15.26
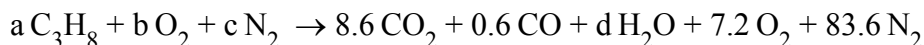
Liquid propane is burned with dry air. A volumetric analysis of the products of combustion yields the following volume percent composition on a dry basis: 8.6% CO_2, 0.6% CO, 7.2% O_2 and 83.6% N_2. Determine the percent of theoretical air used in this combustion process.

$$a\,C_3H_8 + b\,O_2 + c\,N_2 \rightarrow 8.6\,CO_2 + 0.6\,CO + d\,H_2O + 7.2\,O_2 + 83.6\,N_2$$

C balance:　　$3a = 8.6 + 0.6 = 9.2 \Rightarrow a = 3.067$

H_2 balance:　$4a = d \Rightarrow d = 12.267$

N_2 balance:　$c = 83.6$

O_2 balance:　$b = 8.6 + \dfrac{0.6}{2} + \dfrac{12.267}{2} + 7.2 = 22.234$

Air-Fuel ratio $= \dfrac{22.234 + 83.6}{3.067} = 34.51$

Theoretical:

$$C_3H_8 + 5\,O_2 + 18.8\,N_2 \rightarrow 3\,CO_2 + 4\,H_2O + 18.8\,N_2$$
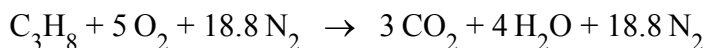
$$\Rightarrow \text{theo. A-F ratio} = \dfrac{5 + 18.8}{1} = 23.8$$

$$\% \text{ theoretical air} = \dfrac{34.51}{23.8} \times 100\,\% = \mathbf{145\,\%}$$

15.31

Butane is burned with dry air at 40°C, 100 kPa with AF = 26 on a mass basis. For complete combustion find the equivalence ratio, % theoretical air and the dew point of the products. How much water (kg/kg fuel) is condensed out, if any, when the products are cooled down to ambient temperature?

Solution:

$$C_4H_{10} + \nu_{O_2}\{O_2 + 3.76\ N_2\} \rightarrow 4\ CO_2 + 5\ H_2O + 3.76\ \nu_{O_2}N_2$$
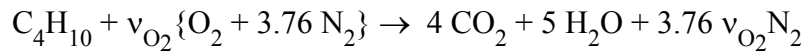
Stoichiometric $\quad \nu_{O_2\ S} = 4 + 5/2 = 6.5; \quad\quad 3.76\ \nu_{O_2} = 24.44$

$$(A/F)_S = 6.5(31.999 + 3.76 \times 28.013)/58.124 = 15.3574$$

Actual: $\quad\quad \nu_{O_2ac} = \dfrac{(A/F)_{ac}}{(A/F)_s}\ \nu_{O_2\ S} = \dfrac{26}{15.3574}\ 6.5 = 11$

$$\% \text{ Theoretical air} = \dfrac{26}{15.3574}\ 100 = \mathbf{169.3\%}$$

$$\text{Equivalence ratio} \quad \Phi = 1/1.693 = \mathbf{0.59}$$

Actual products: $\quad\quad 4\ CO_2 + 5\ H_2O + 4.5\ O_2 + 41.36\ N_2$

The water partial pressure becomes

$$P_v = y_v\ P_{tot} = \dfrac{5}{4 + 5 + 4.5 + 41.36}\ 100 = 9.114 \text{ kPa}$$

$$T_{dew} = \mathbf{43.85°C}$$

$$P_{g\ 40} = 7.348 \text{ kPa} \quad \Rightarrow \quad y_{v\ max} = \dfrac{7.384}{100} = \dfrac{\nu_{H2O}}{4 + \nu_{H2O} + 4.5 + 41.36}$$

Solve for $\nu_{H2O\ vap}$:

$\nu_{H2O\ vap} = 3.975 \quad$ still vapor,

$\nu_{H2O\ LIQ} = 5 - 3.975 = 1.025 \quad$ is liquid

$$\dfrac{m_{H2O\ LIQ}}{m_{Fuel}} = \dfrac{1.025 \times 18.015}{58.124} = \mathbf{0.318}$$

15.69

Natural gas, we assume methane, is burned with 200% theoretical air and the reactants are supplied as gases at the reference temperature and pressure. The products are flowing through a heat exchanger where they give off energy to some water flowing in at 20°C, 500 kPa and out at 700°C, 500 kPa. The products exit at 400 K to the chimney. How much energy per kmole fuel can the products deliver and how many kg water per kg fuel can they heat?

The reaction equation for stoichiometric mixture is:

$$CH_4 + v_{O2} (O_2 + 3.76\ N_2) \rightarrow 2\ H_2O + 1\ CO_2 + c\ N_2$$

O balance: $\quad 2\ v_{O2} = 2 + 2 \quad => \quad v_{O2} = 2$
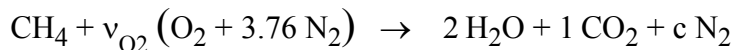
200% theoretical air: $\quad v_{O2} = 2 \times 2 = 4 \quad$ so now more O_2 and N_2

$$CH_4 + 4 (O_2 + 3.76\ N_2) \rightarrow 2\ H_2O + 1\ CO_2 + 15.04\ N_2 + 2\ O_2$$
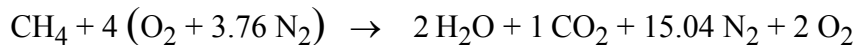
The products are cooled to 400 K (so we do not consider condensation) and the energy equation is

Energy Eq.: $\quad H_R + Q = H_P = H_P^\circ + \Delta H_P = H_R^\circ + Q$

$$Q = H_P^\circ - H_R^\circ + \Delta H_P = H_{RP}^\circ + \Delta H_P$$

From Table 15.3: $\quad H_{RP}^\circ = 16.04\ (-50\ 010) = -802\ 160$ kJ/kmol

$$\Delta H_P = \Delta \bar{h}_{CO2}^* + 2\ \Delta \bar{h}_{H2O}^* + 2\ \Delta \bar{h}_{O2}^* + 15.04\ \Delta \bar{h}_{N2}^*$$

From Table A.9

$\quad \Delta H_{P\ 400} = 4003 + 2 \times 3450 + 2 \times 3027 + 15.04 \times 2971 = 61\ 641$ kJ/kmol

$\quad Q = H_{RP}^\circ + \Delta H_P = -802\ 160 + 61\ 641 = \textbf{-740 519 kJ/kmol}$

$\quad q_{prod} = -Q / M = 740\ 519 / 16.04 = 46\ 167$ kJ/kg fuel

The water flow has a required heat transfer, using B.1.3 and B.1.4 as

$\quad q_{H2O} = h_{out} - h_{in} = 3925.97 - 83.81 = 3842.2$ kJ/kg water

The mass of water becomes

$\quad m_{H2O} / m_{fuel} = q_{prod} / q_{H2O} = \textbf{12.0 kg water / kg fuel}$

15.71

Gasoline, C_7H_{17}, is burned in a steady state burner with stoichiometric air at P_o, T_o. The gasoline is flowing as a liquid at T_o to a carburetor where it is mixed with air to produce a fuel air gas mixture at T_o. The carburetor takes some heat transfer from the hot products to do the heating. After the combustion the products go through a heat exchanger, which they leave at 600 K. The gasoline consumption is 10 kg per hour. How much power is given out in the heat exchanger and how much power does the carburetor need?

Stoichiometric combustion:

$$C_7H_{17} + v_{O_2} \left(O_2 + 3.76\,N_2\right) \rightarrow 8.5\,H_2O + 7\,CO_2 + c\,N_2$$
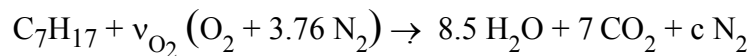
O balance: $\quad 2\,v_{O_2} = 8.5 + 14 = 22.5 \quad \Rightarrow \quad v_{O_2} = 11.25$

N balance: $\quad c = 3.76\,v_{O_2} = 3.76 \times 11.25 = 42.3$

$$M_{FUEL} = 7\,M_C + 17\,M_H = 7 \times 12.011 + 8.5 \times 2.016 = 101.213$$

C.V. Total, heat exchanger and carburetor included, Q out.

Energy Eq.: $\quad H_R = \overset{\circ}{H_R} = \overset{\circ}{H_P} + \Delta H_P + Q_{out}$

From Table A.9

$\quad \Delta H_P = 8.5 \times 10\,499 + 7 \times 12\,906 + 42.3 \times 8894 = 555\,800$ kJ/kmol

From energy equation and Table 15.3

$$Q_{out} = \overset{\circ}{H_R} - \overset{\circ}{H_P} - \Delta H_P = -\overset{\circ}{H_{RP}} - \Delta H_P$$

$$= 101.213\,(44\,506) - 555\,800 = 3\,948\,786 \text{ kJ/kmol}$$

Now the power output is

$$\dot{Q} = \dot{n}\,Q_{out} = Q_{out}\,\dot{m}/M = 3\,948\,786 \times \frac{10}{3600} / 101.213 = \mathbf{108.4\ kW}$$

The carburetor air comes in and leaves at the same T so no change in energy, all we need is to evaporate the fuel, h_{fg} so
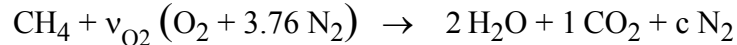
$$\dot{Q} = \dot{m}\,h_{fg} = \frac{10}{3600}\,(44\,886 - 44\,506) = \frac{1}{360} \times 380 = \mathbf{1.06\ kW}$$

Here we used Table 15.3 for fuel liquid and fuel vapor to get h_{fg} and any phase of the water as long as it is the same for the two.
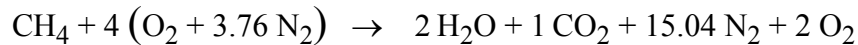
15.79

A gas turbine burns methane with 200% theoretical air. The air and fuel comes in through two separate compressors bringing them from 100 kPa, 298 K to 1400 kPa and after mixing enters the combustion chamber at 600 K. Find the adiabatic flame temperature using constant specific heat for the ΔH_P terms.

The reaction equation for stoichiometric mixture is:

$$CH_4 + v_{O2}\left(O_2 + 3.76\ N_2\right) \rightarrow 2\ H_2O + 1\ CO_2 + c\ N_2$$

O balance: $\quad 2\ v_{O2} = 2 + 2 \quad => \quad v_{O2} = 2$

200% theoretical air: $\quad v_{O2} = 2 \times 2 = 4 \quad$ so now more O_2 and N_2

$$CH_4 + 4\left(O_2 + 3.76\ N_2\right) \rightarrow 2\ H_2O + 1\ CO_2 + 15.04\ N_2 + 2\ O_2$$

The energy equation around the combustion chamber becomes

Energy Eq.: $\qquad H_P - H_R = 0 \quad => \quad \Delta H_P = \overset{\circ}{H_R} + \Delta H_R - \overset{\circ}{H_P} = -\overset{\circ}{H_{RP}} + \Delta H_R$

$\Delta H_R = \Delta H_{Fuel} + \Delta H_{air} = M\ C_P\ \Delta T + 4(\Delta \bar{h}_{O2} + 3.76\ \Delta \bar{h}_{N2})$

$\qquad = 16.043 \times 2.254\ (600{-}298) + 4(9245 + 3.76 \times 8894) = 181\ 666$ kJ/kmol

$-\overset{\circ}{H_{RP}} = 16.043 \times 50\ 010 = 802\ 310$ kJ/kmol

$\Delta H_P = \Delta \bar{h}_{CO2} + 2\ \Delta \bar{h}_{H2O} + 15.04\ \Delta \bar{h}_{N2} + 2\ \Delta \bar{h}_{O2} \approx \Delta T \sum v_i \bar{C}_{Pi}$

$\qquad = 802\ 310 + 181\ 666 = 983\ 976$ kJ/kmol   (from energy Eq.)

$\sum v_i \bar{C}_{Pi} = 0.842 \times 44.01 + 2 \times 1.872 \times 18.015 + 15.04 \times 1.042 \times 28.013$

$\qquad\qquad + 2 \times 0.922 \times 31.999 = 602.52$ kJ/kmol-K

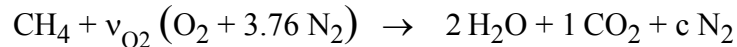$\Delta T = \Delta H_P / \sum v_i \bar{C}_{Pi} = 983\ 976 / 602.52 = 1633.1$ K
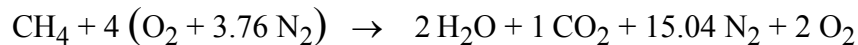
$T = 298 + 1633 = \textbf{1931 K}$

15.80

Extend the solution to the previous problem by using Table A.9 for the ΔH_P terms.

The reaction equation for stoichiometric mixture is:

$$CH_4 + v_{O2}\left(O_2 + 3.76\,N_2\right) \rightarrow 2\,H_2O + 1\,CO_2 + c\,N_2$$

O balance: $2\,v_{O2} = 2 + 2 \Rightarrow v_{O2} = 2$

200% theoretical air: $v_{O2} = 2 \times 2 = 4$ so now more O_2 and N_2

$$CH_4 + 4\left(O_2 + 3.76\,N_2\right) \rightarrow 2\,H_2O + 1\,CO_2 + 15.04\,N_2 + 2\,O_2$$

The energy equation around the combustion chamber becomes

Energy Eq.: $H_P - H_R = 0 \Rightarrow \Delta H_P = \overset{\circ}{H}_R + \Delta H_R - \overset{\circ}{H}_P = -\overset{\circ}{H}_{RP} + \Delta H_R$

$\Delta H_R = \Delta H_{Fuel} + \Delta H_{air} = M\,C_P\,\Delta T + 4(\Delta\bar{h}_{O2} + 3.76\,\Delta\bar{h}_{N2})$

$= 16.043 \times 2.254\,(600{-}298) + 4(9245 + 3.76 \times 8894) = 181\,666$ kJ/kmol

$-\overset{\circ}{H}_{RP} = 16.043 \times 50\,010 = 802\,310$ kJ/kmol

$\Delta H_P = \Delta\bar{h}_{CO2} + 2\,\Delta\bar{h}_{H2O} + 15.04\,\Delta\bar{h}_{N2} + 2\,\Delta\bar{h}_{O2}$

$= 802\,310 + 181\,666 = 983\,976$ kJ/kmol   (from energy Eq.)

Trial and error with $\Delta\bar{h}$ from Table A.9

At 1800 K
$\Delta H_P = 79\,432 + 2 \times 62\,693 + 15.04 \times 48\,979 + 2 \times 51\,674 = 1\,044\,810$ kJ/kmol
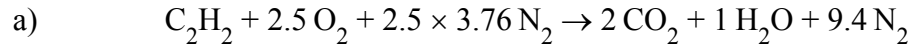At 1700 K
$\Delta H_P = 73\,480 + 2 \times 57\,757 + 15.04 \times 45\,430 + 2 \times 47\,959 = 968\,179$ kJ/kmol

Linear interpolation: $T = 1700 + 100\,\dfrac{983\,976 - 968\,179}{1\,044\,810 - 968\,179} = \mathbf{1721\ K}$

15.83

Acetylene gas at 25°C, 100 kPa is fed to the head of a cutting torch. Calculate the adiabatic flame temperature if the acetylene is burned with

a.   100% theoretical air at 25°C.

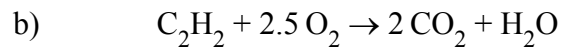b.   100% theoretical oxygen at 25°C.

a) $C_2H_2 + 2.5\,O_2 + 2.5 \times 3.76\,N_2 \rightarrow 2\,CO_2 + 1\,H_2O + 9.4\,N_2$

$H_R = \bar{h}^o_{f\,C2H2} = +226\,731$ kJ/kmol   from table A.10

$H_P = 2(-393\,522 + \Delta\bar{h}^*_{CO2}) + 1(-241\,826 + \Delta\bar{h}^*_{H2O}) + 9.4\,\Delta\bar{h}^*_{N2}$

$Q_{CV} = H_P - H_R = 0 \quad \Rightarrow \quad 2\,\Delta\bar{h}^*_{CO2} + 1\,\Delta\bar{h}^*_{H2O} + 9.4\,\Delta\bar{h}^*_{N2} = 1\,255\,601$ kJ

Trial and Error A.9: $LHS_{2800} = 1\,198\,369,\quad LHS_{3000} = 1\,303\,775$

Linear interpolation: $T_{PROD} = \textbf{2909 K}$

b) $C_2H_2 + 2.5\,O_2 \rightarrow 2\,CO_2 + H_2O$

$H_R = +226\,731$ kJ ; $H_P = 2(-393\,522 + \Delta\bar{h}^*_{CO2}) + 1(-241\,826 + \Delta\bar{h}^*_{H2O})$

$\Rightarrow \quad 2\,\Delta\bar{h}^*_{CO2} + 1\,\Delta\bar{h}^*_{H2O} = 1\,255\,601$ kJ/kmol fuel
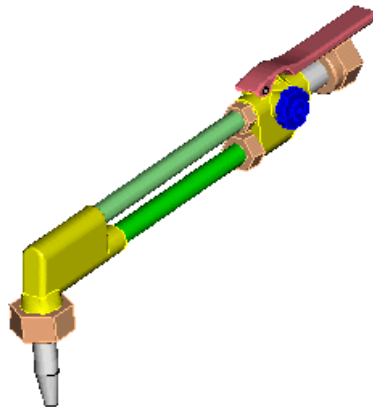
At 6000 K (limit of A.9) $2 \times 343\,782 + 302\,295 = 989\,859$

At 5600 K $2 \times 317\,870 + 278\,161 = 913\,901$

Slope   75 958/400 K change

Extrapolate to cover the difference above 989 859 kJ/kmol fuel

$T_{PROD} \approx 6000 + 400(265\,742/75\,958) \approx \textbf{7400 K}$

15.92

Consider the combustion of methanol, CH_3OH, with 25% excess air. The combustion products are passed through a heat exchanger and exit at 200 kPa, 400 K. Calculate the absolute entropy of the products exiting the heat exchanger assuming all the water is vapor.

$$CH_3OH + 1.25 \times 1.5\,(O_2 + 3.76\,N_2) \rightarrow CO_2 + 2\,H_2O + 0.375\,O_2 + 7.05\,N_2$$
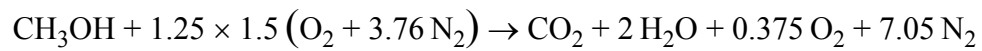
We need to find the mole fractions to do the partial pressures,

$$n = 1 + 2 + 0.375 + 7.05 = 10.425 \quad => \quad y_i = n_i\,/\,n$$

Gas mixture:

|	n_i	y_i	$\bar{s}_i^\circ$	$-\bar{R}\ln\dfrac{y_i P}{P_0}$	$\bar{S}_i$
CO_2	1.0	0.0959	225.314	+13.730	239.044
H_2O	2	0.1918	198.787	+7.967	206.754
O_2	0.375	0.0360	213.873	+20.876	234.749
N_2	7.05	0.6763	200.181	-2.511	197.670

$$S_{GAS\ MIX} = \sum n_i \bar{S}_i = \textbf{2134.5 kJ/K kmol fuel}$$

15.94

Propene, C_3H_6, is burned with air in a steady flow burner with reactants at P_o, T_o. The mixture is lean so the adiabatic flame temperature is 1800 K. Find the entropy generation per kmol fuel neglecting all the partial pressure corrections.

The reaction equation for a mixture with excess air is:

$$C_3H_6 + \nu_{O2}\left(O_2 + 3.76\,N_2\right) \rightarrow 3\,H_2O + 3\,CO_2 + 3.76\nu_{O2}\,N_2 + (\nu_{O2} - 4.5)O_2$$
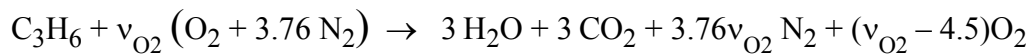
Energy Eq.: $\qquad\qquad H_R = \overset{\circ}{H}_R + \Delta H_R = \overset{\circ}{H}_R = H_P = \overset{\circ}{H}_P + \Delta H_P$

The entropy equation: $\qquad S_R + S_{gen} = S_P \quad => \quad S_{gen} = S_P - S_R = S_P - \overset{\circ}{S}_R$

From table A.9 at reference T

$$\Delta H_R = \Delta h_{Fu} + \nu_{O2}(\Delta h_{O2} + 3.76\,\Delta h_{N2}) = 0$$

From table A.9 at 1800 K:

$$\Delta H_P = 3\,\Delta h_{H2O} + 3\,\Delta h_{CO2} + 3.76\,\nu_{O2}\,\Delta h_{N2} + (\nu_{O2} - 4.5)\,\Delta h_{O2}$$

$$= 3 \times 62\,693 + 3 \times 79432 + 3.76\,\nu_{O2} \times 48\,979 + (\nu_{O2} - 4.5)\,51\,674$$

$$= 193\,842 + 235\,835\,\nu_{O2}$$

From table 15.3: $\quad \overset{\circ}{H}_P - \overset{\circ}{H}_R = \overset{\circ}{H}_{RP} = 42.081(-45\,780) = -1\,926\,468$ kJ/kmol

Now substitute all terms into the energy equation

$$-1\,926\,468 + 193\,842 + 235\,835\,\nu_{O2} = 0$$

Solve for ν_{O2}: $\qquad \nu_{O2} = \dfrac{1\,926\,468 - 193\,842}{235\,835} = 7.3468, \qquad \nu_{N2} = 27.624$

Table A.9-10 contains the entropies at 100 kPa so we get:

$$S_P = 3 \times 259.452 + 3 \times 302.969 + (7.3468 - 4.5)\,264.797 + 27.624 \times 248.304$$

$$= 9300.24 \text{ kJ/kmol-K}$$

$$S_R = 267.066 + 7.3468 \times 205.148 + 27.624 \times 191.609 = 7067.25 \text{ kJ/kmol K}$$

$$S_{gen} = 9300.24 - 7067.25 = \mathbf{2233 \text{ kJ/kmol-K}}$$