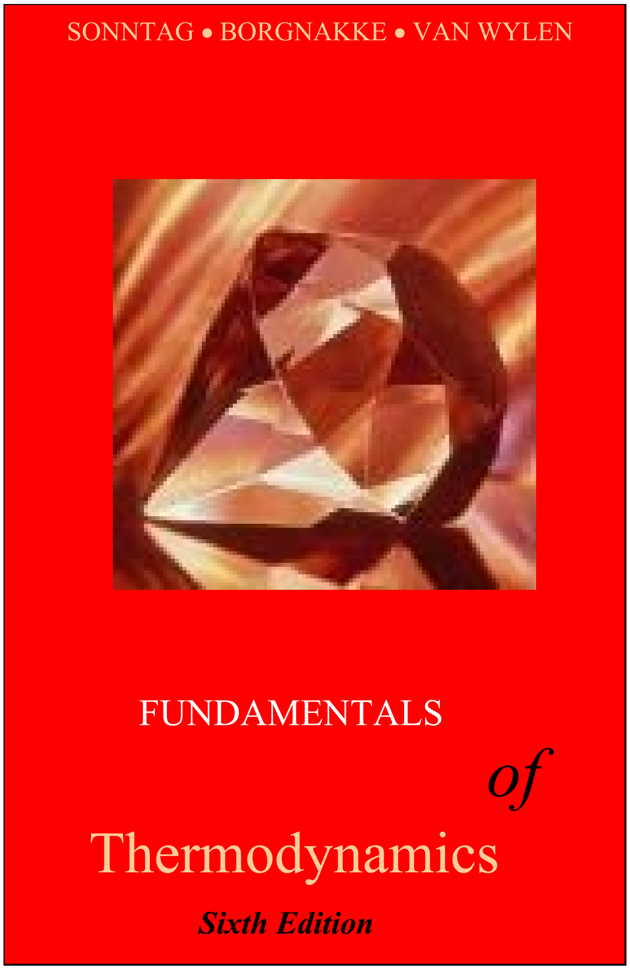
SOLUTION MANUAL
SI UNIT PROBLEMS
CHAPTER 14



SONNTAG • BORGNAKKE • VAN WYLEN

FUNDAMENTALS

of

Thermodynamics

Sixth Edition

CONTENT

Correspondence List

CHAPTER 14 6th edition Sonntag/Borgnakke/Wylen

The correspondence between the new problem set and the previous 5th edition chapter 14 problem set.

The concepts study guide problems 14.1-14.20 are all new

New	**5th**	**New**	**5th**	**New**	**5th**	**New**	**5th**
21	1	51	16	81	37	111	63
22	3	52	30	82	38	112	23
23	new	53	41	83	39	113	new
24	2mod	54	68	84	new	114	36
25	4	55	18	85	48	115	40
26	5	56	new	86	52	116	49
27	6	57	22mod	87	new	117	55
28	new	58	42	88	new	118	71
29	7	59	43	89	53	119	73
30	8	60	24mod	90	new	120	70
31	9	61	44	91	54	121	72
32	new	62	45	92	57		
33	10	63	25mod	93	58		
34	11	64	new	94	60		
35	12	65	46	95	new		
36	13	66	47	96	69		
37	new	67	new	97	21		
38	14	68	new	98	27		
39	15	69	50	99	33		
40	51	70	new	100	56		
41	19	71	new	101	66		
42	20	72	32	102	67		
43	22	73	new	103	74		
44	17	74	34	104	64		
45	24	75	new	105	new		
46	new	76	new	106	65a		
47	25	77	35	107	65mod		
48	26	78	31	108	59		
49	28	79	new	109	61		
50	29	80	new	110	62		

The English unit problems are:

New	**5th**	**SI**	**New**	**5th**	**SI**	**New**	**5th**	**SI**
122	new	11	132	86	52	142	91	83
123	new	13	133	92	53	143	new	87
124	75	34	134	new	64	144	94	93
125	76	39	135	93	66	145	95	94
126	78	41	136	new	71	146	80	97
127	79	42	137	new	73	147	83	98
128	82	45	138	87	78	148	98	104
129	84	49	139	88	77	149	96	110
130	85	50	140	89	81	150	97	111
131	77	51	141	90	82	151	81	112

Concept-Study Guide Problems

14.1

How many kmoles of air are needed to burn 1 kmol of carbon?

Combustion Eq.: $C + O_2 \Rightarrow 1\ CO_2$

One kmol of O_2 is required to form CO_2 . Since air is 21 % O_2 , this means 4.76 kmol of air.

14.2

If I burn 1 kmol of hydrogen H_2 with 6 kmol air what is A/F ratio on a mole basis and what is the percent theoretical air?

Combustion Eq. stoichiometric:

$$H_2 + \nu_{O2}(O_2 + 3.76\ N_2) \Rightarrow 1\ H_2O + 3.76\ \nu_{O2}\ N_2$$

$$\nu_{O2} = 0.5\ ; \qquad (A/F)_S = \nu_{O2} \times (1 + 3.76) / 1 = 2.38$$

Six kmol of air is: $1.26\ O_2 + 4.74\ N_2$.

The A/F mole ratio is 6, so the percent theoretical air is

$$\%\text{Theoretical air} = \frac{(A/F)_{ac}}{(A/F)_S} \times 100 = \frac{6}{2.38} \times 100 = 252\ \%$$

14.3

Why would I sometimes need A/F on a mole basis? on a mass basis?

If you want to meter (measure) the fuel and air flows it can be done as a volume flowrate which is proportional to moles ($P\dot{V} = \dot{n}\bar{R}T$) in which case concentrations on a mole basis are needed.

The fuel and air flows can also be measured with a method that measures mass flow rate $\dot{m}$ or if you are filling up tanks to store the fuel and oxidicer as in a rocket in both cases the concentrations on a mass basis are needed.

14.4

Why is there no significant difference between the number of moles of reactants versus products in combustion of hydrocarbon fuels with air?

In most hydrocarbon fuels, there are approximately twice as many hydrogen atoms as carbon atoms, so the numbers of moles of CO_2 and H_2O in the products are roughly equal, the total of which is not too different from the O_2 required in the reactants. The number of excess O_2 is the same in reactants and products. The total number of moles is dominated by the N_2 in each, especially with excess air.

14.5

For the 110% theoretical air in Eq.14.8 what is the equivalence ratio? Is that mixture rich or lean?

110% Theoretical air means also $AF = 1.1 \, AF_S$

so from the definition in Eq.14.6

$$\Phi = \frac{AF_S}{AF} = \frac{1}{1.10} = 0.909 \qquad \text{a lean mixture}$$

14.6

Why are products measured on a dry basis?

Combustion products have traditionally been measured by passing the gas mixture through a series of solutions that selectively absorb the components one-by-one and measuring the resulting gas volume decreases. The water component is condensed out in these processes, leaving the others – that is, a dry basis. Other and newer instruments measure the concentrations by optical means and these are sensitive to moisture content, which can corrode the surfaces and destroy the sensors. If the water stays in the mixture it typically have to stay hot to prevent condensation at undesirable locations where that would alter the concentrations of the remaining gas mixture components.

14.7

What is the dew point of hydrogen burned with stoichiometric pure oxygen? air?

For H_2 burned with pure oxygen

$$H_2 + 0.5\ O_2 \quad \Rightarrow \quad 1\ H_2O$$

with the only product as water, so the dew-point at 100 kPa pressure is **99.6°C**.
For H_2 burned with stoichiometric air

$$H_2 + 0.5\ (O_2 + 3.76\ N_2) \quad \Rightarrow \quad 1\ H_2O\ + 1.88\ N_2$$

the product mixture is water and nitrogen. The partial pressure of the water at a pressure of 100 kPa is

$$P_v = P_{tot}\ y_v = 100 \times \frac{1}{1 + 1.88} = 34.7 \text{ kPa},$$

corresponding to a dew-point of **72.3°C** .

14.8

How does the dew point change as equivalence ratio goes from 0.9 to 1 to 1.1?

For a given amount of water in the products, the smaller the total number of moles of other gases is (as Φ increases), the higher the partial pressure of the water and therefore the dew-point temperature. As Φ becomes greater than 1.0, there will be incomplete combustion, and the resulting composition will be affected to have some unburned fuel and therefore relative less water. The relative maximum amount of water is then at a stoichiometric mixture $\Phi = 1$ and this is also maximum dew point temperature.

14.9

In most cases combustion products are exhausted above the dew point. Why?

If any water in the products condenses, it will be acidic due to the other gases in the products. There are always minute amounts of unburned or partially burned fuel and intermediate species in the products that can combine with water and create a very corrosive mixture.

14.10

Why does combustion contribute to global warming?

Any combustion of a hydrocarbon fuel produces carbon dioxide, which in the atmosphere is a contributing factor to global warming. Carbon dioxide absorbs radiation over a wide spectrum and thus heats the atmosphere. This is not just man-made, but nature has forest fires and volcanic action that liberates gases into the atmosphere.

14.11

What is the enthalpy of formation for oxygen as O_2? If O? For CO_2?

The enthalpy of formation of O_2 is zero, by choice of the reference base. Relative to this base, the value for the monatomic form O is

$\bar{h}^\circ_{f\,O}$ = +249 170 kJ/kmol (Table A.9),

and the value for CO_2 is

$\bar{h}^\circ_{f\,CO2}$ = –393 522 kJ/kmol (Table A.9 or A.10).

14.12

How is a fuel enthalpy of combustion connected to its enthalpy of formation?

The enthalpy of combustion of a fuel is the difference in enthalpy of the products and reactants for the combustion involving the fuel, these enthalpies include the various enthalpies of formation.

14.13

What is the higher and lower heating value HHV, LHV of n-Butane?

The heating value is the negative of the enthalpy of combustion. From Table 14.3, the HHV of gaseous n-Butane at 25°C is

$$\text{HHV} = 49\,500 \text{ kJ/kg},$$

and the corresponding LHV is

$$\text{LHV} = 45\,714 \text{ kJ/kg}.$$

Notice the table is on a mass basis (per kg fuel).

14.14

What is the value of h_{fg} for n-Octane?

This can be obtained from two places. From Table A.10 we get

$$h_{fg} = (\bar{h}^0_{f\,vap} - \bar{h}^0_{f\,liq}) / M = [(-208\ 600 - (-250\ 105)] / 114.232 = \mathbf{363\ kJ/kg}$$

The h_{fg} of a fuel listed in Table 14.3 is the difference between the first two columns in the table (or the third and fourth). For n-Octane, this value is

$$h_{fg} = -47\ 893 - (-48\ 256) = \mathbf{363\ kJ/kg}$$

To see this remember

$$H^{\circ}_{RP} = H^{\circ}_{P} - H^{\circ}_{R} = H^{\circ}_{P} - \bar{h}^0_{f\ \text{fuel vap or liq}}$$

so when we take the difference between fuel as gas or liquid all other terms will cancel out leaving h_{fg} for the fuel.

14.15

Why do some fuels not have entries for liquid fuel in Table 14.3?

Those fuels cannot exist as liquids at 25°C (above their critical temperature).

14.16

Does it make a difference for the enthalpy of combustion whether I burn with pure oxygen or air? What about the adiabatic flame temperature?

No difference in the enthalpy of combustion – the nitrogen in the air is the same in the reactants and products, and its enthalpy cancels out. The adiabatic flame temperature is much lower for combustion with air, because a significant part of the energy release from the combustion goes into heating the nitrogen (as well as the other products) to the flame temperature.

14.17

What happens to the adiabatic flame temperature if I burn rich or lean?

The higher the percent theoretical air used in combustion (the leaner), the larger the number of moles of products, especially nitrogen, which decreases the adiabatic flame temperature. Burning rich causes incomplete combustion, however, with a smaller release of energy.

Experimentally the highest temperature is reached for slightly rich.

Heavy molecules show up as yellow. Oxygen diffuses in from the air and the fuel evaporates from the wick. As air mixes in, the flame cools.

See the different colors of the flame? They reflect the different temperatures.

14.18

Is the irreversibility in a combustion process significant? Why is that?

A combustion process is highly irreversible with a large increase in entropy. It takes place at a rapid rate, due to large driving forces, and results in stable products of combustion that have little or no tendency to return to their former constituents and states.

14.19

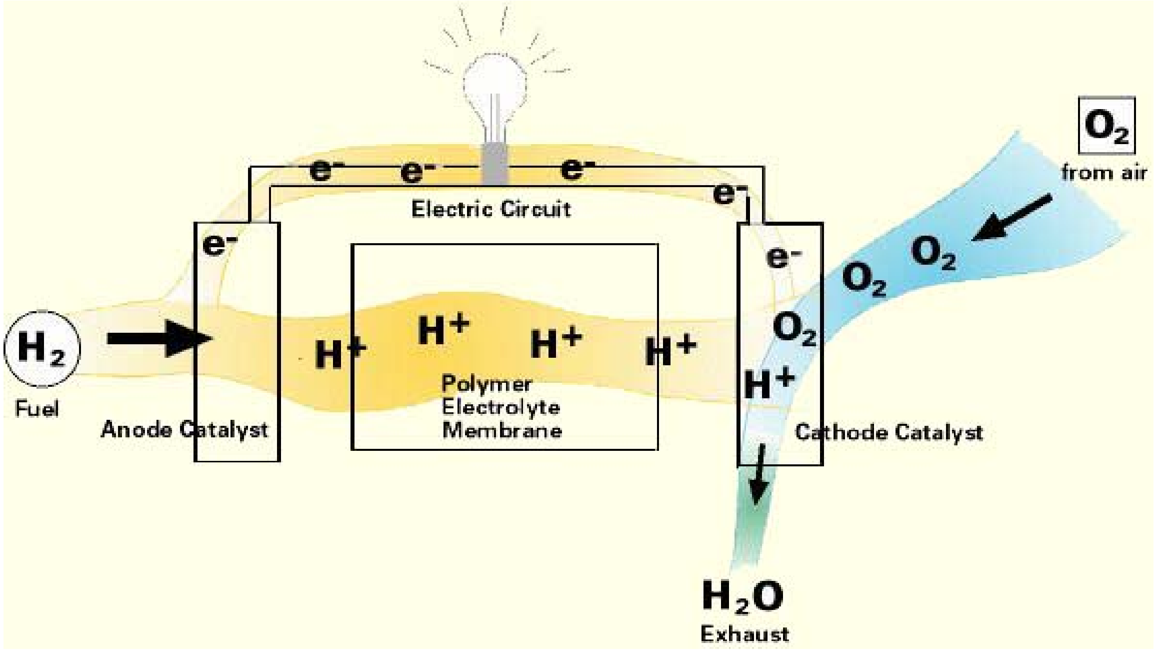
If the A/F ratio is larger than stoichiometric is it more or less reversible?

Less reversible more irreversible. The excess oxydizer (air) is being heated up, Q over a finite temperature difference is an irreversible process. The same is true for A/F smaller than one where the excess fuel is heated up.

14.20

What makes the fuel cell attractive from a power generating point of view?

Fuel cells are attractive for power generation because their direct output is electrical energy. They also have a much higher power density as power per unit volume or power per unit mass and thus can be used in mobile applications.



Fuels and the Combustion Process

14.21

Calculate the theoretical air-fuel ratio on a mass and mole basis for the combustion of ethanol, C_2H_5OH.

Reaction Eq.: $C_2H_5OH + \nu_{O2}(O_2 + 3.76N_2) \Rightarrow aCO_2 + bH_2O + cN_2$

Balance C: $2 = a$ Balance H: $6 = 2b \Rightarrow b = 3$

Balance O: $1 + 2\nu_{O2} = 2a + b = 4 + 3 = 7 \Rightarrow \nu_{O2} = 3$

$(\text{air/fuel})_{\text{mol}} = \nu_{O2}(1 + 3.76)/1 = 3 \times 4.76 = \mathbf{14.28}$

$(\text{air/fuel})_{\text{mass}} = (\nu_{O2}M_{O2} + \nu_{N2}\, M_{N2})/M_{\text{Fuel}}$

$= (3\times31.999 + 11.28\times28.013)/46.069 = \mathbf{8.943}$

14.22

A certain fuel oil has the composition $C_{10}H_{22}$. If this fuel is burned with 150% theoretical air, what is the composition of the products of combustion?

$$C_{10}H_{22} + (1/\phi)\,\nu_{O2}\left(O_2 + 3.76\,N_2\right) \rightarrow a\,H_2O + b\,CO_2 + c\,N_2 + d\,O_2$$

Stoichiometric combustion: $\phi = 1$, $d = 0$,

C balance: $b = 10$

H balance: $a = 22/2 = 11$,

O balance: $2\,\nu_{O_2} = a + 2b = 11 + 20 = 31$ => $\nu_{O2} = 15.5$

Actual case: $1/\phi = 1.5$ => $\nu_{O2} = 1.5 \times 15.5 = 23.25$

H balance: $a = 11$, C balance: $b = 10$,

N balance: $c = 23.25 \times 3.76 = 87.42$

O_2 balance: $d = 23.25 - 10 - 11/2 = 7.75$ (excess oxygen)

14.23

Methane is burned with 200% theoretical air. Find the composition and the dew point of the products.

The reaction equation for stoichiometric mixture is:

$$CH_4 + \nu_{O2}\left(O_2 + 3.76\ N_2\right) \rightarrow a\,H_2O + b\,CO_2 + c\,N_2$$

C balance: $1 = b$; H balance: $4 = 2a$

O balance: $2\ \nu_{O2} = a + 2b = 2 + 2 \times 1$ => $\nu_{O2} = 2$

N_2 balance: $3.76\ \nu_{O2} = c = 7.52$

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

$$CH_4 + \nu_{O2}\left(O_2 + 3.76\ N_2\right) \rightarrow a\,H_2O + b\,CO_2 + c\,N_2 + d\,O_2$$

N_2 balance: $3.76\ \nu_{O2} = c = 15.04$

Extra oxygen: $d = 4 – 1 - 1 = 2$

Products: $2\ H_2O + 1\ CO_2 + 15.04\ N_2 + 2\ O_2$

Water vapor mole fraction: $y_v = \dfrac{2}{1 + 2 + 2 + 15.04} = 0.0998$

Partial water vapor pressure: $P_v = y_v\ P_o = 0.0998 \times 101 = 9.98$ kPa

$$P_g(T_{dew}) = P_v = 9.98 \text{ kPa} \quad \Rightarrow \quad T_{dew} = \mathbf{45.8^oC}$$

14.24

In a combustion process with decane, $C_{10}H_{22}$, and air, the dry product mole fractions are 83.61% N_2, 4.91% O_2, 10.56% CO_2 and 0.92% CO. Find the equivalence ratio and the percent theoretical air of the reactants.

$$x\, C_{10}H_{22} + (1/\phi)\, \nu_{O2} \left(O_2 + 3.76\, N_2\right) \rightarrow a\, H_2O + b\, CO_2 + c\, CO + d\, N_2 + e\, O_2$$

Stoichiometric combustion: $\phi = 1, \quad c = 0, \quad e = 0,$

C balance: $b = 10x$

H balance: $a = 22x/2 = 11x,$

O balance: $2\, \nu_{O2} = a + 2b = 11x + 20x = 31x$

$\nu_{O2} = 15.5x, \quad \nu_{N2} = 58.28x \qquad \Rightarrow \; (A/F)_s = (\nu_{O2} + \nu_{N2})/x = \mathbf{73.78}$

Actual combustion: $d = 83.61 \quad \rightarrow$

N balance: $(1/\phi)\, \nu_{O2} \times 3.76 = 83.61 \quad \rightarrow (1/\phi)\, \nu_{O2} = 22.24$

C balance: $10x = 10.56 + 0.92 = 11.48 \quad \Rightarrow \quad x = 1.148$

$(A/F)_{ac} = (1/\phi)\, \nu_{O2} \times 4.76/1.148 = \mathbf{92.215}$

$\phi = (F/A)_{ac} / (F/A)_s = (A/F)_s / (A/F)_{ac} = 73.78 / 92.215 = \mathbf{0.80}$ **or** $\boldsymbol{\phi = 0.8}$

Percent theoretical air $= 100\, (1/\phi) = \mathbf{125\%}$

14.25

Natural gas B from Table 14.2 is burned with 20% excess air. Determine the composition of the products.

The reaction equation (stoichiometric and complete combustion) with the fuel composition is:

$$60.1\ CH_4 + 14.8\ C_2H_6 + 13.4\ C_3H_8 + 4.2\ C_4H_{10} + 7.5\ N_2$$

$$+ \nu_{O2}\left(O_2 + 3.76\ N_2\right) \;\rightarrow\; a\,H_2O + b\,CO_2 + c\,N_2$$

C balance: $60.1 + 2\times14.8 + 3\times13.4 + 4\times4.2 = b = 146.7$

H balance: $4\times60.1 + 6\times14.8 + 8\times13.4 + 10\times4.2 = 2a = 478.4 \;\Rightarrow\; a = 239.2$

O balance: $2\,\nu_{O2} = a + 2b = 239.2 + 2\times146.7 \;\Rightarrow\; \nu_{O2} = 266.3$

N_2 balance: $7.5 + 3.76\,\nu_{O2} = c = 1008.8$

20% excess air: $\nu_{O2} = 1.2\times266.3 = 319.56$ so now more O_2 and N_2

Extra oxygen: $d = 319.56 - 266.3 = 53.26$, $c = 7.5 + 3.76\times319.56 = 1209$

Products: $239.2\ H_2O + 146.7\ CO_2 + 1209\ N_2 + 53.26\ O_2$

To the expert the color of the flame can tell about the composition. It can also tell about other gases present if they have distinct color emission.

14.26

A Pennsylvania coal contains 74.2% C, 5.1% H, 6.7% O, (dry basis, mass percent) plus ash and small percentages of N and S. This coal is fed into a gasifier along with oxygen and steam, as shown in Fig. P14.26. The exiting product gas composition is measured on a mole basis to: 39.9% CO, 30.8% H_2, 11.4% CO_2, 16.4% H_2O plus small percentages of CH_4, N_2, and H_2S. How many kilograms of coal are required to produce 100 kmol of product gas? How much oxygen and steam are required?

Convert the mass concentrations to number of kmol per 100 kg coal:

C : $n = 74.2/12.01 = 6.178$ H_2: $n = 5.1/2.016 = 2.530$

O_2: $n = 6.7/31.999 = 0.209$

Now the combustion equation reads

$$x(6.178\,C + 2.53\,H_2 + 0.209\,O_2) + y\,H_2O + z\,O_2 \text{ in and}$$

$$39.9\,CO + 30.8\,H_2 + 11.4\,CO_2 + 16.4\,H_2O \text{ out}$$

in 100 kmol of mix out

Now we can do the atom balance to find (x, y, z)

C balance: $6.178\,x = 39.9 + 11.4 \rightarrow x = 8.304$

H_2 balance: $2.53 \times 8.304 + y = 30.8 + 16.4 \rightarrow y = 26.191$

O_2 balance: $0.209 \times 8.304 + \frac{26.191}{2} + z = \frac{39.9}{2} + 11.4 + \frac{16.4}{2} \rightarrow z = 24.719$

Therefore, for 100 kmol of mixture out

require: **830.4 kg** of coal
26.191 kmol of steam
24.719 kmol of oxygen

14.27

Repeat Problem 14.26 for a certain Utah coal that contains, according to the coal analysis, 68.2% C, 4.8% H, 15.7% O on a mass basis. The exiting product gas contains 30.9% CO, 26.7% H_2, 15.9% CO_2 and 25.7% H_2O on a mole basis.

Convert the mass concentrations to number of kmol per 100 kg coal:

C : 68.2/12.01 = 5.679 $\quad$ H_2: 4.8/2.016 = 2.381

O_2: 15.7/32.00 = 0.491

Now the combustion equation reads

$$x(5.679\,C + 2.381\,H_2 + 0.491\,O_2) + y\,H_2O + z\,O_2 \;\text{ in}$$

$$30.9\,CO + 26.7\,H_2 + 15.9\,CO_2 + 25.7\,H_2O \;\text{ out}$$

in 100 kmol of mix out

Now we can do the atom balance to find (x, y, z)

$$C: \quad 5.679x = 30.9 + 15.9 \quad \rightarrow \quad x = 8.241$$

$$H_2: \quad 2.381 \times 8.241 + y = 26.7 + 25.7 \quad \rightarrow \quad y = 32.778$$

$$O_2: \quad 0.491 \times 8.241 + \frac{32.778}{2} + z = \frac{30.9}{2} + 15.9 + \frac{25.7}{2}$$

$$\rightarrow \quad z = 23.765$$

Therefore, for 100 kmol of mixture out,

require: **824.1 kg** of coal

32.778 kmol of steam

23.765 kmol of oxygen

14.28

For complete stoichiometric combustion of gasoline, C_7H_{17}, determine the fuel molecular weight, the combustion products, and the mass of carbon dioxide produced per kg of fuel burned.

Stoichiometric combustion:

$$C_7H_{17} + \nu_{O_2}\left(O_2 + 3.76\ N_2\right) \rightarrow \bullet\ a\ H_2O + b\ CO_2 + c\ N_2$$

C balance: $7 = b$

H balance: $17 = 2a, \quad \Rightarrow \quad a = 8.5$

O balance: $2\ \nu_{O_2} = a + 2b = 8.5 + 14 = 22.5 \quad \Rightarrow \quad \nu_{O_2} = 11.25$

N balance: $c = 3.76\ \ \nu_{O_2} = 3.76 \times 11.25 = 42.3$

$$M_{FUEL} = 7\ M_C + 17\ M_H = 7 \times 12.011 + 17 \times 1.008 = \mathbf{101.213}$$

$$\frac{m_{CO2}}{m_{FUEL}} = \frac{7\ M_{CO2}}{M_{FUEL}} = \frac{7 \times 44.01}{101.213} = \mathbf{3.044\ kg\ CO_2\ per\ kg\ fuel}$$

14.29

A sample of pine bark has the following ultimate analysis on a dry basis, percent by mass: 5.6% H, 53.4% C, 0.1% S, 0.1% N, 37.9% O and 2.9% ash. This bark will be used as a fuel by burning it with 100% theoretical air in a furnace. Determine the air–fuel ratio on a mass basis.

Converting the Bark Analysis from a mass basis:

Substance	S	H_2	C	O_2	N_2
c/M =	0.1/32	5.6/2	53.4/12	37.9/32	0.1/28
kmol / 100 kg coal	0.003	2.80	4.45	1.184	0.004
Product	SO_2	H_2O	CO_2		
oxygen required	0.003	1.40	4.45	--	--

Combustion requires: 0.003 + 1.40 + 4.45 = 5.853 kmol O_2 there is in the bark 1.184 kmol O_2 so the net from air is 4.669 kmol O_2

$$AF = (4.669 + 4.669 \times 3.76) \times \frac{28.97}{100} = \mathbf{6.44} \ \frac{\textbf{kg air}}{\textbf{kg bark}}$$

14.30

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: $b = 8.6 + \frac{0.6}{2} + \frac{12.267}{2} + 7.2 = 22.234$

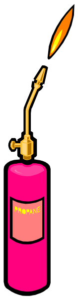
$$\text{Air-Fuel ratio} = \frac{22.234 + 83.6}{3.067} = 34.51$$

Theoretical:

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

$$\Rightarrow \text{theo. A-F ratio} = \frac{5 + 18.8}{1} = 23.8$$

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



14.31

A fuel, C_xH_y, is burned with dry air and the product composition is measured on a dry mole basis to be: 9.6% CO_2, 7.3% O_2 and 83.1% N_2. Find the fuel composition (*x*/*y*) and the percent theoretical air used.

$$\nu_{Fu}C_xH_y + \nu_{O_2}O_2 + 3.76\nu_{O_2}N_2 \rightarrow 9.6\ CO_2 + 7.3\ O_2 + 83.1\ N_2 + \nu_{H_2O}H_2O$$

N_2 balance: $3.76\nu_{O_2} = 83.1 \Rightarrow \nu_{O_2} = \mathbf{22.101}$

O_2 balance: $\nu_{O_2} = 9.6 + 7.3 + \frac{1}{2}\nu_{H_2O} \Rightarrow \nu_{H_2O} = \mathbf{10.402}$

H balance: $\nu_{Fu}\ y = 2\ \nu_{H_2O} = \mathbf{20.804}$

C balance: $\nu_{Fu}\ x = \mathbf{9.6}$

Fuel composition ratio = x/y = 9.6/20.804 = **0.461**

$$\text{Theoretical air} = \frac{\nu_{O_2AC}}{\nu_{O_2stoich}}100 = \frac{22.101}{9.6 + \frac{1}{4}\times 29.804}\ 100 = \mathbf{149.3\%}$$

14.32

For the combustion of methane 150% theoretical air is used at 25°C, 100 kPa and relative humidity of 70%. Find the composition and dew point of the products.

The reaction equation for stoichiometric mixture is:

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

C balance: $\nu_{CO2} = 1$, H balance: $2\nu_{H2O} = 4$, N_2 balance: $\nu_{N2} = 3.76\ \nu_{O2}$

O balance: $2\ \nu_{O2} = \nu_{H2O} + 2\nu_{CO2} = 2 + 2 \times 1 \quad => \quad \nu_{O2} = 2$

150% theoretical air: $\nu_{O2} = 1.5 \times 2 = 3$ so now more O_2 and N_2

$$CH_4 + 3\left(O_2 + 3.76\ N_2\right) \rightarrow 2\ H_2O + 1\ CO_2 + 11.28\ N_2 + 1\ O_2$$

Add water to the dry air from Eq.12.28

$$w = 0.622\frac{\phi P_g}{P_{tot} - \phi P_g} = 0.622\frac{0.7 \times 3.169}{100 - 0.7\ \times 3.169} = 0.0141$$

So the number of moles to add is from Eq.14.9

$$x = 7.655\ w = 7.655 \times 0.0141 = 0.108$$

and the added number of moles is $\nu_{O2}\ x = 0.324$, the products are then

Products: $2.324\ H_2O + 1\ CO_2 + 11.28\ N_2 + 1\ O_2$

The water partial pressure becomes

$$P_v = y_v\ P_{tot} = \frac{2.324}{2.324 + 1 + 11.28 + 1}\ 100 = 14.894\ \text{kPa}$$

$$T_{dew} = \mathbf{53.8^oC}$$

14.33

Many coals from the western United States have a high moisture content. Consider the following sample of Wyoming coal, for which the ultimate analysis on an as-received basis is, by mass:

Component	Moisture	H	C	S	N	O	Ash
% mass	28.9	3.5	48.6	0.5	0.7	12.0	5.8

This coal is burned in the steam generator of a large power plant with 150% theoretical air. Determine the air–fuel ratio on a mass basis.

Converting from mass analysis:

Substance	S	H_2	C	O_2	N_2
c/M =	0.5/32	3.5/2	4.86/12	12/32	0.7/28
kmol / 100 kg coal	0.0156	1.75	4.05	0.375	0.025
Product	SO_2	H_2O	CO_2		
oxygen required	0.0156	0.875	4.05	--	--

Combustion requires then oxygen as: $0.0156 + 0.875 + 4.05 = 4.9406$

The coal does include 0.375 O_2 so only 4.5656 O_2 from air/100 kg coal

AF = $1.5 \times (4.5656 + 4.5656 \times 3.76) \times 28.97/100 =$ **9.444 kg air/kg coal**

14.34

Pentane is burned with 120% theoretical air in a constant pressure process at 100 kPa. The products are cooled to ambient temperature, 20°C. How much mass of water is condensed per kilogram of fuel? Repeat the answer, assuming that the air used in the combustion has a relative humidity of 90%.

$$C_5H_{12} + 1.2 \times 8\,(O_2 + 3.76\,N_2) \rightarrow\ 5\,CO_2 + 6\,H_2O + 0.96\,O_2 + 36.1\,N_2$$

Products cooled to 20°C, 100 kPa, so for H_2O at 20°C: $P_g = 2.339$ kPa

$$y_{H2O\ MAX} = P_g/P = \frac{2.339}{100} = \frac{n_{H2O\ MAX}}{n_{H2O\ MAX} + 42.06} \quad \Rightarrow \quad n_{H2O\ MAX} = 1.007 < n_{H2O}$$

Therefore, $n_{H2O\ VAP} = 1.007$, $n_{H2O\ LIQ} = 6 - 1.007 = 4.993$

$$m_{H2O\ LIQ} = \frac{4.993 \times 18.015}{72.151} = \mathbf{1.247\ kg/kg\ fuel}$$

$$P_{v1} = 0.9 \times 2.339 = 2.105 \text{ kPa} \quad \Rightarrow \quad w_1 = 0.622 \times \frac{2.105}{97.895} = 0.013375$$

$$n_{H2O\ IN} = 0.013375 \times \frac{28.97}{18.015} \times (9.6 + 36.1) = 0.983 \text{ kmol}$$

$$n_{H2O\ OUT} = 0.983 + 6 = 6.983 \quad \Rightarrow \quad n_{H2O\ LIQ} = 6.983 - 1.007 = 5.976 \text{ kmol}$$

$$n_{H2O\ LIQ} = \frac{5.976 \times 18.015}{72.151} = \mathbf{1.492\ kg/kg\ fuel}$$

14.35

The coal gasifier in an integrated gasification combined cycle (IGCC) power plant produces a gas mixture with the following volumetric percent composition:

Product	CH_4	H_2	CO	CO_2	N_2	H_2O	H_2S	NH_3
% vol.	0.3	29.6	41.0	10.0	0.8	17.0	1.1	0.2

This gas is cooled to 40°C, 3 MPa, and the H_2S and NH_3 are removed in water scrubbers. Assuming that the resulting mixture, which is sent to the combustors, is saturated with water, determine the mixture composition and the theoretical air–fuel ratio in the combustors.

CH_4	H_2	CO	CO_2	N_2	n
0.3	29.6	41.0	10.0	0.8	81.7

$y_{H2O} = \frac{n_V}{n_V + 81.7}$, where n_V = number of moles of water vapor

Cool to 40°C $P_G = 7.384$, $P = 3000$ kPa

$$y_{H2O\ MAX} = \frac{7.384}{3000} = \frac{n_V}{n_V + 81.7} \quad \rightarrow \quad n_V = 0.2016$$

a) Mixture composition:

CH_4	H_2	CO	CO_2	N_2	$H_2O(v)$
0.3 kmol	29.6	41.0	10.0	0.8	0.2016

81.9016 kmol (from 100 kmol of the original gas mixture)

$$0.3\ CH_4 + 0.6\ O_2 \rightarrow 0.3\ CO_2 + 0.6\ H_2O$$

$$29.6\ H_2 + 14.8\ O_2 \rightarrow 29.6\ H_2O$$

$$41\ CO + 20.5\ O_2 \rightarrow 41\ CO_2$$

$\Rightarrow$ Number of moles of $O_2 = 0.6 + 14.8 + 20.5 = 35.9$

Number of moles of air = $35.9 + 3.76 \times 35.9\ (N_2)$

$$A/F = \frac{28.97(35.9 + 3.76(35.9))}{0.3(16) + 29.6(2) + 41(28) + 10(44) + 0.8(28) + 0.2016(18)}$$

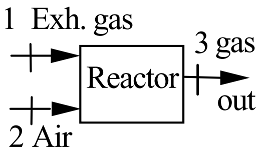
= **2.95 kg air/kg fuel**

14.36

The hot exhaust gas from an internal combustion engine is analyzed and found to have the following percent composition on a volumetric basis at the engine exhaust manifold. 10% CO_2, 2% CO, 13% H_2O, 3% O_2 and 72% N_2. This gas is fed to an exhaust gas reactor and mixed with a certain amount of air to eliminate the carbon monoxide, as shown in Fig. P14.36. It has been determined that a mole fraction of 10% oxygen in the mixture at state 3 will ensure that no CO remains. What must the ratio of flows be entering the reactor?

Exhaust gas at state 1: CO_2 10 %, H_2O 13%,
CO 2%, O_2 3%, N_2 72%

Exhaust gas at state 3: CO = 0 %, O_2 = 10 %



Reaction equation for the carbon monoxide

$$\Rightarrow 0.02\,CO + x\,O_2 + 3.76x\,N_2 \rightarrow 0.02\,CO_2 + (x-0.01)\,O_2 + 3.76x\,N_2$$

At 3: $\nu_{CO_2} = 0.10 + 0.02 = 0.12,$ $\nu_{H_2O} = 0.13$

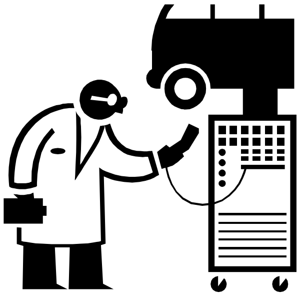
$\nu_{O_2} = (x-0.01) + 0.03 = x + 0.02$ $\nu_{N_2} = 0.72 + 3.76x$

or

$$n_{TOT} = 0.12 + 0.13 + x + 0.02 + 0.72 + 3.76x = 0.99 + 4.76x$$

$$y_{O2} = 0.10 = \frac{x + 0.02}{0.99 + 4.76x} \quad \rightarrow \quad x = 0.151$$

or $$\frac{\text{air}_2}{\text{Exh. Gas}_1} = \frac{4.76x}{1} = \mathbf{0.718}\,\frac{\textbf{kmol air}}{\textbf{kmol Exh. gas}}$$



14.37

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 $\nu_{O_2\ S} = 4 + 5/2 = 6.5;\qquad 3.76\ \nu_{O_2} = 24.44$

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

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

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

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

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

The water partial pressure becomes

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

$$T_{dew} = \mathbf{43.85^oC}$$

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

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

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

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

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

14.38

Methanol, CH_3OH, is burned with 200% theoretical air in an engine and the products are brought to 100 kPa, 30°C. How much water is condensed per kilogram of fuel?

$$CH_3OH + \nu_{O2}\{O_2 + 3.76\ N_2\} \rightarrow CO_2 + 2\ H_2O + 3.76\ \nu_{O2}N_2$$

Stoichiometric $\nu_{O2\ S} = 1.5 \quad \Rightarrow \quad \nu_{O2\ AC} = 3$

Actual products: $CO_2 + 2\ H_2O + 1.5\ O_2 + 11.28\ N_2$

$P_{sat}(30°C) = 4.246$ kPa

$$\Rightarrow y_{H2O} = 0.04246 = \frac{\nu_{H2O}}{1 + \nu_{H2O} + 1.5 + 11.28}$$

$$\Rightarrow \nu_{H2O} = 0.611 \quad \Rightarrow \quad \Delta\nu_{H2O\ cond} = 2 - 0.611 = 1.389$$

$$M_{Fu} = 32.042 \qquad \frac{\Delta M_{H2O}}{M_{Fu}} = \frac{1.389 \times 18}{32.042} = \mathbf{0.781}\ \frac{\mathbf{kg\ H_2O}}{\mathbf{kg\ fuel}}$$

14.39

The output gas mixture of a certain air–blown coal gasifier has the composition of producer gas as listed in Table 14.2. Consider the combustion of this gas with 120% theoretical air at 100 kPa pressure. Determine the dew point of the products and find how many kilograms of water will be condensed per kilogram of fuel if the products are cooled 10°C below the dew-point temperature.

$$\{3\ CH_4 + 14\ H_2 + 50.9\ N_2 + 0.6\ O_2 + 27\ CO + 4.5\ CO_2\}$$

$$+ 31.1\ O_2 + 116.9\ N_2 \rightarrow 34.5\ CO_2 + 20\ H_2O + 5.2\ O_2 + 167.8\ N_2$$

Products:

$$y_{H2O} = y_{H2O\ MAX} = P_G/100 = \frac{20}{34.5 + 20 + 5.2 + 167.8}$$

$$\Rightarrow P_G = 8.79\ \text{kPa} \quad \rightarrow \quad T_{DEW\ PT} = \mathbf{43.2°C}$$

At T = 33.2°C, P_G = 5.13 kPa

$$y_{H2O} = \frac{5.13}{100} = \frac{n_{H2O}}{n_{H2O}+34.5+5.2+167.8} \quad \rightarrow \quad n_{H2O} = 11.22$$

$$m_{H2O\ LIQ} = \frac{8.78(18)}{3(16) + 14(2) + 50.9(28) + 0.6(32) + 27(28) + 4.5(44)} = \mathbf{0.0639\ kg/kg\ fuel}$$

14.40

In an engine liquid octane and ethanol, mole ration 9:1, and stoichiometric air are taken in at 298 K, 100 kPa. After complete combustion, the products run out of the exhaust system where they are cooled to 10°C. Find the dew point of the products and the mass of water condensed per kilogram of fuel mixture.

Reaction equation with 0.9 octane and 0.1 ethanol is

$$0.9\ C_8H_{18} + 0.1\ C_2H_5OH + 11.55\ O_2 + 43.428\ N_2$$

$$\rightarrow 8.4\ H_2O + 7.4\ CO_2 + 43.428\ N_2$$

$$y_{H_2O} = \frac{8.4}{8.4 + 7.4 + 43.428} = 0.1418$$

$$P_{H_2O} = y_{H_2O} P_{tot} = \mathbf{14.3\ kPa} \qquad \Rightarrow\ T_{dew} = \mathbf{52.9\ °C}$$

$$10\ °C \Rightarrow P_{H_2O} = 1.2276 \quad \Rightarrow \quad y_{H_2O} = 0.012276 = \frac{x}{x + 7.4 + 43.428}$$

$$\Rightarrow x = \mathbf{0.6317} \quad \Rightarrow \quad \Delta\nu_{H_2O} = -7.77\ \frac{\text{kmol}}{\text{kmol Fu mix}}$$

$$m_{H_2O\ cond} = \frac{-\Delta\nu_{H_2O} \times 18.015}{107.414} = \mathbf{1.303\ \frac{kmol}{kmol\ Fu\ mix}}$$

Energy Equation, Enthalpy of Formation

14.41

A rigid vessel initially contains 2 kmol of carbon and 2 kmol of oxygen at 25°C, 200 kPa. Combustion occurs, and the resulting products consist of 1 kmol of carbon dioxide, 1 kmol of carbon monoxide, and excess oxygen at a temperature of 1000 K. Determine the final pressure in the vessel and the heat transfer from the vessel during the process.

$$2\,C + 2\,O_2 \;\rightarrow 1\,CO_2 + 1\,CO + \tfrac{1}{2}O_2$$

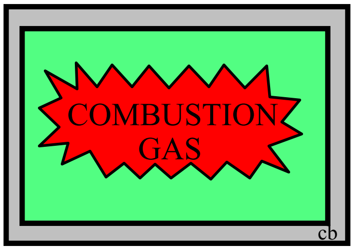
Process V = constant, C: solid, $n_{1(GAS)} = 2$, $n_{2(GAS)} = 2.5$

$$P_2 = P_1 \times \frac{n_2T_2}{n_1T_1} = 200 \times \frac{2.5 \times 1000}{2 \times 298.2} = \mathbf{838.4\ kPa}$$

$$H_1 = 0$$

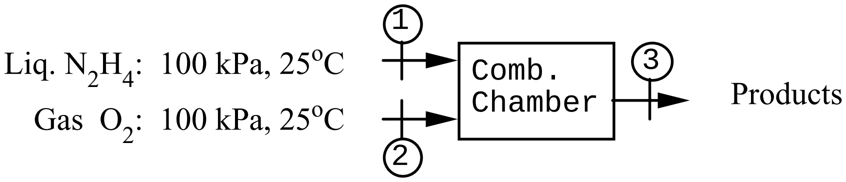
$$H_2 = 1(-393\,522 + 33\,397) + 1(-110\,527 + 21\,686)$$
$$+ (1/2)(0 + 22\,703) = \mathbf{-437\,615\ kJ}$$

$$_1Q_2 = (U_2 - U_1) = (H_2 - H_1) - n_2\bar{R}T_2 + n_1\bar{R}T_1$$
$$= (-437\,615 - 0) - 8.3145(2.5 \times 1000 - 2 \times 298.2) = \mathbf{-453\,442\ kJ}$$



14.42

In a test of rocket propellant performance, liquid hydrazine (N_2H_4) at 100 kPa, 25°C, and oxygen gas at 100 kPa, 25°C, are fed to a combustion chamber in the ratio of 0.5 kg O_2/kg N_2H_4. The heat transfer from the chamber to the surroundings is estimated to be 100 kJ/kg N_2H_4. Determine the temperature of the products exiting the chamber. Assume that only H_2O, H_2, and N_2 are present. The enthalpy of formation of liquid hydrazine is +50 417 kJ/kmol.



$\dot{m}_{O2}/\dot{m}_{N2H4} = 0.5 = 32\dot{n}_{O2}/32\dot{n}_{N2H4}$ and $\dot{Q}/\dot{m}_{N2H4} = -100$ kJ/kg

Energy Eq.: $Q_{CV} = H_P - H_R = -100 \times 32.045 = -3205$ kJ/kmol fuel

Combustion eq.: $1\, N_2H_4 + \frac{1}{2} O_2 \rightarrow H_2O + H_2 + N_2$

$H_R = 1(50417) + \frac{1}{2}(0) = 50417$ kJ

$H_P = -241\,826 + \Delta\bar{h}_{H2O} + \Delta\bar{h}_{H2} + \Delta\bar{h}_{N2}$

Energy Eq. now reads

$$H_P = H_R + Q_{CV} = H^o_P + \Delta H_P$$

$$\Delta H_P = \Delta\bar{h}_{H2O} + \Delta\bar{h}_{H2} + \Delta\bar{h}_{N2} = -H^o_P + H_R + Q_{CV}$$

$$= 241\,826 + 50\,417 - 3205 = 289\,038 \text{ kJ/kmol fuel}$$

Table A.9 : Guess T and read for water, hydrogen and nitrogen

2800 K: $\Delta H_P = 115\,463 + 81\,355 + 85\,323 = 282\,141$ too low

3000 K: $\Delta H_P = 126\,548 + 88\,725 + 92\,715 = 307\,988$ too high

Interpolate to get $T_P =$ **2854 K**

14.43

The combustion of heptane C_7H_{16} takes place in a steady flow burner where fuel and air are added as gases at P_o, T_o. The mixture has 125% theoretical air and the products are going through a heat exchanger where they are cooled to 600 K.Find the heat transfer from the heat exchanger per kmol of heptane burned.

The reaction equation for stoichiometric ratio is:

$$C_7H_{16} + v_{O2} (O_2 + 3.76\ N_2) \quad => \ 7CO_2 + 8\ H_2O + v_{O2} \times 3.76\ N_2$$

So the balance (C and H was done in equation) of oxygen gives $v_{O2} = 7 + 4 = 11$, and actual one is $11 \times 1.25 = 13.75$. Now the actual reaction equation is:

$$C_7H_{16} + 13.75\ O_2 + 51.7\ N_2 \quad => \ 7CO_2 + 8\ H_2O + 51.7\ N_2 + 2.75\ O_2$$

To find the heat transfer take control volume as combustion chamber and heat exchanger

$$H_R + Q = H_P \quad => \quad Q = H^o_P + \Delta H_P - H^o_R$$

Take the enthalpies from Tables A.9 for products and A.10 for fuel

$$Q = 7(-393\,522 + 12\,906) + 8\,(-241\,826 + 10\,499) + 51.7(8894) + 2.75(9245) - (-187\,900)$$

$$= \mathbf{-\,3\,841\,784\ kJ/kmol\ fuel}$$

14.44

Butane gas and 200% theoretical air, both at 25°C, enter a steady flow combustor. The products of combustion exits at 1000 K. Calculate the heat transfer from the combustor per kmol of butane burned.

$C_4H_{10} + (1/\phi)\,\nu_{O2}(O_2 + 3.76\,N_2) \;\rightarrow\; a\,CO_2 + b\,H_2O + c\,N_2 + d\,O_2$

First we need to find the stoichiometric air ($\phi = 1$, $d = 0$)

C balance: $4 = a$, H balance: $10 = 2b \;\Rightarrow\; b = 5$

O balance: $2\nu_{O2} = 2a + b = 8 + 5 = 13 \;\Rightarrow\; \nu_{O2} = 6.5$

Now we can do the actual air: $(1/\phi) = 2 \;\Rightarrow\; \nu_{O2} = 2 \times 6.5 = 13$

N balance: $c = 3.76\,\nu_{O2} = 48.88$, O balance: $d = 13 - 6.5 = 6.5$

Energy Eq.: $q = H_R - H_P = H^o_R - H^o_P - \Delta H_P$

Table A.10: $H^o_R = -126\,200 + 0 + 0 = -126\,200$ kJ/kmol fuel

$H^o_P = 4\,(-393\,522) + 5(-241\,826) + 0 + 0 = -2\,783\,218$ kJ/kmol fuel

The rest of the values are from Table A.9 at 1000 K

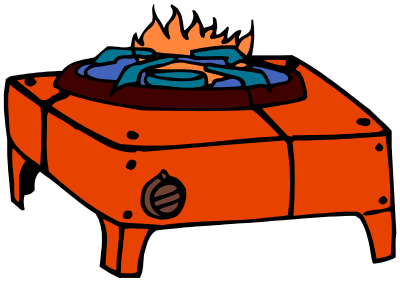
$\Delta\bar{h}_{CO2} = 33397$, $\Delta\bar{h}_{N2} = 21463$, $\Delta\bar{h}_{O2} = 22703$, $\Delta\bar{h}_{H2O} = 26000$ kJ/kmol

$\Delta H_P = 4 \times 33\,397 + 5 \times 26\,000 + 48.88 \times 21\,463 + 6.5 \times 22\,703$

$= 1\,460\,269$ kJ/kmol fuel

From the energy equation we get

$q = -126\,200 - (-2\,783\,218) - 1\,460\,269 =$ **1 196 749 kJ/kmol butane**



14.45

One alternative to using petroleum or natural gas as fuels is ethanol (C_2H_5OH), which is commonly produced from grain by fermentation. Consider a combustion process in which liquid ethanol is burned with 120% theoretical air in a steady flow process. The reactants enter the combustion chamber at 25°C, and the products exit at 60°C, 100 kPa. Calculate the heat transfer per kilomole of ethanol.

$$C_2H_5OH + 1.2 \times 3 \left(O_2 + 3.76\ N_2\right) \rightarrow 2CO_2 + 3H_2O + 0.6O_2 + 13.54N_2$$

Fuel: $\bar{h}_f^0$ = -277 380 kJ/kmol for liquid from Table A.10,

Products at 60°C, 100 kPa, check for condensation of water

$$y_{H2O\ MIX} = \frac{19.94}{100} = \frac{n_{V\ MAX}}{n_{V\ MAX}+2+0.6+13.54} \quad => \quad n_{V\ MAX} = 4.0 > 3 \ \Rightarrow\ \text{No liq.}$$

H_R = 1(-277 380) + 0 + 0 = -277 380 kJ/kmol fuel

H_P = 2(-393 522 + 1327) + 3(-241 826 + 1178)

+ 0.6(0 + 1032) + 13.54(0 + 1020) = -1 491 904 kJ/kmol fuel

$Q_{CV} = H_P - H_R$ = **-1 214 524 kJ/kmol fuel**

14.46

Do the previous problem with the ethanol fuel delivered as a vapor.
One alternative to using petroleum or natural gas as fuels is ethanol (C_2H_5OH), which is commonly produced from grain by fermentation. Consider a combustion process in which liquid ethanol is burned with 120% theoretical air in a steady flow process. The reactants enter the combustion chamber at 25°C, and the products exit at 60°C, 100 kPa. Calculate the heat transfer per kilomole of ethanol.

$$C_2H_5OH + 1.2 \times 3 \left(O_2 + 3.76\ N_2\right) \rightarrow 2CO_2 + 3H_2O + 0.6O_2 + 13.54N_2$$

Fuel: $\bar{h}_f^0 = -235\ 000$ kJ/kmol for IG from Table A.10

Products at 60°C, 100 kPa, check for condensation of water

$$y_{H2O\ MIX} = \frac{19.94}{100} = \frac{n_{V\ MAX}}{n_{V\ MAX}+2+0.6+13.54} \quad => \quad n_{V\ MAX} = 4.0 > 3 \Rightarrow \text{ No liq.}$$

$H_R = 1(-235\ 000) + 0 + 0 = -235\ 000$ kJ/kmol fuel

$H_P = 2(-393\ 522 + 1327) + 3(-241\ 826 + 1178)$

$+ 0.6(0 + 1032) + 13.54(0 + 1020) = -1\ 491\ 904$ kJ/kmol fuel

$Q_{CV} = H_P - H_R =$ **-1 256 904 kJ/kmol fuel**

14.47

Another alternative to using petroleum or natural gas as fuels is methanol, CH_3OH, which can be produced from coal. Both methanol and ethanol have been used in automotive engines. Repeat the previous problem using liquid methanol as the fuel instead of ethanol.

$$CH_3OH + 1.2 \times 1.5\left(O_2 + 3.76\, N_2\right) \rightarrow 1\, CO_2 + 2\, H_2O + 0.3\, O_2 + 6.77\, N_2$$

Reactants at 25 °C, products are at 60 °C = 333.2 K, 100 kPa, check for condensation of water

$$y_{H2O\ MAX} = \frac{19.94}{100} = \frac{n_{V\ MAX}}{n_{V\ MAX}+1+0.3+6.77} \;\; => \;\; n_{V\ MAX} = 2.0 > 2 \Rightarrow \text{No liq.}$$

CH_3OH: $\bar{h}_f^o = -239\ 220$ kJ/kmol from table A.10 for the liquid state

$H_R = 1\ \bar{h}_{LIQ} = -239\ 220$ kJ/kmol fuel

$H_P = 1(-393\ 522 + 1327) + 2(-241\ 826 + 1178)$

$+ 0.3(1032) + 6.77(1020) = -866\ 276$ kJ/kmol fuel

$Q = H_P - H_R =$ **-627 056 kJ/kmol fuel**

14.48

Another alternative fuel to be seriously considered is hydrogen. It can be produced from water by various techniques that are under extensive study. Its biggest problem at the present time are cost, storage, and safety. Repeat Problem 14.45 using hydrogen gas as the fuel instead of ethanol.

$H_2 + 1.2 \times 0.5\ O_2 + 1.2 \times 3.76 \times 0.5\ N_2 \ \rightarrow 1\ H_2O + 0.1\ O_2 + 2.256\ N_2$

Products at 60°C, 100 kPa, check for condensation of water

$$y_{H2O\ MAX} = \frac{19.94}{100} = \frac{n_{V\ MAX}}{n_{V\ MAX} + 0.1 + 2.256}$$

Solving, $n_{V\ MAX} = 0.587 < 1 \ \ => \ n_V = 0.587, \ \ n_{LIQ} = 0.413$

$H_R = 0 + 0 + 0 = 0$

Notice the products are at 60°C so add for water liquid from steam tables

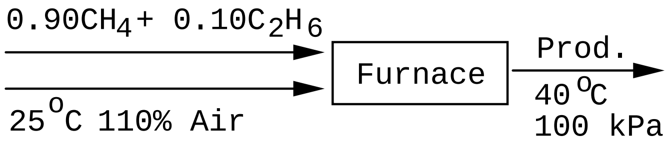
$H_P = 0.413[-285\ 830 + 18.015(251.1 - 104)] + 0.587(-241\ 826 + 1178)$

$\quad + 0.1(0 + 1032) + 2.256(0 + 1020) = -255\ 816$ kJ

$Q_{CV} = H_P - H_R =$ **-255 816 kJ**

14.49

In a new high-efficiency furnace, natural gas, assumed to be 90% methane and 10% ethane (by volume) and 110% theoretical air each enter at 25°C, 100 kPa, and the products (assumed to be 100% gaseous) exit the furnace at 40°C, 100 kPa. What is the heat transfer for this process? Compare this to an older furnace where the products exit at 250°C, 100 kPa.



$$0.9\,CH_4 + 0.1\,C_2H_6 + 1.1 \times 2.15\,O_2 + 3.76 \times 2.365\,N_2$$

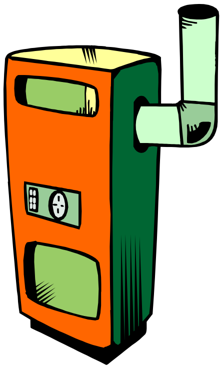
$$\rightarrow 1.1\,CO_2 + 2.1\,H_2O + 0.215\,O_2 + 8.892\,N_2$$

Fuel values from table A.10 and the rest from Table A.9

$H_R = 0.9(-74\ 873) + 0.1(-84\ 740) = -75860$ kJ/kmol fuel

$H_P = 1.1(-393\ 522 + 562) + 2.1(-241\ 826 + 504) + 0.215(441) + 8.892(437)$

$= -935\ 052$ kJ/kmol fuel   assuming all gas



$Q_{CV} = H_P - H_R =$ **-859 192 kJ/kmol fuel**

b) $T_P = 250\ ^oC$

$H_P = 1.1(-393\ 522 + 9346) + 2.1(-241\ 826 + 7740)$

$+ 0.215(6808) + 8.892(6597) \quad = -854\ 050$ kJ

$Q_{CV} = H_P - H_R =$ **-778 190 kJ/kmol fuel**

14.50

Repeat the previous problem, but take into account the actual phase behavior of the products exiting the furnace.

$$0.9\,CH_4 + 0.1\,C_2H_6 + 1.1 \times 2.15\,O_2 + 3.76 \times 2.365\,N_2$$
$$\rightarrow 1.1\,CO_2 + 2.1\,H_2O + 0.215\,O_2 + 8.892\,N_2$$

Same as 14.49, except check products for saturation at 40°C, 100 kPa

$$y_{V\,MAX} = \frac{7.384}{100} = \frac{n_{V\,MAX}}{n_{V\,MAX}+10.207} \quad => \quad \text{Solving, } n_{V\,MAX} = 0.814$$

$$n_V = 0.814, \quad n_{LIQ} = 2.1 - 0.814 = 1.286$$

Fuel values from table A.10 and the rest from Table A.9

$$H_R = 0.9(-74\,873) + 0.1(-84\,740) = -75\,860 \text{ kJ/kmol fuel}$$

For the liquid water add difference (40°C – 25°C) from steam tables

$$H_{LIQ} = 1.286\big[-285\,830 + 18.015(167.6 - 104.9)\big] = -366\,125 \text{ kJ/kmol fuel}$$

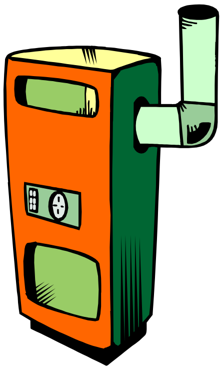
$$H_{GAS} = 1.1(-393\,522 + 562) + 0.814(-241\,826 + 504\,)$$
$$+ 0.215(441) + 8.892(437) = -624\,711 \text{ kJ/kmol fuel}$$

$$Q_{CV} = H_P - H_R = (-366\,125 - 624\,711) + 75\,860$$
$$= \mathbf{-914\,976 \text{ kJ/kmol fuel}}$$

b) $T_P = 250$ °C

$$H_P = 1.1(-393\,522 + 9346) + 2.1(-241\,826 + 7740)$$
$$+ 0.215(6808) + 8.892(6597)$$
$$= -854\,050 \text{ kJ/kmol}$$

$$Q_{CV} = H_P - H_R = \mathbf{-778\,190 \text{ kJ/kmol fuel}}$$



14.51

Pentene, C_5H_{10} is burned with pure oxygen in a steady flow process. The products at one point are brought to 700 K and used in a heat exchanger, where they are cooled to 25°C. Find the specific heat transfer in the heat exchanger.

$$C_5H_{10} + \nu_{O2}O_2 \rightarrow 5\ CO_2 + 5\ H_2O \qquad \Rightarrow \qquad \nu_{O2} = \mathbf{7.5}$$

The heat exchanger cools the products so energy equation is

$$5\ \dot{n}_F\ \bar{h}_{CO2} + 5\ \dot{n}_F\ \bar{h}_{H2O} + \dot{Q} = 5\ \dot{n}_F\ \bar{h}^\circ_{f\ CO2}$$

$$+ (5 - x)\ \dot{n}_F\ \bar{h}^\circ_{liq\ H2O} + (x)\ \dot{n}_F\ \bar{h}^\circ_{vap\ H2O}$$

Check for condensation amount

$$\text{Find x:} \quad y_{H2O\ max} = \frac{P_g(25°)}{P_{tot}} = 0.0313 = \frac{x}{5 + x} \qquad \Rightarrow \qquad x = \mathbf{0.1614}$$

Out of the 5 H_2O only 0.1614 still vapor.

$$\frac{\dot{Q}}{\dot{n}_F} = -5\ \Delta\bar{h}_{CO2,700} + (5\text{-}x)(\bar{h}^\circ_{f\ liq} - \bar{h}^\circ_{f\ vap} - \Delta\bar{h}_{700}) + x(\bar{h}^\circ_{f\ vap} - \bar{h}^\circ_{f\ vap} - \Delta\bar{h}_{700})$$

$$= -5(17\ 761) + 4.84(-44\ 011 - 14\ 184) - 0.16(14\ 184)$$

$$= \mathbf{-372\ 738\ kJ/kmol\ Fu}$$

14.52

Methane, CH_4, is burned in a steady flow process with two different oxidizers: Case **A**: Pure oxygen, O_2 and case **B**: A mixture of $O_2 + x$ Ar. The reactants are supplied at T_0, P_0 and the products for both cases should be at 1800 K. Find the required equivalence ratio in case (A) and the amount of Argon, x, for a stoichiometric ratio in case (B).

a) Stoichiometric has $\nu = 2$, actual has:

$$CH_4 + \nu O_2 \rightarrow CO_2 + 2H_2O + (\nu - 2)O_2$$

Energy eq.: $\overset{\circ}{H}_R = \overset{\circ}{H}_P + \Delta H_{P\ 1800}$

$$\Delta H_{P\ 1800} = \overset{\circ}{H}_R - \overset{\circ}{H}_P = \bar{h}^o_{f\ fuel} + 0 - \bar{h}^o_{f\ CO_2} - 2\bar{h}^o_{f\ H_2O} - 0$$

$$= -74\ 873 - (-393\ 522) - 2(-241\ 826) = 802\ 301 \text{ kJ/kmol}$$

$\Delta\bar{h}_{CO_2} = 79\ 432$, $\Delta\bar{h}_{H_2O} = 62\ 693$, $\Delta\bar{h}_{O_2} = 51\ 674$ all in kJ/kmol

$\Delta H_{P\ 1800} = 101\ 470 + \nu\ 51\ 674 = 802\ 301$ kJ/kmol fuel

$$\Rightarrow \quad \boldsymbol{\nu = 13.56}, \quad \Phi = \frac{AF_S}{AF} = \frac{2}{13.56} = \mathbf{0.1475}$$

b) $CH_4 + 2\ O_2 + 2x\ Ar \rightarrow CO_2 + 2H_2O + 2x\ Ar$

$\Delta H_{P\ 1800} = 79\ 432 + 2 \times 62\ 693$

$+ 2x \times 0.52 \times 39.948(1800 - 298) = 204\ 818 + x\ 62\ 402$

Now the energy equation becomes

$802\ 301 = 204\ 818 + x\ 62\ 402 \quad \Rightarrow \quad x = \mathbf{9.575}$

14.53

A closed, insulated container is charged with a stoichiometric ratio of oxygen and hydrogen at 25°C and 150 kPa. After combustion, liquid water at 25°C is sprayed in such that the final temperature is 1200 K. What is the final pressure?

Combustion reaction: $H_2 + \frac{1}{2}(O_2) \rightarrow H_2O$

Products: $1\ H_2O + x_i H_2O$

$$U_2 - U_1 = x_i \bar{h}_i = x_i h^{\circ}_{f\,liq} = (1 + x_i)H_P - H_R - (1 + x_i)\bar{R}T_P + \tfrac{3}{2}\bar{R}T_R$$

From Table A.9: $H_R = 0$, $H_P = -241\,826 + 34\,506 = -207\,320$ kJ/kmol

From Table A.10: $h^{\circ}_{f\,liq} = -285\,830$ kJ/kmol

Substitute

$$x_i(-285830 + 207320 + 8.3145 \times 1200) =$$

$$-207\,320 - 8.3145\left(1200 - \tfrac{3}{2}\times 298.15\right) = -213\,579$$

$$x_i = 3.116$$

$$P_1V_1 = n_R\bar{R}T_1, \quad P_2V_1 = n_p\bar{R}T_p$$

$$P_2 = \frac{P_1(1 + x_i)T_P}{\frac{3}{2}(T_1)} = \frac{150 \times 4.116 \times 1200}{\frac{3}{2} \times 298.15} = \mathbf{1657\ kPa}$$

14.54

Gaseous propane mixes with air, both supplied at 500 K, 0.1 MPa. The mixture goes into a combustion chamber and products of combustion exit at 1300 K, 0.1 MPa. The products analyzed on a dry basis are 11.42% CO_2, 0.79% CO, 2.68% O_2, and 85.11% N_2 on a volume basis. Find the equivalence ratio and the heat transfer per kmol of fuel.

$$C_3H_8 + \alpha\, O_2 + 3.76\,\alpha\, N_2 \rightarrow \beta\, CO_2 + \gamma\, H_2O + 3.76\,\alpha\, N_2$$

$$\beta = 3, \quad \gamma = 4, \quad \alpha = \beta + \gamma/2 = 5, \qquad (A/F)_S = 4.76\alpha = 23.8$$

The actual combustion reaction is

$$xC_3H_8 + \alpha\, O_2 + 3.76\,\alpha\, N_2 \rightarrow$$

$$11.42\, CO_2 + y\, H_2O + 85.11\, N_2 + 0.79\, CO + 2.68\, O_2$$

C balance: $3x = 11.42 + 0.79 \quad => \quad x = 4.07$

H balance: $8x = 2y \quad => \quad y = 4x = 16.28$

O balance: $2\alpha = 2 \times 11.42 + y + 0.79 + 2 \times 2.68 = 45.27 \quad => \quad \alpha = 22.635$

N balance: $3.76\,\alpha = 85.11 \quad => \quad \alpha = 22.6356$ checks close enough

Rescale the equation by dividing with x to give

$$C_3H_8 + 5.5614\,(O_2 + 3.76\, N_2) \rightarrow$$

$$2.806\, CO_2 + 4\, H_2O + 20.91\, N_2 + 0.194\, CO + 0.6584\, O_2$$

$$A/F = 5.5614\,(1 + 3.76) / 1 = 26.472$$

$\phi = (A/F)_S/(A/F) = 23.8 / 26.472 = \mathbf{0.899}$, %Theo. air = $1/\phi$ = **111%**

$$h_P = h_P^o + \sum \nu_i \Delta h(1300\text{ K})$$

$$q = h_P - h_R = h_P^o + \sum \nu_i \Delta h(1300\text{ K}) - h_R$$

$$h_R = h^o_{f\,fuel} + \Delta h_{fuel} + 5.5614\, \Delta h_{O2} + 20.91\, \Delta h_{N2}$$

$$= -103\,900 + 1.679 \times 44.094\,(500 - 298) + 5.5614\,(6086)$$

$$+ 20.91\,(5911) = 68\,500 \text{ kJ/kmol fuel}$$

$$h_P^o = 2.806\,(-393\,522) + 4(-241\,826) + 0 + 0.194\,(-110\,527) + 0$$

$$= -2\,092\,969 \text{ kJ/kmol fuel}$$

$$\sum \nu_i \Delta h(1300\text{ K}) = 2.806\,(50\,148) + 4(38\,941) + 20.91\,(31\,503)$$

$$+ 0.194\,(31\,867) + 0.6584\,(33\,345)$$

$$= 983\,344 \text{ kJ/kmol fuel}$$

$q = -2\,092\,969 + 983\,344 - 68\,500 =$ **-1 178 125 kJ/kmol fuel**

Enthalpy of Combustion and Heating Value

14.55

Liquid pentane is burned with dry air and the products are measured on a dry basis as: 10.1% CO_2, 0.2% CO, 5.9% O_2 remainder N_2. Find the enthalpy of formation for the fuel and the actual equivalence ratio.

$$\nu_{Fu} C_5H_{12} + \nu_{O2} O_2 + 3.76\ \nu_{O2} N_2 \rightarrow$$

$$x\ H_2O + 10.1\ CO_2 + 0.2\ CO + 5.9\ O_2 + 83.8\ N_2$$

Balance of C: $5\ \nu_{Fu} = 10.1 + 0.2 \ \Rightarrow\ \nu_{Fu} = \mathbf{2.06}$

Balance of H: $12\ \nu_{Fu} = 2\ x \ \Rightarrow\ x = 6\ \nu_{Fu} = \mathbf{12.36}$

Balance of O: $2\ \nu_{O_2} = x + 20.2 + 0.2 + 2 \times 5.9 \ \Rightarrow\ \nu_{O_2} = \mathbf{22.28}$

Balance of N: $2 \times 3.76\ \nu_{O_2} = 83.8 \times 2 \ \Rightarrow\ \nu_{O_2} = \mathbf{22.287} \Rightarrow \mathbf{OK}$

ν_{O_2} for 1 kmol fuel = 10.816

$\phi = 1,\ C_5H_{12} + 8\ O_2 + 8 \times 3.76\ N_2 \rightarrow 6\ H_2O + 5\ CO_2 + 30.08\ N_2$

$H^{\circ}_{RP} = H^{\circ}_P - H^{\circ}_R = 6\ \bar{h}^{\circ}_{f\ H_2O} + 5\ \bar{h}^{\circ}_{f\ CO_2} - \bar{h}^{\circ}_{f\ fuel}$

14.3: $H^{\circ}_{RP} = 44\ 983 \times 72.151 \qquad \Rightarrow \qquad \bar{h}^{\circ}_{f\ fuel} = \mathbf{-172\ 998\ kJ/kmol}$

$\phi = AFs / AF = \nu_{O2\ stoich}/\nu_{O2\ AC} = 8/10.816 = \mathbf{0.74}$

14.56

Phenol has an entry in Table 14.3, but it does not have a corresponding value of the enthalpy of formation in Table A.10. Can you calculate it?

$$C_6H_5OH + \nu_{O2}[\,O_2 + 3.76\,N_2\,] \rightarrow \;3\,H_2O + 6\,CO_2 + 3.76\nu_{O2}\,N_2$$

The C and H balance was introduced (6 C's and 6 H's). At the reference condition the oxygen and nitrogen have zero enthalpy of formation.

Energy Eq.: $H_P = H_R = H_P^o = H_R^o$ since ref. T is assumed.

$$H_{RP}^{\circ} = H_P - H_R = H_P^o - H_R^o = \;3\,\bar{h}_{f\,H2O}^{\circ} + 6\,\bar{h}_{f\,CO_2}^{\circ} - \bar{h}_{f\,fuel}^{\circ}$$

Table 14.3 is on a mass basis and let us chose liquid fuel, so we get the molecular weight from the composition

$$M = 6 \times 12.011 + 3 \times 2.016 \;+ 16 = 94.114$$

$$H_{RP}^{\circ} = 94.114\,(-31\,117) = -2\,928\,545 \text{ kJ/kmol}$$

Solve the energy equation for fuel formation enthalpy

$$\bar{h}_{f\,fuel}^{\circ} = \;3\,\bar{h}_{f\,H2O}^{\circ} + 6\,\bar{h}_{f\,CO_2}^{\circ} - H_{RP}^{\circ}$$

$$= \;3\,(-241\,826) + 6(-393\,522) - (-2\,928\,545)$$

$$= \mathbf{-158\,065\text{ kJ/kmol}}$$

For fuel as vapor we get

$$H_{RP}^{\circ} = 94.114\,(-31\,774) = -2\,990\,378 \text{ kJ/kmol}$$

$$\bar{h}_{f\,fuel}^{\circ} = \;3\,\bar{h}_{f\,H2O}^{\circ} + 6\,\bar{h}_{f\,CO_2}^{\circ} - H_{RP}^{\circ}$$

$$= \;3\,(-241\,826) + 6(-393\,522) - (-2\,990\,378)$$

$$= \mathbf{-96\,232\text{ kJ/kmol}}$$

Notice if I took liquid water in products to do H_{RP}° then I must use liquid value for $\bar{h}_{f\,H2O}^{\circ} = -\,285\,830$ kJ/kmol and the final result is the same.

14.57

Do problem 14.43 using table 14.3 instead of Table A.10 for the solution.

The reaction equation for stiochiometric ratio is:

$$C_7H_{16} + \nu_{O2} (O_2 + 3.76 N_2) \quad => \; 7CO_2 + 8 H_2O + \nu_{O2} \times 3.76 N_2$$

So the balance (C and H was done in equation) of oxygen gives $v_{O2} = 7 + 4 = 11$, and actual one is 11×1.25 = 13.75. Now the actual reaction equation is:

$$C_7H_{16} + 13.75 O_2 + 51.7 N_2 \quad => \; 7CO_2 + 8 H_2O + 51.7 N_2 + 2.75 O_2$$

To find the heat transfer take control volume as combustion chamber and heat exchanger

$$H_R + Q = H_P \quad => \quad Q = H_{Po} + \Delta H_P - H_{Ro} = H_{RPo} + \Delta H_P$$

Now we get the enthalpy of combustion from table 14.3, which is per kg, so scale it with the molecular weight for the fuel. Add all the ΔH_P from A.9

$$Q = M H_{RPo} + 7 \Delta\bar{h}_{CO_2} + 8 \Delta\bar{h}_{H_2O} + 51.7 \Delta\bar{h}_{N_2} + 2.75 \Delta\bar{h}_{O_2}$$

$$= 100.205(-44\,922) + 7(12\,906) + 8(10\,499) + 51.7(8894) + 2.75(9245)$$

$$= -4\,501\,409 + 90\,342 + 83\,922 + 459\,819.8 + 25\,423.75$$

= **- 3 841 831 kJ/kmol fuel**

14.58

Wet biomass waste from a food-processing plant is fed to a catalytic reactor, where in a steady flow process it is converted into a low-energy fuel gas suitable for firing the processing plant boilers. The fuel gas has a composition of 50% methane, 45% carbon dioxide, and 5% hydrogen on a volumetric basis. Determine the lower heating value of this fuel gas mixture per unit volume.

For 1 kmol fuel gas,

$$0.5\,CH_4 + 0.45\,CO_2 + 0.05\,H_2 + 1.025\,O_2$$

$$\rightarrow (0.5 + 0.45)\,CO_2 + 1.05\,H_2O$$

The lower heating value is with water vapor in the products. Since the $0.45\,CO_2$ cancels,

$$\bar{h}_{RP} = 0.5(-393\,522) + 1.05(-241\,826) - 0.5(-74\,873) - 0.05(0)$$

$$= -413242 \text{ kJ/kmol fuel gas}$$

$$\text{With}\quad \frac{n}{V} = P/\bar{R}T = \frac{100}{8.3145 \times 298.2} = 0.04033 \text{ kmol/m}^3$$

$$\text{LHV} = +413\,242 \times 0.04033 = \mathbf{16\,666\ kJ/m^3}$$

14.59

Determine the lower heating value of the gas generated from coal as described in Problem 14.35. Do not include the components removed by the water scrubbers.

The gas from problem 14.35 is saturated with water vapor. Lower heating value LHV has water as vapor.

$$\text{LHV} = -H^{\circ}_{RP} = H^{\circ}_{P} - H^{\circ}_{R}$$

Only CH_4, H_2 and CO contributes. From 14.12 the gas mixture after the scrubbers has $\sum \nu_i = 81.9$ of composition:

$$0.3\ CH_4 + 29.6\ H_2 + 41\ CO + 10\ CO_2 + 0.8\ N_2 + 0.2016\ H_2O$$

$$\text{LHV} = -[0.3\bar{H}^{\circ}_{RP\,CH4} + 29.6\bar{H}^{\circ}_{RP\,H2} + 41\bar{H}^{\circ}_{RP\,CO}]/81.9$$

$$= -[0.3(-50\ 010 \times 16.043) + 29.6(-241\ 826) + 41(-393\ 522 + 110\ 527)]/81.9$$

$$= \mathbf{232\ 009}\ \frac{\mathbf{kJ}}{\mathbf{kmol\ gas}}$$

14.60

Do problem 14.45 using table 14.3 instead of Table A.10 for the solution. One alternative to using petroleum or natural gas as fuels is ethanol (C_2H_5OH), which is commonly produced from grain by fermentation. Consider a combustion process in which liquid ethanol is burned with 120% theoretical air in a steady flow process. The reactants enter the combustion chamber at 25°C, and the products exit at 60°C, 100 kPa. Calculate the heat transfer per kilomole of ethanol.

$$C_2H_5OH + 1.2 \times 3 \left(O_2 + 3.76\ N_2\right) \rightarrow 2\ CO_2 + 3\ H_2O + 0.6\ O_2 + 13.54\ N_2$$

Products at 60°C, 100 kPa, so check for condensation of water

$$y_{H2O\ MIX} = \frac{19.94}{100} = \frac{n_{V\ MAX}}{n_{V\ MAX}+2+0.6+13.54} \quad => \quad n_{V\ MAX} = 4.0 > 3 \ \Rightarrow \ \text{No liq.}$$

Fuel: table 14.3 select (liquid fuel, water vapor) and convert to mole basis

$$\overset{\circ}{H}_{RP} = 46.069\ (-26\ 811) = -1\ 235\ 156\ \text{kJ/kmol}$$

Since the reactants enter at the reference state the energy equation becomes

$$Q_{CV} = H_P - H_R = H_P^o + \Delta H_P - H_R^o = \overset{\circ}{H}_{RP} + \Delta H_P$$

$$\Delta H_P = 2\ \Delta\bar{h}_{CO_2} + 3\ \Delta\bar{h}_{H_2O} + 0.6\ \Delta\bar{h}_{O_2} + 13.54\ \Delta\bar{h}_{N_2}$$

$$= 2(1327) + 3(1178) + 0.6(1032) + 13.54(1020) = 20\ 618\ \text{kJ/kmol}$$

$$Q_{CV} = -1\ 235\ 156 + 20\ 618 = \mathbf{-1\ 214\ 538\ kJ/kmol\ fuel}$$

14.61

Propylbenzene, C_9H_{12}, is listed in Table 14.3, but not in table A.10. No molecular weight is listed in the book. Find the molecular weight, the enthalpy of formation for the liquid fuel and the enthalpy of evaporation.

$$C_9H_{12} + 12\ O_2 \rightarrow\ 9\ CO_2 + 6\ H_2O$$

$\hat{M} = 9 \times 12.011 + 6 \times 2.016 =$ **120.195**

$$\bar{h}^\circ_{RP} = = H^o_P - H^o_R = \sum_P \nu_i \bar{h}^\circ_{f_i} - \bar{h}^\circ_{fFu} \quad \Rightarrow \qquad \bar{h}^\circ_{fFu} = \sum_P \nu_i \bar{h}^\circ_{f_i} - \bar{h}^\circ_{RP}$$

Formation enthalpies from Table A.10 and enthalpy of combustion from Table 14.3

$$\bar{h}^\circ_{fFu} = 9\bar{h}^\circ_{fCO_2} + 6\bar{h}^\circ_{fH_2O\ g} - \hat{M}(-41\ 219)_{liq\ Fu\ H_2O\ vap}$$

$$= 9(-393\ 522) + 6(-241\ 826) – 120.195(-41\ 219)$$

$$= \mathbf{-38\ 336\ kJ/kmol}$$

Take the enthalpy of combustion from Table 14.3 for fuel as a gas and as a vapor, the difference is the enthalpy of evaporation

$$h_{fg} = -(h^\circ_{RP\ gas} - h^\circ_{RP\ liq}) = 41\ 603 – 41\ 219 = \mathbf{384\ kJ/kg}$$

14.62

Determine the higher heating value of the sample Wyoming coal as specified in Problem 14.33.

The higher heating value is for liquid water in the products. We need the composition of the products.

Converting from mass analysis:

Substance	S	H_2	C	O_2	N_2
c/M =	0.5/32	3.5/2	4.86/12	12/32	0.7/28
kmol / 100 kg coal	0.0156	1.75	4.05	0.375	0.025
Product	SO_2	H_2O	CO_2		

So the combustion equation becomes (for 100 kg coal)

$$\text{Fuel} + \text{Air} \rightarrow 1.75\ H_2O + 4.05\ CO_2 + 0.0156\ SO_2$$

The formation enthalpies are from Table A.10. Therefore,

$$\bar{h}_{RP0} = H_P^o - H_R^o = 4.05(-393\ 522) + 1.75(-285\ 830) + 0.0156(-296\ 842)$$

$$= -2\ 098\ 597 \text{ kJ/100 kg coal}$$

So that HHV = **+20 986 kJ/kg coal**

14.63

Do problem 14.47 using table 14.3 instead of Table A.10 for the solution.

Another alternative to using petroleum or natural gas as fuels is methanol, CH_3OH, which can be produced from coal. Both methanol and ethanol have been used in automotive engines. Repeat the previous problem using liquid methanol as the fuel instead of ethanol.

$$CH_3OH + 1.2 \times 1.5\,(O_2 + 3.76\,N_2) \;\rightarrow\; 1\,CO_2 + 2\,H_2O + 0.3\,O_2 + 6.77\,N_2$$

Products at 60°C, 100 kPa, so check for condensation of water

$$y_{H2O\,MAX} = \frac{19.94}{100} = \frac{n_{V\,MAX}}{n_{V\,MAX} + 1 + 0.3 + 6.77}$$

$$\Rightarrow \quad n_{V\,MAX} = 2.0 > 2 \qquad \Rightarrow \text{No liquid is formed}$$

Fuel: table 14.3 select (liquid fuel, water vapor) and convert to mole basis

$$\overset{\circ}{H}_{RP} = 32.042\,(-19\,910) = -637\,956 \text{ kJ/kmol}$$

Since the reactants enter at the reference state the energy equation becomes

$$Q_{CV} = H_P - H_R = H_P^o + \Delta H_P - H_R^o = \overset{\circ}{H}_{RP} + \Delta H_P$$

The enthalpies are from Table A.9

$$\Delta H_P = \Delta\bar{h}_{CO_2} + 2\,\Delta\bar{h}_{H_2O} + 0.3\,\Delta\bar{h}_{O_2} + 6.77\,\Delta\bar{h}_{N_2}$$

$$= 1(1327) + 2(1178) + 0.3(1032) + 6.77(1020) = 10\,898 \text{ kJ/kmol}$$

$$Q_{CV} = -637\,956 + 10\,898 = \mathbf{-627\,058\ kJ}$$

14.64

A burner receives a mixture of two fuels with mass fraction 40% n-butane and 60% methanol, both vapor. The fuel is burned with stoichiometric air. Find the product composition and the lower heating value of this fuel mixture (kJ/kg fuel mix).

Since the fuel mixture is specified on a mass basis we need to find the mole fractions for the combustion equation. From Eq.12.4 we get

$$y_{butane} = (0.4/58.124) / [0.4/58.124 + 0.6/32.042] = 0.26875$$

$$y_{methanol} = 1 - y_{butane} = 0.73125$$

The reaction equation is

$$0.73125\ CH_3OH + 0.26875\ C_4H_{10} + \nu_{O2}\ (O_2 + 3.76\ N_2)$$

$$\rightarrow\ \nu_{CO2}\ CO_2 + \nu_{H2O}H_2O + 3.76\ \nu_{O2}\ N_2$$

C balance: $0.73125 + 4 \times 0.26875 = \nu_{CO2} = 1.80625$

H_2 balance: $2 \times 0.73125 + 5 \times 0.26875 = \nu_{H2O} = 2.80625$

O balance: $0.73125 + 2\ \nu_{O2} = 2\ \nu_{CO2} + \nu_{H2O} = 6.41875 \quad => \nu_{O2} = 2.84375$

Now the products are:

$$1.80625\ CO_2 + 2.80625\ H_2O + 10.6925\ N_2$$

Since the enthalpy of combustion is on a mass basis in table 14.3 (this is also the negative of the heating value) we get

LHV = 0.4 × 45 714 + 0.6 × 21 093 = **30 941 kJ/kg fuel mixture**

Notice we took fuel vapor and water as vapor (lower heating value).

14.65

Consider natural gas A and natural gas D, both of which are listed in Table 14.2. Calculate the enthalpy of combustion of each gas at 25°C, assuming that the products include vapor water. Repeat the answer for liquid water in the products.

Natural Gas A

$$0.939\ CH_4 + 0.036\ C_2H_6 + 0.012\ C_3H_8 + 0.013\ C_4H_{10}$$

$$+ 2.1485\ O_2 + 3.76 \times 2.1485\ N_2 \rightarrow 1.099\ CO_2 + 2.099\ H_2O + 8.0784\ N_2$$

$$H_R = 0.939(-74\ 878) + 0.036(-84\ 740) + 0.012(-103\ 900)$$

$$+ 0.013(-126\ 200) = -76244\ \text{kJ}$$

a) vapor H_2O

$$H_P = 1.099(-393\ 522) + 2.099(-241\ 826) = -940\ 074$$

$\bar{h}_{RP} = H_P - H_R =$ **-863 830 kJ/kmol**

b) Liq. H_2O

$$H_P = 1.099(-393\ 522) + 2.099(-285\ 830) = -1\ 032\ 438$$

$\bar{h}_{RP} =$ **-956 194 kJ/kmol**

Natural Gas D:

$$0.543\ CH_4 + 0.163\ C_2H_6 + 0.162\ C_3H_8 + 0.074\ C_4H_{10}$$

$$+ 0.058\ N_2 + O_2 + N_2 \quad \rightarrow \quad 1.651\ CO_2 + 2.593\ H_2O + N_2$$

$$H_R = 0.543(-74\ 873) + 0.163(-84\ 740) + 0.162(-130\ 900)$$

$$+ 0.074(-126\ 200) = -80\ 639\ \text{kJ}$$

a) vapor H_2O

$$H_P = 1.651(-393\ 522) + 2.593(-241\ 826) = -1\ 276\ 760\ \text{kJ}$$

$\bar{h}_{RP} =$ **-1 196 121 kJ/kmol**

b) Liq. H_2O

$$H_P = 1.651(-393\ 522) + 2.593(-285\ 830) = -1\ 390\ 862\ \text{kJ}$$

$\bar{h}_{RP} =$ **-1 310 223 kJ/kmol**

14.66

Blast furnace gas in a steel mill is available at 250°C to be burned for the generation of steam. The composition of this gas is, on a volumetric basis,

Component	CH_4	H_2	CO	CO_2	N_2	H_2O
Percent by volume	0.1	2.4	23.3	14.4	56.4	3.4

Find the lower heating value (kJ/m^3) of this gas at 250°C and ambient pressure.

Of the six components in the gas mixture, only the first 3 contribute to the heating value. These are, per kmol of mixture:

$$0.024\ H_2, \quad 0.001\ CH_4, \quad 0.233\ CO$$

For these components,

$$0.024\ H_2 + 0.001\ CH_4 + 0.233\ CO + 0.1305\ O_2 \rightarrow 0.026\ H_2O + 0.234\ CO_2$$

The remainder need not be included in the calculation, as the contributions to reactants and products cancel. For the lower HV(water as vapor) at 250°C

$$\bar{h}_{RP} = 0.026(-241\ 826 + 7742) + 0.234(-393\ 522 + 9348)$$

$$- 0.024(0 + 6558) - 0.001(-74\ 873 + 2.254 \times 16.04(250-25))$$

$$- 0.233(-110\ 527 + 6625) - 0.1305(0 + 6810)$$

$$= -72\ 573\ \frac{\text{kJ}}{\text{kmol fuel}}$$

$$\bar{v}_0 = \bar{R}\ T_o/P_o = 8.3145 \times 523.2/100 = 43.5015\ \text{m}^3/\text{kmol}$$

LHV = +72 573 / 43.5015 = **1668 kJ/m^3**

14.67

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:

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

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

200% theoretical air: $\nu_{O2} = 2 \times 2 = 4$ 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 products are cooled to 400 K (so we do not consider condensation) and the energy equation is

Energy Eq.: $H_R + Q = H_P = H^{\circ}_P + \Delta H_P = H^{\circ}_R + Q$

$$Q = H^{\circ}_P - H^{\circ}_R + \Delta H_P = H^{\circ}_{RP} + \Delta H_P$$

From Table 14.3: $H^{\circ}_{RP} = 16.04\ (-50\ 010) = -802\ 160$ 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

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

$$Q = H^{\circ}_{RP} + \Delta H_P = -802\ 160 + 61\ 641 = \mathbf{-740\ 519\ kJ/kmol}$$

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

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

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

The mass of water becomes

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

14.68

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:

$$C_7H_{17} + \nu_{O2}\left(O_2 + 3.76\ N_2\right) \rightarrow\bullet\ 8.5\ H_2O + 7\ CO_2 + c\ N_2$$

O balance: $2\ \nu_{O_2} = 8.5 + 14 = 22.5 \quad \Rightarrow \quad \nu_{O_2} = 11.25$

N balance: $c = 3.76\ \nu_{O_2} = 3.76 \times 11.25 = 42.3$

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

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

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

From Table A.9

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

From energy equation and Table 14.3

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

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

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 14.3 for fuel liquid and fuel vapor to get h_{fg}.

14.69

In an engine a mixture of liquid octane and ethanol, mole ratio 9:1, and stoichiometric air are taken in at T_0, P_0. In the engine the enthalpy of combustion is used so that 30% goes out as work, 30% goes out as heat loss and the rest goes out the exhaust. Find the work and heat transfer per kilogram of fuel mixture and also the exhaust temperature.

$$0.9\ C_8H_{18} + 0.1\ C_2H_5OH + 11.55\ O_2 + 43.428\ N_2$$

$$\rightarrow 8.4\ H_2O + 7.4\ CO_2 + 43.428\ N_2$$

For 0.9 octane + 0.1 ethanol, convert to mole basis

$$\bar{H}^\circ_{RP\ mix} = 0.9\ H^\circ_{RP\ C_8H_{18}} + 0.1\ H^\circ_{RP\ C_2H_5OH}$$

$$= 0.9\ (-44\ 425) \times 114.232 + 0.1\ (-26\ 811) \times 46.069$$

$$= -4\ 690\ 797\ \frac{\text{kJ}}{\text{kmol}}$$

$$\hat{M}_{mix} = 0.9\ \hat{M}_{oct} + 0.1\ \hat{M}_{alc} = 107.414$$

Energy: $\bar{h}^\circ_{in} + q_{in} = \bar{h}_{ex} + \omega_{ex} = \bar{h}^\circ_{ex} + \Delta\bar{h}_{ex} + \omega_{ex}$

$$\bar{h}^\circ_{ex} - \bar{h}^\circ_{in} = \bar{H}^\circ_{RP\ mix} \quad \Rightarrow \quad \omega_{ex} + \Delta\bar{h}_{ex} - q_{in} = -\bar{H}^\circ_{RP\ mix}$$

$$\omega_{ex} = -q_{in} = 0.3\left(-\bar{H}^\circ_{RP}\right) = \mathbf{1\ 407\ 239}\ \frac{\mathbf{kJ}}{\mathbf{kmol}} = \mathbf{13\ 101}\ \frac{\mathbf{kJ}}{\mathbf{kg\ Fu}}$$

$$\Delta\bar{h}_{prod} = \Delta\bar{h}_{ex} = 0.4\left(-\bar{H}^\circ_{RP}\right) = 1\ 876\ 319\ \frac{\text{kJ}}{\text{kmol Fu}}$$

$$\Delta\bar{h}_{prod} = 8.4\ \Delta\bar{h}_{H_2O} + 7.4\Delta\bar{h}_{CO_2} + 43.428\ \Delta\bar{h}_{N_2}$$

$$\Delta\bar{h}_{prod\ 1300} = 8.4\times 38\ 941 + 7.4\times 50\ 148 + 43.428\times 31\ 503 = 2\ 066\ 312$$

$$\Delta\bar{h}_{prod\ 1200} = 8.4\times 34\ 506 + 7.4\times 44\ 473 + 43.428\times 28\ 109 = 1\ 839\ 668$$

Linear interpolation to get the value of $\Delta\bar{h}_{prod} = 1\ 876\ 319$

$\Rightarrow$ satisfied for T = **1216 K**

14.70

Liquid nitromethane is added to the air in a carburetor to make a stoichiometric mixture where both fuel and air are added at 298 K, 100 kPa. After combustion a constant pressure heat exchanger brings the products to 600 K before being exhausted. Assume the nitrogen in the fuel becomes N_2 gas. Find the total heat transfer per kmole fuel in the whole process.

$$CH_3NO_2 + \ \nu_{O2}\,(O_2 + 3.76\ N_2) \ \rightarrow\ 1.5\ H_2O + 1\ CO_2 + a\ N_2$$

C and H balances done in equation. The remaining

O balance: $2 + 2\ \nu_{O2} = 1.5 + 2 \quad \Rightarrow \quad \nu_{O2} = 0.75$

N balance: $1 + 3.76\ \nu_{O2} \times 2 = 2a \quad \Rightarrow \quad a = 3.32$

Energy eq.: $H_R + Q = H_P \quad \Rightarrow \quad Q = H_P - H_R = H^\circ_P - H^\circ_R + \Delta H_P - \Delta H_R$

The reactants enter at the reference state, $\Delta H_R = 0$, and the products at 600 K from table A.9

$$\Delta H_P = 1.5\ \Delta\bar{h}_{H2O} + \Delta\bar{h}_{CO_2} + 3.32\ \Delta\bar{h}_{N2}$$

$$= 1.5\ (10\ 499) + 1\ (12\ 906) + 3.32\ (8894) = 58\ 183 \text{ kJ/kmol fuel}$$

$$H^\circ_P - H^\circ_R = H^\circ_{RP} = 61.04\ (-10\ 537) = -643\ 178 \text{ kJ/kmol}$$

$Q = -643\ 178 + 58\ 183 =$ **$-584\ 995$ kJ/kmol fuel**

Adiabatic Flame Temperature

14.71

Hydrogen gas is burned with pure oxygen in a steady flow burner where both reactants are supplied in a stoichiometric ratio at the reference pressure and temperature. What is the adiabatic flame temperature?

The reaction equation is:

$$H_2 + \nu_{O2}\, O_2 \quad \Rightarrow \quad H_2O$$

The balance of hydrogen is done, now for oxygen we need $\nu_{O2} = 0.5$.

Energy Eq.: $H_R = H_P \;\Rightarrow\; 0 = -241\,826 + \Delta\bar{h}_{H2O}$

$$\Rightarrow \quad \Delta\bar{h}_{H2O} = 241\,826 \text{ kJ/kmol}$$

Interpolate now in table A.9 for the temperature to give this enthalpy

T = 4991 K

14.72

In a rocket, hydrogen is burned with air, both reactants supplied as gases at P_o, T_o. The combustion is adiabatic and the mixture is stoichiometeric (100% theoretical air). Find the products dew point and the adiabatic flame temperature (~2500 K).

The reaction equation is:

$$H_2 + \nu_{O2} (O_2 + 3.76 N_2) \Rightarrow H_2O + 3.76 \nu_{O2} N_2$$

The balance of hydrogen is done, now for oxygen we need $\nu_{O2} = 0.5$ and thus we have 1.88 for nitrogen.

$y_v = 1/(1+1.88) = 0.3472 \Rightarrow P_v = 101.325 \times 0.3472 = 35.18 \text{ kPa} = P_g$

Table B.1.2: $\mathbf{T_{dew} = 72.6\ C.}$

$H_R = H_P \Rightarrow 0 = -241826 + \Delta h_{water} + 1.88\ \Delta h_{nitrogen}$

Find now from table A.9 the two enthalpy terms

At 2400 K : $\Delta H_P = 93741 + 1.88 \times 70640 = 226544$ kJ/kmol fuel

At 2600 K : $\Delta H_P = 104520 + 1.88 \times 77963 = 251090$ kJ/kmol fuel

Then interpolate to hit 241 826 to give **T = 2525 K**

14.73

Carbon is burned with air in a furnace with 150% theoretical air and both reactants are supplied at the reference pressure and temperature. What is the adiabatic flame temperature?

$$C + \nu_{O2}O_2 + 3.76\ \nu_{O2}N_2 \ \rightarrow\ 1\ CO_2 + 3.76\ \nu_{O2}N_2$$

From this we find $\nu_{O2} = 1$ and the actual combustion reaction is

$$C + 1.5\ O_2 + 5.64\ N_2 \ \rightarrow\ 1\ CO_2 + 5.64\ N_2 + 0.5\ O_2$$

$$H_P = H^{\circ}_P + \Delta H_P = H_R = H^{\circ}_R \quad \Rightarrow$$

$$\Delta H_P = H^{\circ}_R - H^{\circ}_P = 0\ - (-393\ 522) = 393\ 522\ \text{kJ/kmol}$$

$$\Delta H_P = \ \Delta\bar{h}_{CO2} + 5.64\ \Delta\bar{h}_{N2} + 0.5\ \Delta\bar{h}_{O2}$$

Find T so ΔH_P takes on the required value. To start guessing assume all products are nitrogen (1 + 5.64 + 0.5 = 7.14) that gives 1900 < T < 2000 K from Table A.9.

$$\Delta H_{P\ 1900} = 85\ 420 + 5.64 \times\bullet 52\ 549 + 0.5\bullet\times\bullet 55\ 414 = 409\ 503 \quad \text{too high}$$

$$\Delta H_{P\ 1800} = 79\ 432 + 5.64 \times\bullet 48\ 979 + 0.5\bullet\times\bullet 51\ 674 = 381\ 511$$

Linear interpolation to find

$$T = 1800 + 100\ \frac{393\ 522 - 381\ 511}{409\ 503 - 381\ 511} = \mathbf{1843\ K}$$

14.74

A stoichiometric mixture of benzene, C_6H_6, and air is mixed from the reactants flowing at 25°C, 100 kPa. Find the adiabatic flame temperature. What is the error if constant specific heat at T_0 for the products from Table A.5 are used?

$$C_6H_6 + \nu_{O2}O_2 + 3.76\ \nu_{O2}N_2 \rightarrow 6CO_2 + 3H_2O + 3.76\ \nu_{O2}N_2$$

$$\nu_{O2} = 6 + 3/2 = 7.5 \ \Rightarrow\ \nu_{N2} = 28.2$$

$$H_P = H^{\circ}_P + \Delta H_P = H_R = H^{\circ}_R \quad \Rightarrow$$

$$\Delta H_P = -H^{\circ}_{RP} = 40576 \times 78.114 = 3\ 169\ 554 \text{ kJ/kmol}$$

$$\Delta H_P = 6\ \Delta\bar{h}_{CO2} + 3\ \Delta\bar{h}_{H2O} + 28.2\ \Delta\bar{h}_{N2}$$

$$\Delta H_{P\ 2600K} = 6(128074) + 3(104\ 520) + 28.2(77\ 963) = 3\ 280\ 600,$$

$$\Delta H_{P\ 2400K} = 6(115\ 779) + 3(93\ 741) + 28.2(70\ 640) = 2\ 968\ 000$$

Linear interpolation $\Rightarrow$ T_{AD} = **2529 K**

$$\sum \nu_i \bar{C}_{Pi} = 6 \times 0.842 \times 44.01 + 3 \times 1.872 \times 18.015 + 28.2 \times 1.042 \times 28.013$$

$$= 1146.66 \text{ kJ/kmol K}$$

$$\Delta T = \Delta H_P / \sum \nu_i \bar{C}_{Pi} = 3\ 169\ 554 / 1146.66 = 2764$$

$\Rightarrow$ T_{AD} = 3062 K, **21% high**

14.75

Hydrogen gas is burned with 200% theoretical air in a steady flow burner where both reactants are supplied at the reference pressure and temperature. What is the adiabatic flame temperature?

The stoichiometric reaction equation is:

$$H_2 + \nu_{O2}\,(O_2 + 3.76\,N_2) \;\Rightarrow\; H_2O + 3.76\,\nu_{O2}\,N_2$$

The balance of hydrogen is done, now for oxygen we need $\nu_{O2} = 0.5$ and thus we have for the actual mixture $\nu_{O2} = 1$. The actual reaction is

$$H_2 + 1\,(O_2 + 3.76\,N_2) \;\Rightarrow\; 1\,H_2O + 3.76\,N_2 + 0.5\,O_2$$

The energy equation with formation enthalpy from A.9 or A.10 for water is

$$H_R = H_P \;\Rightarrow\; 0 = -241\,826 + \Delta h_{H2O} + 3.76\,\Delta h_{N2} + 0.5\,\Delta h_{O2}$$

Find now from table A.9 the two enthalpy terms

At 2000 K : $\Delta H_P = 72\,788 + 3.76 \times 56\,137 + 0.5 \times 59\,176 = 313\,451$

At 1800 K : $\Delta H_P = 62\,693 + 3.76 \times 48\,979 + 0.5 \times 51\,674 = 272\,691$

At 1600 K : $\Delta H_P = 52\,907 + 3.76 \times 41\,904 + 0.5 \times 44\,267 = 232\,600$

At 1700 K : $\Delta H_P = 57\,757 + 3.76 \times 45\,430 + 0.5 \times 47\,959 = 252\,553$

Then interpolate to hit 241 826 to give

$$T = 1600 + 100\,\frac{241\,826 - 232\,600}{252\,553 - 232\,600} = \mathbf{1646\ K}$$

14.76

A gasturbine burns natural gas (assume methane) where the air is supplied to the combustor at 1000 kPa, 500 K and the fuel is at 298 K, 1000 kPa. What is the equivalence ratio and the percent theoretical air if the adiabatic flame temperature should be limited to 1800 K?

The reaction equation for a mixture with excess air is:

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

$$H_P = \overset{o}{H}_P + \Delta H_P = H_R = \overset{o}{H}_R + \Delta H_R$$

From table A.9 at 500 K (notice fuel is at 298 K)

$$\Delta H_R = 0 + \nu_{O2}(\Delta h_{O2} + 3.76\ \Delta h_{N2}) = \nu_{O2}(6086 + 3.76 \times 5911) = 28\ 311.4\ \nu_{O2}$$

From table A.9 at 1800 K:

$$\Delta H_P = 2\ \Delta h_{H2O} + \Delta h_{CO2} + 3.76\ \nu_{O2}\ \Delta h_{N2} + (\nu_{O2} - 2)\ \Delta h_{O2}$$

$$= 2 \times 62\ 693 + 79432 + 3.76\ \nu_{O2} \times 48\ 979 + (\nu_{O2} - 2)\ 51\ 674$$

$$= 101\ 470 + 235\ 835\ \nu_{O2}$$

From table 14.3: $\overset{o}{H}_P - \overset{o}{H}_R = \overset{o}{H}_{RP} = 16.04(-50\ 010) = -802\ 160$ kJ/kmol

Now substitute all terms into the energy equation

$$-802\ 160 + 101\ 470 + 235\ 835\ \nu_{O2} = 28\ 311.4\ \nu_{O2}$$

Solve for ν_{O2}

$$\nu_{O2} = \frac{802\ 160 - 101\ 470}{235\ 835 - 28\ 311.4} = 3.376$$

%Theoretical air = 100 (3.376 / 2) = **168.8 %**

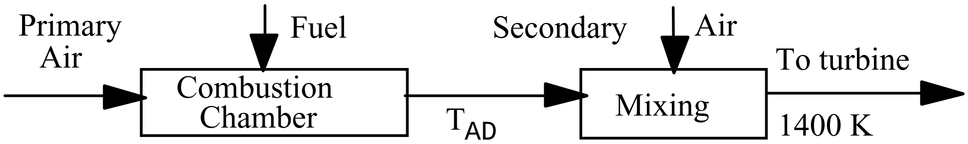
Φ = AFs / AF = 2 / 3.376 = **0.592**

14.77

Liquid *n*-butane at T_0, is sprayed into a gas turbine with primary air flowing at 1.0 MPa, 400 K in a stoichiometric ratio. After complete combustion, the products are at the adiabatic flame temperature, which is too high, so secondary air at 1.0 MPa, 400 K is added, with the resulting mixture being at 1400 K. Show that $T_{ad} >$ 1400 K and find the ratio of secondary to primary air flow.

C.V. Combustion Chamber.

$$C_4H_{10} + 6.5\ O_2 + 6.5 \times 3.76\ N_2 \rightarrow 5\ H_2O + 4\ CO_2 + 24.44\ N_2$$



Energy Eq.: $H_{air} + H_{fuel} = H_R = H_P$

$\overset{o}{H}_P + \Delta H_P = \overset{o}{H}_R + \Delta H_R \quad \Rightarrow \quad \Delta H_P = \overset{o}{H}_R + \Delta H_R - \overset{o}{H}_P = -\overset{o}{H}_{RP} + \Delta H_R$

$\Delta H_P = 45344 \times 58.124 + 6.5(3.76 \times 2971 + 3027) =$ **2 727 861 kJ/kmol**

$\Delta H_{P\ 1400} = 5 \times 43491 + 4 \times 55895 + 24.44 \times 34936 =$ **1 294 871 <** $\mathbf{\Delta H_P}$

Try $T_{AD} > 1400$: $\Delta H_P = 2658263$ @2400 K, $\Delta H_P = 2940312$ @2600 K

C.V. Mixing Chamber. Air Second: $\nu_{O2\ s}O_2 + 3.76\ N_2$

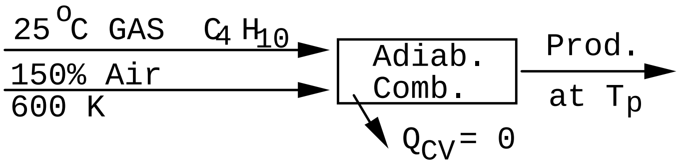
$$\Delta H_P + \nu_{O2\ second}\,\Delta H_{air} = \Delta H_{P\ 1400} + \nu_{O2\ second}\,\Delta H_{air\ 1400}$$

$$\Rightarrow \nu_{O2\ second} = \frac{\Delta H_P - \Delta H_{P\ 1400}}{\Delta H_{air\ 1400} - \Delta H_{air\ 400}} = \frac{1432990}{168317 - 14198} = \mathbf{9.3}$$

ratio = $\nu_{O2\ sec}/\nu_{O2\ prim} = 9.3/6.5 =$ **1.43**

14.78

Butane gas at 25°C is mixed with 150% theoretical air at 600 K and is burned in an adiabatic steady flow combustor. What is the temperature of the products exiting the combustor?



$$C_4H_{10} + 1.5\times6.5\left(O_2 + 3.76\ N_2\right) \rightarrow 4\ CO_2 + 5\ H_2O + 3.25\ O_2 + 36.66\ N_2$$

Energy Eq.: $H_P - H_R = 0 \quad \Rightarrow \quad \Delta H_P = H^\circ_R + \Delta H_R - H^\circ_P$

Reactants: $\Delta H_R = 9.75(9245) + 36.66(8894) = 416\ 193$ kJ ;

$$H^\circ_R = \bar{h}_{C4H10} = \bar{h}^\circ_{f\ IG} = -126\ 200 \text{ kJ} \quad \Rightarrow \quad H_R = +289\ 993 \text{ kJ}$$

$$H^\circ_P = 4(-393522) + 5(-241826) = -2\ 783\ 218 \text{ kJ/kmol}$$

$$\Delta H_P = 4\ \Delta\bar{h}^*_{CO2} + 5\ \Delta\bar{h}^*_{H2O} + 3.25\ \Delta\bar{h}^*_{O2} + 36.66\ \Delta\bar{h}^*_{N2}$$

From the energy equation we then get

$$\Delta H_P = -126\ 200 + 416\ 193 - (-2\ 783\ 218) = 3\ 073\ 211 \text{ kJ/kmol}$$

Trial and Error: $LHS_{2000\ K} = 2\ 980\ 000$, $LHS_{2200\ K} = 3\ 369\ 866$

Linear interpolation to match RHS $\Rightarrow$ $T_P =$ **2048 K**

14.79

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 and then out the exhaust, as in Fig. P14.79. What is the adiabatic flame temperature right after combustion before the heat exchanger?

The reaction equation for stoichiometric mixture is:

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

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

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

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

Energy Eq.: $H_{air} + H_{fuel} = H_R = H_P$

$$\overset{o}{H}_P + \Delta H_P = \overset{o}{H}_R + \Delta H_R \quad \Rightarrow \quad \Delta H_P = \overset{o}{H}_R + \Delta H_R - \overset{o}{H}_P = -\overset{o}{H}_{RP} + 0$$

From Table 14.3: $-\overset{o}{H}_{RP} = -16.04\ (-50\ 010) = 802\ 160$ 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

$$\Delta H_{P\ 1600} = 67\ 659 + 2 \times 52\ 907 + 2 \times 44\ 267 + 15.04 \times 41\ 904 = 892\ 243$$

$$\Delta H_{P\ 1500} = 61\ 705 + 2 \times 48\ 149 + 2 \times 40\ 600 + 15.04 \times 38\ 405 = 816\ 814$$

$$\Delta H_{P\ 1400} = 55\ 895 + 2 \times 43\ 491 + 2 \times 36\ 958 + 15.04 \times 34\ 936 = 742\ 230$$

Linear interpolation to get 802 160

$$T = 1400 + 100 \frac{802\ 160 - 742\ 230}{816\ 814 - 742\ 230} = \mathbf{1480\ K}$$

14.80

Liquid butane at 25°C is mixed with 150% theoretical air at 600 K and is burned in a steady flow burner. Use the enthalpy of combustion from Table 14.3 to find the adiabatic flame temperature out of the burner.

$$C_4H_{10} + 1.5\times6.5\left(O_2 + 3.76\ N_2\right) \rightarrow 4\ CO_2 + 5\ H_2O + 3.25\ O_2 + 36.66\ N_2$$

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

Reactants: $\Delta H_R = 9.75(9245) + 36.66(8894) = 416\ 193$ kJ/kmol;

$$H^{\circ}_{RP} = 58.124(-45\ 344) = -2\ 635\ 575 \text{ kJ/kmol}$$

$$\Delta H_P = 4\Delta\bar{h}^*_{CO2} + 5\Delta\bar{h}^*_{H2O} + 3.25\ \Delta\bar{h}^*_{O2} + 36.66\ \Delta\bar{h}^*_{N2}$$

So the energy equation becomes

$$\Delta H_P = 2\ 635\ 575 + 416\ 193 = 3\ 051\ 768 \text{ kJ/kmol}$$

Trial and Error: $LHS_{2000\ K} = 2\ 980\ 000$, $\quad LHS_{2200\ K} = 3\ 369\ 866$

Linear interpolation to match RHS => $T_P =$ **2037 K**

14.81

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

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

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

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

Linear interpolation: $T_{PROD} =$ **2909 K**

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

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

$\Rightarrow\; 2\,\Delta\bar{h}^*_{CO2} + 1\,\Delta\bar{h}^*_{H2O} = 1\,255\,601$ kJ/kmol fuel

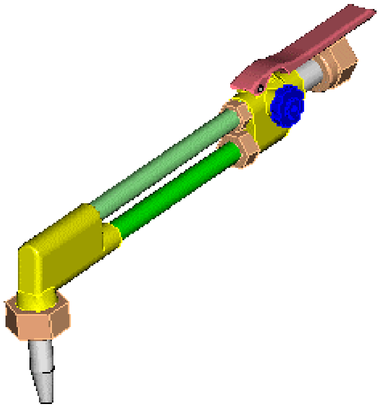
At 6000 K (limit of A.9) $2 \times 343\,782 + 302\,295 = 989\,859$

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

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



14.82

Ethene, C_2H_4, burns with 150% theoretical air in a steady flow constant-pressure process with reactants entering at P_0, T_0. Find the adiabatic flame temperature.

Stoichiometric

$$C_2H_4 + 3(O_2 + 3.76N_2) \rightarrow 2CO_2 + 2H_2O + 11.28N_2$$

Actual

$$C_2H_4 + 4.5(O_2 + 3.76N_2) \rightarrow 2CO_2 + 2H_2O + 1.5\ O_2 + 16.92N_2$$

$$H_P = H^\circ_P + 2\Delta\bar{h}_{CO_2} + 2\Delta\bar{h}_{H_2O} + 1.5\Delta\bar{h}_{O_2} + 16.92\Delta\bar{h}_{N_2}$$

$$H^\circ_R = \bar{h}^\circ_{f\,Fu} \qquad\qquad \Delta H_P + H^\circ_P = H^\circ_R$$

$$\Rightarrow \Delta H_P = -H^\circ_{RP} = 28.054 \times 47158 = 1\ 322\ 970.5\ \frac{kJ}{kmol\ Fu}$$

$$\Delta H_P = 2\Delta\bar{h}_{CO_2} + 2\Delta\bar{h}_{H_2O} + 1.5\Delta\bar{h}_{O_2} + 16.92\Delta\bar{h}_{N_2}$$

Initial guess based on (2+2+1.5+16.92) N_2 from A.9: $T_1 = 2100$ K

$$\Delta H_P(2000) = 1\ 366\ 982, \quad \Delta H_P(1900) = 1\ 278\ 398$$

=> $\mathbf{T_{AD} \cong 1950\ K}$

14.83

Solid carbon is burned with stoichiometric air in a steady flow process. The reactants at T_0, P_0 are heated in a preheater to T_2 = 500 K as shown in Fig. P14.83, with the energy given by the product gases before flowing to a second heat exchanger, which they leave at T_0. Find the temperature of the products T_4, and the heat transfer per kmol of fuel (4 to 5) in the second heat exchanger.

Control volume: Total minus last heat exchanger.

$$C + O_2 + 3.76\ N_2 \rightarrow CO_2 + 3.76\ N_2$$

C.V. Combustion chamber and preheater from 1 to 4, no external Q. For this CV states 2 and 3 are internal and do not appear in equations.

Energy Eq.:

$$H_R = H^{\circ}_R = H_{P_4} = H^{\circ}_P + \Delta H_{P4} = \bar{h}_{f\,CO_2} + \Delta\bar{h}_{CO_2} + 3.76\Delta\bar{h}_{N_2}$$

Table A.9 or A.10: $\bar{h}_{f\,CO_2}$= -393 522 kJ/kmol,

$$\Delta H_{P4\,2400} = 115\ 779 + 3.76 \times 70\ 640 = 381\ 385 \text{ kJ/kmol fuel},$$

$$\Delta H_{P4\,2600} = 128\ 074 + 3.76 \times 77\ 963 = 421\ 215 \text{ kJ/kmol fuel}$$

$$\Rightarrow \quad T_4 = T_{ad.flame} = \mathbf{2461\ K}$$

Control volume: Total. Then energy equation:

$$H^{\circ}_R + \bar{Q} = H^{\circ}_P$$

$$\bar{Q} = \bar{H}^{\circ}_{RP} = \bar{h}^{\circ}_{f\,CO_2} - 0 = \mathbf{-393\ 522\ \frac{kJ}{kmol\ fuel}}$$

14.84

Gaseous ethanol, C_2H_5OH, is burned with pure oxygen in a constant volume combustion bomb. The reactants are charged in a stoichiometric ratio at the reference condition. Assume no heat transfer and find the final temperature (> 5000 K).

$$C_2H_5OH + 3\ O_2 \;\rightarrow\; 2\ CO_2 + 3\ H_2O$$

Energy Eq.:

$$U_P = U_R = \overset{\circ}{H}_R + \Delta H_R - n_R\bar{R}T_R = \overset{\circ}{H}_P + \Delta H_P - n_P\bar{R}T_P$$

Solve for the properties that depends on T_P and recall $\Delta H_R = 0$

$$\Delta H_P - n_P\bar{R}T_P = \overset{\circ}{H}_R - \overset{\circ}{H}_P - n_R\bar{R}T_R = \bar{h}^0_{f\,\text{fuel}} - 2\ \bar{h}^0_{f\,CO2} - 3\ \bar{h}^0_{f\,H2O} - 4\bar{R}T_R$$

Fuel: $\bar{h}^0_{f\,\text{fuel}}$ = -235 000 kJ/kmol for IG from Table A.10 so

$$\Delta H_P - n_P\bar{R}T_P = -235\ 000 - 2(-393\ 522) - 3(-241\ 826)$$

$$- 4 \times 8.31451 \times 298.15 = 1\ 267\ 606 \text{ kJ/kmol}$$

$$\text{LHS} = \Delta H_P - n_P\bar{R}T_P = 2\ \Delta\bar{h}_{CO_2} + 3\ \Delta\bar{h}_{H_2O} - 5 \times 8.31451 \times T_P$$

From Table A.9 we find

$$\text{LHS}_{5600} = 2 \times 317\ 870 + 3 \times 278\ 161 - 41.5726 \times 5600 = 1\ 237\ 417$$

$$\text{LHS}_{6000} = 2 \times 343\ 782 + 3 \times 302\ 295 - 41.5726 \times 6000 = 1\ 345\ 014$$

$$T_{ad.flame} = \mathbf{5712\ K}$$

14.85

The enthalpy of formation of magnesium oxide, MgO(s), is −601827 kJ/kmol at 25°C. The melting point of magnesium oxide is approximately 3000 K, and the increase in enthalpy between 298 and 3000 K is 128 449 kJ/kmol. The enthalpy of sublimation at 3000 K is estimated at 418 000 kJ/kmol, and the specific heat of magnesium oxide vapor above 3000 K is estimated at 37.24 kJ/kmol K.

a. Determine the enthalpy of combustion per kilogram of magnesium.

b. Estimate the adiabatic flame temperature when magnesium is burned with theoretical oxygen.

a) $\mathrm{Mg} + \frac{1}{2}\mathrm{O_2} \rightarrow \mathrm{MgO(s)}$

$$\Delta h_{COMB} = \Delta\bar{h}_{COMB}/M = \bar{h}^{o}_{f}/M = -601827/24.32 = \mathbf{-24746\ kJ/kg}$$

b) assume T_R = 25°C and also that T_P > 3000 K, (MgO = vapor phase)

1st law: $Q_{CV} = H_P - H_R = 0, \quad \text{but } H_R = 0$

$$\Rightarrow H_P = \bar{h}^{o}_{f} + (\bar{h}_{3000} - \bar{h}_{298})_{SOL} + \Delta\bar{h}_{SUB} + \bar{C}_{P\,VAP}(T_P - 3000)$$

$$= -601827 + 128449 + 418000 + 37.24(T_P - 3000) = 0$$

Solving, $T_P = \mathbf{4487\ K}$

Second Law for the Combustion Process

14.86

Calculate the irreversibility for the process described in Problem 14.41.

$$2\,C + 2\,O_2 \;\rightarrow 1\,CO_2 + 1\,CO + \tfrac{1}{2}O_2$$

Process V = constant, C: solid, $n_{1(GAS)} = 2$, $n_{2(GAS)} = 2.5$

$$P_2 = P_1 \times \frac{n_2T_2}{n_1T_1} = 200 \times \frac{2.5 \times 1000}{2 \times 298.2} = 838.4 \text{ kPa}$$

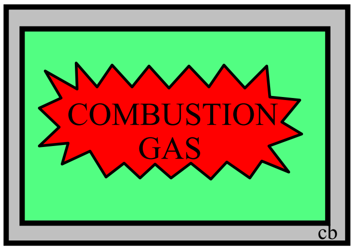
$$H_1 = H_R = 0$$

$$H_2 = H_P = 1(-393522 + 33397) + 1(-110527 + 21686)$$

$$+ (1/2)(0 + 22703) = -437\,615 \text{ kJ}$$

$${}_1Q_2 = (U_2 - U_1) = (H_2 - H_1) - n_2\bar{R}T_2 + n_1\bar{R}T_1$$

$$= (-437\,615 - 0) - 8.3145(2.5 \times 1000 - 2 \times 298.2) = -453\,442 \text{ kJ}$$



Reactants:

$$S_R = 2(5.740) + 2(205.148 - 8.31451 \ln\frac{200}{100}) = 410.250 \text{ kJ/K}$$

Products:

	n_i	y_i	$\bar{s}^\circ_i$	$-\bar{R}\ln\frac{y_iP}{P_0}$	S_i
CO_2	1.0	0.40	269.299	-10.061	259.238
CO	1.0	0.40	234.538	-10.061	224.477
O_2	0.5	0.20	243.579	-4.298	239.281

$$S_P = 1.0(259.238) + 1.0(224.477) + 0.5(239.281) = 603.355 \text{ kJ/K}$$

$$I = T_0(S_P - S_R) - {}_1Q_2$$

$$= 298.15(603.355 - 410.250) - (-453\,442) = \mathbf{+511\,016 \text{ kJ}}$$

14.87

Methane is burned with air, both of which are supplied at the reference conditions. There is enough excess air to give a flame temperature of 1800 K. What are the percent theoretical air and the irreversibility in the process?

The combustion equation with X times theoretical air is

$$CH_4 + 2X(O_2 + 3.76\ N_2) \rightarrow CO_2 + 2H_2O + 2(X-1)O_2 + 7.52X\ N_2$$

Energy Eq.: $H_{air} + H_{fuel} = H_R = H_P = \overset{o}{H}_P + \Delta H_P = \overset{o}{H}_R + \Delta H_R$

$$\Rightarrow \quad \Delta H_P = \overset{o}{H}_R + \Delta H_R - \overset{o}{H}_P = -\overset{o}{H}_{RP} + 0$$

From Table 14.3: $-\overset{o}{H}_{RP} = -16.04\ (-50\ 010) = 802\ 160$ kJ/kmol

$$\Delta H_P = \Delta\bar{h}^*_{CO2} + 2\ \Delta\bar{h}^*_{H2O} + 2(X-1)\ \Delta\bar{h}^*_{O2} + 7.52X\ \Delta\bar{h}^*_{N2}$$

From Table A.9 and the energy equation

$$\Delta H_{P\ 1800} = 79\ 432 + 2 \times 62\ 693 + 2(X-1)\ 51\ 674 + 7.52X \times 48\ 979 = 802\ 160$$

so

$$101\ 470 + 471\ 670\ X = 802\ 160 \quad => \quad X = 1.4856$$

%Theoretical air = **148.6%**

The products are

Products: $CO_2 + 2H_2O + 0.9712\ O_2 + 11.172\ N_2$

The second law

$$S_{gen} = S_P - S_R \quad \text{and} \quad I = T_o\ S_{gen}$$

Reactants: $P_i = 100$ kPa, $P_o = 100$ kPa, $\bar{s}^o_f$ from Table A.9

	n_i	y_i	$\bar{s}^o_f$	$-\bar{R}\ln\frac{y_iP_i}{P_o}$	$\bar{S}_i\ \frac{kJ}{kmol\ K}$
CH_4	1	1	186.251	0	186.251
O_2	2X	0.21	205.148	12.976	218.124
N_2	7.52 X	0.79	191.609	1.96	193.569

$$S_R = \sum n_i\bar{S}_i = 2996.84 \text{ kJ/K kmol fuel}$$

Products: $P_e = 100$ kPa, $P_o = 100$ kPa

	n_i	y_i	$\bar{s}^o_{1800}$	$- \bar{R} \ln \frac{y_i P_e}{P_o}$	$\bar{S}_i$ $\frac{\text{kJ}}{\text{kmol K}}$
CO_2	1	0.06604	302.969	22.595	325.564
H_2O	2	0.13208	259.452	16.831	276.283
O_2	0.9712	0.06413	264.797	22.838	287.635
N_2	11.172	0.73775	248.304	2.529	250.833

$S_P = \sum n_i \bar{S}_i = 3959.72$ kJ/K kmol fuel;

$I = T_o(S_P - S_R) = 298.15(3959.72 - 2996.84) =$ **287 MJ/kmol fuel**

14.88

Consider the combustion of hydrogen with pure oxygen in a stoichiometric ratio under steady flow adiabatic conditions. The reactants enter separately at 298 K, 100 kPa and the product(s) exit at a pressure of 100 kPa. What is the exit temperature and what is the irreversibility?

The reaction equation is:

$$H_2 + \nu_{O2}\, O_2 \;\; => \;\; H_2O$$

The balance of hydrogen is done, now for oxygen we need $\nu_{O2} = 0.5$.

Energy Eq.: $H_R = H_P \;=>\; 0 = -241\,826 + \Delta\bar{h}_{H2O}$

$$=> \quad \Delta\bar{h}_{H2O} = 241\,826 \text{ kJ/kmol}$$

Interpolate now in table A.9 for the temperature to give this enthalpy

T = 4991 K

For this temperature we find from Table A.9, $P = P_o$, so we do not need any pressure correction for the entropy

$$S_P = S^{\circ}_P = \bar{s}^{\circ}_{H2O} = 315.848 \text{ kJ/kmol K}$$

For the reactants we have (again no pressure correction)

$$S_R = \bar{s}^{\circ}_{H2} + 0.5\, \bar{s}^{\circ}_{O2} = 130.678 + 0.5 \times 205.148 = 233.252 \text{ kJ/kmol K}$$

$$S_{gen} = S_P - S_R = 315.848 - 233.252 = 82.596 \text{ kJ/kmol } H_2 \text{ K}$$

$$I = T_o\, S_{gen} = 298.15 \times 82.596 = \mathbf{24\,626 \text{ kJ/kmol } H_2}$$

14.89

Pentane gas at 25°C, 150 kPa enters an insulated steady flow combustion chamber. Sufficient excess air to hold the combustion products temperature to 1800 K enters separately at 500 K, 150 kPa. Calculate the percent theoretical air required and the irreversibility of the process per kmol of pentane burned.

$$C_5H_{12} + 8X(O_2 + 3.76\ N_2) \rightarrow 5\ CO_2 + 6\ H_2O + 8(X-1)\ O_2 + 30.08X\ N_2$$

Energy Eq.: $Q_{CV} + H_R = H_P + W_{CV};\ W_{CV} = 0,\ Q_{CV} = 0$

Reactants: C_5H_{12} : $\bar{h}_f^o$ from A.9 and $\Delta\bar{h}_{500}$ for O_2 and N_2 from A.9

$$H_R = \left(\bar{h}_f^o\right)_{C5H12} + 8X\ \Delta\bar{h}_{O2} + 30.08X\ \Delta\bar{h}_{N2}$$

$$= -146\ 500 + 8X\ 6086 + 30.08\ X\ 5911 = 226\ 491\ X - 146\ 500$$

$$H_P = 5\left(\bar{h}_f^o + \Delta\bar{h}\right)_{CO2} + 6\left(\bar{h}_f^o + \Delta\bar{h}\right)_{H2O} + 8(X-1)\ \Delta\bar{h}_{O2} + 30.08\ X\ \Delta\bar{h}_{N2}$$

$$= 5(-393\ 522 + 79\ 432) + 6(-241\ 826 + 62\ 693) + 8(X-1)\ 51\ 674$$

$$+ 30.08\ X\ 48\ 979 = 1\ 886\ 680\ X - 3\ 058\ 640$$

Energy Eq. solve for X;

$$H_R = H_P = 226\ 491\ X - 146\ 500 = 1\ 886\ 680\ X - 3\ 058\ 640$$

$$\Rightarrow\quad X = \mathbf{1.754}$$

b) Reactants: $P_i = 150$ kPa, $P_o = 100$ kPa, $\bar{s}_f^o$

	n_i	y_i	$\bar{s}_f^o,\ \bar{s}_{500}^o$	$-\bar{R}\ln\frac{y_iP_i}{P_o}$	$\bar{S}_i\ \frac{kJ}{kmol\ K}$
C_5H_{12}	1	1	348.945	-3.371	345.574
O_2	8X	0.21	220.693	9.605	230.298
N_2	30.08 X	0.79	206.74	-1.411	205.329

$$S_R = \sum n_i\bar{S}_i = 14410.34\ \text{kJ/K kmol fuel}$$

Products: $P_e = 150$ kPa, $P_o = 100$ kPa

	n_i	y_i	$\bar{s}^o_{1800}$	$-\bar{R}\ln\frac{y_iP_e}{P_o}$	$\bar{S}_i$ $\frac{kJ}{kmol\ K}$
CO_2	5	0.0716	302.969	18.550	321.519
H_2O	6	0.086	259.452	17.027	276.479
O_2	8(X-1)	0.0864	264.797	16.988	281.785
N_2	30.08X	0.756	248.304	-1.045	247.259

$$S_P = \sum n_i \bar{S}_i = 17\,732.073 \text{ kJ/K kmol fuel;}$$

$$I = T_O(S_P - S_R) = 298.15(17\,732.07 - 14\,410.34)$$

$= \mathbf{990\ MJ/kmol\ fuel}$

14.90

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.

$$CH_3OH + 1.25 \times 1.5 \left(O_2 + 3.76\, N_2\right) \rightarrow CO_2 + 2\, H_2O + 0.375\, O_2 + 7.05\, N_2$$

We need to find the mole fractions to do the partial pressures,

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

Gas mixture:

	n_i	y_i	$\bar{s}_i^\circ$	$-\bar{R}\ln\frac{y_iP}{P_0}$	$\bar{S}_i$
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

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

14.91

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, 40°C. Calculate the absolute entropy of the products exiting the heat exchanger per kilomole of methanol burned, using the proper amounts of liquid and vapor for the water.

$$CH_3OH + 1.25 \times 1.5\left(O_2 + 3.76\,N_2\right) \rightarrow CO_2 + 2\,H_2O + 0.375\,O_2 + 7.05\,N_2$$

Products exit at 40 °C, 200 kPa, check for saturation:

$$y_{V\,MAX} = \frac{P_G}{P} = \frac{7.384}{200} = \frac{n_{V\,MAX}}{n_{V\,MAX} + 1 + 0.375 + 7.05}$$

$$n_V = n_{V\,MAX} = 0.323 \quad n_{LIQ} = 1.677$$

Gas mixture:

	n_i	y_i	$\bar{s}_i^\circ$	$-\bar{R}\ln\frac{y_iP}{P_0}$	$\bar{S}_i$
CO_2	1.0	0.1143	215.633	+12.270	227.903
H_2O	0.323	0.0369	190.485	+21.671	212.156
O_2	0.375	0.0429	206.592	+20.418	227.01
N_2	7.05	0.8059	193.039	-3.969	189.07

$$S_{GAS\,MIX} = \sum n_i\bar{S}_i = 1714.50 \text{ kJ/K kmol fuel}$$

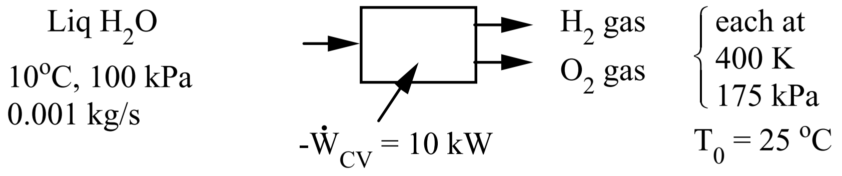
$$\bar{s}_{LIQ} = 69.950 + 18.015(0.5725 - 0.3674) = 73.645 \text{ kJ/kmol}$$

$$S_{LIQ} = 1.677 \times 73.645 = 123.50 \text{ kJ/K kmol fuel}$$

$$S_{PROD} = 1714.50 + 123.50 = \mathbf{1838\ kJ/K\ kmol\ fuel}$$

14.92

An inventor claims to have built a device that will take 0.001 kg/s of water from the faucet at 10°C, 100 kPa, and produce separate streams of hydrogen and oxygen gas, each at 400 K, 175 kPa. It is stated that this device operates in a 25°C room on 10-kW electrical power input. How do you evaluate this claim?



$$H_2O \rightarrow H_2 + \frac{1}{2}O_2$$

$$H_i - H_e = [-285830 + 18.015(42.01 - 104.89)] - 2961 - \frac{1}{2}(3027)$$

$$= -291\,437 \text{ kJ/kmol}$$

$$(S_i - S_e) = [69.950 + 18.015(0.151 - 0.3674)] - (139.219 - 8.3145 \ln 1.75)$$

$$- \frac{1}{2}(213.873 - 8.3145 \ln 1.75) = -173.124 \text{ kJ/kmol K}$$

$$W_{REV} = (H_i - H_e) - T_0(S_i - S_e) = -291\,437 - 298.15(-173.124)$$

$$= -239820 \text{ kJ/kmol}$$

$$\dot{W}_{REV} = (0.001/18.015)(-239\,820) = -13.31 \text{ kW}$$

$$\dot{I} = \dot{W}_{REV} - \dot{W}_{CV} = -13.31 - (-10) < 0$$ **Impossible**

14.93

Two kilomoles of ammonia are burned in a steady flow process with *x* kmol of oxygen. The products, consisting of H_2O, N_2, and the excess O_2, exit at 200°C, 7 MPa.

a. Calculate *x* if half the water in the products is condensed.

b. Calculate the absolute entropy of the products at the exit conditions.

$$2NH_3 + xO_2 \rightarrow 3H_2O + N_2 + (x - 1.5)O_2$$

Products at 200 °C, 7 MPa with $n_{H2O\ LIQ} = n_{H2O\ VAP} = 1.5$

a) $y_{H2O\ VAP} = P_G/P = \frac{1.5538}{7} = \frac{1.5}{1.5 + 1 + x - 1.5}$ => x = **5.757**

b) $S_{PROD} = S_{GAS\ MIX} + S_{H2O\ LIQ}$

Gas mixture:	n_i	y_i	$\bar{s}_i^\circ$	$-\bar{R}\ln(y_iP/P_0)$	$\bar{S}_i$
H_2O	1.5	0.222	204.595	-22.810	181.785
O_2	4.257	0.630	218.985	-31.482	187.503
N_2	1.0	0.148	205.110	-19.439	185.671

$S_{GAS\ MIX} = 1.5(181.785) + 4.257(187.503) + 1.0(185.67) = 1256.55$ kJ/K

$S_{H2O\ LIQ} = 1.5[69.950 + 18.015(2.3223 - 0.3674)] = 157.75$ kJ/K

$S_{PROD} = 1256.55 + 157.75 =$ **1414.3 kJ/K**

14.94

Graphite, C, at P_0, T_0 is burned with air coming in at P_0, 500 K in a ratio so the products exit at P_0, 1200 K. Find the equivalence ratio, the percent theoretical air, and the total irreversibility.

$$C + (1/\phi)b(O_2 + 3.76\ N_2) \;\rightarrow\; CO_2 + \big((1/\phi) - 1\big)O_2 + 3.76\ (1/\phi)\ N_2$$

Energy Eq.: $\quad H_P = H_R \Rightarrow \Delta H_{P\ 1200} - \Delta H_R = H^{\circ}_R - H^{\circ}_P$

$$44\,473 + \big((1/\phi) - 1\big)29\,761 + 3.76(1/\phi)28\,109$$

$$- (1/\phi)\big(6086 + 3.76{\times}5911\big) = 0 - (-393\,522) \;\Rightarrow\; (1/\phi) = 3.536$$

$$S_{gen} = S_P - S_R = \sum_{P-R} \nu\big(\bar{s}^{\circ} - \bar{R}\ \ln(y)\big)$$

R: $y_{O_2} = 0.21$, $y_{N_2} = 0.79$

P: $y_{O_2} = 0.1507$, $y_{N_2} = 0.79$, $y_{CO_2} = 0.0593$

$$S^{\circ}_P = 279.39 + 2.536 \times 250.011 + 13.295 \times 234.227 = 4027.5$$

$$S^{\circ}_R = 5.74 + 3.536(220.693 + 3.76 \times 206.74) = 3534.8$$

For the pressure correction the term with the nitrogen drops out (same y).

$$\bar{R} \sum_{P-R} -\nu\ \ln(y) = 8.3145(2.8235 + 1.8927 - 1.5606) = 26.236$$

$$S_{gen} = 4027.5 - 3534.8 + 26.236 = 518.94 \text{ kJ/kmol carbon-K}$$

$$I = T_0\ S_{gen} = \mathbf{154\ 721\ \frac{kJ}{kmol\ C}}$$

14.95

A flow of hydrogen gas is mixed with a flow of oxygen in a stoichiometric ratio, both at 298 K and 50 kPa. The mixture burns without any heat transfer in complete combustion. Find the adiabatic flame temperature and the amount of entropy generated per kmole hydrogen in the process.

The reaction equation is:

$$H_2 + \nu_{O2}\, O_2 \quad \Rightarrow \; H_2O$$

The balance of hydrogen is done, now for oxygen we need $\nu_{O2} = 0.5$.

Energy Eq.: $H_R = H_P \;\Rightarrow\; 0 = -241\,826 + \Delta\bar{h}_{H2O}$

$\Rightarrow \quad \Delta\bar{h}_{H2O} = 241\,826$ kJ/kmol

Interpolate now in table A.9 for the temperature to give this enthalpy

T = 4991 K

For this temperature we find from Table A.9

$$S_P = \bar{s}^{\circ}_{H2O} - \bar{R}\ln(P/P_o) = 315.848 - 8.31451\ln(0.5) = 321.611 \text{ kJ/kmol K}$$

For the reactants we have

$$S_R = \bar{s}^{\circ}_{H2} - \bar{R}\ln(P/P_o) + 0.5\,[\bar{s}^{\circ}_{O2} - \bar{R}\ln(P/P_o)\,]$$

$$= 130.678 + 0.5 \times 205.148 - 1.5 \times 8.31451\ln(0.5)$$

$$= 241.897 \text{ kJ/kmol K}$$

$S_{gen} = S_P - S_R = 321.611 - 241.897 =$ **79.714 kJ/kmol H_2 K**

Recall that this includes the mixing process.

14.96

A closed rigid container is charged with propene, C_3H_6, and 150% theoretical air at 100 kPa, 298 K. The mixture is ignited and burns with complete combustion. Heat is transferred to a reservoir at 500 K so the final temperature of the products is 700 K. Find the final pressure, the heat transfer per kmole fuel and the total entropy generated per kmol fuel in the process.

$$C_3H_6 + \nu_{O2}\left(O_2 + 3.76\ N_2\right) \rightarrow 3\ CO_2 + 3\ H_2O + x\ N_2$$

Oxygen, O_2, balance: $2\ \nu_{O2} = 6 + 3 = 9 \quad \Rightarrow \quad \nu_{O2} = 4.5$

Actual Combustion: $\phi = 1.5 \Rightarrow \nu_{O_2\ ac} = 1.5 \times 4.5 = 6.75$

$$C_3H_6 + 6.75\ O_2 + 25.38\ N_2 \rightarrow 3\ CO_2 + 3\ H_2O + 25.38\ N_2 + 2.25\ O_2$$

$$P_2 = P_1 \frac{n_P T_2}{n_R T_1} = 100 \times \frac{33.63 \times 700}{33.13 \times 298.15} = \mathbf{238.3\ kPa}$$

Enthalpies from Table A.9

$$\Delta H_{P\ 700} = 3\times 17\ 754 + 3\times 14\ 190 + 25.38\times 11\ 937 + 2.25\times 12\ 499$$

$$= 426\ 916\ \frac{\text{kJ}}{\text{kmol fuel}}$$

Enthalpy of combustion from table 14.3 converted to mole basis

$$\overset{\circ}{H}_{RP} = -45\ 780 \times 42.081 = -1\ 926\ 468\ \text{kJ/kmol fuel}$$

$$U_2 - U_1 = {}_1Q_2 - 0 = H_2 - H_1 - n_2\bar{R}T_2 + n_1\bar{R}T_1$$

$${}_1Q_2 = \overset{\circ}{H}_{RP} + \Delta H_{P\ 700} - n_P\bar{R}T_2 + n_1\bar{R}T_1$$

$$= -1\ 926\ 468 + 426\ 916 - 33.63 \times 8.3145 \times 700$$

$$+ 33.13 \times 8.3145 \times 298.15 = \mathbf{-1.613\times 10^6\ \frac{kJ}{kmol\ fuel}}$$

Entropies from Table A.9 and pressure correction

Reactants:	n_i	y_i	$\bar{s}_i^\circ$	$-\bar{R}\ln(y_iP/P_0)$	S_i
C_3H_8	1.0	0.0302	267.066	29.104	296.17
O_2	6.75	0.2037	205.143	13.228	218.376
N_2	25.38	0.7661	191.609	2.216	189.393

$$S_1 = 296.17 + 6.75 \times 218.376 + 25.38 \times 189.393 = 6577\ \frac{\text{kJ}}{\text{kmol fuel K}}$$

Products:	n_i	y_i	$\bar{s}_i^\circ$	$-\bar{R}\ln(y_iP/P_0)$	S_i
CO_2	3	0.0892	250.752	+12.875	263.627
H_2O	3	0.0892	218.739	+12.875	231.614
O_2	2.25	0.0669	231.465	+15.266	246.731
N_2	25.38	0.7547	216.865	- 4.88	211.985

$$S_2 = 3(263.627 + 231.614) + 2.25 \times 246.731 + 25.38 \times 211.985$$
$$= 7421 \text{ kJ/kmol fuel K}$$

$${}_1S_{2\,\text{gen}} = S_2 - S_1 - {}_1Q_2/T_{\text{res}} = 7421 - 6577 + \frac{1.613\times10^6}{500} = \mathbf{4070}\ \frac{\mathbf{kJ}}{\mathbf{kmol\ fuel\ K}}$$

Problems Involving Generalized Charts or Real Mixtures

14.97

Repeat Problem 14.42, but assume that saturated-liquid oxygen at 90 K is used instead of 25°C oxygen gas in the combustion process. Use the generalized charts to determine the properties of liquid oxygen.

Problem same as 14.42, except oxygen enters at 2 as saturated liquid at 90 K.

$\dot{m}_{O2}/\dot{m}_{N2H4} = 0.5 = 32\dot{n}_{O2}/32\dot{n}_{N2H4}$ and $\dot{Q}/\dot{m}_{N2H4} = -100$ kJ/kg

Energy Eq.: $Q_{CV} = H_P - H_R = -100 \times 32.045 = -3205$ kJ/kmol fuel

Reaction equation: $1\ N_2H_4 + \frac{1}{2}O_2 \rightarrow H_2O + H_2 + N_2$

At 90 K, $T_{r2} = 90/154.6 = 0.582 \Rightarrow \Delta\tilde{h}_f = 5.2$

Figure D.2, $(\bar{h}^* - \bar{h}) = 8.3145 \times 154.6 \times 5.2 = 6684$ kJ/kmol

$\Delta\bar{h}_{AT\ 2} = -6684 + 0.922 \times 32(90 - 298.15) = -12825$ kJ/kmol

$H_R = 50417 + \frac{1}{2}(0 - 12825) = 44005$ kJ, $H_P^\circ = -241826$

1st law: $\Delta\bar{h}_P = \Delta\bar{h}_{H_2O} + \Delta\bar{h}_{H_2} + \Delta\bar{h}_{N_2} = Q_{cv} + H_R - H_P^\circ = 282626$

From Table A.9, $\Delta H_{P\ 2800K} = 282141$, $\Delta H_{P\ 3000K} = 307988$

Therefore, $T_P =$ **2804 K**

14.98

Hydrogen peroxide, H_2O_2, enters a gas generator at 25°C, 500 kPa at the rate of 0.1 kg/s and is decomposed to steam and oxygen exiting at 800 K, 500 kPa. The resulting mixture is expanded through a turbine to atmospheric pressure, 100 kPa, as shown in Fig. P14.98. Determine the power output of the turbine, and the heat transfer rate in the gas generator. The enthalpy of formation of liquid H_2O_2 is −187 583 kJ/kmol.

$$H_2O_2 \rightarrow H_2O + \frac{1}{2} O_2 \qquad \dot{n}_{H2O2} = \frac{\dot{m}_{H2O2}}{M} = \frac{0.1}{34.015} = 0.00294 \text{ kmol/s}$$

$$\dot{n}_{MIX} = \dot{n}_{H2O2} \times 1.5 = 0.00441 \text{ kmol/s}$$

$$\bar{C}_{P0\,MIX} = \frac{2}{3} \times 1.872 \times 18.015 + \frac{1}{3} \times 0.922 \times 31.999 = 32.317$$

$$\bar{C}_{V0\,MIX} = 32.317 - 8.3145 = 24.0 \quad => \quad k_{MIX} = 32.317/24.0 = 1.3464$$

CV: turbine. Assume reversible → $s_3 = s_2$

$$T_3 = T_2\left(\frac{P_3}{P_2}\right)^{\frac{k-1}{k}} = 800\left(\frac{100}{500}\right)^{0.2573} = 528.8 \text{ K}$$

$$w = \bar{C}_{P0}(T_2 - T_3) = 32.317(800 - 528.8) = 8765 \text{ kJ/kmol}$$

$$\dot{W}_{CV} = 0.00441 \times 8765 = \mathbf{38.66\ kW}$$

CV: Gas Generator

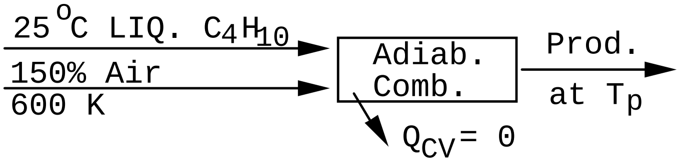
$$\dot{H}_1 = 0.00294(-187\,583 + 0) = -551.49$$

$$\dot{H}_2 = 0.00294(-241\,826 + 18002) + 0.00147(0 + 15836) = -634.76$$

$$\dot{Q}_{CV} = \dot{H}_2 - \dot{H}_1 = -634.76 + 551.49 = \mathbf{-83.27\ kW}$$

14.99

Liquid butane at 25°C is mixed with 150% theoretical air at 600 K and is burned in an adiabatic steady state combustor. Use the generalized charts for the liquid fuel and find the temperature of the products exiting the combustor.



$T_C = 425.2$ K

$T_R = 0.701$

$\bar{h}_{C4H10} = \bar{h}^o_{f\,IG} + (\bar{h}_{LIQ} - \bar{h}^*)$ see Fig. D.2

$= -126200 + (-4.85 \times 8.3145 \times 425.2) = -143\ 346$ kJ

$C_4H_{10} + 1.5 \times 6.5\ O_2 + 3.76 \times 9.75\ N_2$

$\rightarrow 4\ CO_2 + 5\ H_2O + 3.25\ O_2 + 36.66\ N_2$

$\bar{h}_{AIR} = 9.75(9245) + 36.66(8894) = 416\ 193$ kJ

$\Rightarrow \quad H_R = 416\ 193 - 143\ 346 = +272\ 847$ kJ

$H_P = 4(-393522 + \Delta\bar{h}^*_{CO2}) + 5(-241826 + \Delta\bar{h}^*_{H2O}) + 3.25\ \Delta\bar{h}^*_{O2} + 36.66\ \Delta\bar{h}^*_{N2}$

Energy Eq.: $H_P - H_R = 0$

$4\ \Delta\bar{h}^*_{CO2} + 5\ \Delta\bar{h}^*_{H2O} + 3.25\ \Delta\bar{h}^*_{O2} + 36.66\ \Delta\bar{h}^*_{N2} = 3\ 056\ 065$

Trial and Error: $LHS_{2000\ K} = 2\ 980\ 000$, $LHS_{2200\ K} = 3\ 369\ 866$

Linear interpolation to match RHS => $T_P =$ **2039 K**

14.100

Saturated liquid butane enters an insulated constant pressure combustion chamber at 25°C, and *x* times theoretical oxygen gas enters at the same *P* and *T*. The combustion products exit at 3400 K. With complete combustion find *x*. What is the pressure at the chamber exit? and what is the irreversibility of the process?

Butane: $T_1 = T_o = 25^oC$, sat liq., $x_1 = 0$, $T_c = 425.2$ K, $P_c = 3.8$ MPa

Do the properties from the generalized charts

Fig. D.1: $T_{r1} = 0.7$, $P_{r1} = 0.1$, $P_1 = P_{r1}P_c = 380$ kPa

Figs. D.2 and D.3: $\left(\bar{h}_1^* - \bar{h}_1\right)_f = 4.85\ \bar{R}T_c$, $(\bar{s}_1^* - \bar{s}_1)_f = 6.8\ \bar{R}$

Oxygen: $T_2 = T_o = 25^oC$, X - Theoretical O_2

Products: $T_3 = 3400$ K, Assumes complete combustion

$$C_4H_{10} + 6.5X\ O_2 \rightarrow 4\ CO_2 + 5\ H_2O + 6.5(X\text{-}1)\ O_2$$

Energy Eq.: $Q_{CV} + H_R = H_P + W_{CV}$; $Q_{CV} = 0$, $W_{CV} = 0$

$$H_R = n\left(\bar{h}_f^o + \Delta\bar{h}\right)_{C4H10} = 1(-126\ 200 + -17\ 146) = -143\ 346\text{ kJ}$$

Products: CO_2 $n\left(\bar{h}_f^o + \Delta\bar{h}\right)_{CO2} = 4(-393\ 522 + 177\ 836) = -862\ 744$ kJ

H_2O $n\left(\bar{h}_f^o + \Delta\bar{h}\right)_{H2O} = 5(-241\ 826 + 149\ 073) = -463\ 765$ kJ

O_2 $n\left(\bar{h}_f^o + \Delta\bar{h}\right)_{O2} = 6.5(X\text{-}1)(0 + 114\ 101) = (X\text{-}1)741\ 657$ kJ

$H_P = \sum n_i\left(\bar{h}_f^o + \Delta\bar{h}\right)_i = 741\ 657X - 2\ 068\ 166$

Energy Eq.: $H_P = H_R$ solve for X; **X = 2.594**

Assume that the exit pressure equals the inlet pressure: $P_e = P_i = 380$ kPa

$$\bar{s}_{C_4H_{10}} = \left[\bar{s}_f^o - \bar{R}\ln\frac{P_1}{P_o} - (\bar{s}_1^* - \bar{s}_1)_f\right]\ ;\ \ \bar{s}_{O_2} = \left[\bar{s}^o - \bar{R}\ln\frac{P_1}{P_o}\right]$$

$S_R = S_{C4H10} + S_{O2} = [306.647 - 11.10 - 56.539]$

$+ [205.48 - 11.10] \times 6.5 \times 2.594 = 3516.45$ kJ/K

Products:

	n_i	y_i	$\bar{s}_i^o$	$-\bar{R}\ln\frac{y_iP_e}{P_o}$	$\bar{S}_i$ $\frac{\text{kJ}}{\text{kmol K}}$
CO_2	4	0.2065	341.988	2.016	344.004
H_2O	5	0.2582	293.550	0.158	293.708
O_2	10.368	0.5353	289.499	-5.904	283.595

$S_P = \sum n_i\bar{S}_i = 5784.87$ kJ/K;

$I = T_o(S_P - S_R) = 298.15\ (5784.87 - 3516.45) =$ **676 329 kJ**

14.101

A gas mixture of 50% ethane and 50% propane by volume enters a combustion chamber at 350 K, 10 MPa. Determine the enthalpy per kilomole of this mixture relative to the thermochemical base of enthalpy using Kay's rule.

$\bar{h}^*_{MIX\ O} = 0.5(-84740) + 0.5(-103900) = -94320$ kJ/kmol

$\bar{C}_{P0\ MIX} = 0.5 \times 30.07 \times 1.7662 + 0.5 \times 44.097 \times 1.67 = 63.583$

$\bar{h}^*_{350} - \bar{h}^*_{298} = 63.583(350 - 298.2) = 3294$ kJ/kmol

Kay's rule: $T_{C\ MIX} = 0.5 \times 305.4 + 0.5 \times 369.8 = 337.6$ K

$P_{C\ MIX} = 0.5 \times 4.88 + 0.5 \times 4.25 = 4.565$ MPa

$T_r = 350/337.6 = 1.037, \quad P_r = 10/4.565 = 2.19$

From Fig. D.2: $\bar{h}^* - \bar{h} = 8.3145 \times 337.6 \times 3.53 = 9909$ kJ/kmol

$\bar{h}_{MIX\ 350K,10MPa} = -94320 + 3294 - 9909 =$ **-100 935 kJ/kmol**

14.102

A mixture of 80% ethane and 20% methane on a mole basis is throttled from 10 MPa, 65°C, to 100 kPa and is fed to a combustion chamber where it undergoes complete combustion with air, which enters at 100 kPa, 600 K. The amount of air is such that the products of combustion exit at 100 kPa, 1200 K. Assume that the combustion process is adiabatic and that all components behave as ideal gases except the fuel mixture, which behaves according to the generalized charts, with Kay's rule for the pseudocritical constants. Determine the percentage of theoretical air used in the process and the dew-point temperature of the products.

Reaction equation:

Fuel mix: $\bar{h}^0_{f\,FUEL} = 0.2(-74873) + 0.8(-84740) = -82767$ kJ/kmol

$\bar{C}_{P0\,FUEL} = 0.2 \times 2.2537 \times 16.04 + 0.8 \times 1.7662 \times 30.07 = 49.718$

$\Delta\bar{h}^*_{FUEL} = 49.718(65 - 25) = 1989$ kJ/kmol

$T_{CA} = 305.4$ K, $T_{CB} = 190.4$ K $\Rightarrow$ $T_{c\,mix}=282.4$ K

$P_{CA} = 4.88$, $P_{CB}=4.60$ $\Rightarrow$ $P_{c\,mix}= 4.824$ MPa

$T_r = 338.2/282.4 = 1.198$, $P_r = 10/4.824 = 2.073$

$(\bar{h}^* - \bar{h})_{FUEL\,IN} = 8.31451 \times 282.4 \times 2.18 = 5119$

$\Rightarrow \bar{h}_{FUEL\,IN} = -82767 + 1989 - 5119 = -85897 \frac{kJ}{kmol}$

1st law:

$$1.8(-393522 + 44473) + 2.8(-241826 + 34506) + 3.2(x - 1)(29761) + (12.03x)(28109) + 85897 - (3.2x)(9245) - (12.03x)(8894) = 0$$

a) x = 4.104 or **410.4 %**

b) $n_P = 1.8 + 2.8 + 3.2(4.104 - 1) + 12.03 \times 4.104 = 63.904$

$y_{H2O} = 2.8/63.904 = P_V/100$; $P_V = 4.38$ kPa, T = **30.5°C**

14.103

Liquid hexane enters a combustion chamber at 31°C, 200 kPa, at the rate 1 kmol/s 200% theoretical air enters separately at 500 K, 200 kPa, and the combustion products exit at 1000 K, 200 kPa. The specific heat of ideal gas hexane is C_p = 143 kJ/kmol K. Calculate the rate of irreversibility of the process.

Hexane: $T_c = 507.5$ K, $P_c = 3010$ kPa

$T_{r1} = 0.6$, Fig. D.1: $P_{rg} = 0.028$, $P_{g1} = P_{r1}P_c = 84.47$ kPa

Figs D.2 and D.3: $(\bar{h}^*_1 - \bar{h}_1)_f = 5.16\ \bar{R}T_c$, $(\bar{s}^*_1 - \bar{s}_1)_f = 8.56\ \bar{R}$

Air: $T_2 = 500$ K, $P_2 = 200$ kPa, 200% theoretical air

Products: $T_3 = 1000$ K, $P_3 = 200$ kPa

a) $\bar{h}_{C6H14} = \bar{h}^o_f - (\bar{h}^*_1 - \bar{h}_1)_f + (\bar{h}^*_1 - \bar{h}^*_0) + (\bar{h}^*_0 - \bar{h}_0)$

$\bar{h}^*_0 - \bar{h}_0 = 0$, $\bar{h}^*_1 - \bar{h}^*_0 = \bar{C}_P\,(T_1 - T_o) = 858$ kJ/kmol, $\bar{h}^o_f = -167300$ kJ/kmol

$\bar{h}^*_1 - \bar{h}_1 = 5.16 \times 8.3145 \times 507.5 = 21773$ kJ/kmol, $\bar{h}_{C6H14} =$ **-188215 kJ/kmol**

$$\bar{s}_{C6H14} = \bar{s}^o_{T_o} + \bar{C}_P \ln\frac{T_1}{T_o} - \bar{R}\ln\frac{P_1}{P_o} + (\bar{s}_1 - \bar{s}^*_1)$$

$$\bar{s}^o_{T_o} + \bar{C}_P \ln\frac{T_1}{T_o} - \bar{R}\ln\frac{P_1}{P_o} = 387.979 + 2.85 - 5.763 = 385.066 \text{ kJ/kmol-K}$$

$\bar{s}^*_1 - \bar{s}_1 = 8.56 \times 8.3145 = 71.172$ kJ/kmol-K, $\bar{s}_{C6H14} =$ **313.894 kJ/kmol-K**

b) $C_6H_{14} + 19O_2 + 71.44N_2 \rightarrow 6CO_2 + 7H_2O + 9.5O_2 + 71.44N_2$

$T_{c\ prod} = \sum y_i T_{ci} = 179.3$ K, $T_{r3} = \dfrac{T_3}{T_{c\ prod}} =$ **5.58 → Ideal Gas**

c) 1st Law: $Q + H_R = H_P + W$; $W = 0$ => $Q = H_P - H_R$

$H_R = (\bar{h})_{C6H14} + 19\Delta\bar{h}_{O2} + 71.44\ \Delta\bar{h}_{N2}$

$= -188\,215 + 19\ \ 6086 + 71.44\ \ 5911 = 349701$ kJ/kmol fuel

$$H_P = 6(\bar{h}^o_f + \Delta\bar{h})_{CO2} = 7(\bar{h}^o_f + \Delta\bar{h})_{H2O} + 9.5\,(\bar{h}^o_f + \Delta\bar{h})_{O2} + 71.44(\bar{h}^o_f + \Delta\bar{h})_{N2}$$

CO_2 - $(\bar{h}^o_f + \Delta\bar{h}) = (-393522 + 33397) = -360125$ kJ/kmol

H_2O - $(\bar{h}^o_f + \Delta\bar{h}) = (-241826 + 26000) = -215826$ kJ/kmol

O_2 - $(\bar{h}^o_f + \Delta\bar{h}) = (0 + 22703) = 22703$ kJ/kmol

N_2 - $(\bar{h}^o_f + \Delta\bar{h}) = (0 + 21463) = 21463$ kJ/kmol

$H_P = -1922537$ kJ; $\dot{Q} =$ **-2272238 kW**

d) $\dot{I} = T_o\,\dot{n}\,(S_P - S_R) - \dot{Q};\ T_o = 25^oC$

$$S_R = (\bar{s})_{C6H14} + 19\,(\bar{s}^o_{500} - \bar{R}\ln\frac{y_{O2}P_2}{P_o})_{O2} + 71.44(\bar{s}^o_{500} - \bar{R}\ln\frac{y_{N2}P_2}{P_o})_{N2}$$

$(\bar{s})_{C6H14} = 313.894$ kJ/kmol K, $(\bar{s}^o_{500})_{O2} = 220.693$ kJ/kmol K

$(\bar{s}^o_{500})_{N2} = 206.740$ kJ/kmol K, $y_{O2} = 0.21$, $y_{N2} = 0.79$

$\dot{S}_R = 19141.9$ kW/K

Products:

	n_i	y_i	$\bar{s}^o_i$	$-\bar{R}\ln\frac{y_iP_e}{P_o}$	$\bar{S}_i$ (kJ/kmol-K)
CO_2	6	0.0639	269.299	17.105	286.404
H_2O	7	0.0745	232.739	15.829	248.568
O_2	9.5	0.1011	243.579	13.291	256.87
N_2	71.44	0.7605	228.171	-3.487	224.684

$S_P = \sum n_i\bar{s}_i = 21950.1$ kJ/K;

$\dot{I} = T_o\,\dot{n}\,(S_P - S_R) - \dot{Q} =$ **3 109 628 kW**

Fuel Cells

14.104

In Example 14.16, a basic hydrogen–oxygen fuel cell reaction was analyzed at 25°C, 100 kPa. Repeat this calculation, assuming that the fuel cell operates on air at 25°C, 100 kPa, instead of on pure oxygen at this state.

Anode: $2\ H_2 \rightarrow 4\ e^- + 4\ H^+$

Cathode: $4\ H^+ + 4\ e^- \rightarrow 1\ O_2 + 2\ H_2O$

Overall: $2\ H_2 + 1\ O_2 \rightarrow 2\ H_2O$

Example 14.16: $\Delta G_{25°C} = -474\ 283$ kJ (for pure O_2)

For $P_{O2} = 0.21 \times 0.1$ MPa:

$\bar{S}_{O_2} = 205.148 - 8.3145 \ln 0.21 = 218.124$ kJ/kmol

$\Delta S = 2(69.950) - 2(130.678) - 1(218.124) = -339.58$ kJ/kmol K

$\Delta G_{25°C} = -571\ 660 - 298.15(-339.58) = -470\ 414$ kJ/kmol

$$E° = \frac{470414}{96487 \times 4} = \mathbf{1.219\ V}$$

14.105

Assume that the basic hydrogen-oxygen fuel cell operates at 600 K instead of 298 K as in example 14.16. Find the change in the Gibbs function and the reversible EMF it can generate.

Reaction: $2\ H_2 + O_2 \Rightarrow 2\ H_2O$

At a temperature of 600 K the water formed is in a vapor state. We can thus find the change in the enthalpy as

$$\Delta H^0_{600\ K} = 2\left(\bar{h}^o_f + \Delta\bar{h}\right)_{H2O\ g} - 2\left(\bar{h}^o_f + \Delta\bar{h}\right)_{H2} - \left(\bar{h}^o_f + \Delta\bar{h}\right)_{O2}$$

$$= 2(-241\ 826 + 10\ 499) - 2(0 + 8799) - 0 - 9245$$

$$= -489\ 497\ \text{kJ/4 kmol e-}$$

$$\Delta S^0_{600\ K} = 2\ \bar{s}^o_{f\,H2O\ g} - 2\ \bar{s}^o_{f\,H2} - \bar{s}^o_{f\,O2}$$

$$= 2 \times 213.051 - 2 \times 151.078 - 226.45$$

$$= -102.504\ \text{kJ/4 kmol e- K}$$

$$\Delta G^0_{600\ K} = \Delta H^0_{600\ K} - T\Delta S^0_{600\ K} = -489\ 497 - 600(-102.504)$$

$$= -\ 427\ 995\ \text{kJ/4 kmol e-}$$

$$W^{rev} = -\Delta\bar{G}^0 = 427\ 995\ \text{kJ/4 kmol e-}$$

$$E^0 = \frac{-\Delta\bar{G}^0}{96485 \times 8} = \frac{427\ 995}{96\ 485 \times 4} = \mathbf{1.109\ V}$$

14.106

Consider a methane-oxygen fuel cell in which the reaction at the anode is

$$CH_4 + 2H_2O \rightarrow CO_2 + 8e^- + 8H^+$$

The electrons produced by the reaction flow through the external load, and the positive ions migrate through the electrolyte to the cathode, where the reaction is

$$8\ e^- + 8\ H^+ + 2\ O_2 \rightarrow 4\ H_2O$$

Calculate the reversible work and the reversible EMF for the fuel cell operating at 25°C, 100 kPa.

$$CH_4 + 2H_2O \rightarrow CO_2 + 8e^- + 8H^+$$

and $8e^- + 8H^+ + 2CO_2 \rightarrow 4H_2O$

Overall $CH_4 + 2O_2 \rightarrow CO_2 + 2H_2O$

a) 25 °C assume all liquid H_2O and all comp. at 100 kPa

$$\Delta H^0_{25\ C} = -393\ 522 + 2(-285\ 830) - (-74\ 873) - 0 = -890\ 309\ \text{kJ}$$

$$\Delta S^0_{25\ C} = 213.795 + 2(69.950) - 186.251 - 2(205.148) = -\ 242.852\ \text{kJ/K}$$

$$\Delta G^0_{25\ C} = -890\ 309 - 298.15(-242.852) = -\ 817\ 903\ \text{kJ}$$

$$W^{rev} = -\Delta\bar{G}^0 = +817903\ \text{kJ}$$

$$E^0 = \frac{-\Delta\bar{G}^0}{96485 \times 8} = \frac{+817903}{96485 \times 8} = \mathbf{1.06\ V}$$

14.107

Consider a methane-oxygen fuel cell in which the reaction at the anode is

$$CH_4 + 2H_2O \rightarrow CO_2 + 8e^- + 8H^+$$

The electrons produced by the reaction flow through the external load, and the positive ions migrate through the electrolyte to the cathode, where the reaction is

$$8\ e^- + 8\ H^+ + 2\ O_2 \rightarrow 4\ H_2O$$

Assume that the fuel cell operates at 1200 K instead of at room temperature.

$$CH_4 + 2H_2O \rightarrow CO_2 + 8e^- + 8H^+$$

and $8e^- + 8H^+ + 2CO_2 \rightarrow 4H_2O$

Overall $\quad CH_4 + 2O_2 \rightarrow CO_2 + 2H_2O$

1200 K assume all gas H_2O and all comp. at 100 kPa

$$\Delta H^0_{1200\text{ K}} = 1(-393522 + 44473) + 2(-241826 + 34506) - 2(0 + 29761)$$
$$- 1[-74873 + 16.043 \times 2.254(1200 - 298.2)] = -780\,948\text{ kJ}$$

$$\Delta S^0_{1200\text{ K}} = 1(279.390) + 2(240.485)$$
$$- 1(186.251 + 16.043 \times 2.254 \ln\frac{1200}{298.2}) - 2(250.011)$$
$$= 23.7397\text{ kJ/K}$$

$$\Delta G^0_{1200\text{ K}} = \Delta H^0_{1200\text{ K}} - T\Delta S^0_{1200\text{ K}} = -780\,948 - 1200(23.7397)$$
$$= -809\,436\text{ kJ}$$

$W^{rev} = \mathbf{+809\ 436\ kJ} \qquad E^0 = \dfrac{+809\,436}{96\,485 \times 8} = \mathbf{1.049\ V}$

Combustion Efficiency

14.108

Consider the steady combustion of propane at 25°C with air at 400 K. The products exit the combustion chamber at 1200 K. It may be assumed that the combustion efficiency is 90%, and that 95% of the carbon in the propane burns to form carbon dioxide; the remaining 5% forms carbon monoxide. Determine the ideal fuel–air ratio and the heat transfer from the combustion chamber.

Ideal combustion process, assumed adiabatic, excess air to keep 1200 K out.

$$C_3H_8 + 5x\ O_2 + 18.8x\ N_2 \ \rightarrow\ 3\ CO_2 + 4\ H_2O + 5(x - 1)\ O_2 + 18.8x\ N_2$$

$$H_R = -103900 + 5x(0 + 3027) + 18.8x(0 + 2971) = -103900 + 70990x$$

$$H_P = 3(-393522 + 44473) + 4(-241826 + 34506)$$

$$+ 5(x - 1)(0 + 29761) + 18.8x(0 + 28109) = -2025232 + 677254x$$

1st law: $H_P - H_R = 0$ Solving, $x = 3.169$

$FA_{IDEAL} = 1/(23.8 \times 3.169) =$ **0.01326**

b) $FA_{ACTUAL} = 0.01326/0.90 = 0.01473$

$$C_3H_8 + 14.26\ O_2 + 53.62\ N_2$$

$$\rightarrow 2.85\ CO_2 + 0.15\ CO + 4\ H_2O + 9.335\ O_2 + 53.62\ N_2$$

$$H_R = -103900 + 14.26(0 + 3027) + 53.62(0 + 2971) = +98570 \text{ kJ}$$

$$H_P = 2.85(-393522 + 44473) + 0.15(-110527 + 28427) + 4(-241826 + 34506)$$

$$+ 9.335(0 + 29761) + 53.62(0 + 28109) = -51361 \text{ kJ}$$

$Q_{CV} = H_P - H_R =$ **-149931 kJ**

14.109

A gasoline engine is converted to run on propane. Assume the propane enters the engine at 25°C, at the rate 40 kg/h. Only 90% theoretical air enters at 25°C such that 90% of the C burns to form CO_2, and 10% of the C burns to form CO. The combustion products also include H_2O, H_2 and N_2, exit the exhaust at 1000 K. Heat loss from the engine (primarily to the cooling water) is 120 kW. What is the power output of the engine? What is the thermal efficiency?

Propane: $T_1 = 25^oC$, $\dot{m} = 40$ kg/hr, M = 44.094 kg/kmol

Air: $T_2 = 25^oC$, 90% theoretical Air produces 90% CO_2, 10% CO

Products: $T_3 = 1000$ K, CO_2, CO, H_2O, H_2, N_2

$$C_3H_8 + 4.5O_2 + 16.92N_2 \rightarrow 2.7\,CO_2 + 0.3CO + 3.3H_2O + 0.7H_2 + 16.92N_2$$

$\dot{n}_{C3H8} = \dot{m}/(M\times3600) = 0.000252$ kmol/s

1^{st} Law: $\dot{Q} + H_R = H_P + \dot{W}$; $\dot{Q} = -120$ kW

$H_R = n_{C3H8}\,\bar{h}_f^o = -103\;900$ kJ

Products:

CO_2 - $n_{CO2}\left(\bar{h}_f^o + \Delta\bar{h}\right) = 2.7(-393522 + 33397) = -972337.5$ kJ

CO - $n_{CO}\left(\bar{h}_f^o + \Delta\bar{h}\right) = 0.3(-110527 + 21686) = -26652$ kJ

H_2O - $n_{H2O}\left(\bar{h}_f^o + \Delta\bar{h}\right) = 3.3(-241826 + 26000) = -712226$ kJ

H_2 - $n_{H2}\left(\bar{h}_f^o + \Delta\bar{h}\right) = 0.7(0 + 20663) = 14464.1$ kJ

N_2 - $n_{N2}\left(\bar{h}_f^o + \Delta\bar{h}\right) = 16.92(0 + 21463) = 363154$ kJ

$H_P = \sum n_i\left(\bar{h}_f^o + \Delta\bar{h}\right)_i = -1\;333\;598$ kJ

$\dot{W} = \dot{Q} + \dot{n}(H_R - H_P) = \mathbf{189.9\ kW}$

C_3H_8: Table 14.3 $H_{RPo} = -50343$ kJ/kg

$H\dot{H}V = \dot{n}_{C3H8}\,M(-H_{RPo}) = 559.4$ kW

$\eta_{th} = \dot{W}/H\dot{H}V = \mathbf{0.339}$

14.110

A small air-cooled gasoline engine is tested, and the output is found to be 1.0 kW. The temperature of the products is measured to 600 K. The products are analyzed on a dry volumetric basis, with the result: 11.4% CO_2, 2.9% CO, 1.6% O_2 and 84.1% N_2. The fuel may be considered to be liquid octane. The fuel and air enter the engine at 25°C, and the flow rate of fuel to the engine is 1.5×10^{-4} kg/s. Determine the rate of heat transfer from the engine and its thermal efficiency.

$$a\,C_8H_{18} + b\,O_2 + 3.76b\,N_2$$

$$\rightarrow 11.4\,CO_2 + 2.9\,CO + c\,H_2O + 1.6\,O_2 + 84.1\,N_2$$

$$b = \frac{84.1}{3.76} = 22.37, \qquad a = \frac{1}{8}(11.4 + 2.9) = 1.788$$

$$c = 9a = 16.088$$

$$C_8H_{18} + 12.5\,O_2 + 47.1\,N_2$$

$$\rightarrow 6.38\,CO_2 + 1.62\,CO + 9\,H_2O + 0.89\,O_2 + 47.1\,N_2$$

$$H_R = \bar{h}^0_{f\,C_8H_{18}} = -250\ 105 \text{ kJ/kmol}$$

$$H_P = 6.38(-393\ 522 + 15\ 788) + 1.62(-110527 + 10\ 781)$$

$$+ 9(-241\ 826 + 12\ 700) + 0.89(0 + 11187)$$

$$+ 47.1(0 + 10712) = -4\ 119\ 174 \text{ kJ/kmol}$$

$$H_P - H_R = -4\ 119\ 174 - (-250\ 105) = -3\ 869\ 069 \text{ kJ/kmol}$$

$$\dot{H}_P - \dot{H}_R = (0.00015/114.23)(-3\ 869\ 069) = -5.081 \text{ kW}$$

$$\dot{Q}_{CV} = -5.081 + 1.0 = \mathbf{-4.081\ kW}$$

Fuel heating value from table 14.3

$$\dot{Q}_H = 0.00015\ (47\ 893) = 7.184 \text{ kW}$$

$$\eta_{TH} = \dot{W}_{NET}/\dot{Q}_H = 1.0/7.184 = \mathbf{0.139}$$

14.111

A gasoline engine uses liquid octane and air, both supplied at P_0, T_0, in a stoichiometric ratio. The products (complete combustion) flow out of the exhaust valve at 1100 K. Assume that the heat loss carried away by the cooling water, at 100°C, is equal to the work output. Find the efficiency of the engine expressed as (work/lower heating value) and the second law efficiency.

$$C_8H_{18} + \nu_{O2}\left(O_2 + 3.76\ N_2\right) \rightarrow 8\ CO_2 + 9\ H_2O + 47\ N_2$$

$$2\ \nu_{O2} = 16 + 9 \Rightarrow \nu_{O2} = 12.5$$

$$LHV = 44425\ \frac{kJ}{kg\ fuel} \ \Rightarrow\ LHV = 5.07476\times10^6\ \frac{kJ}{kmol\ fuel}$$

$$\Delta H_{P\ 1100} = 8 \times 38885 + 9 \times 30190 + 47 \times 24760 = 1746510$$

C.V. Total engine

$$H_{in} = H_{ex} + W + Q_{loss} = H_{ex} + 2\ W$$

$$\Rightarrow 2\ W = H_{in} - H_{ex} = H_R - H_\nu = -\overset{\circ}{H}_{RP} + \Delta H_R - \Delta H_{P\ 1100}$$

$$= 5.07476\times10^6 + 0 - 1746510 = 3328250$$

$$W = 1.664\times10^6\ \frac{kJ}{kmol\ fuel}$$

$$\eta_{th} = \frac{W}{LHV} = \frac{1.664\times10^6}{5.07476\times10^6} = \mathbf{0.328}$$

Find entropies in and out:

inlet: $\bar{S}^\circ_{Fu} = 360.575$

$$\bar{S}^\circ_{O2} = 205.148 - 8.3145\ \ln\frac{1}{4.76} = 218.12$$

$$\bar{S}^\circ_{N2} = 191.609 - 8.3145\ \ln\frac{3.76}{4.76} = 193.57$$

$$\bar{S}^\circ_{in} = 360.575 + 12.5 \times 218.12 + 47 \times 193.57 = 12185$$

exit: $\bar{S}^\circ_{CO_2} = 275.528 - 8.3145\ \ln\frac{8}{64} = 292.82$

$$\bar{S}^\circ_{H_2O} = 236.732 - 8.3145\ \ln\frac{9}{64} = 253.04$$

$$\bar{S}^\circ_{N_2} = 231.314 - 8.1345\ \ln\frac{47}{64} = 233.88$$

$$\bar{S}^\circ_{ex} = 8 \times 292.82 + 9 \times 253.04 + 47 \times 233.88 = 15612$$

Assume the same Q_{loss} out to 100°C reservoir in the reversible case and compute Q_0^{rev}:

$$\bar{S}_{in} + Q_0^{rev}/T_0 = \bar{S}_{ex} + Q_{loss}/T_{res}$$

$$Q_0^{rev} = T_0\left(\bar{S}_{ex} - \bar{S}_{in}\right) + Q_{loss}\, T_0/T_{res}$$

$$= 298.15(15612 - 12185) + 1.664\times10^6 \times 298.15/373.15$$

$$= 2.351\times10^6 \,\frac{\text{kJ}}{\text{kmol fuel}}$$

$$H_{in} + Q_0^{rev} = H_{ex} + W^{rev} + Q_{loss}$$

$$\Rightarrow W^{rev} = H_{in} - H_{ex} - Q_{loss} + Q_0^{rev} = W_{ac} + Q_0^{rev} = 4.015\times10^6 \,\frac{\text{kJ}}{\text{kmol fuel}}$$

$$\eta_{II} = W_{ac}/\, W^{rev} = 1.664\times10^6/4.015\times10^6 = \mathbf{0.414}$$

Review Problems

14.112

Ethene, C_2H_4, and propane, C_3H_8, in a 1:1 mole ratio as gases are burned with 120% theoretical air in a gas turbine. Fuel is added at 25°C, 1 MPa and the air comes from the atmosphere, 25°C, 100 kPa through a compressor to 1 MPa and mixed with the fuel. The turbine work is such that the exit temperature is 800 K with an exit pressure of 100 kPa. Find the mixture temperature before combustion, and also the work, assuming an adiabatic turbine.

$\phi = 1$: $C_2H_4 + C_3H_8 + 8\ O_2 + 30.08\ N_2 \rightarrow 5\ CO_2 + 6\ H_2O + 30.08\ N_2$

$\phi = 1.2$: $C_2H_4 + C_3H_8 + 9.6\ O_2 + 36.096\ N_2$

$\rightarrow 5\ CO_2 + 6\ H_2O + 1.6\ O_2 + 36.096\ N_2$

45.696 kmol air per 2 kmol fuel

<u>C.V. Compressor</u> (air flow)

Energy Eq.: $w_c = h_2 - h_1$ Entropy Eq.: $s_2 = s_1 \Rightarrow$

$P_{r2} = P_{r1} \times P_2/P_1 = 13.573 \Rightarrow T_{2\ air} = \mathbf{570.8\ K}$

$w_c = 576.44 - 298.34 = \mathbf{278.1\ kJ/kg = 8056.6\ kJ/kmol\ air}$

<u>C.V. Mixing Chanber</u> (no change in composition)

$\dot{n}_{air}\bar{h}_{air\ in} + \dot{n}_{Fu1}\bar{h}_{1\ in} + \dot{n}_{Fu2}\bar{h}_{2\ in} = (\text{SAME})_{exit}$

$\left(\bar{C}_{P\ F1} + \bar{C}_{P\ F2}\right)\left(T_{exit} - T_0\right) = 45.696\ \bar{C}_{P\ air}\left(T_{2\ air} - T_{exit}\right)$

C_2H_4: $\bar{C}_{P\ F1} = 43.43$, C_3H_8: $\bar{C}_{P\ F2} = 74.06$, $\bar{C}_{P\ air} = 29.07$

$$T_{exit} = \frac{\left(45.696\bar{C}_{P\ air}T_2 + \left(\bar{C}_{P\ F1} + \bar{C}_{P\ F2}\right)T_0\right)}{\bar{C}_{P\ F1} + \bar{C}_{P\ F2} + 45.696\ \bar{C}_{P\ air}} = \mathbf{548.7\ K}$$

Dew Point Products: $y_{H_2O} = \frac{6}{5 + 6 + 1.6 + 36.096} = 0.1232$

$P_{H_2O} = y_{H_2O}P_{tot} = 123.2\text{ kPa} \Rightarrow T_{dew} = \mathbf{105.5°C}$

C.V. Turb. + combustor + mixer + compressor (no Q)

$w_{net} = H_{in} - H_{out} = H_R - H_{P\ 800}$ (800°K out so no liquid H_2O)

$= \bar{h}^{\circ}_{fC_2H_4} + \bar{h}^{\circ}_{fC_3H_8} - 5\ \bar{h}_{CO_2} - 6\ \bar{h}_{H_2O} - 1.6\ \bar{h}_{O_2} - 36.096\ \bar{h}_{N_2}$

$= \mathbf{2\ 576\ 541}\ \frac{\mathbf{kJ}}{\mathbf{2\ kmol\ Fu}}$

$w_T = w_{net} + w_{comp} = \mathbf{2\ 944\ 695}\ \frac{\mathbf{kJ}}{\mathbf{2\ kmol\ Fu}}$

14.113

Carbon monoxide, CO, is burned with 150% theoretical air and both gases are supplied at 150 kPa and 600 K. Find the reference enthalpy of reaction and the adiabatic flame temperature.

$$CO + \nu_{O2}\left(O_2 + 3.76\ N_2\right) \rightarrow CO_2 + \nu_{N2}\ N_2$$

$$\text{O balance:} \quad 1 + 2\nu_{O2} = 1 \quad \Rightarrow \quad \nu_{O2} = 0.5 \quad \Rightarrow \nu_{O2\ \text{actual}} = 0.75$$

Now the actual reaction equation has excess oxygen as

$$CO + 0.75\left(O_2 + 3.76\ N_2\right) \rightarrow CO_2 + 2.82\ N_2 + 0.25\ O_2$$

From the definition of enthalpy of combustion, Eq.14.14 or 14.15

$$H^\circ_{RP} = H^\circ_P - H^\circ_R = \bar{h}^o_{f\,CO2} + 0 - \bar{h}^o_{f\,CO}$$

$$= -393\,522 - (-110\,527) = \mathbf{-282\,995\ kJ/kmol\ CO}$$

$$= \mathbf{-\ 10\,103\ kJ/kg\ CO} \qquad \text{(as for Table 14.3)}$$

$$\text{Actual energy Eq.:} \quad H_R = H_P = H^\circ_P + \Delta H_P = H^\circ_R + \Delta H_R$$

$$\Delta H_P = H^\circ_R + \Delta H_R - H^\circ_P = -H^\circ_{RP} + \Delta\bar{h}_{CO} + 0.75\ \Delta\bar{h}_{O2} + 2.82\ \Delta\bar{h}_{N2}$$

$$= 282\,995 + 8942 + 0.75 \times 9245 + 2.82 \times 8894$$

$$= 323\,952\ \text{kJ/kmol}$$

The left hand side is

$$\Delta H_P = \Delta\bar{h}_{CO2} + 0.25\ \Delta\bar{h}_{O2} + 2.82\ \Delta\bar{h}_{N2}$$

$$\Delta H_{P\,2600} = 128\,074 + 0.25 \times 82\,225 + 2.82 \times 77\,963 = 368\,486$$

$$\Delta H_{P\,2400} = 115\,779 + 0.25 \times 74\,453 + 2.82 \times 70\,640 = 333\,597$$

$$\Delta H_{P\,2200} = 103\,562 + 0.25 \times 66\,770 + 2.82 \times 63\,362 = 298\,935$$

Now we can do a linear interpolation for the adiabatic flame temperature

$$T = 2200 + 200\,\frac{323\,952 - 298\,935}{333\,597 - 298\,935} = \mathbf{2344\ K}$$

14.114

Consider the gas mixture fed to the combustors in the integrated gasification combined cycle power plant, as described in Problem 14.12. If the adiabatic flame temperature should be limited to 1500 K, what percent theoretical air should be used in the combustors?

Product	CH_4	H_2	CO	CO_2	N_2	H_2O	H_2S	NH_3
% vol.	0.3	29.6	41.0	10.0	0.8	17.0	1.1	0.2

Mixture may be saturated with water so the gases are (H_2S and NH_3 out)

CH_4	H_2	CO	CO_2	N_2	n
0.3	29.6	41.0	10.0	0.8	81.7

$y_{V\,MAX} = 7.384/3000 = n_V/(n_V + 81.7)$

Solving, $n_V = 0.2$ kmol, rest condensed

$$\{0.3\,CH_4 + 29.6\,H_2 + 41.0\,CO + 10.0\,CO_2 + 0.8\,N_2$$
$$+ 0.2\,H_2O + 35.9x\,O_2 + 3.76 \times 35.9x\,N_2\} \rightarrow$$
$$51.3\,CO_2 + 30.4\,H_2O + 35.9(x - 1)\,O_2 + (135.0x + 0.8)\,N_2$$

For the fuel gas mixture,

$$n\bar{C}_{P0\,MIX} = 0.3 \times 16.04 \times 2.2537 + 29.6 \times 2.016 \times 14.2091$$
$$+ 41.0 \times 28.01 \times 1.0413 + 10.0 \times 44.01 \times 0.8418$$
$$+ 0.8 \times 28.013 \times 1.0416 + 0.2 \times 18.015 \times 1.8723 \quad = 2455.157$$

$$n\bar{h}^0_{f\,MIX} = 0.3(-74873) + 29.6(0) + 41.0(-110527)$$
$$+ 10.0(-393522) + 0.8(0) + 0.2(-241826) \quad = -8537654 \text{ kJ}$$

At 40°C, for the fuel mixture:

$$H_{MIX} = -8537654 + 2455.157(40 - 25) = -8500827 \text{ kJ}$$

Assume air enters at 25°C: $\bar{h}_{AIR} = 0$

Products at 1500 K:

$$H_P = 51.3(-393522 + 61705) + 30.4(-241826 + 48149)$$
$$+ 35.9(x - 1)(0 + 40600) + (135x + 0.8)(0 + 38405)$$
$$= -24336806 + 6642215x$$

1st law: $H_P = H_R = H_{MIX}$

$$x = \frac{+24336809 - 8500827}{6642215} = 2.384 \quad \text{or} \quad \textbf{238 \% theo. air}$$

14.115

A study is to be made using liquid ammonia as the fuel in a gas-turbine engine. Consider the compression and combustion processes of this engine.

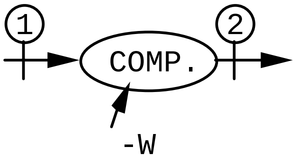
a. Air enters the compressor at 100 kPa, 25°C, and is compressed to 1600 kPa, where the isentropic compressor efficiency is 87%. Determine the exit temperature and the work input per kilomole.

b. Two kilomoles of liquid ammonia at 25°C and *x* times theoretical air from the compressor enter the combustion chamber. What is *x* if the adiabatic flame temperature is to be fixed at 1600 K?

Air

$P_1 = 100$ kPa

$T_1 = 25\ ^oC$



$P_2 = 1600$ kPa

$\eta_{S\ COMP} = 0.87$

a) ideal compressor process (adiabatic reversible):

$$s_{2S} = s_1 \Rightarrow T_{2S} = T_1\left(\frac{P_2}{P_1}\right)^{\frac{k-1}{k}} = 298.2\left(\frac{1600}{100}\right)^{0.286} = 659 \text{ K}$$

$$-w_S = C_{P0}(T_{2S} - T_1) = 1.004(659 - 298.2) = 362.2$$

Real process:

$$-w = -w_S/\eta_S = 362.2/0.87 = 416.3 \text{ kJ/kg}$$

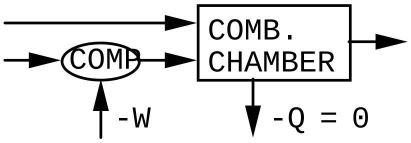
$$T_2 = T_1 - w/C_{P0} = 298.2 + 416.3/1.004 = \mathbf{713\ K}$$

Also $-w = 416.3 \times 28.97 =$ **12060 kJ/kmol**

b)

2 liq NH_3, 25 °C

Air 100 kPa, 25 °C



Prod.

$P_P = 1600$ kPa

$T_P = 1600$ K

$$2\ NH_3 + 1.5x\ O_2 + 5.64x\ N_2 \rightarrow 3\ H_2O + 1.5(x - 1)\ O_2 + (5.64x + 1)\ N_2$$

Using Tables 14.3, A.10 and A.2,

$$\bar{h}_{NH3} = -45\ 720 + 17.031(298.36 - 1530.04) = -66\ 697 \text{ kJ/kmol}$$

$$H_R = 2(-66\ 697) + 0 = -133\ 394 \text{ kJ}$$

$$-W = 12\ 060 \times 7.14x = 86\ 108\ x \text{ kJ}$$

$$H_P = 3(-241\ 826 + 52\ 907) + 1.5(x - 1)(0 + 44267) + (5.64x + 1)(0 + 41904)$$
$$= 302\ 739x - 591\ 254$$

Energy Eq.: $H_R = H_P + W$

$$-133\ 394 = 302\ 739\ x - 591\ 254 - 86\ 108\ x$$
$$\Rightarrow \quad x = \mathbf{2.11}$$

14.116

A rigid container is charged with butene, C_4H_8, and air in a stoichiometric ratio at P_0, T_0. The charge burns in a short time with no heat transfer to state 2. The products then cool with time to 1200 K, state 3. Find the final pressure, P_3, the total heat transfer, ${}_1Q_3$, and the temperature immediately after combustion, T_2.

The reaction equation is, having used C and H atom balances:

$$C_4H_8 + \nu_{O2}\left(O_2 + 3.76\ N_2\right) \rightarrow 4\ CO_2 + 4\ H_2O + 3.76\ \nu_{O2}N_2$$

Counting now the oxygen atoms we get $\nu_{O_2} = 6$.

C.V. analysis gives:

$$U_2 - U_1 = Q - W = Q = H_2 - H_1 - P_2V_2 + P_1V_1$$

$$= H_2 - H_1 - \bar{R}\left(n_2T_2 - n_1T_1\right)$$

$$H_2 - H_1 = H_{P\ 1200} - H^{\circ}_R = H^{\circ}_P - H^{\circ}_R + \Delta H_P = \hat{M}\ H^{\circ}_{RP} + \Delta H_P$$

$$= -2542590 + 950055 = \mathbf{-1592535}$$

Where $\hat{M} = 56.108$ and $n_1 = 1 + 6 \times 4.76 = 29.56$,

$n_2 = 4 + 4 + 6 \times 3.76 = 30.56$,

Table A.9 at 1200 K: $\Delta h_{CO_2} = 44473$, $\Delta h_{H2O} = 34506$, $\Delta h_{N_2} = 28109$.

Now solving for the heat transfer:

$$Q = -1592535 - 8.3145(30.56 \times 1200 - 29.56 \times 298.15) = -1824164\ \frac{\text{kJ}}{\text{kmol fuel}}$$

To get the pressure, assume ideal gases:

$$P_2 = \frac{n_2\bar{R}T_2}{V_2} = P_1\frac{n_2T_2}{n_1T_1} = 421.6\ \text{kPa}$$

Before heat transfer takes place we have constant U so:

$$U_{1_a} - U_1 = 0 = H_{1_a} - H_1 - n_2\bar{R}T_{1_a} + n_1\bar{R}T_1$$

Now split the enthalpy $H_{1_a} = H^{\circ}_P + \Delta H_P\left(T_{1_a}\right)$ and arrange things with the unknowns on LHS and knowns on RHS:

$$\Delta H_P - n_2\bar{R}T = H_R - H^{\circ}_P - n_1\bar{R}T_1 = 2\ 542\ 590 - 73278 = 2\ 469\ 312$$

Trial and error leads to:

$$\text{LHS (3000 K)} = 3\ 209\ 254 - 30.56 \times 8.31451 \times 3000 = 2\ 446\ 980$$

$$\text{LHS (3200 K)} = 3\ 471\ 331 - 30.56 \times 8.31451 \times 3200 = 2\ 658\ 238$$

linear interpolation T = **3021 K**

14.117

The turbine in Problem 14.112 is adiabatic. Is it reversible, irreversible, or impossible?

Inlet to the turbine is the exit from the mixing of air and fuel at 1 MPa. From solution to 14.112, we have:

$\bar{C}_{P\,C_2H_2} = 43.43$, $\bar{C}_{P\,C_3H_8} = 74.06$, $T_{\text{turbine,in}} = 548.7$ K

$$C_2H_4 + C_3H_8 + 9.6\ O_2 + 36.096\ N_2 \rightarrow 5\ CO_2 + 6\ H_2O + 1.6\ O_2 + 36.096\ N_2$$

$$S_{ex} - S_{in} = \int \frac{dQ}{T} + S_{gen} = S_{gen} \geq \phi$$

Inlet: 1 MPa, 548.7 K $\qquad \bar{S}_{Fu} = \bar{S}_i^{\circ} + \bar{C}_{P\,Fu} \ln(T/T_0)$

	n_i	y_i	$\bar{s}_i^{\circ}$	$-\bar{R}\ln\frac{y_iP}{P_0}$	$\bar{S}_i$
C_2H_4	1	0.02097	245.82	12.989	258.809
C_3H_8	1	0.02097	315.09	12.989	328.079
O_2	9.6	0.2013	223.497	-5.816	217.681
N_2	36.096	0.7568	209.388	-16.828	192.56

$$S_{in} = 258.809 + 328.079 + 9.6 \times 217.681 + 36.096 \times 192.56 = 9627.3$$

	n_i	y_i	$\bar{s}_i^{\circ}$	$-\bar{R}\ln\frac{y_iP}{P_0}$	S_i
CO_2	5	0.1027	257.496	18.925	276.421
H_2O	6	0.1232	223.826	17.409	241.235
O_2	1.6	0.0329	235.92	28.399	264.319
N_2	36.096	0.7413	221.016	2.489	223.505

$$S_{ex} = 5 \times 276.421 + 6 \times 241.235\ 1.6 \times 264.319 + 36.096 \times 223.505 = 11320\ \frac{\text{kJ}}{\text{2kmol Fu K}}$$

$$S_{gen} = S_{ex} - S_{in} = \mathbf{1693\ \frac{kJ}{2kmol\ Fu\ K} > 0}$$

Possible, but one should check the state after combustion to account for generation by combustion alone and then the turbine expansion separately.

14.118

Consider the combustion process described in Problem 14.102.

a. Calculate the absolute entropy of the fuel mixture before it is throttled into the combustion chamber.

b. Calculate the irreversibility for the overall process.

From solution to 14.102, fuel mixture $0.8\ C_2H_6 + 0.2\ CH_4$ at 65°C, 10 MPa

$\bar{C}_{P0\ FUEL} = 49.718$ kJ/kmol K. Using Kay's rule: $T_{r1} = 1.198$, $P_{r1} = 2.073$

and $x = 410.4$ % theoretical air

or $13.13\ O_2 + 49.36\ N_2$ in at 600 K, 100 kPa

and $1.8CO_2 + 2.8H_2O + 9.93O_2 + 49.36N_2$ out at 100 kPa, 1200 K

$$\text{a) } \bar{s}^*_{0\ FUEL} = 0.2(186.251) + 0.8(229.597)$$

$$- 8.3145(0.2 \ln 0.2 + 0.8 \ln 0.8) = 225.088$$

$$\Delta s^*_{TP} = 49.718 \ln \frac{338.2}{298.2} - 8.3145 \ln \frac{10}{0.1} = -32.031$$

From Fig. D.3: $(\bar{s}^* - \bar{s})_{FUEL} = 1.37 \times 8.3145 = 11.391$

$\bar{s}_{FUEL} = 225.088 - 32.031 - 11.391 =$ **181.66 kJ/kmol K**

b) Air at 600 K, 100 kPa

	n_i	y_i	$\bar{s}^o_i$	$-\bar{R}\ln(y_iP/P_0)$	$\bar{S}_i$
O_2	13.13	0.21	226.45	+12.976	239.426
N_2	49.36	0.79	212.177	+1.96	214.137

$S_{AIR} = \sum n_i\bar{S}_i = 13713.47$ kJ/K

$S_R = 181.66 + 13713.47 = 13895.1$ kJ/K

Products at 1200 K, 100 kPa

PROD	n_i	y_i	$\bar{s}^o_i$	$-\bar{R}\ln(y_iP/P_0)$	$\bar{S}_i$
CO_2	1.8	0.0282	279.390	+29.669	309.059
H_2O	2.8	0.0438	240.485	+26.008	266.493
O_2	9.93	0.1554	250.011	+15.479	265.490
N_2	49.36	0.7726	234.227	+2.145	236.372

$S_P = \sum n_i\bar{S}_i = 15606.1$ kJ/K

$I = T_0(S_P - S_R) - Q_{CV} = 298.15(15\ 606.1 - 13\ 895.1) + 0 =$ **510 132 kJ**

14.119

Natural gas (approximate it as methane) at a ratio of 0.3 kg/s is burned with 250% theoretical air in a combustor at 1 MPa where the reactants are supplied at T_0. Steam at 1 MPa, 450°C at a rate of 2.5 kg/s is added to the products before they enter an adiabatic turbine with an exhaust pressure of 150 kPa. Determine the turbine inlet temperature and the turbine work assuming the turbine is reversible.

$$CH_4 + \nu_{O2}(O_2 + 3.76\ N_2) \rightarrow CO_2 + 2\ H_2O + 7.52\ N_2$$

$$2\ \nu_{O_2} = 2 + 2 \ \Rightarrow\ \nu_{O_2} = 2 \quad => \quad \text{Actual } \nu_{O_2} = 2 \times 2.5 = 5$$

$$CH_4 + 5\ O_2 + 18.8\ N_2 \rightarrow CO_2 + 2\ H_2O + 3\ O_2 + 18.8\ N_2$$

C.V. combustor and mixing chamber

$$H_R + n_{H2O}\bar{h}_{H2O\ in} = H_{P\ ex}$$

$$n_{H2O} = \frac{\dot{n}_{H2O}}{\dot{n}_{Fu}} = \frac{\dot{m}_{H2O}M_{Fu}}{\dot{m}_{Fu}M_{H2O}} = \frac{2.5 \times 16.043}{0.3 \times 18.015} = 7.421\ \frac{\text{kmol steam}}{\text{kmol fuel}}$$

Energy equation becomes

$$n_{H2O}(\bar{h}_{ex} - \bar{h}_{in})_{H2O} + \left(\Delta\bar{h}_{CO2} + 2\Delta\bar{h}_{H2O} + 3\Delta\bar{h}_{O2} + 18.8\Delta\bar{h}_{N2}\right)_{ex}$$

$$= -\overset{o}{H}_{RP} = 50\ 010 \times 16.043 = 802\ 310$$

$$(\bar{h}_{ex} - \bar{h}_{in})_{H2O} = \Delta\bar{h}_{H2O\ ex} - 15072.5, \text{ so then:}$$

$$\left(\Delta\bar{h}_{CO2} + 9.421\Delta\bar{h}_{H2O} + 3\Delta\bar{h}_{O2} + 18.8\Delta\bar{h}_{N2}\right)_{ex} = 914\ 163\ \frac{\text{kJ}}{\text{kmol fuel}}$$

Trial and error on T_{ex}

$T_{ex} = 1000$ K $\Rightarrow$ LHS = 749 956 ; $T_{ex} = 1100$ K $\Rightarrow$ LHS = 867429

$T_{ex} = 1200$ K $\Rightarrow$ LHS = 987 286 $\Rightarrow$ $T_{ex} \cong$ **1139 K = $T_{in\ turbine}$**

If air then $T_{ex\ turbine} \approx 700$ K and $T_{avg} \approx 920$ K. Find $\bar{C}_{P\ mix}$ between 900 and 1000 K. From Table A.9:

$$\bar{C}_{P\ mix} = \sum n_i\bar{C}_{P_i}/\sum n_i = \frac{53.67 + 9.421(40.63) + 3(34.62) + 18.8(32.4)}{32.221}$$

$$= 35.673 \text{ kJ/kmol K}$$

$$\bar{C}_{V\ mix} = \bar{C}_{P\ mix} - \bar{R} = 27.3587 \text{ kJ/kmol}, \quad k_{mix} = 1.304$$

$$T_{ex\ turbine} = 1139\ (150 / 1000)^{0.2331} = 732 \text{ K}$$

$\Delta H_{732} = 19370.6 + 9.421(15410) + 3(13567) + 18.8(12932)$

$= 448\ 371$ kJ/kmol

$w_T = H_{in} - H_{ex} = \Delta H_{in} - \Delta H_{ex} = 914\ 163 - 448\ 371 = 465\ 792\ \frac{\text{kJ}}{\text{kmol fuel}}$

$\dot{W}_T = \dot{n}_{Fu} w_T = \dot{m}_{Fu} w_T / \hat{M}_{Fu} = (0.3 \times 465\ 792)16.043 =$ **8710 kW**

14.120

Consider one cylinder of a spark-ignition, internal-combustion engine. Before the compression stroke, the cylinder is filled with a mixture of air and methane. Assume that 110% theoretical air has been used, that the state before compression is 100 kPa, 25°C. The compression ratio of the engine is 9 to 1.

a. Determine the pressure and temperature after compression, assuming a reversible adiabatic process.

b. Assume that complete combustion takes place while the piston is at top dead center (at minimum volume) in an adiabatic process. Determine the temperature and pressure after combustion, and the increase in entropy during the combustion process.

c. What is the irreversibility for this process?

$1\ CH_4 + 1.1 \times 2\ O_2 + 3.76 \times 2.2\ N_2 \rightarrow 1\ CO_2 + 2\ H_2O + 0.2\ O_2 + 2\ H_2O$

$P_1 = 100$ kPa, $T_1 = 298.2$ K, $V_2/V_1 = 1/8$, Rev. Ad. $s_2 = s_1$

Assume $T_2 \sim 650\text{ K} \rightarrow T_{AVE} \sim 475$ K

Table A.6: $\bar{C}_{P0\ CH4} = 44.887$, $\bar{C}_{P0\ O2} = 30.890$, $\bar{C}_{P0\ N2} = 29.415$

$\bar{C}_{P0\ MIX} = (1 \times 44.887 + 2.2 \times 30.890 + 8.27 \times 29.415)/11.47 = 31.047$

$\bar{C}_{V0\ MIX} = \bar{C}_{P0} - \bar{R} = 22.732$, $k = \bar{C}_{P0}/\bar{C}_{V0} = 1.366$

a) $T_2 = T_1(V_1/V_2)^{k-1} = 298.2\ (9)^{0.366} = 666.4$ K (avg OK)

$P_2 = P_1(V_1/V_2)^k = 100\ (9)^{1.366} =$ **2011 kPa**

b) comb. 2-3 const. vol., Q = 0

$_2Q_3 = 0 = (H_3 - H_2) - \bar{R}(n_3T_3 - n_2T_2)$

$H_2 = 1\ \bar{h}^0_{f\ CH4} + n_2\ \bar{C}_{P0\ MIX}\ (T_2 - T_1)$

$H_2 = -74873 + 11.47 \times 31.047(666.4 - 298.2) = +56246$ kJ

$H_3 = 1(-393522 + \Delta\bar{h}^*_{CO2}) + 2(-241826 + \Delta\bar{h}^*_{H2O}) + 0.2\ \Delta\bar{h}^*_{O2} + 8.27\ \Delta\bar{h}^*_{N2}$

Substituting,

$1\ \Delta\bar{h}^*_{CO2} + 2\ \Delta\bar{h}^*_{H2O} + 0.2\ \Delta\bar{h}^*_{O2} + 8.27\ \Delta\bar{h}^*_{N2} - 95.366\ T_3 - 869868 = 0$

Trial & error: $T_3 =$ **2907 K**

$1 \times 147072 + 2 \times 121377 + 0.2 \times 94315 + 8.27 \times 89274$

$- 95.366 \times 2907 - 869868 \approx 0$ **OK**

$P_3 = P_2 \dfrac{n_3T_3}{n_2T_2} = P_2 \dfrac{T_3}{T_2} = 2011 \times \dfrac{2907}{666.4} =$ **8772 kPa**

c) state 1

REAC	n_i	y_i	$\bar{s}^\circ_i$	$-\bar{R}\ln(y_iP/P_0)$	$\bar{S}_i$
CH_4	1	0.0872	186.251	+20.283	206.534
O_2	2.2	0.1918	205.148	+13.730	218.878
N_2	8.27	0.7210	191.609	+2.720	194.329
	11.47				

$S_2 = S_1 = \sum n_i\bar{S}_i = 2295.17$ kJ/K

state 3

PROD	n_i	y_i	$\bar{s}^\circ_i$	$-\bar{R}\ln(y_iP/P_0)$	$\bar{S}_i$
CO_2	1	0.0872	332.213	-16.916	315.297
H_2O	2	0.1744	284.753	-22.680	262.073
O_2	0.2	0.0174	283.213	-3.516	279.697
N_2	8.27	0.7210	265.726	-34.480	231.246
	11.47				

$S_3 = \sum n_i\bar{S}_i = 2807.79$ kJ/K

$I = T_0(S_3 - S_2) = 298.2(2807.79 - 2295.17) =$ **152860 kJ**

14.121

Liquid acetylene, C_2H_2, is stored in a high-pressure storage tank at ambient temperature, 25°C. The liquid is fed to an insulated combustor/steam boiler at the steady rate of 1 kg/s, along with 140% theoretical oxygen, O_2, which enters at 500 K, as shown in Fig. P14.72. The combustion products exit the unit at 500 kPa, 350 K. Liquid water enters the boiler at 10°C, at the rate of 15 kg/s, and superheated steam exits at 200 kPa.

a. Calculate the absolute entropy, per kmol, of liquid acetylene at the storage tank state.

b. Determine the phase(s) of the combustion products exiting the combustor boiler unit, and the amount of each, if more than one.

c. Determine the temperature of the steam at the boiler exit.

a) C_2H_2: $\bar{S}^\circ_{IG\ 25°C} = 200.958$

$T_{R_1} = 298.2/308.3 = 0.967$ => From Fig. D.1: $P_{R_1} = 0.82$

$P_1 = 0.82 \times 6.14 = 5.03$ MPa, $\left(\bar{S}^* - \bar{S}\right)_1 = 3.33\bar{R} = 27.687$

$$\bar{S}_{liq\ T_1\ P_1} = \bar{S}^\circ_{T_0\ P_0} + \Delta T - \bar{R}\ln\left(P_1/P^\circ\right) + \left(\bar{S} - \bar{S}^*\right)_{P_1\ T_1} = \mathbf{140.695}\ \frac{\mathbf{kJ}}{\mathbf{kmol\ K}}$$

b) $1\ C_2H_2 + 1.4 \times 2.5\ O_2 \rightarrow 2\ CO_2 + 1\ H_2O + 1\ O_2$

$H_1 = 226731 + (-3.56 \times \bar{R} \times 308.3) = 217605$ kJ

$H_2 = 3.5(0 + 6086) = 21301$ kJ

Products $T_3 = 350\text{ K} = 76.8°C \Rightarrow P_G = 41.8$ kPa

$$y_{V\,max} = \frac{P_G}{P} = \frac{41.8}{500} = 0.0836 = \frac{n_{V\,max}}{n_{V\,max} + 2 + 1} \Rightarrow n_{V\,max} = 0.2737 = n_{V\ gas\ mix}$$

$\Rightarrow n_{liq} = 1 - 0.2737 = \mathbf{0.7263}$

Gas Mix = $\mathbf{2\ CO_2 + 0.2737\ H_2O + 1\ O_2}$

c) $H_{liq3} = 0.7263(-285830 + 18.015(321.5 - 104.9)) = -204764$ kJ

$H_{gas\ mix3} = 2(-393522 + 2036) + 0.2737(-241826 + 1756) + 1541$

$= -847138$ kJ

$H_3 = H_{liq3} + H_{gas\ mix3} = -204\ 764 - 847\ 138 = -1\ 051\ 902$ kJ

$H_3 - H_1 - H_2 = -1\ 290\ 808$ kJ

or $\dot{H}_3 - \dot{H}_1 - \dot{H}_2 = -1\ 290\ 808/26.038 = -49\ 574\text{ kW} = \dot{m}_{H_2O}\left(h_4 - h_5\right)$

$$h_5 = 42.01 + \frac{49574}{15} = 3346.9 \quad \Rightarrow \quad T_5 = \mathbf{433.4°C}$$