 0.3258$$
$$\text{S balance:} \quad z = 0.00496$$
$$\text{O}_2 \text{ balance:}$$
$$0.0890 + a_{\text{th}} = 0.95x + 0.5 \times 0.05x + 0.5y + z$$
$$a_{\text{th}} = 0.95 \times 0.5758 + 0.5 \times 0.05 \times 0.5758 + 0.5 \times 0.3258 + 0.00496 - 0.0890 = 0.6403$$
$$\text{N}_2 \text{ balance:} \quad k = 0.00438 + 3.76a_{\text{th}} = 0.00438 + 3.76 \times 0.6403 = 2.412$$

Substituting,

$$0.5758\text{C} + 0.3258\text{H}_2 + 0.0890\text{O}_2 + 0.00438\text{N}_2 + 0.00496\text{S} + 0.7293(\text{O}_2 + 3.76\text{N}_2)$$
$$\longrightarrow 0.5470\text{CO}_2 + 0.0288\text{CO} + 0.3258\text{H}_2\text{O} + 0.00496\text{SO}_2 + 2.412\text{N}_2$$

The mass fractions of the products are

$$m_{\text{total}} = 0.5470 \times 44 + 0.0288 \times 28 + 0.3258 \times 18 + 0.00496 \times 64 + 2.412 \times 28 = 98.6 \text{ kg}$$

$$\text{mf}_{\text{CO2}} = \frac{m_{\text{CO2}}}{m_{\text{total}}} = \frac{(0.5470 \times 44) \text{ kg}}{98.6 \text{ kg}} = \mathbf{0.2441}$$

$$\text{mf}_{\text{CO}} = \frac{m_{\text{CO}}}{m_{\text{total}}} = \frac{(0.0288 \times 28) \text{ kg}}{98.6 \text{ kg}} = \mathbf{0.0082}$$

$$\text{mf}_{\text{H2O}} = \frac{m_{\text{H2O}}}{m_{\text{total}}} = \frac{(0.3258 \times 18) \text{ kg}}{98.6 \text{ kg}} = \mathbf{0.0595}$$

$$\text{mf}_{\text{SO2}} = \frac{m_{\text{SO2}}}{m_{\text{total}}} = \frac{(0.00496 \times 64) \text{ kg}}{98.6 \text{ kg}} = \mathbf{0.0032}$$

$$\text{mf}_{\text{N2}} = \frac{m_{\text{N2}}}{m_{\text{total}}} = \frac{(2.412 \times 28) \text{ kg}}{98.6 \text{ kg}} = \mathbf{0.6849}$$

The total mole number of the products is

$$N_m = 0.5470 + 0.0288 + 0.3258 + 0.00496 + 2.712 = 3.319 \text{ kmol}$$

The apparent molecular weight of the product gas is

$$M_m = \frac{m_m}{N_m} = \frac{98.6 \text{ kg}}{3.319 \text{ kmol}} = \mathbf{29.71 \text{ kg/kmol}}$$

The air-fuel mass ratio is then

$$\text{AF} = \frac{m_{\text{air}}}{m_{\text{fuel}}} = \frac{(0.6403 \times 4.76 \times 29) \text{ kg}}{(0.5758 \times 12 + 0.3258 \times 2 + 0.0890 \times 32 + 0.00438 \times 28 + 0.00496 \times 32) \text{ kg}}$$
$$= \frac{88.39 \text{ kg}}{10.69 \text{ kg}}$$
$$= \mathbf{8.27 \text{ kg air/kg fuel}}$$

15-32 Propane is burned with 30 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 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

$$\mathrm{C_3H_8} + 1.3a_{\mathrm{th}}\left[\mathrm{O_2} + 3.76\mathrm{N_2}\right] \longrightarrow 3\mathrm{CO_2} + 4\mathrm{H_2O} + 0.3a_{\mathrm{th}}\mathrm{O_2} + (1.3\times 3.76)a_{\mathrm{th}}\mathrm{N_2}$$

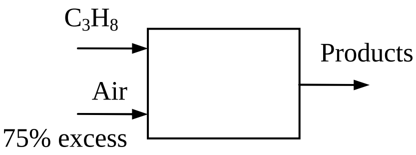
where a_{th} is the stoichiometric coefficient for air. We have automatically accounted for the 30% excess air by using the factor $1.3a_{\mathrm{th}}$ instead of a_{th} for air. The stoichiometric amount of oxygen ($a_{\mathrm{th}}O_2$) will be used to oxidize the fuel, and the remaining excess amount ($0.3a_{\mathrm{th}}O_2$) will appear in the products as free oxygen. The coefficient a_{th} is determined from the O_2 balance,

O_2 balance: $$1.3a_{\mathrm{th}} = 3 + 2 + 0.3a_{\mathrm{th}} \longrightarrow a_{\mathrm{th}} = 5$$

Substituting, $$\mathrm{C_3H_8} + 6.5\left[\mathrm{O_2} + 3.76\mathrm{N_2}\right] \longrightarrow 3\mathrm{CO_2} + 4\mathrm{H_2O} + 1.5\mathrm{O_2} + 24.44\mathrm{N_2}$$

The mole fractions of the products are

$$N_m = 3 + 4 + 1.5 + 24.44 = 32.94\ \text{kmol}$$



$$y_{\mathrm{CO2}} = \frac{N_{\mathrm{CO2}}}{N_m} = \frac{3\ \text{kmol}}{32.94\ \text{kmol}} = \mathbf{0.0911}$$

$$y_{\mathrm{H2O}} = \frac{N_{\mathrm{H2O}}}{N_m} = \frac{4\ \text{kmol}}{32.94\ \text{kmol}} = \mathbf{0.1214}$$

$$y_{\mathrm{O2}} = \frac{N_{\mathrm{O2}}}{N_m} = \frac{1.5\ \text{kmol}}{32.94\ \text{kmol}} = \mathbf{0.0455}$$

$$y_{\mathrm{N2}} = \frac{N_{\mathrm{N2}}}{N_m} = \frac{24.44\ \text{kmol}}{32.94\ \text{kmol}} = \mathbf{0.7420}$$

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

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

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

$$\mathrm{AF} = \frac{m_{\mathrm{air}}}{m_{\mathrm{fuel}}} = \frac{(6.5\times 4.76\ \text{kmol})(29\ \text{kg/kmol})}{(3\ \text{kmol})(12\ \text{kg/kmol}) + (4\ \text{kmol})(2\ \text{kg/kmol})} = \mathbf{20.39\ kg\ air/kg\ fuel}$$

15-33 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.

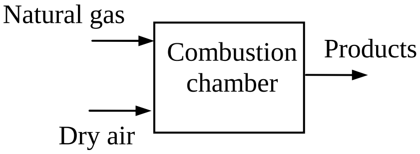
Properties The molar masses of C, H_2, N_2, O_2, and air are 12 kg/kmol, 2 kg/kmol, 28 kg/kmol, 32 kg/kmol, and 29 kg/kmol, respectively (Table A-1).

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

$$(0.65\mathrm{CH_4}+0.08\mathrm{H_2}+0.18\mathrm{N_2}+0.03\mathrm{O_2}+0.06\mathrm{CO_2})+a_{\mathrm{th}}(\mathrm{O_2}+3.76\mathrm{N_2})\longrightarrow x\mathrm{CO_2}+y\mathrm{H_2O}+z\mathrm{N_2}$$

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

$$\mathrm{C:}\ \ 0.65+0.06=x \qquad \longrightarrow x=0.71$$
$$\mathrm{H:}\ \ 0.65\times4+0.08\times2=2y \qquad \longrightarrow y=1.38$$
$$\mathrm{O_2:}\ \ 0.03+0.06+a_{\mathrm{th}}=x+y/2 \qquad \longrightarrow a_{\mathrm{th}}=1.31$$
$$\mathrm{N_2:}\ \ 0.18+3.76a_{\mathrm{th}}=z \qquad \longrightarrow z=5.106$$



Thus,

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

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,

$$m_{\mathrm{air}}=(1.31\times4.76\ \mathrm{kmol})(29\ \mathrm{kg/kmol})=180.8\ \mathrm{kg}$$
$$m_{\mathrm{fuel}}=(0.65\times16+0.08\times2+0.18\times28+0.03\times32+0.06\times44)\mathrm{kg}=19.2\ \mathrm{kg}$$

and

$$\mathrm{AF_{th}}=\frac{m_{\mathrm{air,th}}}{m_{\mathrm{fuel}}}=\frac{180.8\ \mathrm{kg}}{19.2\ \mathrm{kg}}=\mathbf{9.42\ \ kg\ air/kg\ fuel}$$

15-34 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

$$(0.65\mathrm{CH_4} + 0.08\mathrm{H_2} + 0.18\mathrm{N_2} + 0.03\mathrm{O_2} + 0.06\mathrm{CO_2}) + a_{\mathrm{th}}(\mathrm{O_2} + 3.76\mathrm{N_2}) \longrightarrow x\mathrm{CO_2} + y\mathrm{H_2O} + z\mathrm{N_2}$$

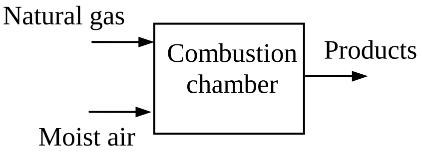
The unknown coefficients in the above equation are determined from mass balances,

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

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

$$\mathrm{O_2}:\ 0.03 + 0.06 + a_{\mathrm{th}} = x + y/2 \longrightarrow a_{\mathrm{th}} = 1.31$$

$$\mathrm{N_2}:\ 0.18 + 3.76 a_{\mathrm{th}} = z \longrightarrow z = 5.106$$



Thus,

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

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

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

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

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

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,

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

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,

$$m_{\mathrm{air}} = (1.31 \times 4.76\ \mathrm{kmol})(29\ \mathrm{kg/kmol}) + (0.17\ \mathrm{kmol} \times 18\ \mathrm{kg/kmol}) = 183.9\ \mathrm{kg}$$

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

and

$$\mathrm{AF_{th}} = \frac{m_{\mathrm{air,th}}}{m_{\mathrm{fuel}}} = \frac{183.9\mathrm{kg}}{19.2\mathrm{kg}} = \mathbf{9.58\ kg\ air/kg\ fuel}$$

15-35 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

$$(0.60\text{CH}_4 + 0.30\text{H}_2 + 0.10\text{N}_2) + 1.3a_{\text{th}}(\text{O}_2 + 3.76\text{N}_2) \longrightarrow x\text{CO}_2 + y\text{H}_2\text{O} + 0.3a_{\text{th}}\text{O}_2 + z\text{N}_2$$

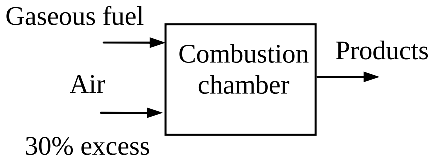
The unknown coefficients in the above equation are determined from mass balances,

$$\text{C}: \ 0.60 = x \quad \longrightarrow x = 0.60$$

$$\text{H}: \ 0.60 \times 4 + 0.30 \times 2 = 2y \quad \longrightarrow y = 1.50$$

$$\text{O}_2: \ 1.3a_{\text{th}} = x + y/2 + 0.3a_{\text{th}} \quad \longrightarrow a_{\text{th}} = 1.35$$

$$\text{N}_2: \ 0.10 + 3.76 \times 1.3a_{\text{th}} = z \quad \longrightarrow z = 6.70$$



Thus,

$$(0.60\text{CH}_4 + 0.30\text{H}_2 + 0.10\text{N}_2) + 1.755(\text{O}_2 + 3.76\text{N}_2) \longrightarrow 0.6\text{CO}_2 + 1.5\text{H}_2\text{O} + 0.405\text{O}_2 + 6.7\text{N}_2$$

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,

$$m_{\text{air}} = (1.755 \times 4.76 \text{ kmol})(29 \text{ kg/kmol}) = 242.3 \text{ kg}$$

$$m_{\text{fuel}} = (0.6 \times 16 + 0.3 \times 2 + 0.1 \times 28)\text{kg} = 13.0 \text{ kg}$$

and

$$\text{AF} = \frac{m_{\text{air}}}{m_{\text{fuel}}} = \frac{242.3 \text{ kg}}{13.0 \text{ kg}} = \mathbf{18.6 \text{ kg air/kg fuel}}$$

(*b*) For each kmol of fuel burned, 0.6 + 1.5 + 0.405 + 6.7 = 9.205 kmol of products are formed, including 1.5 kmol of H_2O. Assuming that the dew-point temperature of the products is above 20°C, some of the water vapor will condense as the products are cooled to 20°C. If N_w kmol of H_2O condenses, there will be 1.5 - N_w kmol of water vapor left in the products. The mole number of the products in the gas phase will also decrease to 9.205 - 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,

$$\frac{N_v}{N_{\text{prod,gas}}} = \frac{P_v}{P_{\text{prod}}} \longrightarrow \frac{1.5 - N_w}{9.205 - N_w} = \frac{2.3392 \text{ kPa}}{101.325 \text{ kPa}} \longrightarrow N_w = 1.32 \text{ kmol}$$

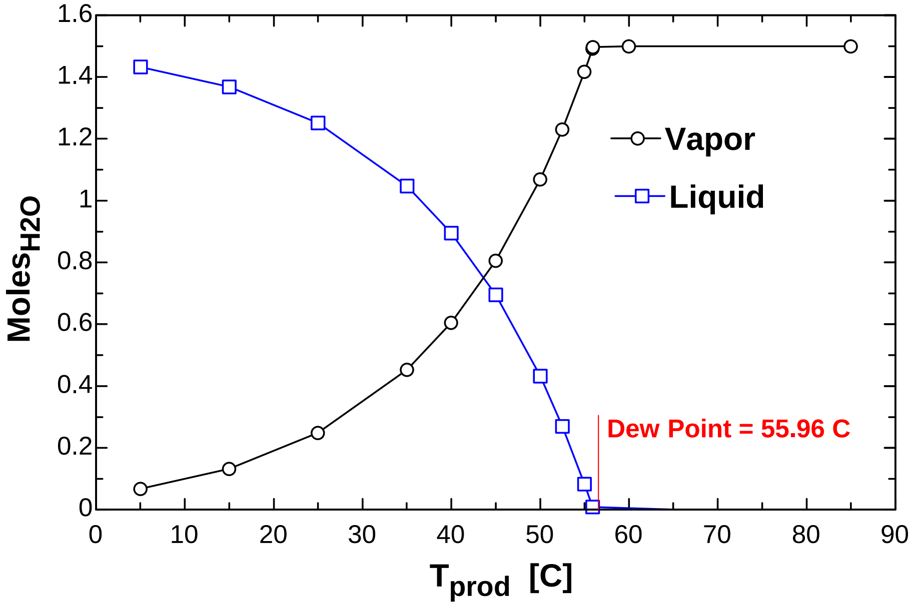
since $P_v = P_{\text{sat @ 20°C}}$ = 2.3392 kPa. Thus the fraction of water vapor that condenses is 1.32/1.5 = 0.88 or **88%**.

15-36 EES Problem 15-35 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,Re
sult$)
P_v = Moles_H2O/(M_other+Moles_H2O)*P_prod
T_DewPoint = temperature(steam,P=P_v,x=0)
IF T_DewPoint <= T_prod then
Moles_H2O_vap = Moles_H2O
Moles_H2O_liq=0
Result$='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=0.6
b=0.3
c=0.1
T_prod = 20 [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,Re
sult$)
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]
18.61	95.54	1.433	0.06692	5
18.61	91.21	1.368	0.1319	15
18.61	83.42	1.251	0.2487	25
18.61	69.8	1.047	0.453	35
18.61	59.65	0.8947	0.6053	40
18.61	46.31	0.6947	0.8053	45
18.61	28.75	0.4312	1.069	50
18.61	17.94	0.2691	1.231	52.5
18.61	5.463	0.08195	1.418	55
18.61	0.5077	0.007615	1.492	55.9
18.61	0.1679	0.002518	1.497	55.96
18.61	0	0	1.5	60
18.61	0	0	1.5	85



15-37 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.

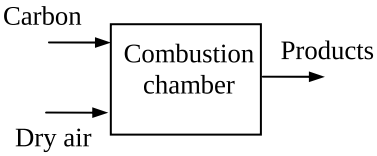
Assumptions **1** Combustion is complete. **2** The combustion products contain CO_2, CO, O_2, and N_2 only.

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,



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

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

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

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,

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

15-38 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.

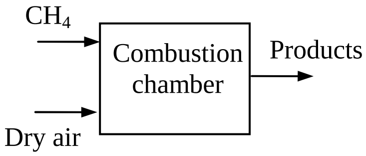
Assumptions **1** Combustion is complete. **2** The combustion products contain CO_2, CO, H_2O, O_2, and N_2 only.

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,

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

(*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,

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

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

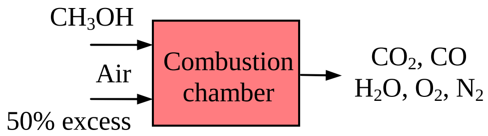
Then,

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

15-39 Methyl alcohol is burned with 50% excess air. The combustion is incomplete. The mole fraction of carbon monoxide and the apparent molar mass of 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 balanced reaction equation for stoichiometric air is

$$CH_3OH + 1.5[O_2 + 3.76N_2] \longrightarrow CO_2 + 2\,H_2O + 5.64\,N_2$$

The reaction with 50% excess air and incomplete combustion can be written as

$$CH_3OH + 1.5\times1.5[O_2 + 3.76N_2] \longrightarrow 0.90\,CO_2 + 0.10\,CO + 2\,H_2O + x\,O_2 + 1.5\times1.5\times3.76\,N_2$$

The coefficient for CO is determined from a mass balance,

O_2 balance: $$0.5 + 1.5\times1.5 = 0.9 + 0.05 + 1 + x \longrightarrow x = 0.8$$

Substituting,

$$CH_3OH + 2.25[O_2 + 3.76N_2] \longrightarrow 0.90\,CO_2 + 0.10\,CO + 2\,H_2O + 0.8\,O_2 + 8.46\,N_2$$

The total moles of the products is

$$N_m = 0.9 + 0.10 + 2 + 0.8 + 8.46 = 12.26\text{ kmol}$$

The mole fraction of carbon monoxide in the products is

$$y_{CO} = \frac{N_{CO}}{N_m} = \frac{0.8\text{ kmol}}{12.26\text{ kmol}} = \mathbf{0.0653}$$

The apparent molecular weight of the product gas is

$$M_m = \frac{m_m}{N_m} = \frac{(0.9\times44 + 0.10\times28 + 2\times18 + 0.8\times32 + 8.46\times28)\text{ kg}}{12.26\text{ kmol}} = \mathbf{27.80\ kg/kmol}$$

15-40 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

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

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

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

The coefficient for CO is determined from a mass balance,

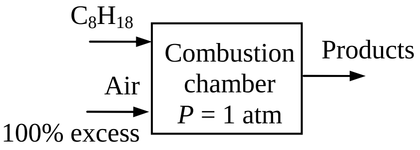
O_2 balance: $\quad 25 = 0.85 \times 8 + 0.5 \times 0.15 \times 8 + 0.5 \times 9 + x \longrightarrow x = 13.1$

Substituting,

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

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,

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

Enthalpy of Formation and Enthalpy of Combustion

15-41C 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.

15-42C 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.

15-43C 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.

15-44C 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.

15-45C Yes.

15-46C No. The enthalpy of formation of N_2 is simply assigned a value of zero at the standard reference state for convenience.

15-47C 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.

15-48 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,

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

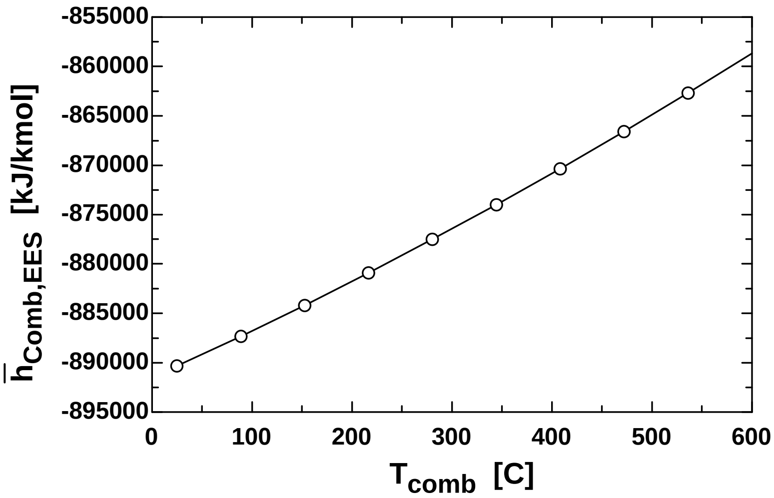
The listed value in Table A-27 is -890,868 kJ/kmol, which is almost identical to the calculated value. Since the water in the products is assumed to be in the liquid phase, this h_c value corresponds to the higher heating value of CH_4.

15-49 EES Problem 15-48 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



15-50 The enthalpy of combustion of gaseous ethane 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

$$\mathrm{C_2H_6} + 3.5[\mathrm{O_2} + 3.76\mathrm{N_2}] \longrightarrow 2\mathrm{CO_2} + 3\mathrm{H_2O}(\ell) + 13.16\mathrm{N_2}$$

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 C_2H_6 becomes

$$h_C = H_P - H_R = \sum N_P \bar{h}^\circ_{f,P} - \sum N_R \bar{h}^\circ_{f,R} = \left(N\bar{h}^\circ_f\right)_{\mathrm{CO_2}} + \left(N\bar{h}^\circ_f\right)_{\mathrm{H_2O}} - \left(N\bar{h}^\circ_f\right)_{\mathrm{C_2H_6}}$$

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

$$\begin{aligned} h_C &= (2\ \text{kmol})(-393{,}520\ \text{kJ/kmol}) + (3\ \text{kmol})(-285{,}830\ \text{kJ/kmol}) \\ &\quad - (1\ \text{kmol})(-84{,}680\ \text{kJ/kmol}) \\ &= \mathbf{-1{,}559{,}850\ \ kJ}\ (\text{per kmol}\mathrm{C_2H_6}) \end{aligned}$$

The listed value in Table A-27 is -1,560,633 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 C_2H_6.

15-51 The enthalpy of combustion of liquid octane 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

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

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 C_8H_{18} becomes

$$h_C = H_P - H_R = \sum N_P \bar{h}^\circ_{f,P} - \sum N_R \bar{h}^\circ_{f,R} = \left(N\bar{h}^\circ_f\right)_{\mathrm{CO_2}} + \left(N\bar{h}^\circ_f\right)_{\mathrm{H_2O}} - \left(N\bar{h}^\circ_f\right)_{\mathrm{C_8H_{18}}}$$

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

$$\begin{aligned} h_C &= (8\ \text{kmol})(-393{,}520\ \text{kJ/kmol}) + (9\ \text{kmol})(-285{,}830\ \text{kJ/kmol}) \\ &\quad - (1\ \text{kmol})(-249{,}950\ \text{kJ/kmol}) \\ &= \mathbf{-5{,}470{,}680\ \ kJ} \end{aligned}$$

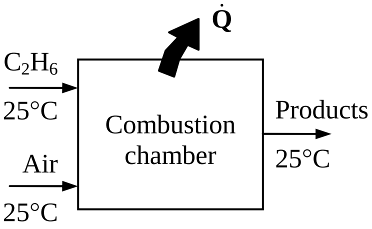
The listed value in Table A-27 is -5,470,523 kJ/kmol, which is almost identical to the calculated value. Since the water in the products is assumed to be in the liquid phase, this h_c value corresponds to the higher heating value of C_8H_{18}.

15-52 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,

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

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

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

15-53 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

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

The minimum pressure of the products is then

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

15-54 The higher and lower heating values of coal from Utah 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, N_2, S, and air are 12, 2, 32, 28, 32, and 29 kg/kmol, respectively (Table A-1).

Analysis The mass fractions of the constituent of the coal when the ash is substituted are

$$\mathrm{mf}_{\mathrm{C}} = \frac{m_{\mathrm{C}}}{m_{\mathrm{total}}} = \frac{61.40\ \mathrm{kg}}{(100-5.00)\ \mathrm{kg}} = \frac{61.40\ \mathrm{kg}}{95.00\ \mathrm{kg}} = 0.6463$$

$$\mathrm{mf}_{\mathrm{H2}} = \frac{m_{\mathrm{H2}}}{m_{\mathrm{total}}} = \frac{5.79\ \mathrm{kg}}{95.00\ \mathrm{kg}} = 0.06095$$

$$\mathrm{mf}_{\mathrm{O2}} = \frac{m_{\mathrm{O2}}}{m_{\mathrm{total}}} = \frac{25.31\ \mathrm{kg}}{95.00\ \mathrm{kg}} = 0.2664$$

$$\mathrm{mf}_{\mathrm{N2}} = \frac{m_{\mathrm{N2}}}{m_{\mathrm{total}}} = \frac{1.09\ \mathrm{kg}}{95.00\ \mathrm{kg}} = 0.01147$$

$$\mathrm{mf}_{\mathrm{S}} = \frac{m_{\mathrm{S}}}{m_{\mathrm{total}}} = \frac{1.41\ \mathrm{kg}}{95.00\ \mathrm{kg}} = 0.01484$$

61.40% C
5.79% H_2
25.31% O_2
1.09% N_2
1.41% S
5.00% ash
(by mass)

We now consider 100 kg of this mixture. Then the mole numbers of each component are

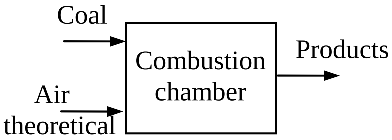
$$N_{\mathrm{C}} = \frac{m_{\mathrm{C}}}{M_{\mathrm{C}}} = \frac{64.63\ \mathrm{kg}}{12\ \mathrm{kg/kmol}} = 5.386\ \mathrm{kmol}$$

$$N_{\mathrm{H2}} = \frac{m_{\mathrm{H2}}}{M_{\mathrm{H2}}} = \frac{6.095\ \mathrm{kg}}{2\ \mathrm{kg/kmol}} = 3.048\ \mathrm{kmol}$$

$$N_{\mathrm{O2}} = \frac{m_{\mathrm{O2}}}{M_{\mathrm{O2}}} = \frac{26.64\ \mathrm{kg}}{32\ \mathrm{kg/kmol}} = 0.8325\ \mathrm{kmol}$$

$$N_{\mathrm{N2}} = \frac{m_{\mathrm{N2}}}{M_{\mathrm{N2}}} = \frac{1.147\ \mathrm{kg}}{28\ \mathrm{kg/kmol}} = 0.04096\ \mathrm{kmol}$$

$$N_{\mathrm{S}} = \frac{m_{\mathrm{S}}}{M_{\mathrm{S}}} = \frac{1.484\ \mathrm{kg}}{32\ \mathrm{kg/kmol}} = 0.04638\ \mathrm{kmol}$$



The mole number of the mixture and the mole fractions are

$$N_m = 5.386 + 3.048 + 0.8325 + 0.04096 + 0.04638 = 9.354\ \mathrm{kmol}$$

$$y_{\mathrm{C}} = \frac{N_{\mathrm{C}}}{N_m} = \frac{5.386\ \mathrm{kmol}}{9.354\ \mathrm{kmol}} = 0.5758$$

$$y_{\mathrm{H2}} = \frac{N_{\mathrm{H2}}}{N_m} = \frac{3.048\ \mathrm{kmol}}{9.354\ \mathrm{kmol}} = 0.3258$$

$$y_{\mathrm{O2}} = \frac{N_{\mathrm{O2}}}{N_m} = \frac{0.8325\ \mathrm{kmol}}{9.354\ \mathrm{kmol}} = 0.0890$$

$$y_{\mathrm{N2}} = \frac{N_{\mathrm{N2}}}{N_m} = \frac{0.04096\ \mathrm{kmol}}{9.354\ \mathrm{kmol}} = 0.00438$$

$$y_{\mathrm{S}} = \frac{N_{\mathrm{S}}}{N_m} = \frac{0.04638\ \mathrm{kmol}}{9.354\ \mathrm{kmol}} = 0.00496$$

Then, the combustion equation in this case may be written as

$$0.5758\text{C}+0.3258\text{H}_2+0.0890\text{O}_2+0.00438\text{N}_2+0.00496\text{S}+a_{\text{th}}(\text{O}_2+3.76\text{N}_2)$$
$$\longrightarrow 0.5758\text{CO}_2+0.3258\text{H}_2\text{O}+0.00496\text{SO}_2+k\text{N}_2$$

According to the species balances,

$$\text{O}_2\text{ balance: } 0.0890+a_{\text{th}}=0.5758+0.5\times 0.3258+0.00496 \longrightarrow a_{\text{th}}=0.6547$$
$$\text{N}_2\text{ balance: } k=0.00438+3.76\times 0.6547=2.466$$

Substituting,

$$0.5758\text{C}+0.3258\text{H}_2+0.0890\text{O}_2+0.00438\text{N}_2+0.00496\text{S}+0.6547(\text{O}_2+3.76\text{N}_2)$$
$$\longrightarrow 0.5758\text{CO}_2+0.3258\text{H}_2\text{O}+0.00496\text{SO}_2+2.466\text{N}_2$$

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,

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

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

$$\begin{aligned}h_C&=(0.5758\text{ kmol})(-393{,}520\text{ kJ/kmol})+(0.3258\text{ kmol})(-285{,}830\text{ kJ/kmol})\\&\quad+(0.00496\text{ kmol})(-297{,}100\text{ kJ/kmol})\\&=-321{,}200\text{ kJ/kmol coal}\end{aligned}$$

The apparent molecular weight of the coal is

$$\begin{aligned}M_m=\frac{m_m}{N_m}&=\frac{(0.5758\times 12+0.3258\times 2+0.0890\times 32+0.00438\times 28+0.00496\times 32)\text{ kg}}{(0.5758+0.3258+0.0890+0.00438+0.00496)\text{ kmol}}\\&=\frac{10.69\text{ kg}}{1.000\text{ kmol}}=10.69\text{ kg/kmol coal}\end{aligned}$$

The HHV of the coal is then

$$\text{HHV}=\frac{-h_C}{M_m}=\frac{321{,}200\text{ kJ/kmol coal}}{10.69\text{ kg/kmol coal}}=\mathbf{30{,}000\text{ kJ/kg coal}}$$

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

$$\begin{aligned}h_C&=(0.5758\text{ kmol})(-393{,}520\text{ kJ/kmol})+(0.3258\text{ kmol})(-241{,}820\text{ kJ/kmol})\\&\quad+(0.00496\text{ kmol})(-297{,}100\text{ kJ/kmol})\\&=-306{,}850\text{ kJ/kmol coal}\end{aligned}$$

The LHV of the coal is then

$$\text{LHV}=\frac{-h_C}{M_m}=\frac{306{,}850\text{ kJ/kmol coal}}{10.69\text{ kg/kmol coal}}=\mathbf{28{,}700\text{ kJ/kg coal}}$$

First Law Analysis of Reacting Systems

15-55C 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.

15-56C 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.

15-57C 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.

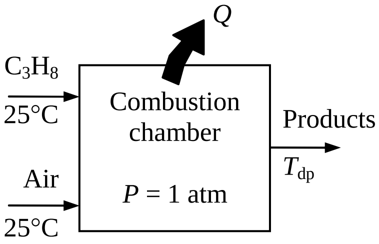
15-58 Propane is burned with an air-fuel ratio of 18. The heat transfer per kilogram of fuel burned when the temperature of the products is such that liquid water just begins to form in the products is to be determined.

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

$$\begin{aligned} m_{\text{air}} &= (\text{AF})m_{\text{fuel}} \\ &= (18\text{ kg air/kg fuel})(1\times 44\text{ kg/kmol fuel}) \\ &= 792\text{ kg air/kmol fuel} \end{aligned}$$



The mole number of air per kmol of fuel is then

$$N_{\text{air}} = \frac{m_{\text{air}}}{M_{\text{air}}} = \frac{792\text{ kg air/kmol fuel}}{29\text{ kg air/kmol air}} = 27.31\text{ kmol air/kmol fuel}$$

The combustion equation can be written as

$$C_3H_8 + (27.31/4.76)(O_2 + 3.76N_2) \longrightarrow 3CO_2 + 4H_2O + xO_2 + (27.31/4.76)\times 3.76N_2$$

The coefficient for CO is obtained from O_2 balance:

$$(27.31/4.76) = 3 + 2 + x \longrightarrow x = 0.7374$$

Substituting,

$$C_3H_8 + 5.7374(O_2 + 3.76N_2) \longrightarrow 3CO_2 + 4H_2O + 0.7374O_2 + 21.57N_2$$

The mole fraction of water in the products is

$$y_v = \frac{N_{\text{H2O}}}{N_{\text{prod}}} = \frac{4\text{ kmol}}{(3+4+0.7374+21.57)\text{ kmol}} = \frac{4\text{ kmol}}{29.31\text{ kmol}} = 0.1365$$

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

$$P_v = y_v P = (0.1365)(101.325\text{ kPa}) = 13.83\text{ kPa}$$

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

$$T_{\text{dp}} = T_{\text{sat@13.83 kPa}} = 52.3°\text{C} \cong 325\text{ K}$$

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	$\bar{h}_{325\text{K}}$ kJ/kmol
C_3H_8	-103,850	---	---
O_2	0	8682	9473
N_2	0	8669	9452
H_2O (*g*)	-241,820	9904	10,808
CO_2	-393,520	9364	10,378

Substituting,

$$\begin{aligned}-\bar{Q}_{\text{out}} &= (3)(-393{,}520+10{,}378-9364)+(4)(-241{,}820+10{,}808-9904)+(0.7374)(0+9473-8682)\\ &\quad +(21.57)(0+9452-8669)-(1)(-103{,}850)-0\\ &= -2{,}019{,}860\text{ kJ/kmol C}_3\text{H}_8\end{aligned}$$

or

$$\bar{Q}_{\text{out}} = 2{,}019{,}860\text{ kJ/kmol C}_3\text{H}_8$$

Then the heat transfer per kg of fuel is

$$Q_{\text{out}} = \frac{\bar{Q}_{\text{out}}}{M_{\text{fuel}}} = \frac{2{,}019{,}860\text{ kJ/kmol fuel}}{44\text{ kg/kmol}} = \mathbf{45{,}900\ kJ/kg\ C_3H_8}$$

15-59 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.

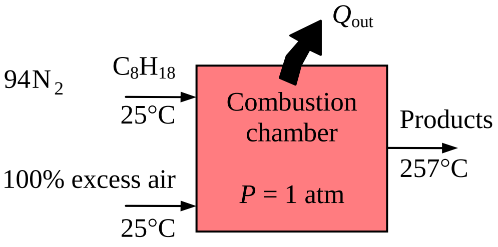
Properties The molar masses of propane and air are 44 kg/kmol and 29 kg/kmol, respectively (Table A-1).

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

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

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	$\bar{h}_{530K}$ 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,

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

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

Then the heat transfer per kg of fuel is

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

15-60 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 The mass fractions of the constituent of the coal when the ash is substituted are

$$\mathrm{mf}_\mathrm{C} = \frac{m_\mathrm{C}}{m_\mathrm{total}} = \frac{39.25\ \mathrm{kg}}{(100-11.20)\ \mathrm{kg}} = \frac{39.25\ \mathrm{kg}}{88.80\ \mathrm{kg}} = 0.4420$$

$$\mathrm{mf}_\mathrm{H2} = \frac{m_\mathrm{H2}}{m_\mathrm{total}} = \frac{6.93\ \mathrm{kg}}{88.80\ \mathrm{kg}} = 0.07804$$

$$\mathrm{mf}_\mathrm{O2} = \frac{m_\mathrm{O2}}{m_\mathrm{total}} = \frac{41.11\ \mathrm{kg}}{88.80\ \mathrm{kg}} = 0.4630$$

$$\mathrm{mf}_\mathrm{N2} = \frac{m_\mathrm{N2}}{m_\mathrm{total}} = \frac{0.72\ \mathrm{kg}}{88.80\ \mathrm{kg}} = 0.00811$$

$$\mathrm{mf}_\mathrm{S} = \frac{m_\mathrm{S}}{m_\mathrm{total}} = \frac{0.79\ \mathrm{kg}}{88.80\ \mathrm{kg}} = 0.00890$$

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

We now consider 100 kg of this mixture. Then the mole numbers of each component are

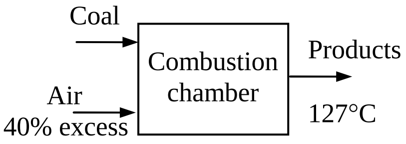
$$N_\mathrm{C} = \frac{m_\mathrm{C}}{M_\mathrm{C}} = \frac{44.20\ \mathrm{kg}}{12\ \mathrm{kg/kmol}} = 3.683\ \mathrm{kmol}$$

$$N_\mathrm{H2} = \frac{m_\mathrm{H2}}{M_\mathrm{H2}} = \frac{7.804\ \mathrm{kg}}{2\ \mathrm{kg/kmol}} = 3.902\ \mathrm{kmol}$$

$$N_\mathrm{O2} = \frac{m_\mathrm{O2}}{M_\mathrm{O2}} = \frac{46.30\ \mathrm{kg}}{32\ \mathrm{kg/kmol}} = 1.447\ \mathrm{kmol}$$

$$N_\mathrm{N2} = \frac{m_\mathrm{N2}}{M_\mathrm{N2}} = \frac{0.811\ \mathrm{kg}}{28\ \mathrm{kg/kmol}} = 0.0290\ \mathrm{kmol}$$

$$N_\mathrm{S} = \frac{m_\mathrm{S}}{M_\mathrm{S}} = \frac{0.890\ \mathrm{kg}}{32\ \mathrm{kg/kmol}} = 0.0278\ \mathrm{kmol}$$



The mole number of the mixture and the mole fractions are

$$N_m = 3.683 + 3.902 + 1.447 + 0.0290 + 0.0278 = 9.089\ \mathrm{kmol}$$

$$y_\mathrm{C} = \frac{N_\mathrm{C}}{N_m} = \frac{3.683\ \mathrm{kmol}}{9.089\ \mathrm{kmol}} = 0.4052$$

$$y_\mathrm{H2} = \frac{N_\mathrm{H2}}{N_m} = \frac{3.902\ \mathrm{kmol}}{9.089\ \mathrm{kmol}} = 0.4293$$

$$y_\mathrm{O2} = \frac{N_\mathrm{O2}}{N_m} = \frac{1.447\ \mathrm{kmol}}{9.089\ \mathrm{kmol}} = 0.1592$$

$$y_\mathrm{N2} = \frac{N_\mathrm{N2}}{N_m} = \frac{0.0290\ \mathrm{kmol}}{9.089\ \mathrm{kmol}} = 0.00319$$

$$y_\mathrm{S} = \frac{N_\mathrm{S}}{N_m} = \frac{0.0278\ \mathrm{kmol}}{9.089\ \mathrm{kmol}} = 0.00306$$

Then, the combustion equation in this case 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 $\bar{Q}_{\text{out}}=253{,}244\text{ kJ/kmol C}_8\text{H}_{18}$

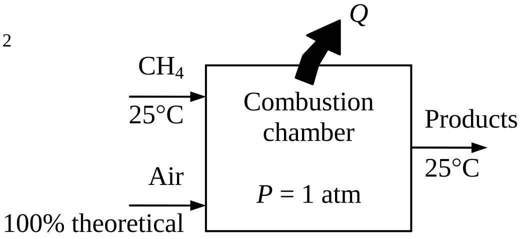
Then the heat transfer per kg of fuel is

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

15-61 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,

$$-Q_{\mathrm{out}} = (1)(-393{,}520) + (2)(-285{,}830) + 0 - (1)(-74{,}850) - 0 - 0 = -890{,}330 \text{ kJ / kmol CH}_4$$

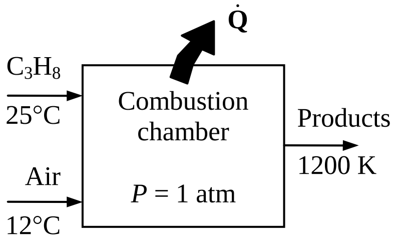
or

$$Q_{\mathrm{out}} = \mathbf{890{,}330 \text{ kJ / kmol CH}_4}$$

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.

15-62 Liquid propane is burned with 150 percent excess air during a steady-flow combustion process. The mass flow rate of air and the rate of heat transfer from the combustion chamber 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** 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 fuel is burned completely with 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

$$C_3H_8(\ell)+2.5a_{th}(O_2+3.76N_2)\longrightarrow 3CO_2+4H_2O+1.5a_{th}O_2+(2.5)(3.76a_{th})N_2$$

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

$$2.5a_{th}=3+2+1.5a_{th} \quad\longrightarrow\quad a_{th}=5$$

Thus,

$$C_3H_8(\ell)+12.5(O_2+3.76N_2)\longrightarrow 3CO_2+4H_2O+7.5O_2+47N_2$$

(*a*) The air-fuel ratio for this combustion process is

$$\mathrm{AF}=\frac{m_{air}}{m_{fuel}}=\frac{(12.5\times 4.76\text{ kmol})(29\text{ kg/kmol})}{(3\text{ kmol})(12\text{ kg/kmol})+(4\text{ kmol})(2\text{ kg/kmol})}=39.22\text{ kg air/kg fuel}$$

Thus, $\dot{m}_{air}=(\mathrm{AF})(\dot{m}_{fuel})=(39.22\text{ kg air/kg fuel})(1.2\text{ kg fuel/min})=\mathbf{47.1\text{ kg air/min}}$

(*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

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

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**	$\bar{\mathbf{h}}_{\mathbf{285\ K}}$ **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

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

Substituting,

$$\begin{aligned}-Q_{out}&=(3)(-393{,}520+53{,}848-9364)+(4)(-241{,}820+44{,}380-9904)+(7.5)(0+38{,}447-8682)\\&\quad+(47)(0+36{,}777-8669)-(1)(-118{,}910+h_{298}-h_{298})-(12.5)(0+8296.5-8682)\\&\quad-(47)(0+8286.5-8669)\\&=-190{,}464\text{ kJ/kmol }C_3H_8\end{aligned}$$

or $\qquad Q_{out}=190{,}464\text{ kJ/kmol }C_3H_8$

Then the rate of heat transfer for a mass flow rate of 1.2 kg/min for the propane becomes

$$\dot{Q}_{out}=\dot{N}Q_{out}=\left(\frac{\dot{m}}{N}\right)Q_{out}=\left(\frac{1.2\text{ kg/min}}{44\text{ kg/kmol}}\right)(190{,}464\text{ kJ/kmol})=\mathbf{5194\text{ kJ/min}}$$

15-63E Liquid octane is burned with 180 percent theoretical air during a steady-flow combustion process. The AF ratio and the heat transfer from the combustion chamber 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** Combustion is complete.

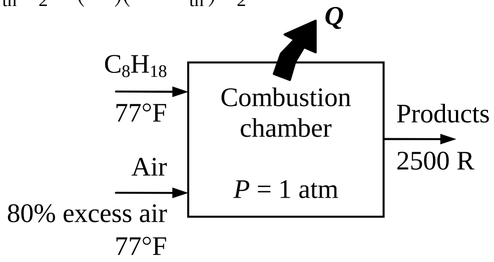
Properties The molar masses of C_3H_{18} and air are 54 kg/kmol and 29 kg/kmol, respectively (Table A-1).

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 of C_8H_{18}, the combustion equation can be written as

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

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

$$1.8a_{\mathrm{th}}=8+4.5+0.8a_{\mathrm{th}}\quad\longrightarrow\quad a_{\mathrm{th}}=12.5$$



Thus,

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

(*a*)
$$\mathrm{AF}=\frac{m_{\mathrm{air}}}{m_{\mathrm{fuel}}}=\frac{(22.5\times 4.76\ \mathrm{lbmol})(29\ \mathrm{lbm/lbmol})}{(8\ \mathrm{lbmol})(12\ \mathrm{lbm/lbmol})+(9\ \mathrm{lbmol})(2\ \mathrm{lbm/lbmol})}=\mathbf{27.2\ \ lbmair/lbmfuel}$$

(*b*) 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

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

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	$\bar{\mathbf{h}}_{\mathbf{f}}^\circ$ **Btu/lbmol**	$\bar{\mathbf{h}}_{\mathbf{537\ R}}$ **Btu/lbmol**	$\bar{\mathbf{h}}_{\mathbf{2500\ R}}$ **Btu/lbmol**
C_8H_{18} (ℓ)	-107,530	---	---
O_2	0	3725.1	19,443
N_2	0	3729.5	18,590
CO_2	-169,300	4027.5	27,801
H_2O (*g*)	-104,040	4258.0	22,735

Thus,

$$\begin{aligned}-Q_{\mathrm{out}}&=(8)(-169{,}300+27{,}801-4027.5)+(9)(-104{,}040+22{,}735-4258)+(10)(0+19{,}443-3725.1)\\&\quad+(84.6)(0+18{,}590-3729.5)-(1)(-107{,}530)-0-0\\&=-412{,}372\ \ \mathrm{Btu/lbmol\ C_8H_{18}}\end{aligned}$$

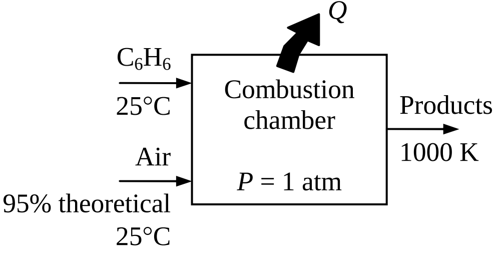
or $\qquad Q_{\mathrm{out}}=\mathbf{412{,}372\ Btu/lbmol\ C_8H_{18}}$

15-64 Benzene gas is burned with 95 percent theoretical air during a steady-flow combustion process. The mole fraction of the CO in the products and the heat transfer from the combustion chamber 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.

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

$$C_6H_6 + 0.95 \times 7.5(O_2 + 3.76N_2) \longrightarrow xCO_2 + (6-x)CO + 3H_2O + 26.79N_2$$

O_2 balance: $\quad 0.95 \times 7.5 = x + (6-x)/2 + 1.5 \longrightarrow x = 5.25$

Thus, $\quad C_6H_6 + 7.125(O_2 + 3.76N_2) \longrightarrow 5.25CO_2 + 0.75CO + 3H_2O + 26.79N_2$

The mole fraction of CO in the products is

$$y_{CO} = \frac{N_{CO}}{N_{total}} = \frac{0.75}{5.25 + 0.75 + 3 + 26.79} = 0.021 \quad \text{or} \quad \mathbf{2.1\%}$$

(*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

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

since all of the reactants are at 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**	$\bar{h}_{298\ K}$ **kJ/kmol**	$\bar{h}_{1000\ K}$ **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

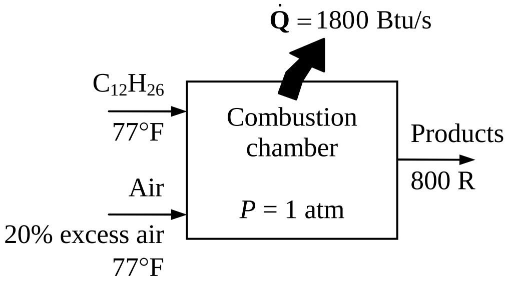
Thus,

$$\begin{aligned}-Q_{out} &= (5.25)(-393{,}520 + 42{,}769 - 9364) + (0.75)(-110{,}530 + 30{,}355 - 8669)\\ &\quad + (3)(-241{,}820 + 35{,}882 - 9904) + (26.79)(0 + 30{,}129 - 8669) - (1)(82{,}930) - 0 - 0\\ &= -2{,}112{,}779 \text{ kJ / kmol } C_6H_6\end{aligned}$$

or $\quad \dot{Q}_{out} = \mathbf{2{,}112{,}800\ \ kJ/kmol\ \ C_6H_6}$

15-65E 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

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

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

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

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 $E_{in} - E_{out} = \Delta E_{system}$ applied on the combustion chamber with W = 0. It reduces to

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

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	$\mathbf{\bar{h}_f^\circ}$ **Btu/lbmol**	$\mathbf{\bar{h}_{537\ R}}$ **Btu/lbmol**	$\mathbf{\bar{h}_{800\ R}}$ **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,

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

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

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

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

15-66 [*Also solved by EES on enclosed CD***]** 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

$$\mathrm{C_8H_{18}}(\mathrm{g}) + 1.3a_{\mathrm{th}}(\mathrm{O_2} + 3.76\mathrm{N_2}) \longrightarrow 8\mathrm{CO_2} + 9\mathrm{H_2O} + 0.3a_{\mathrm{th}}\mathrm{O_2} + (1.3)(3.76a_{\mathrm{th}})\mathrm{N_2}$$

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

$$\mathrm{O_2}\text{ balance:}\quad 1.3a_{\mathrm{th}} = 8 + 4.5 + 0.3a_{\mathrm{th}} \longrightarrow a_{\mathrm{th}} = 12.5$$

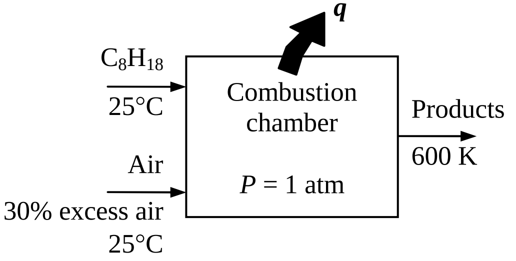
Thus,

$$\mathrm{C_8H_{18}}(\mathrm{g}) + 16.25(\mathrm{O_2} + 3.76\mathrm{N_2}) \longrightarrow 8\mathrm{CO_2} + 9\mathrm{H_2O} + 3.75\mathrm{O_2} + 61.1\mathrm{N_2}$$

Therefore, 16.25 × 4.76 = 77.35 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

$$P_{v,\mathrm{in}} = \phi_{\mathrm{air}} P_{\mathrm{sat@25^\circ C}} = (0.60)(3.1698\ \mathrm{kPa}) = 1.902\ \mathrm{kPa}$$

Assuming ideal gas behavior, the number of moles of the moisture that accompanies 77.35 kmol of incoming dry air is determined to be



$$N_{v,\mathrm{in}} = \left(\frac{P_{v,\mathrm{in}}}{P_{\mathrm{total}}}\right) N_{\mathrm{total}} = \left(\frac{1.902\ \mathrm{kPa}}{101.325\ \mathrm{kPa}}\right)(77.35 + N_{v,\mathrm{in}}) \longrightarrow N_{v,\mathrm{in}} = 1.48\ \mathrm{kmol}$$

The balanced combustion equation is obtained by adding 1.48 kmol of H_2O to both sides of the equation,

$$\mathrm{C_8H_{18}}(\mathrm{g}) + 16.25(\mathrm{O_2} + 3.76\mathrm{N_2}) + 1.48\mathrm{H_2O} \longrightarrow 8\mathrm{CO_2} + 10.48\mathrm{H_2O} + 3.75\mathrm{O_2} + 61.1\mathrm{N_2}$$

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

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

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**	$\bar{\mathbf{h}}_{\mathbf{298\ K}}$ **kJ/kmol**	$\bar{\mathbf{h}}_{\mathbf{600\ K}}$ **kJ/kmol**
C_8H_{18} (*g*)	-208,450	---	---
O_2	0	8682	17,929
N_2	0	8669	17,563
H_2O (*g*)	-241,820	9904	20,402
CO_2	-393,520	9364	22,280

Substituting,

$$\begin{aligned}-Q_{\mathrm{out}} &= (8)(-393{,}520 + 22{,}280 - 9364) + (10.48)(-241{,}820 + 20{,}402 - 9904)\\ &\quad + (3.75)(0 + 17{,}929 - 8682) + (61.1)(0 + 17{,}563 - 8669)\\ &\quad - (1)(-208{,}450) - (1.48)(-241{,}820) - 0 - 0\\ &= -4{,}324{,}643\ \mathrm{kJ/kmol\ C_8H_{18}}\end{aligned}$$

Thus 4,324,643 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

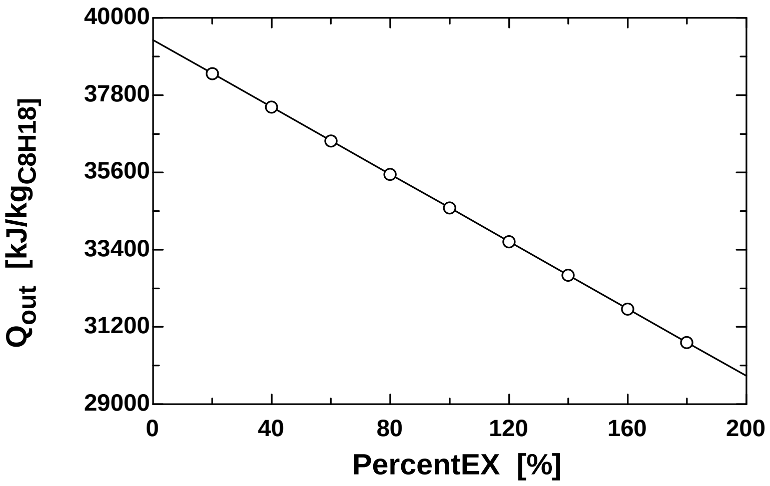
$$q = \frac{Q_{\mathrm{out}}}{M} = \frac{4{,}324{,}643\ \mathrm{kJ}}{114\mathrm{kg}} = \mathbf{37{,}935\ kJ/kg\ \ C_8H_{18}}$$

15-67 EES Problem 15-66 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 = 30 "[%]"}
Ex = PercentEX/100 "[%Excess air/100]"
P_air1 = 101.3 [kPa]
T_air1 = 25+273 "[K]"
RH_1 = 60/100  "[%]"
T_prod = 600 [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."
```

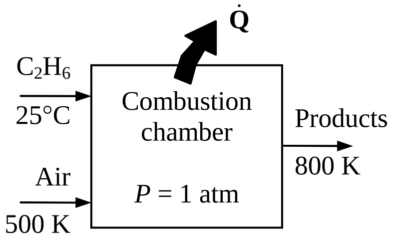
Q_{out} [kJ/kgC8H18]	PercentEX [%]
39374	0
38417	20
37460	40
36503	60
35546	80
34588	100
33631	120
32674	140
31717	160
30760	180
29803	200



15-68 Ethane gas is burned with stoichiometric amount of 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** Combustion is complete.

Properties The molar mass of C_2H_6 is 30 kg/kmol (Table A-1).



Analysis The theoretical combustion equation of C_2H_6 is

$$C_2H_6 + a_{\text{th}}(O_2 + 3.76N_2) \longrightarrow 2CO_2 + 3H_2O + 3.76a_{\text{th}}N_2$$

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

$$a_{\text{th}} = 2 + 1.5 = 3.5$$

Then the actual combustion equation can be written as

$$C_2H_6 + 3.5(O_2 + 3.76N_2) \longrightarrow 1.9CO_2 + 0.1CO + 3H_2O + 0.05O_2 + 13.16N_2$$

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	$\bar{h}_{500\text{ K}}$ kJ/kmol	$\bar{h}_{298\text{ K}}$ kJ/kmol	$\bar{h}_{800\text{ K}}$ kJ/kmol
C_2H_6 (*g*)	-84,680	---	---	---
O_2	0	14,770	8682	24,523
N_2	0	14,581	8669	23,714
H_2O (*g*)	-241,820	---	9904	27,896
CO	-110,530	---	8669	23,844
CO_2	-393,520	---	9364	32,179

Thus,

$$\begin{aligned}
-Q_{\text{out}} &= (1.9)(-393{,}520 + 32{,}179 - 9364) + (0.1)(-110{,}530 + 23{,}844 - 8669) \\
&\quad + (3)(-241{,}820 + 27{,}896 - 9904) + (0.05)(0 + 24{,}523 - 8682) + (13.16)(0 + 23{,}714 - 8669) \\
&\quad - (1)(-84{,}680 + h_{298} - h_{298}) - (3.5)(0 + 14{,}770 - 8682) - (13.16)(0 + 14{,}581 - 8669) \\
&= -1{,}201{,}005 \text{ kJ / kmol } C_2H_6
\end{aligned}$$

or $\quad Q_{\text{out}} = 1{,}201{,}005 \text{ kJ / kmol } C_2H_6$

Then the rate of heat transfer for a mass flow rate of 3 kg/h for the ethane becomes

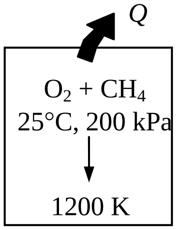
$$\dot{Q}_{\text{out}} = \dot{N}Q_{\text{out}} = \left(\frac{\dot{m}}{M}\right)Q_{\text{out}} = \left(\frac{5 \text{ kg/h}}{30 \text{ kg/kmol}}\right)(1{,}201{,}005 \text{ kJ/kmol}) = \mathbf{200{,}170\ kJ/h}$$

15-69 [*Also solved by EES on enclosed CD***]** 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



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

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

Then the combustion equation can be written as

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

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,

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

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

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

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**	$\bar{\mathbf{h}}_{\mathbf{1200\ K}}$ **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,

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

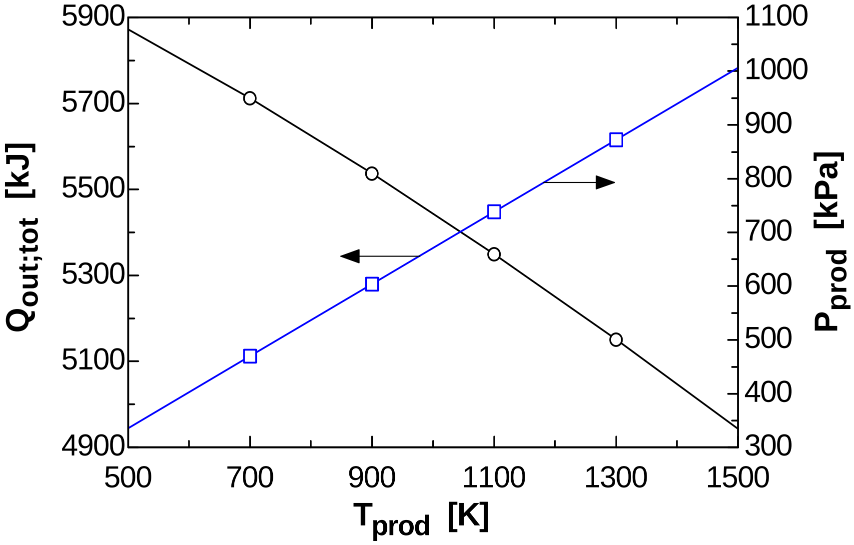
Thus $Q_{\text{out}} = \mathbf{5252\,kJ}$ of heat is transferred from the combustion chamber as 120 g of CH_4 burned in this combustion chamber.

15-70 EES Problem 15-69 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



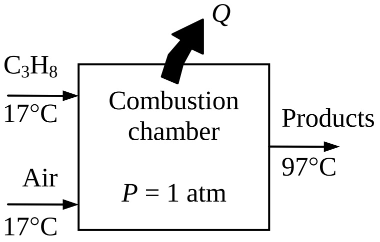
15-71 Propane is burned with 50% excess air. The heat transfer per kmol of fuel is to be determined for given reactants and products temperatures.

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 17°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 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 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	$\bar{h}_f^\circ$ kJ/kmol	$\bar{h}_{290K}$ kJ/kmol	$\bar{h}_{298K}$ kJ/kmol	$\bar{h}_{370K}$ kJ/kmol
C_3H_8	-103,850	---	---	---
O_2	0	8443	8682	10,809
N_2	0	8432	8669	10,763
H_2O (*g*)	-241,820	-	9904	12,331
CO_2	-393,520	-	9364	12,148

Substituting,

$$\begin{aligned}-Q_{out} &= (3)(-393{,}520 + 12{,}148 - 9364) + (4)(-241{,}820 + 12{,}331 - 9904) + (2.5)(0 + 10{,}809 - 8682) \\ &\quad + (28.2)(0 + 10{,}763 - 8669) - (1)(-103{,}850) - (7.5)(0 + 8443 - 8682) - (28.2)(0 + 8432 - 8669) \\ &= -1{,}953{,}000 \text{ kJ/kmol } C_3H_8\end{aligned}$$

or $\quad Q_{out} = \mathbf{1{,}953{,}000\ kJ/kmol\ C_3H_8}$

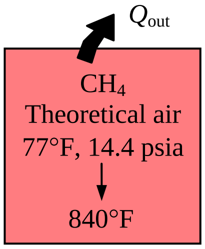
Discussion We neglected the enthalpy difference of propane between the standard temperature 25°C and the given temperature 17°C since the data is not available. It may be shown that the contribution of this would be small.

15-72E 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 $E_{\text{in}} - E_{\text{out}} = \Delta E_{\text{system}}$ 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

$$-Q_{\text{out}} = \sum N_P\left(\bar{h}_f^\circ + \bar{h}_{1300\,\mathrm{R}} - \bar{h}_{537\,\mathrm{R}} - R_uT\right)_P - \sum N_R\left(\bar{h}_f^\circ - R_uT\right)_R$$

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**	$\bar{\mathbf{h}}_{1300\mathrm{R}}$ **Btu/lbmol**
CH_4	-32,210	---	---
O_2	0	3725.1	9412.9
N_2	0	3729.5	9153.9
H_2O (*g*)	-104,040	4258.0	10,714.5
CO_2	-169,300	4027.5	12,136.9

Thus,

$$\begin{aligned}
-Q_{\text{out}} &= (1)(-169{,}300 + 12{,}136.9 - 4027.5 - 1.9858\times 1300)\\
&\quad + (2)(-104{,}040 + 10{,}714.5 - 4258.0 - 1.9858\times 1300)\\
&\quad + (7.52)(0 + 9153.9 - 3729.5 - 1.9858\times 1300)\\
&\quad - (1)(-32{,}210 - 1.9858\times 537) - (2)(-1.9858\times 537) - (7.52)(-1.9858\times 537)\\
&= -299{,}300\ \text{Btu/lbmol}\ \mathrm{CH_4}
\end{aligned}$$

Thus

$$Q_{\text{out}} = \mathbf{299{,}300\ Btu/lbmol\ CH_4}$$

15-73 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

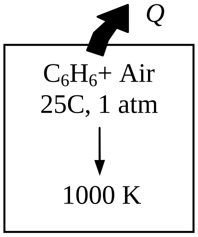
$$\text{HV} = \frac{Q}{m} = \frac{180\ \text{kJ}}{0.010\ \text{kg}} = \mathbf{18{,}000\ kJ/kg}$$

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

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

15-74 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

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

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

It yields

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

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

Substance	$\overline{\mathbf{h}}_\mathbf{f}^\circ$ **kJ/kmol**	$\overline{\mathbf{h}}_{\mathbf{298\ K}}$ **kJ/kmol**	$\overline{\mathbf{h}}_{\mathbf{1000\ K}}$ **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,

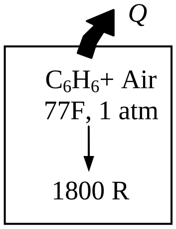
$$\begin{aligned}-Q_{out}&=(5.52)(-393{,}520+42{,}769-9364-8.314\times1000)\\&\quad+(0.48)(-110{,}530+30{,}355-8669-8.314\times1000)\\&\quad+(3)(-241{,}820+35{,}882-9904-8.314\times1000)\\&\quad+(2.49)(0+31{,}389-8682-8.314\times1000)\\&\quad+(36.66)(0+30{,}129-8669-8.314\times1000)\\&\quad-(1)(82{,}930-8.314\times298)-(9.75)(4.76)(-8.314\times298)\\&=-2{,}200{,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

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

15-75E 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

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

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

It yields

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

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

Substance	$\overline{\mathbf{h}}_{\mathbf{f}}^\circ$ **Btu/lbmol**	$\overline{\mathbf{h}}_{\mathbf{537\ R}}$ **Btu/lbmol**	$\overline{\mathbf{h}}_{\mathbf{1800\ R}}$ **Btu/lbmol**
C_6H_6 (*g*)	35,6860	---	---
O_2	0	3725.1	13,485.8
N_2	0	3729.5	12,956.3
H_2O (*g*)	-104,040	4258.0	15,433.0
CO	-47,540	3725.1	13,053.2
CO_2	-169,300	4027.5	18,391.5

Thus,

$$\begin{aligned} -Q_{out} &= (5.52)(-169{,}300 + 18{,}391.5 - 4027.5 - 1.986 \times 1800) \\ &\quad + (0.48)(-47{,}540 + 13{,}053.2 - 3725.1 - 1.986 \times 1800) \\ &\quad + (3)(-104{,}040 + 15{,}433.0 - 4258.0 - 1.986 \times 1800) \\ &\quad + (2.49)(0 + 13{,}485.8 - 3725.1 - 1.986 \times 1800) \\ &\quad + (36.66)(0 + 12{,}956.3 - 3729.5 - 1.986 \times 1800) \\ &\quad - (1)(35{,}680 - 1.986 \times 537) - (9.75)(4.76)(-1.986 \times 537) \\ &= -946{,}870 \text{ Btu} \end{aligned}$$

or $\quad Q_{out} = \mathbf{946{,}870\ Btu}$

15-76 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.

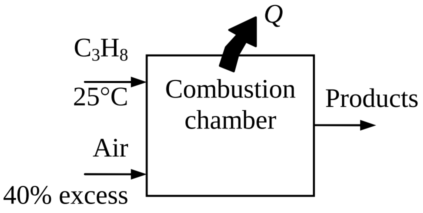
Assumptions **1** Combustion is complete. **2** The combustion products contain CO_2, H_2O, O_2, and N_2 only.

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

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

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



Carbon balance: $B = 3$

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

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

$0.4a_{\mathrm{th}} = E$

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

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

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

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

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

The vapor mole fraction is

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

The kmoles of water condensed is determined from

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

The steady-flow energy balance is expressed as

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

where

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

$$\begin{aligned} H_R &= \bar{h}^{o}_{f\ \mathrm{fuel@25°C}} + 7\bar{h}_{\mathrm{O2@25°C}} + 26.32\bar{h}_{\mathrm{N2@25°C}} \\ &= (-103{,}847\,\mathrm{kJ/kmol}) + 7(0) + 26.32(0) = -103{,}847\,\mathrm{kJ/kmol} \end{aligned}$$

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

Substituting into the energy balance equation,

$$
\begin{aligned}
\dot{N}_{\mathrm{fuel}} H_R &= \dot{Q}_{\mathrm{fuel}} + \dot{N}_{\mathrm{fuel}} H_P \\
\dot{N}_{\mathrm{fuel}}(-103{,}847\ \mathrm{kJ/kmol}) &= 32{,}969\ \mathrm{kJ/h} + \dot{N}_{\mathrm{fuel}}(-2.577\times 10^{6}\ \mathrm{kJ/kmol}) \\
\dot{N}_{\mathrm{fuel}} &= 0.01333\ \mathrm{kmol/h}
\end{aligned}
$$

The molar and mass flow rates of the liquid water are

$$
\dot{N}_{\mathrm{w}} = N_{\mathrm{w}}\dot{N}_{\mathrm{fuel}} = (1.503\ \ \mathrm{kmol/kmol\,fuel})(0.01333\ \ \mathrm{kmol\,fuel/h}) = 0.02003\ \ \mathrm{kmol/h}
$$

$$
\dot{m}_{\mathrm{w}} = \dot{N}_{\mathrm{w}} M_{\mathrm{w}} = (0.02003\ \ \mathrm{kmol/h})(18\ \ \mathrm{kg/kmol}) = 0.3608\ \ \mathrm{kg/h}
$$

The volume flow rate of liquid water is

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

Adiabatic Flame Temperature

15-77C 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.

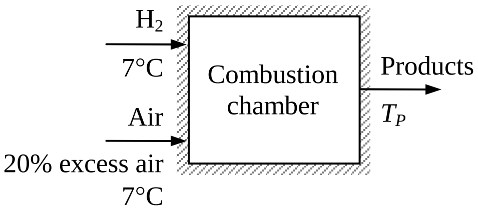
15-78C Under the conditions of complete combustion with stoichiometric amount of air.

15-79 [*Also solved by EES on enclosed CD*] Hydrogen is burned with 20 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 20% excess air is

$$H_2 + 0.6(O_2 + 3.76N_2) \longrightarrow H_2O + 0.1O_2 + 2.256N_2$$

From the tables,

Substance	$\bar{h}_f^\circ$ **kJ/kmol**	$\bar{h}_{280\text{ K}}$ **kJ/kmol**	$\bar{h}_{298\text{ K}}$ **kJ/kmol**
H_2	0	7945	8468
O_2	0	8150	8682
N_2	0	8141	8669
H_2O (*g*)	-241,820	9296	9904

Thus,

$$\begin{aligned}(1)\left(-241{,}820 + \bar{h}_{H_2O} - 9904\right) + (0.1)\left(0 + \bar{h}_{O_2} - 8682\right) + (2.256)\left(0 + \bar{h}_{N_2} - 8669\right)\\ = (1)(0 + 7945 - 8468) + (0.6)(0 + 8150 - 8682) + (2.256)(0 + 8141 - 8669)\end{aligned}$$

It yields

$$\bar{h}_{H_2O} + 0.1\bar{h}_{O_2} + 2.256\bar{h}_{N_2} = 270{,}116 \text{ kJ}$$

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 270,116/(1 + 0.1 + 2.256) = 80,488 kJ/kmol. This enthalpy value corresponds to about 2400 K for N_2. Noting that the majority of the moles are N_2, T_P will be close to 2400 K, but somewhat under it because of the higher specific heat of H_2O.

At 2300 K: $$\begin{aligned}\bar{h}_{H_2O} + 0.1\bar{h}_{O_2} + 2.256\bar{h}_{N_2} &= (1)(98{,}199) + (0.1)(79{,}316) + (2.256)(75{,}676)\\ &= 276{,}856 \text{ kJ (Higher than } 270{,}116 \text{ kJ)}\end{aligned}$$

At 2250 K: $$\begin{aligned}\bar{h}_{H_2O} + 0.1\bar{h}_{O_2} + 2.256\bar{h}_{N_2} &= (1)(95{,}562) + (0.1)(77{,}397) + (2.256)(73{,}856)\\ &= 269{,}921 \text{ kJ (Lower than } 270{,}116 \text{ kJ)}\end{aligned}$$

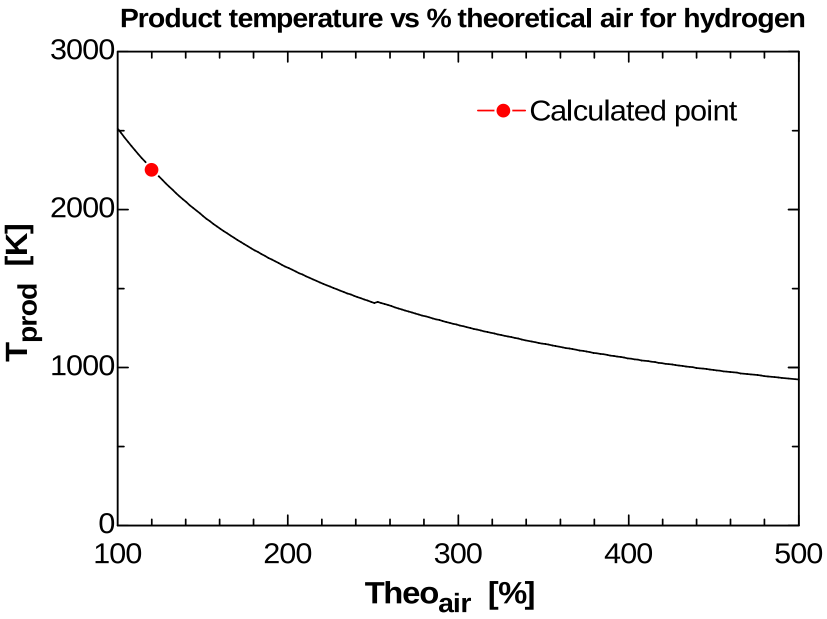
By interpolation, T_P = **2251.4 K**

15-80 EES Problem 15-79 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 = 280 [K]
T_air = 280  [K]
Theo_air = 200  "%"
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"
```

$Theo_{air}$ [%]	T_{prod} [K]
100	2512
144.4	2008
188.9	1693
233.3	1476
277.8	1318
322.2	1197
366.7	1102
411.1	1025
455.6	960.9
500	907.3

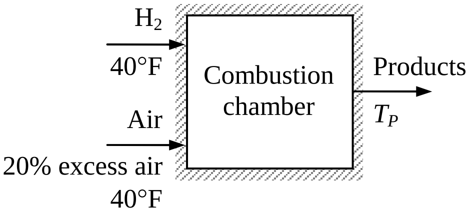


15-81E Hydrogen is burned with 20 percent excess air during a steady-flow combustion process. The exit temperature of product gases is to be determined.

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 $E_{in} - E_{out} = \Delta E_{system}$ 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 20% excess air is

$$H_2 + 0.6(O_2 + 3.76N_2) \longrightarrow H_2O + 0.1O_2 + 2.256N_2$$

From the tables,

Substance	$\bar{h}_f^\circ$ **Btu/lbmol**	$\bar{h}_{500\ R}$ **Btu/lbmol**	$\bar{h}_{537\ R}$ **Btu/lbmol**
H_2	0	3386.1	3640.3
O_2	0	3466.2	3725.1
N_2	0	3472.2	3729.5
H_2O (g)	-104,040	3962.0	4258.0

Thus,

$$(1)\left(-104{,}040 + \bar{h}_{H_2O} - 4258\right) + (0.1)\left(0 + \bar{h}_{O_2} - 3725.1\right) + (2.256)\left(0 + \bar{h}_{N_2} - 3729.5\right)$$
$$= (1)(0 + 3386.1 - 3640.3) + (0.6)(0 + 3466.2 - 3725.1) + (2.256)(0 + 3472.2 - 3729.5)$$

It yields

$$\bar{h}_{H_2O} + 0.1\bar{h}_{O_2} + 2.256\bar{h}_{N_2} = 116{,}094 \text{ Btu}$$

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 116,094/(1 + 0.1+ 2.256) = 34,593 Btu/lbmol. This enthalpy value corresponds to about 4400 R for N_2. Noting that the majority of the moles are N_2, T_P will be close to 4400 R, but somewhat under it because of the higher specific heat of H_2O.

At 4020 R: $\bar{h}_{H_2O} + 0.1\bar{h}_{O_2} + 2.256\bar{h}_{N_2} = (1)(40{,}740) + (0.1)(32{,}989) + (2.256)(31{,}503)$
$= 115{,}110 \text{ Btu (Lower than } 116{,}094 \text{ Btu)}$

At 4100 R: $\bar{h}_{H_2O} + 0.1\bar{h}_{O_2} + 2.256\bar{h}_{N_2} = (1)(41{,}745) + (0.1)(33{,}722) + (2.256)(32{,}198)$
$= 117{,}756 \text{ Btu (Higher than } 116{,}094 \text{ Btu)}$

By interpolation, T_P = **4054 R**

15-82 Methane is burned with 30 percent excess air. 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,

$$\mathrm{CH_4} + 1.3\times a_{\text{th}}(\mathrm{O_2} + 3.76N_2) \longrightarrow \mathrm{CO_2} + 2\mathrm{H_2O} + 0.3\mathrm{O_2} + 1.3\times a_{\text{th}} \times 3.76\mathrm{N_2}$$

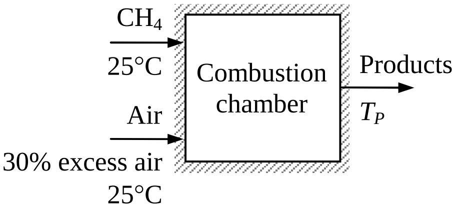
where a_{th} is the stoichiometric coefficient and is determined from the O_2 balance,

$$1.3a_{\text{th}} = 1 + 1 + 0.3a_{\text{th}} \longrightarrow a_{\text{th}} = 2$$

Thus, $\mathrm{CH_4} + 2.6(\mathrm{O_2} + 3.76N_2) \longrightarrow \mathrm{CO_2} + 2\mathrm{H_2O} + 0.3\mathrm{O_2} + 9.776\mathrm{N_2}$

From the tables,

Substance	$\bar{\mathbf{h}}_{\mathbf{f}}^\circ$ kJ/kmol	$\bar{\mathbf{h}}_{\mathbf{298K}}$ kJ/kmol
H_4 (*g*)	-74,850	---
O_2	0	8682
N_2	0	8669
H_2O (*g*)	-241,820	9904
CO_2	-393,520	9364



Thus,

$$\begin{aligned}&(1)\left(-393{,}520 + \bar{h}_{\mathrm{CO2}} - 9364\right) + (2)\left(-241{,}820 + \bar{h}_{\mathrm{H2O}} - 9904\right) + (0.3)\left(0 + \bar{h}_{\mathrm{O2}} - 8682\right)\\ &+ (9.776)\left(0 + \bar{h}_{\mathrm{N2}} - 8669\right) = (1)(-74{,}850) + 0 + 0\end{aligned}$$

It yields $\bar{h}_{\mathrm{CO2}} + 2\bar{h}_{\mathrm{H2O}} + 0.3\bar{h}_{\mathrm{O2}} + 9.776\bar{h}_{\mathrm{N2}} = 918{,}835\ \text{kJ}$

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 918,835/(1 + 2 + 0.3 + 9.776) = 70,269 kJ/kmol. This enthalpy value corresponds to about 2150 K for N_2. Noting that the majority of the moles are N_2, T_P will be close to 2150 K, but somewhat under it because of the higher specific heat of H_2O.

At 2000 K:

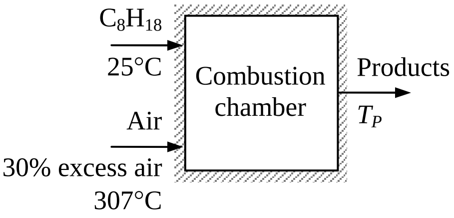
$$\begin{aligned}\bar{h}_{\mathrm{CO2}} + 2\bar{h}_{\mathrm{H2O}} + 0.3\bar{h}_{\mathrm{O2}} + 9.776\bar{h}_{\mathrm{N2}} &= 100{,}804 + 2\times 82{,}593 + 0.3\times 67{,}881 + 9.776\times 64{,}810\\ &= 919{,}936\ \text{kJ (Higher than 918,835 kJ)}\end{aligned}$$

At 1980 K:

$$\begin{aligned}\bar{h}_{\mathrm{CO2}} + 2\bar{h}_{\mathrm{H2O}} + 0.3\bar{h}_{\mathrm{O2}} + 9.776\bar{h}_{\mathrm{N2}} &= 99{,}606 + 2\times 81{,}573 + 0.3\times 67{,}127 + 9.776\times 64{,}090\\ &= 909{,}434\ \text{kJ (Lower than 918,835 kJ)}\end{aligned}$$

By interpolation, T_P = **1998 K = 1725°C**

15-83 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,

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

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

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,

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

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

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**

15-84 A certain coal is burned with 50 percent excess air adiabatically during a steady-flow combustion process. The 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.

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 The mass fractions of the constituent of the coal when the ash is substituted are

$$\mathrm{mf}_C = \frac{m_C}{m_{\text{total}}} = \frac{79.61\,\text{kg}}{(100-8.62)\,\text{kg}} = \frac{79.61\,\text{kg}}{91.38\,\text{kg}} = 0.8712$$

$$\mathrm{mf}_{H2} = \frac{m_{H2}}{m_{\text{total}}} = \frac{4.66\,\text{kg}}{91.38\,\text{kg}} = 0.05100$$

$$\mathrm{mf}_{O2} = \frac{m_{O2}}{m_{\text{total}}} = \frac{4.76\,\text{kg}}{91.38\,\text{kg}} = 0.05209$$

$$\mathrm{mf}_{N2} = \frac{m_{N2}}{m_{\text{total}}} = \frac{1.83\,\text{kg}}{91.38\,\text{kg}} = 0.02003$$

$$\mathrm{mf}_S = \frac{m_S}{m_{\text{total}}} = \frac{0.52\,\text{kg}}{91.38\,\text{kg}} = 0.00569$$

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

We now consider 100 kg of this mixture. Then the mole numbers of each component are

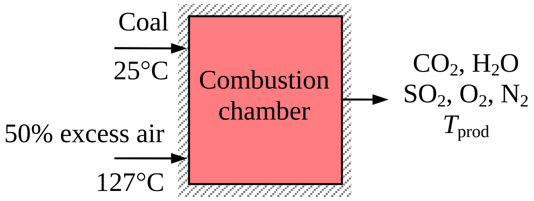
$$N_C = \frac{m_C}{M_C} = \frac{87.12\,\text{kg}}{12\,\text{kg/kmol}} = 7.26\,\text{kmol}$$

$$N_{H2} = \frac{m_{H2}}{M_{H2}} = \frac{5.10\,\text{kg}}{2\,\text{kg/kmol}} = 2.55\,\text{kmol}$$

$$N_{O2} = \frac{m_{O2}}{M_{O2}} = \frac{5.209\,\text{kg}}{32\,\text{kg/kmol}} = 0.1628\,\text{kmol}$$

$$N_{N2} = \frac{m_{N2}}{M_{N2}} = \frac{2.003\,\text{kg}}{28\,\text{kg/kmol}} = 0.07154\,\text{kmol}$$

$$N_S = \frac{m_S}{M_S} = \frac{0.569\,\text{kg}}{32\,\text{kg/kmol}} = 0.01778\,\text{kmol}$$



The mole number of the mixture and the mole fractions are

$$N_m = 7.26 + 2.55 + 0.1628 + 0.07154 + 0.01778 = 10.062\,\text{kmol}$$

$$y_C = \frac{N_C}{N_m} = \frac{7.26\,\text{kmol}}{10.06\,\text{kmol}} = 0.7215$$

$$y_{H2} = \frac{N_{H2}}{N_m} = \frac{2.55\,\text{kmol}}{10.06\,\text{kmol}} = 0.2535$$

$$y_{O2} = \frac{N_{O2}}{N_m} = \frac{0.1628\,\text{kmol}}{10.06\,\text{kmol}} = 0.01618$$

$$y_{N2} = \frac{N_{N2}}{N_m} = \frac{0.07154\,\text{kmol}}{10.06\,\text{kmol}} = 0.00711$$

$$y_S = \frac{N_S}{N_m} = \frac{0.01778\,\text{kmol}}{10.06\,\text{kmol}} = 0.00177$$

Then, the combustion equation in this case may be written as

$$0.7215\text{C} + 0.2535\text{H}_2 + 0.01618\text{O}_2 + 0.00711\text{N}_2 + 0.00177\text{S} + 1.5a_{\text{th}}(\text{O}_2 + 3.76\text{N}_2)$$
$$\longrightarrow x\text{CO}_2 + y\text{H}_2\text{O} + z\text{SO}_2 + k\text{N}_2 + 0.5a_{th}\text{O}_2$$

According to the species balances,

$$\text{C balance: } x = 0.7215$$

$$\text{H}_2 \text{ balance: } y = 0.2535$$

$$\text{S balance: } z = 0.00177$$

$$\text{O}_2 \text{ balance:}$$

$$0.01618 + 1.5a_{\text{th}} = x + 0.5y + z + 0.5a_{\text{th}}$$

$$1.5a_{\text{th}} - 0.5a_{\text{th}} = 0.7215 + 0.5(0.2535) + 0.00177 - 0.01617 \longrightarrow a_{\text{th}} = 0.8339$$

$$\text{N}_2 \text{ balance: } 0.00711 + 1.5 \times 3.76a_{\text{th}} = k \longrightarrow k = 0.00711 + 1.5 \times 3.76 \times 0.8339 = 4.710$$

Substituting,

$$0.7215\text{C} + 0.2535\text{H}_2 + 0.01617\text{O}_2 + 0.00711\text{N}_2 + 0.00177\text{S} + 1.2509(\text{O}_2 + 3.76\text{N}_2)$$
$$\longrightarrow 0.7215\text{CO}_2 + 0.2535\text{H}_2\text{O} + 0.0018\text{SO}_2 + 4.71\text{N}_2 + 0.4170\text{O}_2$$

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**	$\bar{\mathbf{h}}_{\mathbf{400K}}$ **kJ/kmol**
O_2	0	8682	11,711
N_2	0	8669	11,640
H_2O (*g*)	-241,820	9904	---
CO_2	-393,520	9364	---

Thus,

$$(0.7215)\left(-393{,}520 + \bar{h}_{\text{CO2}} - 9364\right) + (0.2535)\left(-241{,}820 + \bar{h}_{\text{H2O}} - 9904\right) + (0.4170)\left(0 + \bar{h}_{\text{O2}} - 8682\right)$$
$$+ (4.71)\left(0 + \bar{h}_{\text{N2}} - 8669\right) = (1.2509)\left(0 + 11{,}711 - 8682\right) + (4.703)(11{,}640 - 8669)$$

It yields $0.7215\bar{h}_{\text{CO2}} + 0.2535\bar{h}_{\text{H2O}} + 0.4170\bar{h}_{\text{O2}} + 4.71\bar{h}_{\text{N2}} = 416{,}706 \text{ kJ}$

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 416,706/(0.7215+0.2535+0.4170+4.71) = 68,290 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:

$$0.7215\bar{h}_{\text{CO2}} + 0.2535\bar{h}_{\text{H2O}} + 0.4170\bar{h}_{\text{O2}} + 4.71\bar{h}_{\text{N2}} = (0.7215)(100{,}804) + (0.2535)(82{,}593)$$
$$+ (0.4170)(67{,}881) + (4.71)(64{,}810)$$
$$= 427{,}229 \text{ kJ} \quad (\text{Higher than } 416{,}706 \text{ kJ})$$

At 1980 K:

$$0.7215\bar{h}_{\text{CO2}} + 0.2535\bar{h}_{\text{H2O}} + 0.4170\bar{h}_{\text{O2}} + 4.71\bar{h}_{\text{N2}} = (0.7215)(99{,}606) + (0.2535)(81{,}573)$$
$$+ (0.4170)(67{,}127) + (4.71)(64{,}090)$$
$$= 422{,}400 \text{ kJ} \quad (\text{Higher than } 416{,}706 \text{ kJ})$$

By extrapolation, T_P = **1956 K = 1683°C**

15-85 A certain coal is burned with 50 percent excess air adiabatically during a steady-flow combustion process. The combustion gases are expanded in an isentropic turbine. The work produced by this turbine is to be determined.

Assumptions **1** Steady operating conditions exist. **2** Combustion gases are ideal gases. **3** Kinetic and potential energies are negligible. 4 The effect of sulfur in the energy and entropy balances is negligible.

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

Analysis The balanced combustion equation from the previous problem is

$$\begin{aligned}&0.7215\text{C}+0.2535\text{H}_2+0.01617\text{O}_2+0.00711\text{N}_2+0.00177\text{S}+1.2509(\text{O}_2+3.76\text{N}_2)\\&\qquad\longrightarrow 0.7215\text{CO}_2+0.2535\text{H}_2\text{O}+0.0018\text{SO}_2+4.71\text{N}_2+0.4170\text{O}_2\end{aligned}$$

The change of the entropy of this mixture during the expansion is given by

$$\Delta\bar{s}=\sum y\left(\bar{s}_2^{\circ}-\bar{s}_1^{\circ}-R\ln\frac{P_2}{P_1}\right)$$

Since there is no entropy change during an isentropic process, this equation becomes

$$\sum y\left(\bar{s}_2^{\circ}-\bar{s}_1^{\circ}\right)=R_u\ \ln\frac{P_2}{P_1}\sum y=(8.314\text{ kJ/kmol}\cdot\text{K})\ln\frac{140\text{ kPa}}{1380\text{ kPa}}(6.1038\text{ kmol})=-116.1\text{ kJ/K}$$

Here, the total moles of the products are 6.1038 kmol. The exit temperature of the combustion gases may be determined by a trial-error method. The inlet temperature is 1956 K (≅ 1960 K) from the previous problem. Guessing the exit temperature to be 1000 K gives (Tables A-18 through A-23)

$$\begin{aligned}\sum y\left(\bar{s}_2^{\circ}-\bar{s}_1^{\circ}\right)&=(0.7215)(269.215-307.992)+(0.2535)(232.597-263.542)\\&\quad+(4.71)(228.057-251.242)+(0.417)(243.471-267.891)\\&=-155.2\text{ kJ/K}\end{aligned}$$

For the exit temperature of 1200 K,

$$\begin{aligned}\sum y\left(\bar{s}_2^{\circ}-\bar{s}_1^{\circ}\right)&=(0.7215)(297.307-307.992)+(0.2535)(240.333-263.542)\\&\quad+(4.71)(234.115-251.242)+(0.417)(249.906-267.891)\\&=-101.8\text{ kJ/K}\end{aligned}$$

By interpolation, T_2 = 1146 K

The work produced during the isentropic expansion of combustion gases from 1956 K to 1146 K is (Tables A-18 through A-23)

$$\begin{aligned}\bar{w}_{\text{out}}&=\sum y(\bar{h}_1-\bar{h}_2)=(0.7215)(98{,}160-50{,}819)+(0.2535)(80{,}352-42{,}037)\\&\quad+(4.71)(63{,}236-34{,}961)+(0.417)(66{,}223-36{,}527)\\&=189{,}428\text{ kJ/kmol fuel}\end{aligned}$$

The molar mass of the product gases is

$$M_m=\frac{m_m}{N_m}=\frac{(0.7215\times 44+0.2535\times 18+0.0018\times 64+4.71\times 28+0.4170\times 32)\text{ kg}}{(0.7215+0.2535+0.0018+4.71+0.417)\text{ kmol}}=29.76\text{ kg/kmol}$$

The work produced per kg of the fuel is then

$$w_{\text{out}}=\frac{\bar{w}_{\text{out}}}{M_m}=\frac{189{,}428\text{ kJ/kmol fuel}}{29.76\text{ kg/kmol}}=\mathbf{6365\ kJ/kg\ fuel}$$

15-86 Ethyl alcohol is burned with 200 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

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

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

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

Substituting,

$$C_2H_5OH + 9[O_2 + 3.76N_2] \longrightarrow 2CO_2 + 3H_2O + 33.84\,N_2 + 6O_2$$

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

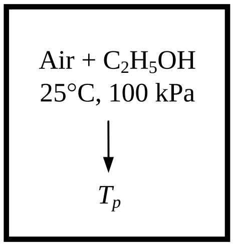
$$\sum N_P\left(\bar{h}_f^\circ + \bar{h} - \bar{h}^\circ - P\bar{v}\right)_P = \sum N_R\left(\bar{h}_f^\circ + \bar{h} - \bar{h}^\circ - P\bar{v}\right)_R$$

Since both the reactants and products are assumed to be 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

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

From the tables,

Substance	$\bar{h}_f^\circ$ kJ/kmol	$\bar{h}_{298K}$ kJ/kmol
C_2H_5 OH (*g*)	-235,310	---
O_2	0	8682
N_2	0	8669
H_2O (*g*)	-241,820	9904
CO_2	-393,520	9364



Thus,

$$\begin{aligned}&(2)(-393{,}520 + \bar{h}_{CO2} - 9364 - 8.314 \times T_p) + (3)(-241{,}820 + \bar{h}_{H2O} - 9904 - 8.314 \times T_p)\\ &+ (6)(0 + \bar{h}_{O2} - 8682 - 8.314 \times T_p) + (33.84)(0 + \bar{h}_{N2} - 8669 - 8.314 \times T_p)\\ &= (1)(-235{,}310 - 8.314 \times 298) + (9)(0 - 8.314 \times 298) + (33.84)(-8.314 \times 298)\end{aligned}$$

It yields

$$2\bar{h}_{CO2} + 3\bar{h}_{H2O} + 6\bar{h}_{O2} + 33.84\bar{h}_{N2} - 372.8T_p = -343{,}927 + 1{,}906{,}391 = 1{,}562{,}464 \text{ kJ}$$

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,

$$44.84\bar{h}_{N2} - 372.8T_p = 1{,}562{,}464 \text{ kJ}$$

An investigation of Table A-18 shows that this equation is satisfied at a temperature close to 1600 K but it will be somewhat under it because of the higher specific heat of H_2O.

At 1500 K:

$$(2)(71{,}078)+(3)(57{,}999)+(6)(49{,}292)+(33.84)(47{,}073)-(372.8)(1500)=1{,}645{,}655$$

$$\text{(Higher than } 1{,}562{,}464\text{ kJ)}$$

At 1400 K:

$$(2)(65{,}271)+(3)(53{,}351)+(6)(45{,}648)+(33.84)(43{,}605)-(372.8)(1400)=1{,}518{,}156$$

$$\text{(Lower than } 1{,}562{,}464\text{ kJ)}$$

By interpolation,

$$T_P = \mathbf{1435\ K}$$

The volume of reactants when 1 kmol of fuel is burned is

$$\mathcal{V} = \mathcal{V}_{\text{fuel}} + \mathcal{V}_{\text{air}} = (N_{\text{fuel}} + N_{\text{air}})\frac{R_u T}{P} = (1+42.84)\text{ kmol})\frac{(8.314\text{ kJ/kmol}\cdot\text{K})(298\text{ K})}{100\text{ kPa}} = 1086\text{ m}^3$$

The final pressure is then

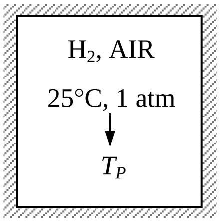
$$P = N_{\text{prod}}\frac{R_u T}{\mathcal{V}} = (44.84\text{ kmol})\frac{(8.314\text{ kJ/kmol}\cdot\text{K})(1435\text{ K})}{1086\text{ m}^3} = \mathbf{493\ kPa}$$

15-87 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 $E_{\text{in}} - E_{\text{out}} = \Delta E_{\text{system}}$ for reacting closed systems under adiabatic conditions (Q = 0) with no work interactions (W = 0),

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

Since both the reactants and the products behave as 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

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

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: $\bar{h}_{H_2O} + 1.88\bar{h}_{N_2} - 23.94\times T_P = (1)(139{,}051) + (1.88)(103{,}260) - (23.94)(3050)$

$$= 260{,}163\text{ kJ (Higher than } 259{,}648\text{ kJ)}$$

At 3000 K: $\bar{h}_{H_2O} + 1.88\bar{h}_{N_2} - 23.94\times T_P = (1)(136{,}264) + (1.88)(101{,}407) - (23.94)(3000)$

$$= 255{,}089\text{ kJ (Lower than } 259{,}648\text{ kJ)}$$

By interpolation, T_P = **3045 K**

Entropy Change and Second Law Analysis of Reacting Systems

15-88C Assuming the system exchanges heat with the surroundings at T_0, the increase-in-entropy principle can be expressed as

$$S_{\text{gen}} = \sum N_P \bar{s}_P - \sum N_R \bar{s}_R + \frac{Q_{\text{out}}}{T_0}$$

15-89C 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.

15-90C It represents the reversible work associated with the formation of that compound.

15-91 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

$$H_2 + 0.5O_2 \longrightarrow H_2O.$$

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,

$$\begin{aligned} W_{\text{rev}} &= \sum N_R \bar{g}^{\circ}_{f,R} - \sum N_P \bar{g}^{\circ}_{f,P} = N_{\text{H}_2} \cancelto{0}{\bar{g}^{\circ}_{f,\text{H}_2}} + N_{\text{O}_2} \cancelto{0}{\bar{g}^{\circ}_{f,\text{O}_2}} - N_{\text{H}_2\text{O}} \bar{g}^{\circ}_{f,\text{H}_2\text{O}} = -N_{\text{H}_2\text{O}} \bar{g}^{\circ}_{f,\text{H}_2\text{O}} \\ &= -(1\text{ kmol})(-237{,}180\text{ kJ/kmol}) = \mathbf{237{,}180\text{ kJ}} \quad (\text{per kmol of } \text{H}_2) \end{aligned}$$

since the g°_f of stable elements at 25°C and 1 atm is zero. Therefore, 237,180 kJ of work could be done as 1 kmol of H_2 is burned with 0.5 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,

$$X_{\text{destruction}} = \mathbf{237{,}180\text{ kJ}} \quad (\text{per kmol of H}_2)$$

We could also determine the reversible work without involving the Gibbs function,

$$\begin{aligned} W_{\text{rev}} &= \sum N_R \left(\bar{h}^{\circ}_f + \bar{h} - \bar{h}^{\circ} - T_0 \bar{s}\right)_R - \sum N_P \left(\bar{h}^{\circ}_f + \bar{h} - \bar{h}^{\circ} - T_0 \bar{s}\right)_P \\ &= \sum N_R \left(\bar{h}^{\circ}_f - T_0 \bar{s}\right)_R - \sum N_P \left(\bar{h}^{\circ}_f - T_0 \bar{s}\right)_P \\ &= N_{\text{H}_2} \left(\bar{h}^{\circ}_f - T_0 \bar{s}^{\circ}\right)_{\text{H}_2} + N_{\text{O}_2} \left(\bar{h}^{\circ}_f - T_0 \bar{s}^{\circ}\right)_{\text{O}_2} - N_{\text{H}_2\text{O}} \left(\bar{h}^{\circ}_f - T_0 \bar{s}^{\circ}\right)_{\text{H}_2\text{O}} \end{aligned}$$

Substituting,

$$W_{\text{rev}} = (1)(0 - 298 \times 130.58) + (0.5)(0 - 298 \times 205.03) - (1)(-285{,}830 - 298 \times 69.92) = 237{,}204\text{ kJ}$$

which is almost identical to the result obtained before.

15-92 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,

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

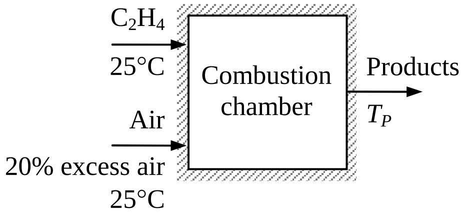
Thus, $$C_2H_4(g)+3.6(O_2+3.76N_2)\longrightarrow 2CO_2+2H_2O+0.6O_2+13.54N_2$$

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



Substituting,

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

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,

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

The entropy calculations can be presented in tabular form as

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

Thus,

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

and

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

15-93 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

$$\mathrm{C_8H_{18}}(\ell) + 1.5a_{\mathrm{th}}(\mathrm{O_2} + 3.76\mathrm{N_2}) \longrightarrow 8\mathrm{CO_2} + 9\mathrm{H_2O} + 0.5a_{\mathrm{th}}\mathrm{O_2} + (1.5)(3.76)a_{\mathrm{th}}\mathrm{N_2}$$

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

$$1.5a_{\mathrm{th}} = 8 + 4.5 + 0.5a_{\mathrm{th}} \quad\longrightarrow\quad a_{\mathrm{th}} = 12.5$$

Thus,

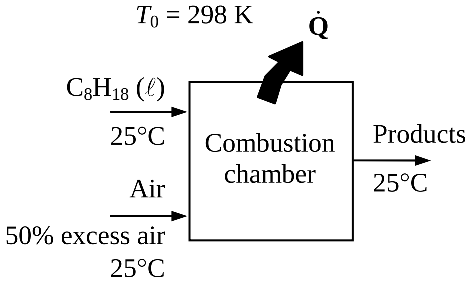
$$\mathrm{C_8H_{18}}(\ell) + 18.75(\mathrm{O_2} + 3.76\mathrm{N_2}) \longrightarrow 8\mathrm{CO_2} + 9\mathrm{H_2O} + 6.25\mathrm{O_2} + 70.5\mathrm{N_2}$$

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

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

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,

$$-Q_{\mathrm{out}} = (8)(-393{,}520) + (9)(-285{,}830) + 0 + 0 - (1)(-249{,}950) - 0 - 0 = -5{,}470{,}680 \text{ kJ/kmol of } \mathrm{C_8H_{18}}$$

or

$$Q_{\mathrm{out}} = 5{,}470{,}680 \text{ kJ/kmol of } \mathrm{C_8H_{18}}$$

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,

$$\dot{Q}_{\mathrm{out}} = \dot{N}Q_{\mathrm{out}} = (2.193\times10^{-3}\text{ kmol/min})(5{,}470{,}680\text{ kJ/kmol}) = \mathbf{11{,}997\ kJ/min}$$

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 $P_i = y_i P_{\text{total}}$, 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

	$\mathbf{N_i}$	$\mathbf{y_i}$	$\mathbf{\bar{s}_i^\circ(T,1atm)}$	$\mathbf{R_u \ln(y_i P_m)}$	$\mathbf{N_i \bar{s}_i}$
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,

$$S_{\text{gen}} = S_P - S_R + \frac{Q_{\text{surr}}}{T_{\text{surr}}} = 17{,}531 - 18{,}097 + \frac{5{,}470{,}523 \text{ kJ}}{298 \text{ K}} = 17{,}798 \text{ kJ/kmol}\cdot\text{K}$$

and

$$\dot{S}_{\text{gen}} = \dot{N} S_{\text{gen}} = \left(2.193 \times 10^{-3} \text{ kmol/min}\right)\left(17{,}798 \text{ kJ/kmol}\cdot\text{K}\right) = \mathbf{39.03 \text{ kJ/min}\cdot\text{K}}$$

(*c*) The exergy destruction rate associated with this process is determined from

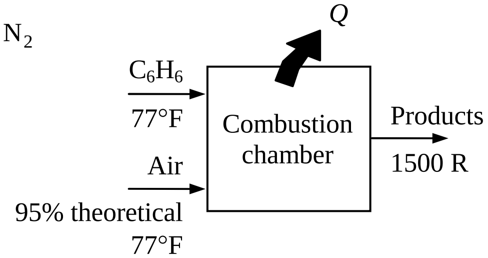
$$\dot{X}_{\text{destroyed}} = T_0 \dot{S}_{\text{gen}} = (298 \text{ K})(39.03 \text{ kJ/min}\cdot\text{K}) = 11{,}632 \text{ kJ/min} = \mathbf{193.9 \text{ kW}}$$

15-94E Benzene gas is burned steadily with 95 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

$$C_6H_6 + (0.95)(7.5)(O_2 + 3.76N_2) \longrightarrow xCO_2 + (6-x)CO + 3H_2O + 26.79N_2$$

The value of *x* is determined from an O_2 balance,

$$(0.95)(7.5) = x + (6-x)/2 + 1.5 \longrightarrow x = 5.25$$

Thus,

$$C_6H_6 + 7.125(O_2 + 3.76N_2) \longrightarrow 5.25CO_2 + 0.75CO + 3H_2O + 26.79N_2$$

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	$\bar{\mathbf{h}}_\mathbf{f}^\circ$ **Btu/lbmol**	$\bar{\mathbf{h}}_{\mathbf{537\ R}}$ **Btu/lbmol**	$\bar{\mathbf{h}}_{\mathbf{1500\ R}}$ **Btu/lbmol**
C_6H_6 (*g*)	35,680	---	---
O_2	0	3725.1	11,017.1
N_2	0	3729.5	10,648.0
H_2O (*g*)	-104,040	4258.0	12,551.4
CO	-47,540	3725.1	10,711.1
CO_2	-169,300	4027.5	14,576.0

Thus,

$$\begin{aligned}-Q_{out} &= (5.25)(-169{,}300 + 14{,}576 - 4027.5) + (0.75)(-47{,}540 + 10{,}711.1 - 3725.1)\\ &\quad + (3)(-104{,}040 + 12{,}551.4 - 4258) + (26.79)(0 + 10{,}648 - 3729.5) - (1)(35{,}680) - 0 - 0\\ &= \mathbf{-1{,}001{,}434\ Btu/lbmol\ of\ C_6H_6}\end{aligned}$$

(*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 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

	$\mathbf{N_i}$	$\mathbf{y_i}$	$\mathbf{\bar{s}_i^\circ(T,1atm)}$	$\mathbf{R_u \ln(y_i P_m)}$	$\mathbf{N_i \bar{s}_i}$
C_6H_6	1	1.00	64.34	---	64.34
O_2	7.125	0.21	49.00	-3.10	371.21
N_2	26.79	0.79	45.77	-0.47	1238.77
					S_R = 1674.32 Btu/R
CO_2	5.25	0.1467	61.974	-3.812	345.38
CO	0.75	0.0210	54.665	-7.672	46.75
H_2O (*g*)	3	0.0838	53.808	-4.924	176.20
N_2	26.79	0.7485	53.071	-0.575	1437.18
					S_P = 2005.51 Btu/R

Thus,

$$S_{\text{gen}} = S_P - S_R + \frac{Q_{\text{out}}}{T_{\text{surr}}} = 2005.51 - 1674.32 + \frac{+1{,}001{,}434}{537} = 2196.1 \text{ Btu/R}$$

Then the exergy destroyed is determined from

$$X_{\text{destroyed}} = T_0 S_{\text{gen}} = (537\text{ R})(2196.1\text{ Btu/lbmol}\cdot\text{R}) = \mathbf{1{,}179{,}306\ Btu/R}\ \ (\text{per lbmol } C_6H_6)$$

15-95 [*Also solved by EES on enclosed CD***]** 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

$$C_3H_8(\ell)+2.5a_{th}(O_2+3.76N_2)\longrightarrow 3CO_2+4H_2O+1.5a_{th}O_2+(2.5)(3.76)a_{th}N_2$$

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

$$2.5a_{th}=3+2+1.5a_{th}\quad\longrightarrow\quad a_{th}=5$$

Substituting,

$$C_3H_8(\ell)+12.5(O_2+3.76N_2)\longrightarrow 3CO_2+4H_2O+7.5O_2+47N_2$$

The air-fuel ratio for this combustion process is

$$\mathrm{AF}=\frac{m_{\text{air}}}{m_{\text{fuel}}}=\frac{(12.5\times 4.76\text{ kmol})(29\text{ kg/kmol})}{(3\text{ kmol})(12\text{ kg/kmol})+(4\text{ kmol})(2\text{ kg/kmol})}=39.2\text{ kg air/kg fuel}$$

Thus,

$$\dot{m}_{\text{air}}=(\mathrm{AF})(\dot{m}_{\text{fuel}})=(39.2\text{ kg air/kg fuel})(0.4\text{ kg fuel/min})=\mathbf{15.7\ kg\ air/min}$$

(*b*) 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

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

Assuming the air and the combustion products to be ideal gases, we have $h = h(T)$. From the tables, (The $\overline{h}_f^\circ$ of liquid propane is obtained by adding the h_{fg} at 25°C to $\overline{h}_f^\circ$ of gaseous propane).

Substance	$\overline{\mathbf{h}}_\mathbf{f}^\circ$ **kJ/kmol**	$\overline{\mathbf{h}}_{\mathbf{285\ K}}$ **kJ/kmol**	$\overline{\mathbf{h}}_{\mathbf{298\ K}}$ **kJ/kmol**	$\overline{\mathbf{h}}_{\mathbf{1200\ K}}$ **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,

$$\begin{aligned}-Q_{\text{out}}&=(3)(-393{,}520+53{,}848-9364)+(4)(-241{,}820+44{,}380-9904)\\&\quad+(7.5)(0+38{,}447-8682)+(47)(0+36{,}777-8669)-(1)(-118{,}910+h_{298}-h_{298})\\&\quad-(12.5)(0+8296.5-8682)-(47)(0+8286.5-8669)\\&=-190{,}464\text{ kJ/kmol of }C_3H_8\end{aligned}$$

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 $\bar{s}_{C_3H_8} = 269.91\,\text{kJ/kmol·K}$ (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

	$\mathbf{N_i}$	$\mathbf{y_i}$	$\mathbf{\bar{s}_i^\circ(T,1atm)}$	$\mathbf{R_u \ln(y_i P_m)}$	$\mathbf{N_i \bar{s}_i}$
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

$$\dot{S}_{\text{gen}} = (\dot{N})(S_{\text{gen}}) = \left(\frac{0.4}{44}\ \text{kmol/min}\right)(3760.1\ \text{kJ/kmol}\cdot\text{K}) = \mathbf{34.2\ kJ/min\cdot K}$$

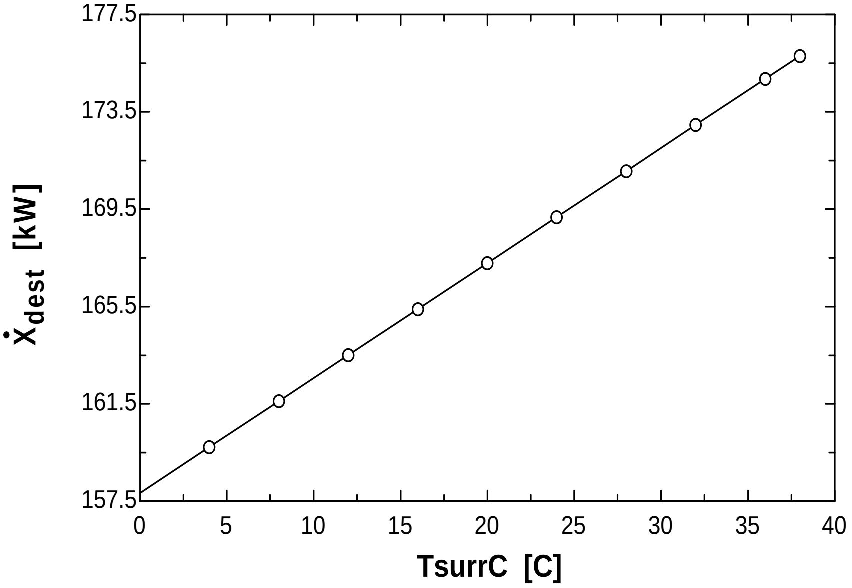
15-96 EES Problem 15-95 is reconsidered. The effect of the surroundings temperature on the rate of exergy destruction is to be studied.

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



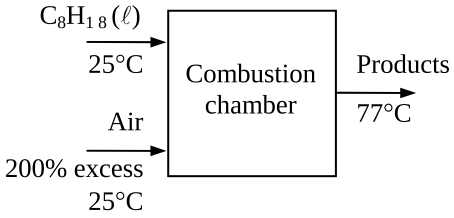
15-97 Liquid octane is burned steadily with 200 percent excess air in a automobile engine. The maximum amount of work that can be produced by this engine is to be determined.

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 C_8H_{18}, the combustion equation can be written as

$$\mathrm{C_8H_{18}}(\ell) + 3a_{\mathrm{th}}(\mathrm{O_2} + 3.76\mathrm{N_2}) \longrightarrow 8\mathrm{CO_2} + 9\mathrm{H_2O} + 2a_{\mathrm{th}}\mathrm{O_2} + (3)(3.76)a_{\mathrm{th}}\mathrm{N_2}$$

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



$$3a_{\mathrm{th}} = 8 + 4.5 + 2a_{\mathrm{th}} \longrightarrow a_{\mathrm{th}} = 12.5$$

Thus,

$$\mathrm{C_8H_{18}}(\ell) + 37.5(\mathrm{O_2} + 3.76\mathrm{N_2}) \longrightarrow 8\mathrm{CO_2} + 9\mathrm{H_2O} + 25\mathrm{O_2} + 141\mathrm{N_2}$$

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
C_8H_{18} (ℓ)	-249,950	---	---
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,

$$\begin{aligned} -Q_{\mathrm{out}} &= (8)(-393{,}520 + 11{,}351 - 9364) + (9)(-241{,}820 + 11{,}652 - 9904) \\ &\quad + (25)(0 + 10{,}213 - 8682) + (141)(0 + 10{,}180 - 8669) - (1)(-249{,}950) \\ &= -4{,}791{,}636 \text{ kJ/kmol of } \mathrm{C_8H_{18}} \end{aligned}$$

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 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

	$\mathbf{N_i}$	$\mathbf{y_i}$	$\mathbf{\bar{s}_i^\circ(T,1atm)}$	$\mathbf{R_u \ln(y_i P_m)}$	$\mathbf{N_i \bar{s}_i}$
C_8H_{18}	1	---	466.73	---	466.73
O_2	37.5	0.21	205.04	-12.98	8175.75
N_2	141	0.79	191.61	-1.960	27,293.37
					S_R = 35,936 kJ/K
CO_2	8	0.0437	219.831	-26.03	1966.89
H_2O (*g*)	9	0.0492	194.125	-25.04	1972.49
O_2	25	0.1366	209.765	-16.55	5657.88
N_2	141	0.7705	196.173	-2.17	27,966.36
					S_P = 37,565 kJ/K

Thus,

$$S_{\text{gen}} = S_P - S_R + \frac{Q_{\text{out}}}{T_{\text{surr}}} = 37{,}565 - 35{,}936 + \frac{4{,}791{,}636}{298} = 17{,}708 \text{ kJ/K } (\text{per kmol } C_8H_{18})$$

The maximum work is equal to the exergy destruction

$$W_{\max} = X_{\text{dest}} = T_0 S_{\text{gen}} = (298)(17{,}708 \text{ kJ/K}) = 5{,}276{,}984 \text{ kJ/K } (\text{per kmol } C_8H_{18})$$

Per unit mass basis,

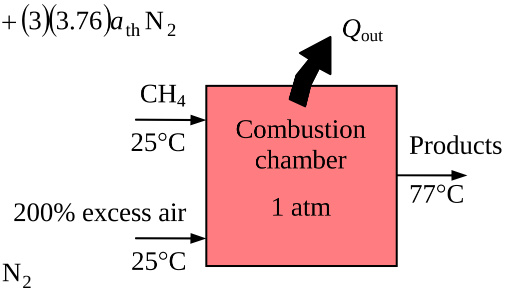
$$W_{\max} = \frac{5{,}276{,}984 \text{ kJ/K}\cdot\text{kmol}}{114 \text{ kg/kmol}} = \mathbf{46{,}290 \text{ kJ/kg fuel}}$$

15-98 Methane is burned steadily with 200 percent excess air in a 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_4, the combustion equation can be written as

$$CH_4 + 3a_{th}(O_2 + 3.76N_2) \longrightarrow CO_2 + 2H_2O + 2a_{th}O_2 + (3)(3.76)a_{th}N_2$$



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

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

Thus,

$$CH_4 + 6(O_2 + 3.76N_2) \longrightarrow CO_2 + 2H_2O + 4O_2 + 22.56N_2$$

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	$\bar{\mathbf{h}}_\mathbf{f}^\circ$ **kJ/kmol**	$\bar{\mathbf{h}}_{\mathbf{298K}}$ **kJ/kmol**	$\bar{\mathbf{h}}_{350\mathbf{K}}$ **kJ/kmol**
CH_4	-74,850	---	---
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,

$$\begin{aligned} -Q_{out} &= (1)(-393{,}520 + 11{,}351 - 9364) + (2)(-241{,}820 + 11{,}652 - 9904) \\ &\quad + (4)(0 + 10{,}213 - 8682) + (22.56)(0 + 10{,}180 - 8669) - (1)(-74{,}850) \\ &= -756{,}615 \text{ kJ/kmol of } CH_4 \end{aligned}$$

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

	$\mathbf{N_i}$	$\mathbf{y_i}$	$\mathbf{\bar{s}_i^\circ(T,1atm)}$	$\mathbf{R_u \ln(y_i P_m)}$	$\mathbf{N_i \bar{s}_i}$
CH_4	1	---	186.16	---	186.16
O_2	6	0.21	205.04	-12.98	1308.12
N_2	22.56	0.79	191.61	-1.960	4366.94
					S_R = 5861.22 kJ/K
CO_2	1	0.0338	219.831	-28.16	247.99
H_2O (*g*)	2	0.0677	194.125	-22.39	433.03
O_2	4	0.1353	209.765	-16.63	905.58
N_2	22.56	0.7632	196.173	-2.25	4476.42
					S_P = 6063.02 kJ/K

Thus,

$$S_{\text{gen}} = S_P - S_R + \frac{Q_{\text{out}}}{T_{\text{surr}}} = 6063.02 - 5861.22 + \frac{756{,}615}{298} = 2741 \text{ kJ/K } (\text{per kmol CH}_4)$$

The maximum work is equal to the exergy destruction

$$W_{\max} = X_{\text{dest}} = T_0 S_{\text{gen}} = (298)(2741 \text{ kJ/K}) = 816{,}818 \text{ kJ/K } (\text{per kmol CH}_4)$$

Per unit mass basis,

$$W_{\max} = \frac{816{,}818 \text{ kJ/K}\cdot\text{kmol}}{16 \text{ kg/kmol}} = \mathbf{51{,}050 \text{ kJ/kg fuel}}$$

15-99 Liquid octane is burned steadily with 30 percent excess air. The entropy generation and exergy destruction per unit mass of the fuel are to be determined.

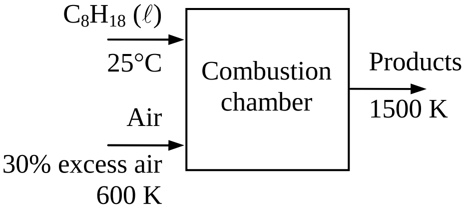
Assumptions **1** Combustion is complete. **2** Steady operating conditions exist. **3** Air and the combustion gases are ideal gases. **4** Changes in kinetic and potential energies are negligible.

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

$$C_8H_{18}(\ell)+1.3a_{\text{th}}(O_2+3.76N_2)\longrightarrow 8CO_2+9H_2O+0.3a_{\text{th}}O_2+(1.3)(3.76)a_{\text{th}}N_2$$

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



$$1.3a_{\text{th}}=8+4.5+0.3a_{\text{th}}\longrightarrow a_{\text{th}}=12.5$$

Thus,

$$C_8H_{18}(\ell)+16.25(O_2+3.76N_2)\longrightarrow 8CO_2+9H_2O+3.75O_2+61.1N_2$$

(*b*) 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	$\bar{h}_f^\circ$ kJ/kmol	$\bar{h}_{298K}$ kJ/kmol	$\bar{h}_{600K}$ kJ/kmol	$\bar{h}_{1500K}$ 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,

$$\begin{aligned}-Q_{\text{out}}&=(8)(-393{,}520+71{,}078-9364)+(9)(-241{,}820+57{,}999-9904)\\&\quad+(3.75)(0+49{,}292-8682)+(61.1)(0+47{,}073-8669)-(1)(-249{,}950)\\&\quad-(16.25)(0+17{,}929-8682)-(61.1)(0+17{,}563-8669)\\&=-2{,}342{,}938\text{ kJ/kmol of }C_8H_{18}\end{aligned}$$

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 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 $P_i = y_i P_{total}$, 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

	$\mathbf{N_i}$	$\mathbf{y_i}$	$\mathbf{\bar{s}_i^\circ(T,1atm)}$	$\mathbf{R_u \ln(y_i P_m)}$	$\mathbf{N_i \bar{s}_i}$
C_8H_{18}	1	---	466.73	14.79	451.94
O_2	16.25	0.21	226.35	1.81	3648.78
N_2	61.1	0.79	212.07	12.83	12,173.56
					S_R = 16,274.28 kJ/K
CO_2	8	0.0977	292.11	-4.552	2373.30
H_2O (*g*)	9	0.1100	250.45	-3.566	2286.14
O_2	3.75	0.0458	257.97	-10.85	1008.08
N_2	61.1	0.7465	241.77	12.35	14,017.56
					S_P = 19,685.08 kJ/K

Thus,

$$S_{gen} = S_P - S_R + \frac{Q_{out}}{T_{surr}} = 19{,}685.08 - 16{,}274.28 + \frac{2{,}342{,}938}{298} = 11{,}273 \text{ kJ/K } (\text{per kmol } C_8H_{18})$$

The exergy detstuction is

$$X_{dest} = T_0 S_{gen} = (298)(11{,}273 \text{ kJ/K}) = 3{,}359{,}354 \text{ kJ/K } (\text{per kmol } C_8H_{18})$$

The entropy generation and exergy destruction per unit mass of the fuel are

$$S_{gen} = \frac{\bar{S}_{gen}}{M_{fuel}} = \frac{11{,}273 \text{ kJ/K}\cdot\text{kmol}}{114 \text{ kg/kmol}} = \mathbf{98.89\ kJ/K\cdot kg\ C_3H_8}$$

$$X_{dest} = \frac{\bar{X}_{dest}}{M_{fuel}} = \frac{3{,}359{,}954 \text{ kJ/K}\cdot\text{kmol}}{114 \text{ kg/kmol}} = \mathbf{29{,}470\ kJ/kg\ C_3H_8}$$

Review Problems

15-100 A sample of a certain fluid is burned in a bomb calorimeter. The heating value of the fuel is to be determined.

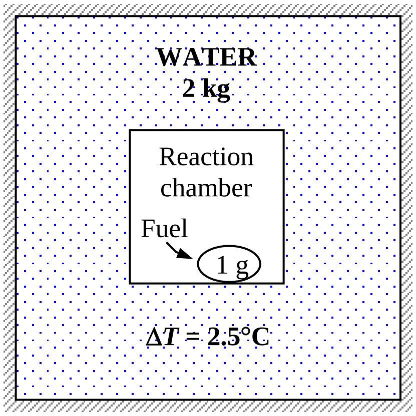
Properties The specific heat of water is 4.18 kJ/kg.°C (Table A-3).

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.

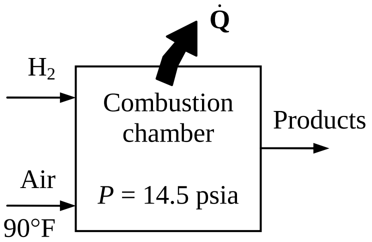
15-101E Hydrogen is burned with 100 percent excess air. The AF ratio and the volume flow rate of air are to be determined.

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

$$\mathrm{H_2} + 2a_{\mathrm{th}}(\mathrm{O_2} + 3.76\mathrm{N_2}) \longrightarrow \mathrm{H_2O} + a_{\mathrm{th}}\mathrm{O_2} + (2)(3.76)a_{\mathrm{th}}\mathrm{N_2}$$



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

O_2 balance: $\quad 2a_{\mathrm{th}} = 0.5 + a_{\mathrm{th}} \longrightarrow a_{\mathrm{th}} = 0.5$

Substituting, $\quad \mathrm{H_2} + (\mathrm{O_2} + 3.76\mathrm{N_2}) \longrightarrow \mathrm{H_2O} + 0.5\mathrm{O_2} + 3.76\mathrm{N_2}$

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

$$P_{v,\mathrm{in}} = \phi_{\mathrm{air}} P_{\mathrm{sat@90°F}} = (0.60)(0.69904\ \mathrm{psi}) = 0.419\ \mathrm{psia}$$

The number of moles of the moisture that accompanies 4.76 lbmol of incoming dry air ($N_{v,\ \mathrm{in}}$) is determined to be

$$N_{v,\mathrm{in}} = \left(\frac{P_{v,\mathrm{in}}}{P_{\mathrm{total}}}\right) N_{\mathrm{total}} = \left(\frac{0.419\ \mathrm{psia}}{14.5\ \mathrm{psia}}\right)(4.76 + N_{v,\mathrm{in}}) \longrightarrow N_{v,\mathrm{in}} = 0.142\ \mathrm{lbmol}$$

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,

$$\mathrm{H_2} + (\mathrm{O_2} + 3.76\mathrm{N_2}) + 0.142\mathrm{H_2O} \longrightarrow 1.142\mathrm{H_2O} + 0.5\mathrm{O_2} + 3.76\mathrm{N_2}$$

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

$$\mathrm{AF} = \frac{m_{\mathrm{air}}}{m_{\mathrm{fuel}}} = \frac{(4.76\ \mathrm{lbmol})(29\ \mathrm{lbm/lbmol}) + (0.142\ \mathrm{lbmol})(18\ \mathrm{lbm/lbmol})}{(1\ \mathrm{lbmol})(2\ \mathrm{lbm/lbmol})} = \mathbf{70.3\ lbm\ air/lbmfuel}$$

(*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

$$\dot{m}_{\mathrm{air}} = (\mathrm{AF})(\dot{m}_{\mathrm{fuel}}) = (70.3)(25\ \mathrm{lbm/h}) = 1758\ \mathrm{lbm/h}$$

The mole fractions of water vapor and the dry air in the incoming air are

$$y_{\mathrm{H_2O}} = \frac{N_{\mathrm{H_2O}}}{N_{\mathrm{total}}} = \frac{0.142}{4.76 + 0.142} = 0.029 \ \text{ and } \ y_{\mathrm{dryair}} = 1 - 0.029 = 0.971$$

Thus,

$$M = (yM)_{\mathrm{H_2O}} + (yM)_{\mathrm{dryair}} = (0.029)(18) + (0.971)(29) = 28.7\ \mathrm{lbm/lbmol}$$

$$\boldsymbol{v} = \frac{RT}{P} = \frac{(10.73/28.7\ \mathrm{psia \cdot ft^3/lbm \cdot R})(550\ \mathrm{R})}{14.5\ \mathrm{psia}} = 14.18\ \mathrm{ft^3/lbm}$$

$$\dot{\boldsymbol{V}} = \dot{m}\boldsymbol{v} = (1758\ \mathrm{lbm/h})(14.18\ \mathrm{ft^3/lbm}) = \mathbf{24{,}928\ ft^3/h}$$

15-102 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.

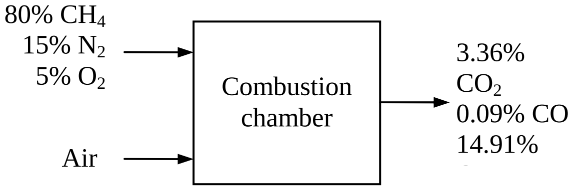
Assumptions **1** Combustion is complete. **2** Air and the combustion gases are ideal gases.

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

$$x(0.80\mathrm{CH_4}+0.15\mathrm{N_2}+0.05\mathrm{O_2})+a(\mathrm{O_2}+3.76\mathrm{N_2})$$
$$\longrightarrow 3.36\mathrm{CO_2}+0.09\mathrm{CO}+14.91\mathrm{O_2}+81.64\mathrm{N_2}+b\mathrm{H_2O}$$

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



$$\mathrm{C}:\ 0.80x=3.36+0.09\ \longrightarrow x=4.31$$
$$\mathrm{H}:\ 3.2x=2b\ \longrightarrow b=6.90$$
$$\mathrm{N_2}:\ 0.15x+3.76a=81.64\ \longrightarrow a=21.54$$

$$\left[\text{Check }\ \mathrm{O_2}:\ 0.05x+a=3.36+0.045+14.91+b/2\longrightarrow a=21.54\right]$$

Thus,

$$4.31(0.80\mathrm{CH_4}+0.15\mathrm{N_2}+0.05\mathrm{O_2})+21.54(\mathrm{O_2}+3.76\mathrm{N_2})$$
$$\longrightarrow 3.36\mathrm{CO_2}+0.09\mathrm{CO}+14.91\mathrm{O_2}+81.64\mathrm{N_2}+6.9\mathrm{H_2O}$$

The combustion equation for 1 kmol of fuel is obtained by dividing the above equation by 4.31,

$$(0.80\mathrm{CH_4}+0.15\mathrm{N_2}+0.05\mathrm{O_2})+5.0(\mathrm{O_2}+3.76\mathrm{N_2})$$
$$\longrightarrow 0.78\mathrm{CO_2}+0.02\mathrm{CO}+3.46\mathrm{O_2}+18.95\mathrm{N_2}+1.6\mathrm{H_2O}$$

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

$$\mathrm{AF}=\frac{m_{\mathrm{air}}}{m_{\mathrm{fuel}}}=\frac{(5.0\times4.76\ \mathrm{kmol})(29\ \mathrm{kg/kmol})}{0.8\times16+0.15\times28+0.05\times32}=\mathbf{37.1\ \ kg\ air/kg\ fuel}$$

(*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,

$$(0.80\mathrm{CH_4}+0.15\mathrm{N_2}+0.05\mathrm{O_2})+a_{\mathrm{th}}(\mathrm{O_2}+3.76\mathrm{N_2})\longrightarrow 0.8\mathrm{CO_2}+1.6\mathrm{H_2O}+(0.15+3.76a_{\mathrm{th}})\mathrm{N_2}$$

$$\mathrm{O_2}:\ 0.05+a_{\mathrm{th}}=0.8+0.8\ \longrightarrow\ a_{\mathrm{th}}=1.55$$

Then, $\text{Percent theoretical air}=\dfrac{m_{\mathrm{air,act}}}{m_{\mathrm{air,th}}}=\dfrac{N_{\mathrm{air,act}}}{N_{\mathrm{air,th}}}=\dfrac{(5.0)(4.76)\ \mathrm{kmol}}{(1.55)(4.76)\ \mathrm{kmol}}=\mathbf{323\%}$

(*c*) The specific volume, mass flow rate, and the volume flow rate of air at the inlet conditions are

$$v=\frac{RT}{P}=\frac{(0.287\ \mathrm{kPa\cdot m^3/kg\cdot K})(298\ \mathrm{K})}{100\ \mathrm{kPa}}=0.855\ \mathrm{m^3/kg}$$
$$\dot{m}_{\mathrm{air}}=(\mathrm{AF})\dot{m}_{\mathrm{fuel}}=(37.1\ \mathrm{kg\ air/kg\ fuel})(1.4\ \mathrm{kg\ fuel/min})=51.94\ \mathrm{m^3/min}$$
$$\dot{V}_{\mathrm{air}}=(\dot{m}v)_{\mathrm{air}}=(51.94\ \mathrm{kg/min})(0.855\ \mathrm{m^3/kg})=\mathbf{44.4\ m^3/min}$$

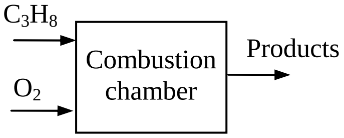
15-103 Propane is burned with stoichiometric amount of oxygen. The mass fractions of each of the products and the mass of water in the products per unit mass of fuel burned are to be determined.

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, 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 is

$$\mathrm{C_3H_8 + 5O_2 \longrightarrow 3CO_2 + 4H_2O}$$



The mass of each product and the total mass are

$$m_{\mathrm{CO2}} = N_{\mathrm{CO2}} M_{\mathrm{CO2}} = (3\,\mathrm{kmol})(44\,\mathrm{kg/kmol}) = 132\,\mathrm{kg}$$
$$m_{\mathrm{H2O}} = N_{\mathrm{H2O}} M_{\mathrm{H2O}} = (4\,\mathrm{kmol})(18\,\mathrm{kg/kmol}) = 72\,\mathrm{kg}$$
$$m_{\mathrm{total}} = m_{\mathrm{CO2}} + m_{\mathrm{H2O}} = 132 + 72 = 204\,\mathrm{kg}$$

Then the mass fractions are

$$\mathrm{mf}_{\mathrm{CO2}} = \frac{m_{\mathrm{CO2}}}{m_{\mathrm{total}}} = \frac{132\,\mathrm{kg}}{204\,\mathrm{kg}} = \mathbf{0.6471}$$
$$\mathrm{mf}_{\mathrm{H2O}} = \frac{m_{\mathrm{H2O}}}{m_{\mathrm{total}}} = \frac{72\,\mathrm{kg}}{204\,\mathrm{kg}} = \mathbf{0.3529}$$

The mass of water in the products per unit mass of fuel burned is determined from

$$\frac{m_{\mathrm{H2O}}}{m_{\mathrm{C3H8}}} = \frac{72\,\mathrm{kg}}{44\,\mathrm{kg}} = \mathbf{1.64\,kg\,H_2O/kg\,C_3H_8}$$

15-104 A coal from Utah is burned with 20% excess air. The mass of water in the products per unit mass of coal burned 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, SO_2, O_2, and N_2. **3** Combustion gases are ideal gases.

Properties The molar masses of C, O_2, H_2, S, and N_2 are 12, 32, 2, 32, and 28 lbm/lbmol, respectively (Table A-1E).

Analysis The mass fractions of the constituent of the coal when the ash is substituted are

$$\mathrm{mf}_{\mathrm{C}} = \frac{m_{\mathrm{C}}}{m_{\mathrm{total}}} = \frac{61.40\ \mathrm{lbm}}{(100-5.00)\ \mathrm{lbm}} = \frac{61.40\ \mathrm{lbm}}{95.00\ \mathrm{lbm}} = 0.6463$$

$$\mathrm{mf}_{\mathrm{H2}} = \frac{m_{\mathrm{H2}}}{m_{\mathrm{total}}} = \frac{5.79\ \mathrm{lbm}}{95.00\ \mathrm{lbm}} = 0.06095$$

$$\mathrm{mf}_{\mathrm{O2}} = \frac{m_{\mathrm{O2}}}{m_{\mathrm{total}}} = \frac{25.31\ \mathrm{lbm}}{95.00\ \mathrm{lbm}} = 0.2664$$

$$\mathrm{mf}_{\mathrm{N2}} = \frac{m_{\mathrm{N2}}}{m_{\mathrm{total}}} = \frac{1.09\ \mathrm{lbm}}{95.00\ \mathrm{lbm}} = 0.01147$$

$$\mathrm{mf}_{\mathrm{S}} = \frac{m_{\mathrm{S}}}{m_{\mathrm{total}}} = \frac{1.41\ \mathrm{lbm}}{95.00\ \mathrm{lbm}} = 0.01484$$

61.40% C
5.79% H_2
25.31% O_2
1.09% N_2
1.41% S
5.00% ash
(by mass)

We now consider 100 lbm of this mixture. Then the mole numbers of each component are

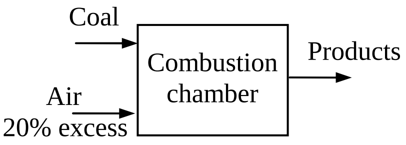
$$N_{\mathrm{C}} = \frac{m_{\mathrm{C}}}{M_{\mathrm{C}}} = \frac{64.63\ \mathrm{lbm}}{12\ \mathrm{lbm/lbmol}} = 5.386\ \mathrm{lbmol}$$

$$N_{\mathrm{H2}} = \frac{m_{\mathrm{H2}}}{M_{\mathrm{H2}}} = \frac{6.095\ \mathrm{lbm}}{2\ \mathrm{lbm/lbmol}} = 3.048\ \mathrm{lbmol}$$

$$N_{\mathrm{O2}} = \frac{m_{\mathrm{O2}}}{M_{\mathrm{O2}}} = \frac{26.64\ \mathrm{lbm}}{32\ \mathrm{lbm/lbmol}} = 0.8325\ \mathrm{lbmol}$$

$$N_{\mathrm{N2}} = \frac{m_{\mathrm{N2}}}{M_{\mathrm{N2}}} = \frac{1.147\ \mathrm{lbm}}{28\ \mathrm{lbm/lbmol}} = 0.04096\ \mathrm{lbmol}$$

$$N_{\mathrm{S}} = \frac{m_{\mathrm{S}}}{M_{\mathrm{S}}} = \frac{1.484\ \mathrm{lbm}}{32\ \mathrm{lbm/lbmol}} = 0.04638\ \mathrm{lbmol}$$



The mole number of the mixture and the mole fractions are

$$N_m = 5.386 + 3.048 + 0.8325 + 0.04096 + 0.04638 = 9.354\ \mathrm{lbmol}$$

$$y_{\mathrm{C}} = \frac{N_{\mathrm{C}}}{N_m} = \frac{5.386\ \mathrm{lbmol}}{9.354\ \mathrm{lbmol}} = 0.5758$$

$$y_{\mathrm{H2}} = \frac{N_{\mathrm{H2}}}{N_m} = \frac{3.048\ \mathrm{lbmol}}{9.354\ \mathrm{lbmol}} = 0.3258$$

$$y_{\mathrm{O2}} = \frac{N_{\mathrm{O2}}}{N_m} = \frac{0.8325\ \mathrm{lbmol}}{9.354\ \mathrm{lbmol}} = 0.0890$$

$$y_{\mathrm{N2}} = \frac{N_{\mathrm{N2}}}{N_m} = \frac{0.04096\ \mathrm{lbmol}}{9.354\ \mathrm{lbmol}} = 0.00438$$

$$y_{\mathrm{S}} = \frac{N_{\mathrm{S}}}{N_m} = \frac{0.04638\ \mathrm{lbmol}}{9.354\ \mathrm{lbmol}} = 0.00496$$

Then, the combustion equation in this case may be written as

$$0.5758\mathrm{C} + 0.3258\mathrm{H}_2 + 0.0890\mathrm{O}_2 + 0.00438\mathrm{N}_2 + 0.00496\mathrm{S} + 1.2a_{\mathrm{th}}(\mathrm{O}_2 + 3.76\mathrm{N}_2)$$
$$\longrightarrow 0.5758\mathrm{CO}_2 + 0.3258\mathrm{H}_2\mathrm{O} + 0.00496\mathrm{SO}_2 + 0.2a_{\mathrm{th}}\mathrm{O}_2 + 1.2a_{\mathrm{th}} \times 3.76k\mathrm{N}_2$$

According to the oxygen balance,

$$\mathrm{O}_2 \text{ balance}:\ 0.0890 + 1.2a_{\mathrm{th}} = 0.5758 + 0.5 \times 0.3258 + 0.00496 + 0.2a_{\mathrm{th}} \longrightarrow a_{\mathrm{th}} = 0.6547$$

Substituting,

$$0.5758\mathrm{C} + 0.3258\mathrm{H}_2 + 0.0890\mathrm{O}_2 + 0.00438\mathrm{N}_2 + 0.00496\mathrm{S} + 0.7856(\mathrm{O}_2 + 3.76\mathrm{N}_2)$$
$$\longrightarrow 0.5758\mathrm{CO}_2 + 0.3258\mathrm{H}_2\mathrm{O} + 0.00496\mathrm{SO}_2 + 0.1309\mathrm{O}_2 + 2.954\mathrm{N}_2$$

The mass of water in the products per unit mass of fuel burned is determined from

$$\frac{m_{\mathrm{H2O}}}{m_{\mathrm{coal}}} = \frac{(0.3258 \times 18)\ \mathrm{lbm}}{(0.5758 \times 12 + 0.3258 \times 2 + 0.0890 \times 32 + 0.00438 \times 28 + 0.00496 \times 32)\ \mathrm{lbm}}$$
$$= \frac{5.864\ \mathrm{kg}}{10.69\ \mathrm{kg}} = \mathbf{0.5486\ lbm\ H_2O/lbm\ coal}$$

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,

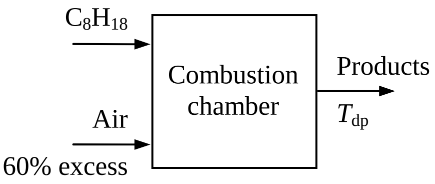
$$P_v = \left(\frac{N_v}{N_{\mathrm{prod}}}\right)P_{\mathrm{prod}} = \left(\frac{0.3258\ \mathrm{lbmol}}{(0.5758 + 0.3258 + 0.00496 + 0.1309 + 2.954)\ \mathrm{lbmol}}\right)(14.696\ \mathrm{psia}) = 1.20\ \mathrm{psia}$$

Thus,

$$T_{\mathrm{dp}} = T_{\mathrm{sat@1.20\ psia}} = \mathbf{108°F} \qquad \text{(from EES)}$$

15-105 Octane is burned steadily with 60 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 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

$$C_8H_{18} + 1.6a_{\text{th}}(O_2 + 3.76N_2) \longrightarrow 8CO_2 + 9H_2O + 0.6a_{\text{th}}O_2 + (1.6)(3.76)a_{\text{th}}N_2$$

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

$$1.6a_{\text{th}} = 8 + 4.5 + 0.6a_{\text{th}} \longrightarrow a_{\text{th}} = 12.5$$

Thus,

$$C_8H_{18} + 20(O_2 + 3.76N_2) \longrightarrow 8CO_2 + 9H_2O + 7.5O_2 + 75.2N_2$$

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,

$$P_v = \left(\frac{N_v}{N_{\text{prod}}}\right)P_{\text{prod}} = \left(\frac{9\text{ kmol}}{(8+9+7.5+75.2)\text{ kmol}}\right)(101.325\text{ kPa}) = 9.147\text{ kPa}$$

Thus,

$$T_{\text{dp}} = T_{\text{sat@9.147 kPa}} = \mathbf{44.1^\circ C} \qquad \text{(from EES)}$$

15-106 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

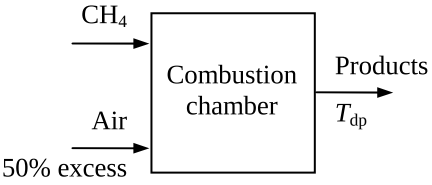
$$\mathrm{CH_4} + 1.5a_{\mathrm{th}}(\mathrm{O_2} + 3.76\mathrm{N_2}) \longrightarrow \mathrm{CO_2} + 2\mathrm{H_2O} + 0.5a_{\mathrm{th}}\mathrm{O_2} + (1.5)(3.76)a_{\mathrm{th}}\mathrm{N_2}$$

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,

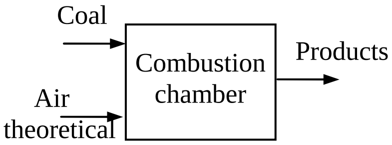
$$P_v = \left(\frac{N_v}{N_{\mathrm{prod}}}\right)P_{\mathrm{prod}} = \left(\frac{2\ \mathrm{kmol}}{(1+2+1+11.28)\ \mathrm{kmol}}\right)(101.325\ \mathrm{kPa}) = 13.26\ \mathrm{kPa}$$

Thus,

$$T_{\mathrm{dp}} = T_{\mathrm{sat@13.26\ kPa}} = \mathbf{51.4°C} \qquad \text{(from EES)}$$

15-107 The higher and lower heating values of gaseous methane 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

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

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,

$$q = h_C = H_P - H_R = \sum N_P \bar{h}^\circ_{f,P} - \sum N_R \bar{h}^\circ_{f,R} = \left(N\bar{h}^\circ_f\right)_{\mathrm{CO2}} + \left(N\bar{h}^\circ_f\right)_{\mathrm{H2O}} - \left(N\bar{h}^\circ_f\right)_{\mathrm{CH4}}$$

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

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

The HHV of the methane is

$$\text{HHV} = \frac{-h_C}{M_m} = \frac{890{,}330\,\text{kJ/kmol CH}_4}{16.043\,\text{kg/kmol CH}_4} = \mathbf{55{,}500\ kJ/kg\ CH_4}$$

The listed value from Table A-27 is **55,530 kJ/kg**. For the LHV, the water in the products is taken to be vapor. Then,

$$\begin{aligned} h_C &= (1\,\text{kmol})(-393{,}520\,\text{kJ/kmol}) + (2\,\text{kmol})(-241{,}820\,\text{kJ/kmol}) - (1\,\text{kmol})(-74{,}850\,\text{kJ/kmol}) \\ &= -802{,}310\,\text{kJ/kmol methane} \end{aligned}$$

The LHV of the coal is then

$$\text{LHV} = \frac{-h_C}{M_m} = \frac{\text{-802,310}\,\text{kJ/kmol CH}_4}{16.043\,\text{kg/kmol CH}_4} = \mathbf{50{,}010\ kJ/kg\ CH_4}$$

The listed value from Table A-27 is **50,050 kJ/kg**. The calculated and listed values are practically identical.

15-108 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

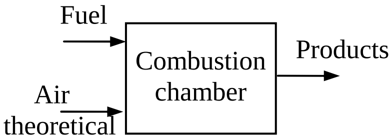
$$\mathrm{C}_n\mathrm{H}_{2n+2} + \frac{3n+1}{2}\left(\mathrm{O}_2 + 3.76\mathrm{N}_2\right) \longrightarrow n\mathrm{CO}_2 + (n+1)\mathrm{H}_2\mathrm{O} + \frac{3n+1}{2}(3.76)\mathrm{N}_2$$

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,

$$\mathrm{LHV} = \frac{n(-393{,}520) + (n+1)(-241{,}820) - \left(\bar{h}^\circ_f\right)_{\mathrm{fuel}}}{M_{\mathrm{fuel}}}$$

15-109 A coal from Colorado is burned with 10 percent excess air during a steady-flow combustion process. The fraction of the water in the combustion products that is liquid and the fraction that is vapor 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.

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 The mass fractions of the constituent of the coal when the ash is substituted are

$$\text{mf}_\text{C} = \frac{m_\text{C}}{m_\text{total}} = \frac{79.61\text{ kg}}{(100-8.62)\text{ kg}} = \frac{79.61\text{ kg}}{91.38\text{ kg}} = 0.8712$$

$$\text{mf}_\text{H2} = \frac{m_\text{H2}}{m_\text{total}} = \frac{4.66\text{ kg}}{91.38\text{ kg}} = 0.05100$$

$$\text{mf}_\text{O2} = \frac{m_\text{O2}}{m_\text{total}} = \frac{4.76\text{ kg}}{91.38\text{ kg}} = 0.05209$$

$$\text{mf}_\text{N2} = \frac{m_\text{N2}}{m_\text{total}} = \frac{1.83\text{ kg}}{91.38\text{ kg}} = 0.02003$$

$$\text{mf}_\text{S} = \frac{m_\text{S}}{m_\text{total}} = \frac{0.52\text{ kg}}{91.38\text{ kg}} = 0.00569$$

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

We now consider 100 kg of this mixture. Then the mole numbers of each component are

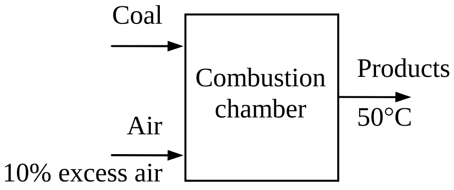
$$N_\text{C} = \frac{m_\text{C}}{M_\text{C}} = \frac{87.12\text{ kg}}{12\text{ kg/kmol}} = 7.26\text{ kmol}$$

$$N_\text{H2} = \frac{m_\text{H2}}{M_\text{H2}} = \frac{5.10\text{ kg}}{2\text{ kg/kmol}} = 2.55\text{ kmol}$$

$$N_\text{O2} = \frac{m_\text{O2}}{M_\text{O2}} = \frac{5.209\text{ kg}}{32\text{ kg/kmol}} = 0.1628\text{ kmol}$$

$$N_\text{N2} = \frac{m_\text{N2}}{M_\text{N2}} = \frac{2.003\text{ kg}}{28\text{ kg/kmol}} = 0.07154\text{ kmol}$$

$$N_\text{S} = \frac{m_\text{S}}{M_\text{S}} = \frac{0.569\text{ kg}}{32\text{ kg/kmol}} = 0.01778\text{ kmol}$$



The mole number of the mixture and the mole fractions are

$$N_m = 7.26 + 2.55 + 0.1628 + 0.07154 + 0.01778 = 10.062\text{ kmol}$$

$$y_\text{C} = \frac{N_\text{C}}{N_m} = \frac{7.26\text{ kmol}}{10.06\text{ kmol}} = 0.7215$$

$$y_\text{H2} = \frac{N_\text{H2}}{N_m} = \frac{2.55\text{ kmol}}{10.06\text{ kmol}} = 0.2535$$

$$y_\text{O2} = \frac{N_\text{O2}}{N_m} = \frac{0.1628\text{ kmol}}{10.06\text{ kmol}} = 0.01618$$

$$y_\text{N2} = \frac{N_\text{N2}}{N_m} = \frac{0.07154\text{ kmol}}{10.06\text{ kmol}} = 0.00711$$

$$y_\text{S} = \frac{N_\text{S}}{N_m} = \frac{0.01778\text{ kmol}}{10.06\text{ kmol}} = 0.00177$$

Then, the combustion equation in this case may be written as

$$0.7215\text{C} + 0.2535\text{H}_2 + 0.01618\text{O}_2 + 0.00711\text{N}_2 + 0.00177\text{S} + 1.1a_{\text{th}}(\text{O}_2 + 3.76\text{N}_2)$$
$$\longrightarrow 0.7215\text{CO}_2 + 0.2535\text{H}_2\text{O} + 0.00177\text{SO}_2 + 1.1a_{\text{th}} \times 3.76\text{N}_2 + 0.1a_{th}\text{O}_2$$

According to the oxygen balance,

$$\text{O}_2 \text{ balance}: \; 0.01618 + 1.1a_{\text{th}} = 0.7215 + (0.5)(0.2535) + (0.00177) + 0.1a_{\text{th}} \longrightarrow a_{\text{th}} = 0.8339$$

Substituting,

$$0.7215\text{C} + 0.2535\text{H}_2 + 0.01617\text{O}_2 + 0.00711\text{N}_2 + 0.00177\text{S} + 0.9173(\text{O}_2 + 3.76\text{N}_2)$$
$$\longrightarrow 0.7215\text{CO}_2 + 0.2535\text{H}_2\text{O} + 0.0018\text{SO}_2 + 3.449\text{N}_2 + 0.0834\text{O}_2$$

If the mixture leaves the boiler saturated with water, the partial pressure of water in the product mixture would be

$$P_v = P_{\text{sat@50°C}} = 12.352 \text{ kPa}$$

and the water mole fraction would be

$$y_v = \frac{P_v}{P} = \frac{12.352 \text{ kPa}}{101.325 \text{ kPa}} = 0.1219$$

The actual mole fraction of water in the products is

$$y_v = \frac{N_{\text{H2O}}}{N_{\text{prod}}} = \frac{0.2535 \text{ kmol}}{(0.7215 + 0.2535 + 0.0018 + 3.449 + 0.0834)\text{kmol}} = 0.0562$$

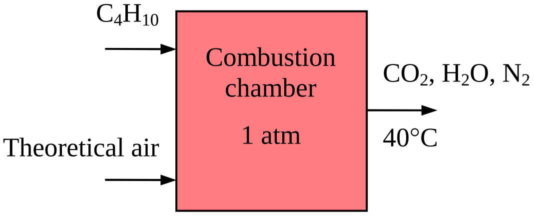
Since the actual mole fraction of the water is less than that when the mixture is saturated, the mixture never becomes saturated and hence, no condensate is formed. All of the water in the mixture is then in the vapor form at this temperature.

15-110 Butane is burned with stoichiometric amount of air. The fraction of the water in the products that is liquid 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_4H_{10}, the combustion equation can be written as

$$C_4H_{10} + 6.5(O_2 + 3.76N_2) \longrightarrow 4CO_2 + 5H_2O + 24.44N_2$$



The mole fraction of water in the products is

$$y_v = \frac{N_{\text{H2O}}}{N_{\text{prod}}} = \frac{5\,\text{kmol}}{(4+5+24.44)\text{kmol}} = 0.1495$$

The saturation pressure for the water vapor is

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

When the combustion gases are saturated, the mole fraction of the water vapor will be

$$y_g = \frac{P_v}{P} = \frac{7.3851\,\text{kPa}}{101.325\,\text{kPa}} = 0.0729$$

The mole fraction of the liquid water is then

$$y_f = y_v - y_g = 0.1495 - 0.0729 = 0.0766$$

Thus, the fraction of liquid water in the combustion products is

$$f_{\text{liquid}} = \frac{y_f}{y_g} = \frac{0.0766}{0.1495} = \mathbf{0.512}$$

15-111 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.

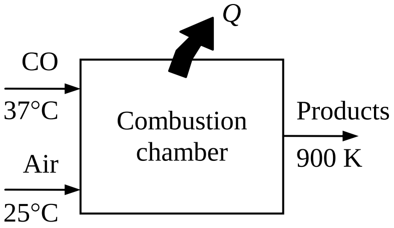
Assumptions **1** Steady operating conditions exist. **2** Air and combustion gases are ideal gases. **3** Kinetic and potential energies are negligible. **4** There are no work interactions. **5** Combustion is complete.

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,

$$\boldsymbol{v}_{\mathrm{CO}} = \frac{RT}{P} = \frac{(0.2968\ \mathrm{kPa \cdot m^3/kg \cdot K})(310\ \mathrm{K})}{(110\ \mathrm{kPa})} = 0.836\ \mathrm{m^3/kg}$$

$$\dot{m}_{\mathrm{CO}} = \frac{\dot{\boldsymbol{V}}_{\mathrm{CO}}}{\boldsymbol{v}_{\mathrm{CO}}} = \frac{0.4\ \mathrm{m^3/min}}{0.836\ \mathrm{m^3/kg}} = 0.478\ \mathrm{kg/min}$$



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	$\bar{\mathbf{h}}_{\mathbf{f}}^\circ$ **kJ/kmol**	$\bar{\mathbf{h}}_{\mathbf{298\ K}}$ **kJ/kmol**	$\bar{\mathbf{h}}_{\mathbf{310\ K}}$ **kJ/kmol**	$\bar{\mathbf{h}}_{\mathbf{900\ K}}$ **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

$$\dot{Q}_{\mathrm{out}} = \dot{N} Q_{\mathrm{out}} = \left(\frac{\dot{m}}{N}\right) Q_{\mathrm{out}} = \left(\frac{0.478\ \mathrm{kg/min}}{28\ \mathrm{kg/kmol}}\right)(208{,}927\ \mathrm{kJ/kmol}) = \mathbf{3567\ kJ/min}$$

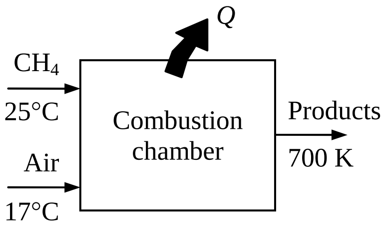
15-112 Methane gas is burned steadily with dry air. The volumetric analysis of the products is given. The percentage of theoretical air used and the heat transfer from the combustion chamber are to be determined.

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.

Analysis (*a*) 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

$$\mathrm{CH_4} + 4[\mathrm{O_2} + 3.76\mathrm{N_2}] \longrightarrow 0.94\mathrm{CO_2} + 0.06\mathrm{CO} + 2.03\mathrm{O_2} + 15.05\mathrm{N_2} + 2\mathrm{H_2O}$$

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

(*b*) Under steady-flow conditions, energy balance applied on the combustion chamber 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**	$\bar{h}_{290\ \mathrm{K}}$ **kJ/kmol**	$\bar{h}_{298\ \mathrm{K}}$ **kJ/kmol**	$\bar{h}_{700\ \mathrm{K}}$ **kJ/kmol**
$\mathrm{CH_4}$ (*g*)	-74,850	---	---	---
$\mathrm{O_2}$	0	8443	8682	21,184
$\mathrm{N_2}$	0	8432	8669	20,604
$\mathrm{H_2O}$ (*g*)	-241,820	---	9904	24,088
CO	-110,530	---	8669	20,690
$\mathrm{CO_2}$	-393,520	---	9364	27,125

Thus,

$$\begin{aligned}-Q_{\mathrm{out}} &= (0.94)(-393{,}520 + 27{,}125 - 9364) + (0.06)(-110{,}530 + 20{,}690 - 8669) \\ &\quad + (2)(-241{,}820 + 24{,}088 - 9904) + (2.03)(0 + 21{,}184 - 8682) + (15.04)(0 + 20{,}604 - 8669) \\ &\quad - (1)(-74{,}850 + h_{298} - h_{298}) - (4)(0 + 8443 - 8682) - (15.04)(0 + 8432 - 8669) \\ &= -530{,}022\ \mathrm{kJ/kmol\,CH_4}\end{aligned}$$

or $Q_{\mathrm{out}} = \mathbf{530{,}022\ \ kJ/kmol\,CH_4}$

15-113 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

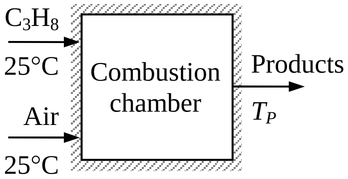
$$\sum N_P\left(\bar{h}_f^\circ + \bar{h} - \bar{h}^\circ\right)_P = \sum N_R\left(\bar{h}_f^\circ + \bar{h} - \bar{h}^\circ\right)_R \longrightarrow \sum N_P\left(\bar{h}_f^\circ + \bar{h}_T - \bar{h}^\circ\right)_P = \left(N\bar{h}_f^\circ\right)_{\mathrm{C_3H_8}}$$

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

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

From the tables,



Substance	$\bar{h}_f^\circ$ kJ/kmol	$\bar{h}_{298\ \mathrm{K}}$ 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,

$$(3)\left(-393{,}520 + \bar{h}_{\mathrm{CO_2}} - 9364\right) + (4)\left(-241{,}820 + \bar{h}_{\mathrm{H_2O}} - 9904\right) + (18.8)\left(0 + \bar{h}_{\mathrm{N_2}} - 8669\right) = (1)(-103{,}850)$$

It yields

$$3\bar{h}_{\mathrm{CO_2}} + 4\bar{h}_{\mathrm{H_2O}} + 18.8\bar{h}_{\mathrm{N_2}} = 2{,}274{,}675 \text{ kJ}$$

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: $$3\bar{h}_{\mathrm{CO_2}} + 4\bar{h}_{\mathrm{H_2O}} + 18.8\bar{h}_{\mathrm{N_2}} = (3)(125{,}152) + (4)(103{,}508) + (18.8)(79{,}320)$$
$$= 2{,}280{,}704 \text{ kJ (Higher than 2,274,675 kJ)}$$

At 2350 K: $$3\bar{h}_{\mathrm{CO_2}} + 4\bar{h}_{\mathrm{H_2O}} + 18.8\bar{h}_{\mathrm{N_2}} = (3)(122{,}091) + (4)(100{,}846) + (18.8)(77{,}496)$$
$$= 2{,}226{,}582 \text{ kJ (Lower than 2,274,675 kJ)}$$

By interpolation, T_P = **2394 K**

(*b*) The balanced combustion equation for complete combustion with 300% theoretical air is

$$C_3H_8(g)+15(O_2+3.76N_2)\longrightarrow 3CO_2+4H_2O+10O_2+56.4N_2$$

Substituting known numerical values,

$$(3)(-393{,}520+\bar{h}_{CO_2}-9364)+(4)(-241{,}820+\bar{h}_{H_2O}-9904)$$
$$+(10)(0+\bar{h}_{O_2}-8682)+(56.4)(0+\bar{h}_{N_2}-8669)=(1)(-103{,}850)$$

which yields

$$3\bar{h}_{CO_2}+4\bar{h}_{H_2O}+10\bar{h}_{O_2}+56.4\bar{h}_{N_2}=2{,}687{,}450\text{ kJ}$$

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 2,687,449 / (3 + 4 + 10 + 56.4) = 36,614 kJ/kmol. This enthalpy value corresponds to about 1200 K for N_2. Noting that the majority of the moles are N_2, T_P will be close to 1200 K, but somewhat under it because of the higher specific heats of CO_2 and H_2O.

At 1160 K:

$$3\bar{h}_{CO_2}+4\bar{h}_{H_2O}+10\bar{h}_{O_2}+56.4\bar{h}_{N_2}=(3)(51{,}602)+(4)(42{,}642)+(10)(37{,}023)+(56.4)(35{,}430)$$
$$=2{,}693{,}856\text{ kJ (Higher than 2,687,450 kJ)}$$

At 1140 K:

$$3\bar{h}_{CO_2}+4\bar{h}_{H_2O}+10\bar{h}_{O_2}+56.4\bar{h}_{N_2}=(3)(50{,}484)+(4)(41{,}780)+(10)(36{,}314)+(56.4)(34{,}760)$$
$$=2{,}642{,}176\text{ kJ (Lower than 2,687,450 kJ)}$$

By interpolation, T_P = **1158 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**

15-114 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

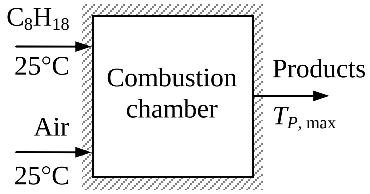
$$\sum N_P\left(\bar{h}_f^\circ + \bar{h} - \bar{h}^\circ\right)_P = \sum N_R\left(\bar{h}_f^\circ + \bar{h} - \bar{h}^\circ\right)_R \longrightarrow \sum N_P\left(\bar{h}_f^\circ + \bar{h}_T - \bar{h}^\circ\right)_P = \left(N\bar{h}_f^\circ\right)_{\mathrm{C_8H_{18}}}$$

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

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

From the tables,

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

$$(8)\left(-393{,}520 + \bar{h}_{\mathrm{CO_2}} - 9364\right) + (9)\left(-241{,}820 + \bar{h}_{\mathrm{H_2O}} - 9904\right) + (47)\left(0 + \bar{h}_{\mathrm{N_2}} - 8669\right) = (1)(-249{,}950)$$

It yields $\quad 8\bar{h}_{\mathrm{CO_2}} + 9\bar{h}_{\mathrm{H_2O}} + 47\bar{h}_{\mathrm{N_2}} = 5{,}646{,}081\ \mathrm{kJ}$

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: $\quad 8\bar{h}_{\mathrm{CO_2}} + 9\bar{h}_{\mathrm{H_2O}} + 47\bar{h}_{\mathrm{N_2}} = (8)(125{,}152) + (9)(103{,}508) + (47)(79{,}320)$
$= 5{,}660{,}828\ \mathrm{kJ}$ (Higher than 5,646,081 kJ)

At 2350 K: $\quad 8\bar{h}_{\mathrm{CO_2}} + 9\bar{h}_{\mathrm{H_2O}} + 47\bar{h}_{\mathrm{N_2}} = (8)(122{,}091) + (9)(100{,}846) + (47)(77{,}496)$
$= 5{,}526{,}654\ \mathrm{kJ}$ (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

$$\mathrm{C_8H_{18}} + 12.5\mathrm{O_2} \longrightarrow 8\mathrm{CO_2} + 9\mathrm{H_2O}$$

Thus, $\quad (8)\left(-393{,}520 + \bar{h}_{\mathrm{CO_2}} - 9364\right) + (9)\left(-241{,}820 + \bar{h}_{\mathrm{H_2O}} - 9904\right) = (1)(-249{,}950)$

It yields $\quad 8\bar{h}_{\mathrm{CO_2}} + 9\bar{h}_{\mathrm{H_2O}} = 5{,}238{,}638\ \mathrm{kJ}$

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: $\quad 8\bar{h}_{\mathrm{CO_2}} + 9\bar{h}_{\mathrm{H_2O}} = (8)(174{,}695) + (9)(147{,}457) = 2{,}724{,}673\ \mathrm{kJ}$

At 3250 K: $\quad 8\bar{h}_{\mathrm{CO_2}} + 9\bar{h}_{\mathrm{H_2O}} = (8)(177{,}822) + (9)(150{,}272) = 2{,}775{,}024\ \mathrm{kJ}$

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.

15-115 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.

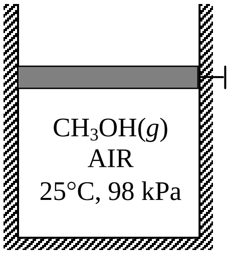
Assumptions **1** Combustion is complete. **2** Air and the combustion gases are ideal gases. **4** Changes in kinetic and potential energies are negligible.

Analysis (*a*) The combustion equation of $CH_3OH(g)$ with stoichiometric amount of air is

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

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

$$1 + 2a_{\text{th}} = 2 + 2 \quad \longrightarrow \quad a_{\text{th}} = 1.5$$



Thus,

$$CH_3OH + 1.5(O_2 + 3.76N_2) \longrightarrow CO_2 + 2H_2O + 5.64N_2$$

The final temperature in the tank is determined from the energy balance relation $E_{\text{in}} - E_{\text{out}} = \Delta E_{\text{system}}$ 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	$\bar{h}_f^\circ$ **kJ/kmol**	$\bar{h}_{298\text{ K}}$ **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,

$$\begin{aligned}&(1)(-393{,}520 + \bar{h}_{CO_2} - 9364 - 8.314 \times T_P) + (2)(-241{,}820 + \bar{h}_{H_2O} - 9904 - 8.314 \times T_P)\\&+ (5.64)(0 + \bar{h}_{N_2} - 8669 - 8.314 \times T_P) = (1)(-200{,}670 - 8.314 \times 298) + (1.5)(0 - 8.314 \times 298)\\&\qquad + (5.64)(0 - 8.314 \times 298)\end{aligned}$$

It yields

$$\bar{h}_{CO_2} + 2\bar{h}_{H_2O} + 5.64\bar{h}_{N_2} - 71.833 \times T_P = 734{,}388 \text{ kJ}$$

The temperature of the product gases is obtained from a trial and error solution,

At 2850 K:

$$\bar{h}_{CO_2} + 2\bar{h}_{H_2O} + 5.64\bar{h}_{N_2} - 71.833 \times T_P = (1)(152{,}908) + (2)(127{,}952) + (5.64)(95{,}859) - (71.833)(2850)$$
$$= 744{,}733 \text{ kJ (Higher than 734,388 kJ)}$$

At 2800 K:

$$\bar{h}_{CO_2} + 2\bar{h}_{H_2O} + 5.64\bar{h}_{N_2} - 71.833 \times T_P = (1)(149{,}808) + (2)(125{,}198) + (5.64)(94{,}014) - (71.833)(2800)$$
$$= 729{,}311 \text{ kJ (Lower than 734,388 kJ)}$$

By interpolation $T_P = 2816$ 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

$$\frac{P_1\mathbf{V}}{P_2\mathbf{V}} = \frac{N_1 R_u T_1}{N_2 R_u T_2} \longrightarrow P_2 = \frac{N_2 T_2}{N_1 T_1} P_1 = \frac{(8.64 \text{ kmol})(2816 \text{ K})}{(8.14 \text{ kmol})(298 \text{ K})}(98 \text{ kPa}) = \mathbf{983 \text{ kPa}}$$

(*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: $\bar{h}_{CO_2} + 2\bar{h}_{H_2O} + 5.64\bar{h}_{N_2} = (1)(122{,}091) + (2)(100{,}846) + (5.64)(77{,}496)$
$= 760{,}860$ kJ (Higher than 754,555 kJ)

At 2300 K: $\bar{h}_{CO_2} + 2\bar{h}_{H_2O} + 5.64\bar{h}_{N_2} = (1)(119{,}035) + (2)(98{,}199) + (5.64)(75{,}676)$
$= 742{,}246$ kJ (Lower than 754,555 kJ)

By interpolation, $T_P = 2333$ 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

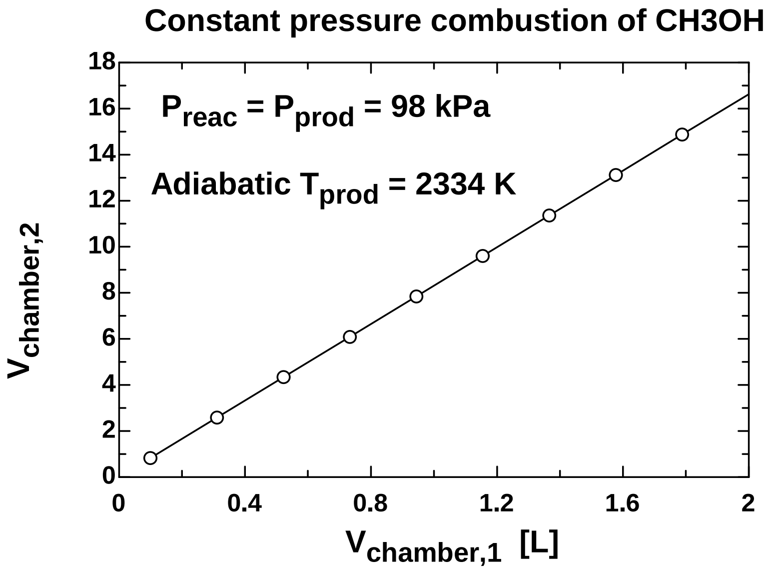
$$\frac{P\mathbf{V}_1}{P\mathbf{V}_2} = \frac{N_1 R_u T_1}{N_2 R_u T_2} \longrightarrow \mathbf{V}_2 = \frac{N_2 T_2}{N_1 T_1}\mathbf{V}_1 = \frac{(8.64 \text{ kmol})(2333 \text{ K})}{(8.14 \text{ kmol})(298 \text{ K})}(0.8 \text{ L}) = \mathbf{6.65 \text{ L}}$$

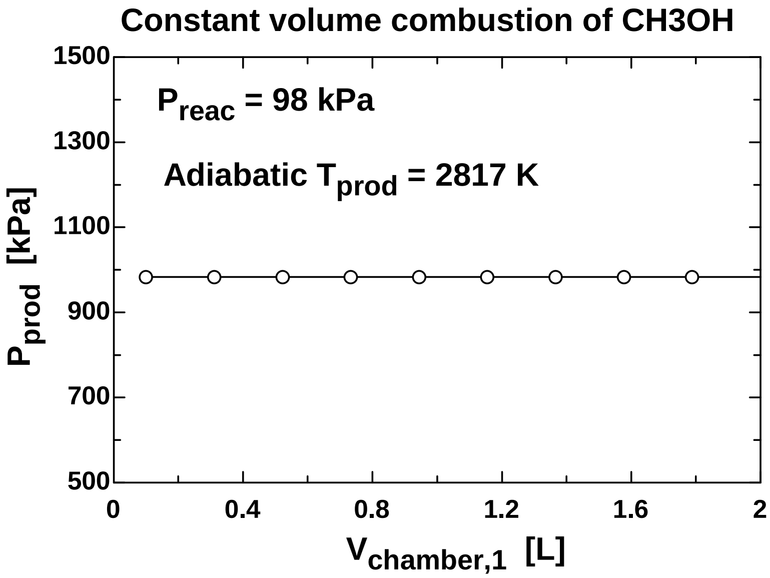
15-116 Problem 15–115 is reconsidered. The effect of the initial volume of the combustion chamber on the maximum pressure of the chamber for constant volume combustion or the maximum volume of the chamber for constant pressure combustion is to be investigated.

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

```
"Input Data"
T_reac = (25+273) "[K]"  "reactant mixture temperature"
P_reac = 98 [kPa] "reactant mixture pressure"
{V_chamber_1 = 0.8 [L]}
h_CH3OH = -200670 [kJ/kmol]
Mw_O2 = 32 [kg/kmol]
Mw_N2 = 28 [kg/kmol]
Mw_CH3OH=(3*12+1*32+4*1) "[kg/kmol]"
R_u = 8.314 [kJ/kmol-K] "universal gas constant"

"For  theoretical oxygen, the complete combustion equation is"
"CH3OH + A_th O2=1 CO2+2 H2O  "
1+ 2*A_th=1*2+2*1"theoretical O balance"

"The balanced complete combustion equation with theroetical air is"
"CH3OH + A_th (O2+3.76 N2)=1 CO2+ 2 H2O + A_th(3.76) N2 "

"now to find the actual moles of reactants and products per mole of fuel"
N_Reac = 1 + A_th*4.76
N_Prod=1+2+A_th*3.76

"Apply First Law to the closed system combustion chamber and assume ideal gas
behavior. Assume the water formed in the combustion process exists in the gas phase."

"The following is the constant volume, adiabatic solution:"
E_in - E_out = DELTAE_sys
E_in = 0  "No heat transfer for the adiabatic combustion process"
E_out = 0"kJ/kmol_CH3OH" "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*(h_CH3OH - R_u*T_reac) +A_th*(enthalpy(O2,T=T_reac) -
R_u*T_reac)+A_th*3.76*(enthalpy(N2,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)+A_th*3.76*(enthalpy(N2,T=T_prod) - R_u*T_prod)
V_chamber_2 = V_chamber_1

"The final pressure and volume of the tank are those of the product gases.  Assuming
ideal gas behavior for the gases in the constant volume tank, the ideal gas law gives:"
P_reac*V_chamber_1*convert(L,m^3) =N_reac*N_fuel* R_u *T_reac
P_prod*V_chamber_2*convert(L,m^3) = N_prod*N_fuel* R_u * T_prod

{ "The following is the constant pressure, adiabatic solution:"
P_prod = P_Reac
H_reac=H_prod
H_reac = 1*h_CH3OH  +A_th*enthalpy(O2,T=T_reac) +A_th*3.76*enthalpy(N2,T=T_reac)
H_prod = 1*enthalpy(CO2,T=T_prod)+2*enthalpy(H2O,T=T_prod)
+A_th*3.76*enthalpy(N2,T=T_prod) }
```

N_{fuel} [kmol]	P_{prod} [kPa]	T_{prod} [K]	$V_{chamber,1}$ [L]
4.859E-07	983.5	2817	0.1
0.000001512	983.5	2817	0.3111
0.000002538	983.5	2817	0.5222
0.000003564	983.5	2817	0.7333
0.000004589	983.5	2817	0.9444
0.000005615	983.5	2817	1.156
0.000006641	983.5	2817	1.367
0.000007667	983.5	2817	1.578
0.000008693	983.5	2817	1.789
0.000009719	983.5	2817	2



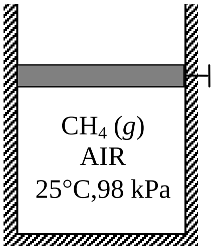


15-117 Methane is burned with the stoichiometric amount of air in a combustion chamber. The maximum pressure that can occur in the combustion chamber if the combustion takes place at constant volume and the maximum volume of the combustion chamber if the combustion occurs at constant pressure are to 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

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



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

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

Thus,

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

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),

$$0 = \sum N_P\left(\bar{h}_f^\circ + \bar{h} - \bar{h}^\circ - P\bar{v}\right)_P - \sum N_R\left(\bar{h}_f^\circ + \bar{h} - \bar{h}^\circ - P\bar{v}\right)_R$$

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

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

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

Substance	$\bar{h}_f^\circ$ **kJ/kmol**	$\bar{h}_{298\ \mathrm{K}}$ **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,

$$\begin{aligned}&(1)\left(-393{,}520 + \bar{h}_{\mathrm{CO_2}} - 9364 - 8.314 \times T_P\right) + (2)\left(-241{,}820 + \bar{h}_{\mathrm{H_2O}} - 9904 - 8.314 \times T_P\right)\\ &+ (7.52)\left(0 + \bar{h}_{\mathrm{N_2}} - 8669 - 8.314 \times T_P\right) = (1)(-74{,}850 - 8.314 \times 298) + (2)(0 - 8.314 \times 298)\\ &\qquad\qquad\qquad\qquad\qquad\qquad + (7.52)(0 - 8.314 \times 298)\end{aligned}$$

It yields

$$\bar{h}_{\mathrm{CO_2}} + 2\bar{h}_{\mathrm{H_2O}} + 7.52\bar{h}_{\mathrm{N_2}} - 87.463 \times T_P = 870{,}609 \text{ kJ}$$

The temperature of the product gases is obtained from a trial and error solution,

At 2850 K:

$$\bar{h}_{CO_2} + 2\bar{h}_{H_2O} + 7.52\bar{h}_{N_2} - 87.463\times T_P = (1)(152{,}908) + (2)(127{,}952) + (7.52)(95{,}859) - (87.463)(2850)$$
$$= 880{,}402 \text{ kJ (Higher than } 870{,}609 \text{ kJ)}$$

At 2800 K:

$$\bar{h}_{CO_2} + 2\bar{h}_{H_2O} + 7.52\bar{h}_{N_2} - 87.463\times T_P = (1)(149{,}808) + (2)(125{,}198) + (7.52)(94{,}014) - (87.463)(2800)$$
$$= 862{,}293 \text{ kJ (Lower than } 870{,}609 \text{ kJ)}$$

By interpolation, T_P = 2823 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

$$\frac{P_1 V}{P_2 V} = \frac{N_1 R_u T_1}{N_2 R_u T_2} \longrightarrow P_2 = \frac{N_2 T_2}{N_1 T_1} P_1 = \frac{(10.52 \text{ kmol})(2823 \text{ K})}{(10.52 \text{ kmol})(298 \text{ K})}(98 \text{ kPa}) = \mathbf{928 \text{ kPa}}$$

(*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

$$(1)(-393{,}520 + \bar{h}_{CO_2} - 9364) + (2)(-241{,}820 + \bar{h}_{H_2O} - 9904) + (7.52)(0 + \bar{h}_{N_2} - 8669)$$
$$= (1)(-74{,}850) + (2)(0) + (7.52)(0)$$

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: $\bar{h}_{CO_2} + 2\bar{h}_{H_2O} + 7.52\bar{h}_{N_2} = (1)(122{,}091) + (2)(100{,}846) + (7.52)(77{,}496)$
$= 906{,}553 \text{ kJ (Higher than } 896{,}673 \text{ kJ)}$

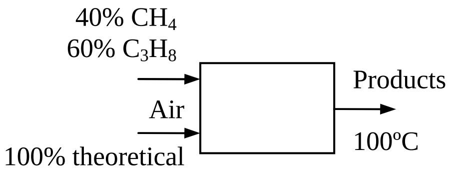
At 2300 K: $\bar{h}_{CO_2} + 2\bar{h}_{H_2O} + 7.52\bar{h}_{N_2} = (1)(119{,}035) + (2)(98{,}199) + (7.52)(75{,}676)$
$= 884{,}517 \text{ kJ (Lower than } 896{,}673 \text{ kJ)}$

By interpolation, T_P = 2328 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

$$\frac{P V_1}{P V_2} = \frac{N_1 R_u T_1}{N_2 R_u T_2} \longrightarrow V_2 = \frac{N_2 T_2}{N_1 T_1} V_1 = \frac{(10.52 \text{ kmol})(2328 \text{ K})}{(10.52 \text{ kmol})(298 \text{ K})}(0.8 \text{ L}) = \mathbf{6.25 \text{ L}}$$

15-118 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

$$P_{v,prod} = y_v P_{prod} = (0.1625)(100\text{ kPa}) = 16.25\text{ kPa}$$

The dew point temperature of the products is

$$T_{dp} = T_{sat@16.25\text{ kPa}} = 55.64°C$$

The partial pressure of the water vapor remaining in the products at the product temperature is

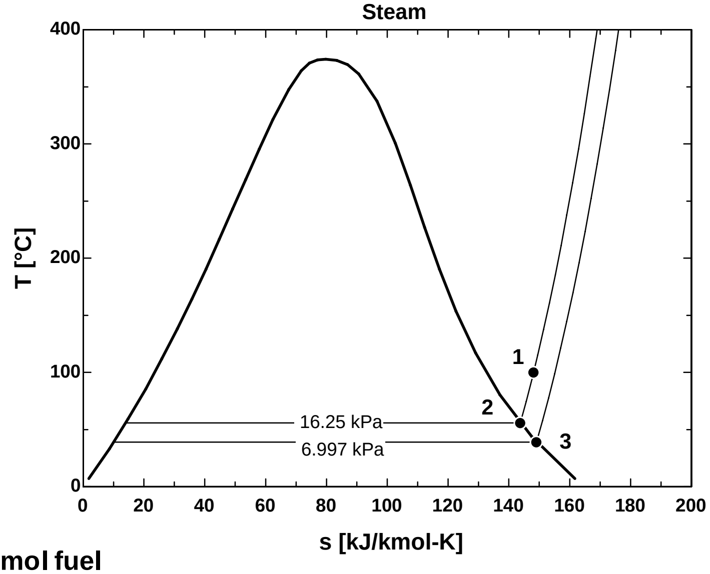
$$P_v = P_{sat\,@\,39°C} = 7.0\text{ kPa}$$

The kmol of water vapor in the products at the product temperature is

$$P_v = \frac{N_v}{N_{total,product}} P_{prod}$$

$$7.0\text{ kPa} = \frac{N_v}{2.2 + N_v + 14.29}$$

$$N_v = 1.241\text{ kmol}$$



The kmol of water condensed is

$$N_w = 3.2 - 1.241 = \textbf{1.96 kmol water/kmol fuel}$$

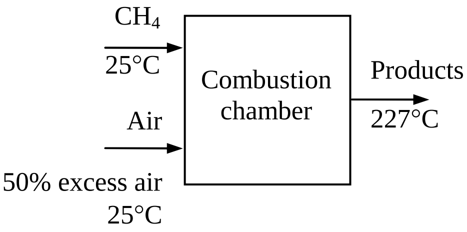
15-119 Methane is burned steadily with 50 percent excess air in a steam boiler. The amount of steam generated per unit of fuel mass burned, the change in the exergy of the combustion streams, the change in the exergy of the steam stream, and the lost work potential are to be determined.

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

$$\mathrm{CH_4} + 3(\mathrm{O_2} + 3.76\mathrm{N_2}) \longrightarrow \mathrm{CO_2} + 2\mathrm{H_2O} + \mathrm{O_2} + 11.28\mathrm{N_2}$$



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	$\bar{h}_f^\circ$ kJ/kmol	$\bar{h}_{298\text{K}}$ kJ/kmol	$\bar{h}_{500\text{K}}$ 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,

$$\begin{aligned}-Q_{\text{out}} &= (1)(-393{,}520 + 17{,}678 - 9364) + (2)(-241{,}820 + 16{,}828 - 9904)\\ &\quad + (1)(0 + 14{,}770 - 8682) + (11.28)(0 + 14{,}581 - 8669) - (1)(-74{,}850)\\ &= -707{,}373 \text{ kJ/kmol of fuel}\end{aligned}$$

The heat loss per unit mass of the fuel is

$$Q_{\text{out}} = \frac{707{,}373 \text{ kJ/kmol of fuel}}{16 \text{ kg/kmol of fuel}} = 44{,}211 \text{ kJ/kg fuel}$$

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)

$$\frac{m_s}{m_f} = \frac{Q_{\text{out}}}{\Delta h_s} = \frac{44{,}211 \text{ kJ/kg fuel}}{(3214.5 - 852.26) \text{ kJ/kg steam}} = \mathbf{18.72\ kg\ steam/kg\ fuel}$$

(*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 $P_i = y_i P_{total}$, 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_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,

$$S_{\text{gen}} = S_P - S_R + \frac{Q_{\text{out}}}{T_{\text{surr}}} = 3306.65 - 3023.69 + \frac{707{,}373}{298} = 2657 \text{ kJ/K (per kmol fuel)}$$

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:

$$\begin{aligned} X_{\text{dest}} &= -\left(\frac{m_s}{m_f}\Delta X_{\text{steam}} + \Delta X_{\text{gases}}\right) \\ &= -[(18.72 \text{ kg steam/kg fuel})(1039 \text{ kJ/kg steam}) + (-49{,}490 \text{ kJ/kg fuel})] \\ &= \mathbf{30{,}040 \text{ kJ/kg fuel}} \end{aligned}$$

15-120 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*) The mass fractions of the constituent of the coal when the ash is substituted are

$$\mathrm{mf}_{\mathrm{C}}=\frac{m_{\mathrm{C}}}{m_{\mathrm{total}}}=\frac{61.40\text{ kg}}{(100-5.00)\text{ kg}}=\frac{61.40\text{ kg}}{95.00\text{ kg}}=0.6463$$

$$\mathrm{mf}_{\mathrm{H2}}=\frac{m_{\mathrm{H2}}}{m_{\mathrm{total}}}=\frac{5.79\text{ kg}}{95.00\text{ kg}}=0.06095$$

$$\mathrm{mf}_{\mathrm{O2}}=\frac{m_{\mathrm{O2}}}{m_{\mathrm{total}}}=\frac{25.31\text{ kg}}{95.00\text{ kg}}=0.2664$$

$$\mathrm{mf}_{\mathrm{N2}}=\frac{m_{\mathrm{N2}}}{m_{\mathrm{total}}}=\frac{1.09\text{ kg}}{95.00\text{ kg}}=0.01147$$

$$\mathrm{mf}_{\mathrm{S}}=\frac{m_{\mathrm{S}}}{m_{\mathrm{total}}}=\frac{1.41\text{ kg}}{95.00\text{ kg}}=0.01484$$

61.40% C
5.79% H_2
25.31% O_2
1.09% N_2
1.41% S
5.00% ash
(by mass)

We now consider 100 kg of this mixture. Then the mole numbers of each component are

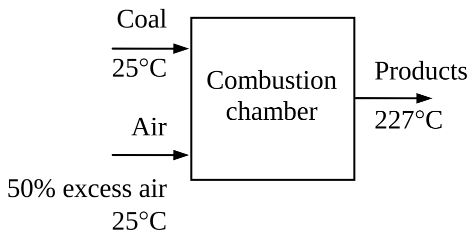
$$N_{\mathrm{C}}=\frac{m_{\mathrm{C}}}{M_{\mathrm{C}}}=\frac{64.63\text{ kg}}{12\text{ kg/kmol}}=5.386\text{ kmol}$$

$$N_{\mathrm{H2}}=\frac{m_{\mathrm{H2}}}{M_{\mathrm{H2}}}=\frac{6.095\text{ kg}}{2\text{ kg/kmol}}=3.048\text{ kmol}$$

$$N_{\mathrm{O2}}=\frac{m_{\mathrm{O2}}}{M_{\mathrm{O2}}}=\frac{26.64\text{ kg}}{32\text{ kg/kmol}}=0.8325\text{ kmol}$$

$$N_{\mathrm{N2}}=\frac{m_{\mathrm{N2}}}{M_{\mathrm{N2}}}=\frac{1.147\text{ kg}}{28\text{ kg/kmol}}=0.04096\text{ kmol}$$

$$N_{\mathrm{S}}=\frac{m_{\mathrm{S}}}{M_{\mathrm{S}}}=\frac{1.484\text{ kg}}{32\text{ kg/kmol}}=0.04638\text{ kmol}$$



The mole number of the mixture and the mole fractions are

$$N_m=5.386+3.048+0.8325+0.04096+0.04638=9.354\text{ kmol}$$

$$y_{\mathrm{C}}=\frac{N_{\mathrm{C}}}{N_m}=\frac{5.386\text{ kmol}}{9.354\text{ kmol}}=0.5758$$

$$y_{\mathrm{H2}}=\frac{N_{\mathrm{H2}}}{N_m}=\frac{3.048\text{ kmol}}{9.354\text{ kmol}}=0.3258$$

$$y_{\mathrm{O2}}=\frac{N_{\mathrm{O2}}}{N_m}=\frac{0.8325\text{ kmol}}{9.354\text{ kmol}}=0.0890$$

$$y_{\mathrm{N2}}=\frac{N_{\mathrm{N2}}}{N_m}=\frac{0.04096\text{ kmol}}{9.354\text{ kmol}}=0.00438$$

$$y_{\mathrm{S}}=\frac{N_{\mathrm{S}}}{N_m}=\frac{0.04638\text{ kmol}}{9.354\text{ kmol}}=0.00496$$

Then, the combustion equation in this case 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

$$M_m=\frac{m_m}{N_m}=\frac{(0.5758\times 12+0.3258\times 2+0.0890\times 32+0.00438\times 28+0.00496\times 32)\text{ kg}}{(0.5758+0.3258+0.0890+0.00438+0.0496)\text{ kmol}}$$
$$=\frac{10.69\text{ kg}}{1.045\text{ kmol}}=10.23\text{ kg/kmol coal}$$

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

$$Q_{\text{out}}=\frac{274{,}505\text{ kJ/kmol of fuel}}{10.23\text{ kg/kmol of fuel}}=26{,}833\text{ kJ/kg fuel}$$

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)

$$\frac{m_s}{m_f}=\frac{Q_{\text{out}}}{\Delta h_s}=\frac{26{,}833\text{ kJ/kg fuel}}{(3214.5-852.26)\text{ kJ/kg steam}}=\mathbf{11.36\text{ kg steam/kg fuel}}$$

(*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 $P_i = y_i P_{total}$, 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

	$\mathbf{N_i}$	$\mathbf{y_i}$	$\mathbf{\bar{s}_i^\circ(T,1atm)}$	$\mathbf{R_u \ln(y_i P_m)}$	$\mathbf{N_i \bar{s}_i}$
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,

$$S_{gen} = S_P - S_R + \frac{Q_{out}}{T_{surr}} = 1071.58 - 1001.61 + \frac{274{,}505}{298} = 991.1 \text{ kJ/K (per kmol fuel)}$$

The exergy change of the combustion streams is equal to the exergy destruction since there is no actual work output. That is,

$$\Delta X_{gases} = -X_{dest} = -T_0 S_{gen} = -(298 \text{ K})(991.1 \text{ kJ/K}) = -295{,}348 \text{ kJ/kmol fuel}$$

Per unit mass basis,

$$\Delta X_{gases} = \frac{-295{,}348 \text{ kJ/K}}{10.23 \text{ kg/kmol}} = \mathbf{-28{,}870 \text{ kJ/kg fuel}}$$

Note that the exergy change is negative since the exergy of combustion gases decreases.

(*c*) The exergy change of the steam stream is

$$\Delta X_{steam} = \Delta h - T_0 \Delta s = (3214.5 - 852.26) - (298)(6.7714 - 2.3305) = \mathbf{1039 \text{ kJ/kg steam}}$$

(*d*) The lost work potential is the negative of the net exergy change of both streams:

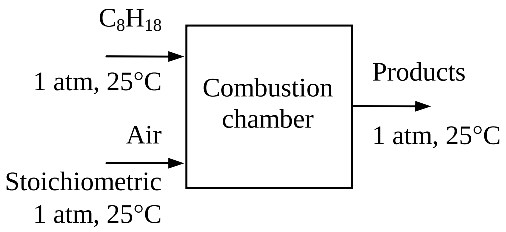
$$\begin{aligned} X_{dest} &= -\left(\frac{m_s}{m_f} \Delta X_{steam} + \Delta X_{gases} \right) \\ &= -\left[(11.36 \text{ kg steam/kg fuel})(1039 \text{ kJ/kg steam}) + (-28{,}870 \text{ kJ/kg fuel}) \right] \\ &= \mathbf{17{,}070 \text{ kJ/kg fuel}} \end{aligned}$$

15-121 Octane is burned with stoichiometric amount of air. The maximum work that can be produced 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 combustion equation is

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



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,

$$\begin{aligned} W_{\text{rev}} &= \sum N_R \bar{g}^\circ_{f,R} - \sum N_P \bar{g}^\circ_{f,P} \\ &= (1)(16{,}530) - (8)(-394{,}360) - (9)(-228{,}590) \\ &= 5{,}228{,}720 \text{ kJ} \quad (\text{per kmol of fuel}) \end{aligned}$$

since the g_f° of stable elements at 25°C and 1 atm is zero. Per unit mass basis,

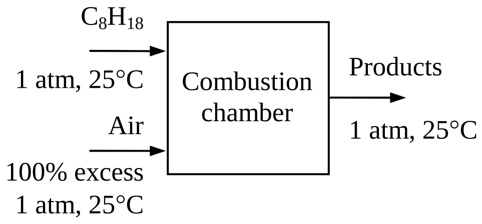
$$W_{\text{rev}} = \frac{5{,}228{,}720 \text{ kJ/kmol}}{114 \text{ kg/kmol}} = \mathbf{45{,}870 \text{ kJ/kg fuel}}$$

15-122 Octane is burned with 100% excess air. The maximum work that can be produced 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 combustion equation is

$$C_8H_{18} + 25(O_2 + 3.76N_2) \longrightarrow 8CO_2 + 9H_2O + 12.5O_2 + 94N_2$$



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,

$$\begin{aligned} W_{\text{rev}} &= \sum N_R \bar{g}^\circ_{f,R} - \sum N_P \bar{g}^\circ_{f,P} \\ &= (1)(16{,}530) - (8)(-394{,}360) - (9)(-228{,}590) \\ &= 5{,}228{,}720 \text{ kJ} \quad (\text{per kmol of fuel}) \end{aligned}$$

since the g_f° of stable elements at 25°C and 1 atm is zero. Per unit mass basis,

$$W_{\text{rev}} = \frac{5{,}228{,}720 \text{ kJ/kmol}}{114 \text{ kg/kmol}} = \mathbf{45{,}870 \text{ kJ/kg fuel}}$$

15-123E Methane 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 is

$$\mathrm{CH_4} + 4(\mathrm{O_2} + 3.76\mathrm{N_2}) \longrightarrow 0.90\mathrm{CO_2} + 0.10\mathrm{CO} + 2\mathrm{H_2O} + x\mathrm{O_2} + 15.04\mathrm{N_2}$$

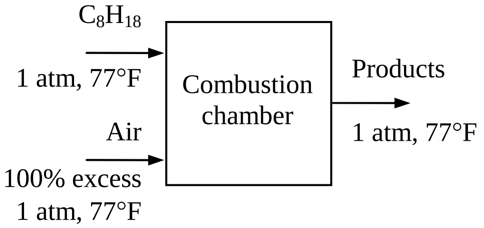
The coefficient for O_2 is determined from its mass balance as

$$4 = 0.9 + (0.5)(0.10) + 1 + x \longrightarrow x = 2.05$$

Substituting,

$$\mathrm{CH_4} + 4(\mathrm{O_2} + 3.76\mathrm{N_2}) \longrightarrow 0.90\mathrm{CO_2} + 0.10\mathrm{CO} + 2\mathrm{H_2O} + 2.05\mathrm{O_2} + 15.04\mathrm{N_2}$$

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,



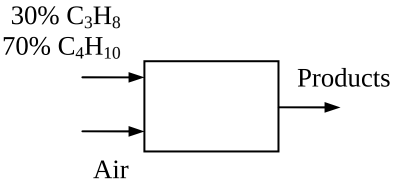
$$\begin{aligned} W_{\text{rev}} &= \sum N_R \bar{g}^\circ_{f,R} - \sum N_P \bar{g}^\circ_{f,P} \\ &= (1)(-21{,}860) - (0.9)(-169{,}680) - (0.10)(-59{,}010) - (2)(-98{,}350) \\ &= 333{,}453\ \text{Btu} \quad \text{(per lbmol of fuel)} \end{aligned}$$

since the g°_f of stable elements at 77°F and 1 atm is zero. Per unit mass basis,

$$W_{\text{rev}} = \frac{333{,}453\ \text{Btu/lbmol}}{16\ \text{lbm/lbmol}} = \mathbf{20{,}840\ Btu/lbm\ fuel}$$

15-124 A gaseous fuel mixture of 30% propane, C_3H_8, and 70% butane, C_4H_{10}, on a volume basis is burned with an air-fuel ratio of 20. 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

$$0.3\,C_3H_8 + 0.7\,C_4H_{10} + a_{th}[O_2 + 3.76N_2] \longrightarrow B\,CO_2 + D\,H_2O + F\,N_2$$

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: $B = 3\times 0.3 + 4\times 0.7 = 3.7$

Hydrogen balance: $2D = 8\times 0.3 + 10\times 0.7 = 2D \longrightarrow D = 4.7$

Oxygen balance: $2a_{th} = 2B + D \longrightarrow 2a_{th} = 2\times 3.7 + 4.7 \longrightarrow a_{th} = 6.05$

Nitrogen balance: $3.76a_{th} = F \longrightarrow 3.76\times 6.05 = F \longrightarrow F = 22.75$

Then, we write the balanced theoretical reaction equation as

$$0.3\,C_3H_8 + 0.7\,C_4H_{10} + 6.05[O_2 + 3.76N_2] \longrightarrow 3.7\,CO_2 + 4.7\,H_2O + 22.75\,N_2$$

The air-fuel ratio for the theoretical reaction is determined from

$$\mathrm{AF_{th}} = \frac{m_{\mathrm{air}}}{m_{\mathrm{fuel}}} = \frac{(6.05\times 4.75\ \mathrm{kmol})(29\ \mathrm{kg/kmol})}{(0.3\times 44 + 0.7\times 58)\ \mathrm{kg}} = 15.47\ \mathrm{kg\ air/kg\ fuel}$$

The percent theoretical air is

$$\mathrm{PercentTH_{air}} = \frac{\mathrm{AF_{actual}}}{\mathrm{AF_{th}}} = \frac{20}{15.47}\times 100 = 129.3\%$$

The moles of nitrogen supplied is

$$N_{N2} = \frac{\mathrm{PercentTH_{air}}}{100}\times a_{th}\times 3.76 = \frac{129.3}{100}(6.05)(3.76) = \mathbf{29.41\ kmol}\ \text{per kmol fuel}$$

The moles of water formed in the combustion process is

$$N_{H2O} = D = \mathbf{4.7\ kmol}\ \text{per kmol fuel}$$

The moles of oxygen in the product gases is

$$N_{O2} = \left(\frac{\mathrm{PercentTH_{air}}}{100} - 1\right)a_{th} = \left(\frac{129.3}{100} - 1\right)(6.05) = \mathbf{1.77\ kmol}\ \text{per kmol fuel}$$

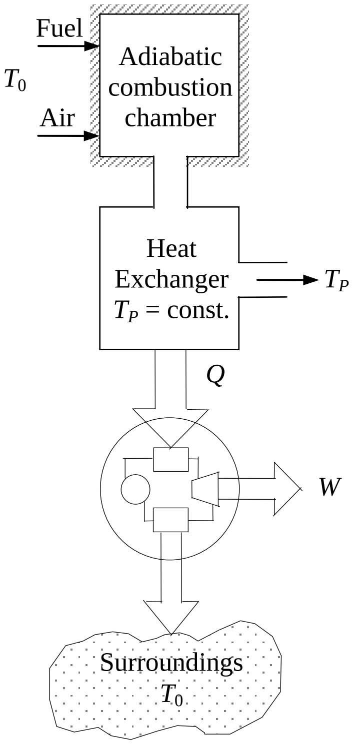
15-125 It is to be shown that the work output of the Carnot engine will be maximum when $T_p = \sqrt{T_0 T_{af}}$. It is also to be shown that the maximum work output of the Carnot engine in this case becomes

$$w = CT_{\text{af}}\left(1 - \frac{\sqrt{T_0}}{\sqrt{T_{\text{af}}}}\right)^2 .$$

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

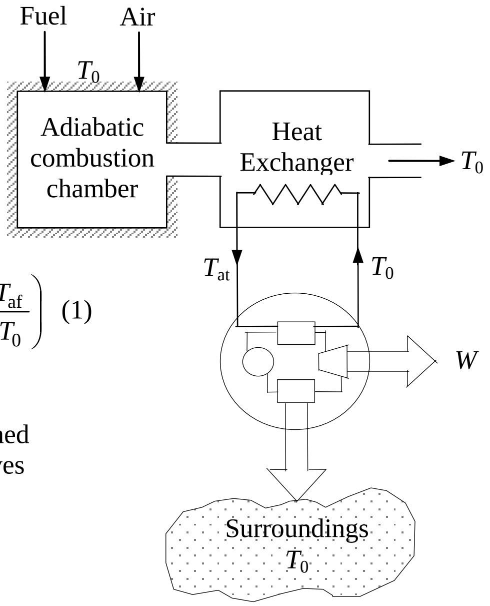
$$w = CT_{\text{af}}\left(1 - \frac{\sqrt{T_0}}{\sqrt{T_{\text{af}}}}\right)^2$$

which is the desired relation.

15-126 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

$$\dot{W}_{\text{rev}} = \dot{m}CT_0\left(\frac{T_{\text{af}}}{T_0} - 1 - \ln\frac{T_{\text{af}}}{T_0}\right) \quad \text{or} \quad w_{\text{rev}} = CT_0\left(\frac{T_{\text{af}}}{T_0} - 1 - \ln\frac{T_{\text{af}}}{T_0}\right) \qquad (1)$$

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

$$\dot{W} = \dot{Q}_H \eta_{\text{th,Carnot}} = \dot{m}C(T_{\text{af}} - T_0)\left(1 - \frac{T_0}{T_e}\right) \qquad (2)$$

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.

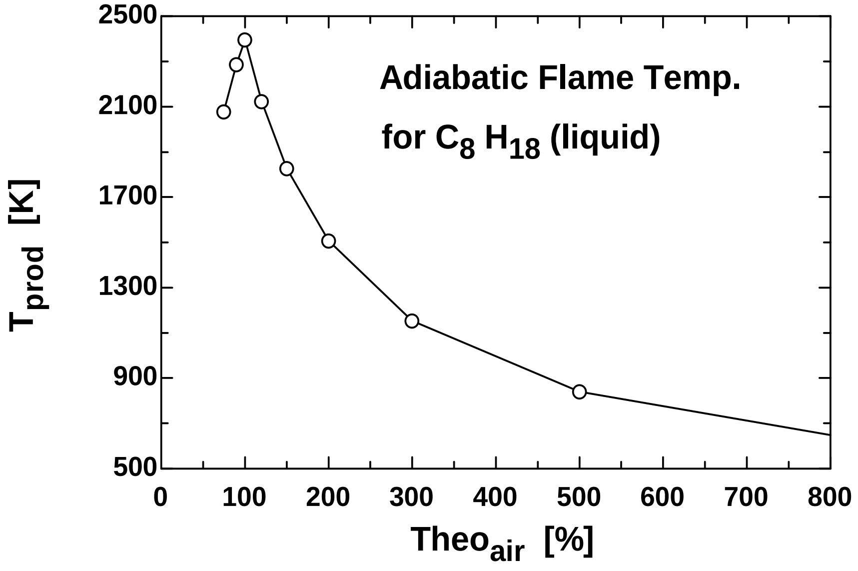
15-127 EES The effect of the amount of air on the adiabatic flame temperature of liquid octane (C_8H_{18}) is to be investigated.

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



15-128 EES 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 = 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
```

SOLUTION for the sample calculation

```
A_th=5                          fuel$='C3H8(l)'          HP=-119067 [kJ/kg]
HR=-119067 [kJ/kg]              h_fuel=-118910           Moles_CO=0.000
Moles_CO2=3.000                 Moles_H2O=4              Moles_N2=28.200
Moles_O2=2.500                  MolO2=2.5                Name$='propane(liq)'
SolMeth$='>= 100%, the solution assumes complete combustion.'
Theo_air=150 [%]                Th_air=1.500             T_air=298 [K]
T_fuel=298 [K]                  T_prod=1820 [K]          w=0
x=3                             y=8                      z=0
```

15-129 EES 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
```

15-130 EES 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
```

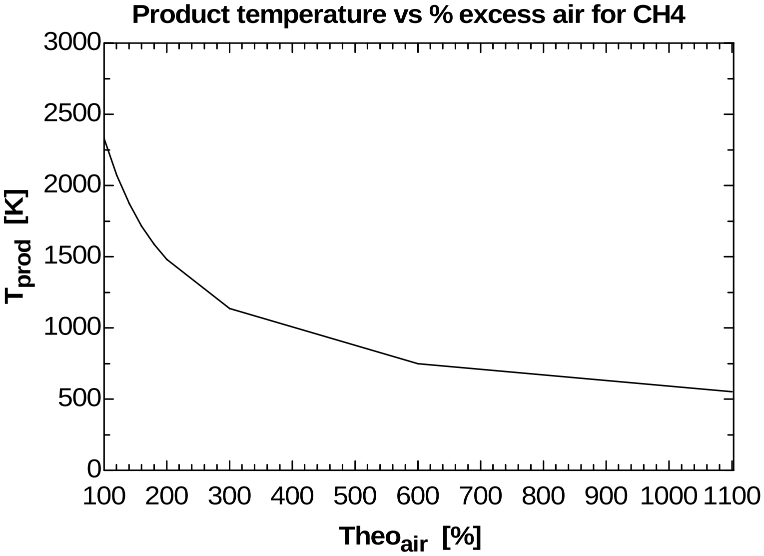
15-131 EES The adiabatic flame temperature of CH_4(g) is to be determined when both the fuel and the air enter the combustion chamber at 25°C for the cases of 0, 20, 40, 60, 80, 100, 200, 500, and 1000 percent excess air.

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



15-132 EES 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

15-133 A fuel is burned with 90 percent theoretical air. This is equivalent to

(a) 10% excess air
(b) 90% excess air
(c) 10% deficiency of air
(d) 90% deficiency of air
(e) stoichiometric amount of air

Answer (c) 10% deficiency of air

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=0.9
"air_th=air_access+1"
air_th=1-air_deficiency
```

15-134 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
```

15-135 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 2 kmol of free O_2 in the products, the air-fuel mass ratio is

(a) 34.3	(b) 17.2	(c) 19.0	(d) 14.9	(e) 12.1

Answer (a) 34.3

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=2  "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"
```

15-136 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.

15-137 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"
```

15-138 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"
```

15-139 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"
```

15-140 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"
```

15-141 Benzene gas (C_6H_6) is burned with 90 percent theoretical air during a steady-flow combustion process. The mole fraction of the CO in the products is

(a) 1.6% (b) 4.4% (c) 2.5% (d) 10.0% (e) 16.7%

Answer (b) 4.4%

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=6
n_H=6
a_th=n_C+n_H/4
coeff=0.90  "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"
```

15-142 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"
```

15-143 ··· 15-147 Design and Essay Problems

15-143 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

$$\text{C}_2\text{H}_5\text{OH}(\ell) + a_{\text{th}}(\text{O}_2 + 3.76\text{N}_2) \longrightarrow 2\text{CO}_2 + 3\text{H}_2\text{O} + 3.76a_{\text{th}}\text{N}_2$$

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,

$$\text{C}_2\text{H}_5\text{OH}(\ell) + 3(\text{O}_2 + 3.76\text{N}_2) \longrightarrow 2\text{CO}_2 + 3\text{H}_2\text{O} + 11.28\text{N}_2$$

Noting that 10.222 kmol of liquid water accompanies each kmol of ethanol, the actual combustion equation can be written as

$$\text{C}_2\text{H}_5\text{OH}(\ell) + 3(\text{O}_2 + 3.76\text{N}_2) + 10.222\text{H}_2\text{O}(\ell) \longrightarrow 2\text{CO}_2 + 3\text{H}_2\text{O}(\text{g}) + 11.28\text{N}_2 + 10.222\text{H}_2\text{O}(l)$$

The heat transfer for this combustion process is determined from the steady-flow energy balance equation with $W = 0$,

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

Assuming the air and the combustion products to be ideal gases, we have $h = h(T)$. 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	$\overline{\mathbf{h}}_\mathbf{f}^\circ$ **kJ/kmol**	$\overline{\mathbf{h}}_{\mathbf{298\ K}}$ **kJ/kmol**	$\overline{\mathbf{h}}_{\mathbf{1400\ K}}$ **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.

* * *