Instructor’s Background

- Meng. (2000): Institut Teknologi Bandung
- PhD. (2006): Universiti Teknologi Malaysia

Specialization:
- Catalyst Design for Energy Conversion
- Process Design for Energy Conversion
- Combustion Engineering
- Computational Fluid Dynamic (CFD)
SYLLABUS

1. Dasar-dasar Teknik Pembakaran (150')
2. Konsep Fundamental Kimia Sistem Pembakaran (2x150')
3. Konsep Perpindahan Panas dalam Sistem pembakaran (150')
4. Flame Impingement (150')
5. Perancangan Sistem Burner (2x150')
6. Troubleshooting Sistem Pembakaran (150')
7. Bahan Bakar untuk Proses Pembakaran (2x150')
8. Permodelan Proses Pembakaran (2x150')
9. Pengendalian Proses Pembakaran (150')
10. Keselamatan Proses Dalam Sistem Pembakaran (2x150')
11. Sistem Flare di Teknologi Pembakaran (150')

Introduction

- Combustion is the controlled release of heat from the chemical reaction between a fuel and an oxidizer.
- The fuels in the refining, petrochemical, and power generation industries are almost exclusively hydrocarbons ($C_xH_y$).
- Hydrocarbons comprise only hydrogen (H) and carbon (C) in their molecular structure.
- Natural gas and fuel oil are examples of hydrocarbon fuels.
Definitions of Natural Gas, Gas Reservoir, Gas Drilling, and Gas Production

- **What is Natural Gas?**
  - The gas obtained from natural underground reservoirs either as free gas or gas associated with crude oil.
  - Contains large amounts of methane (CH4) along with decreasing amounts of other hydrocarbons.
  - Impurities such as H2S, N2, and CO2 are often found with the gas.
  - Generally comes saturated with water vapor.

Why are Oil and Gas so Useful?

- **Oil is liquid.** Meaning that oil may be transported and delivered through pipe.
- The primary use of natural gas is as a fuel, it is also a source of hydrocarbons for petrochemical feedstocks and a major source of elemental sulfur.
- Natural gas presents many environmental advantages over petroleum and coal.
- Carbon dioxide, a greenhouse gas linked to global warming, is produced from oil and coal at a rate of about 1.4 to 1.75 times higher than from natural gas.
## Typical gas Compositions

<table>
<thead>
<tr>
<th></th>
<th>Canada (Alberta)</th>
<th>Western Colorado</th>
<th>Southwest Kansas</th>
<th>Bach Ho Field*</th>
<th>Miskar Field Tunisia</th>
<th>Rio Arriba County, New Mexico</th>
<th>Cliffside Field, Amarillo, Texas</th>
</tr>
</thead>
<tbody>
<tr>
<td>Helium</td>
<td>0.0</td>
<td>0.0</td>
<td>0.45</td>
<td>0.00</td>
<td>0.00</td>
<td>0.0</td>
<td>1.8</td>
</tr>
<tr>
<td>Nitrogen</td>
<td>3.2</td>
<td>26.10</td>
<td>14.85</td>
<td>0.21</td>
<td>16.903</td>
<td>0.68</td>
<td>25.6</td>
</tr>
<tr>
<td>Carbon</td>
<td>1.7</td>
<td>42.66</td>
<td>0.0</td>
<td>0.06</td>
<td>13.588</td>
<td>0.82</td>
<td>0.0</td>
</tr>
<tr>
<td>Hydrogen sulfide</td>
<td>3.3</td>
<td>0.0</td>
<td>0.0</td>
<td>0.00</td>
<td>0.092</td>
<td>0.0</td>
<td>0.0</td>
</tr>
<tr>
<td>Methane</td>
<td>77.1</td>
<td>29.98</td>
<td>72.89</td>
<td>70.85</td>
<td>63.901</td>
<td>96.91</td>
<td>65.8</td>
</tr>
<tr>
<td>Ethane</td>
<td>6.6</td>
<td>0.55</td>
<td>6.27</td>
<td>13.41</td>
<td>3.349</td>
<td>1.33</td>
<td>3.8</td>
</tr>
<tr>
<td>Propane</td>
<td>3.1</td>
<td>0.28</td>
<td>3.74</td>
<td>7.5</td>
<td>0.960</td>
<td>0.19</td>
<td>1.7</td>
</tr>
<tr>
<td>Butanes</td>
<td>2.0</td>
<td>0.21</td>
<td>1.38</td>
<td>4.02</td>
<td>0.544</td>
<td>0.05</td>
<td>0.8</td>
</tr>
<tr>
<td>Pentanes</td>
<td>3.0</td>
<td>0.25</td>
<td>0.62</td>
<td>2.64</td>
<td>0.630</td>
<td>0.02</td>
<td>0.5</td>
</tr>
</tbody>
</table>

* Tabular mol% data is on a wet basis (1.3 mol% water)

Source: U.S. Bureau of Mines (1972) and Jonas et al. (1999).

## Typical gas Composition in Indonesia

<table>
<thead>
<tr>
<th>Component</th>
<th>Terenggau (Malaysia) (^1)</th>
<th>Natuna (Indonesia) (^2)</th>
<th>Terrell County (Texas USA) (^3)</th>
<th>Arun (Indonesia) (^4)</th>
</tr>
</thead>
<tbody>
<tr>
<td>Methane (CH(_4))</td>
<td>80.93</td>
<td>28.0(^5)</td>
<td>45.7</td>
<td>75</td>
</tr>
<tr>
<td>Ethane (C(_2)H(_2))</td>
<td>5.54</td>
<td>-</td>
<td>0.2</td>
<td>5.5</td>
</tr>
<tr>
<td>Propane (C(_3)H(_8))</td>
<td>2.96</td>
<td>-</td>
<td>-</td>
<td>3.4(^b)</td>
</tr>
<tr>
<td>Butane (C(_4)H(_10))</td>
<td>1.40</td>
<td>-</td>
<td>-</td>
<td>-</td>
</tr>
<tr>
<td>(&gt;)C(_5)</td>
<td>-</td>
<td>-</td>
<td>-</td>
<td>0.8</td>
</tr>
<tr>
<td>Nitrogen (N(_2))</td>
<td>0.10</td>
<td>0.5</td>
<td>0.2</td>
<td>0.3</td>
</tr>
<tr>
<td>Carbon dioxide (CO(_2))</td>
<td>8.48</td>
<td>71.9</td>
<td>53.9</td>
<td>15</td>
</tr>
<tr>
<td>Hydrogen Sulfide (H(_2)S)</td>
<td>-</td>
<td>0.5</td>
<td>-</td>
<td>0.01</td>
</tr>
</tbody>
</table>

\(^1\) CH\(_4\) = low C\(_2\)-hydrocarbons  
\(^2\) C\(_3\), C\(_4\) hydrocarbons  
\(^3\) Gordon et al. (2001)  
\(^4\) Suhartanto et al. (2001)  
\(^5\) Bitter (1997)  
\(^6\) Centi et al. (2001)
In the process industries, combustion powers gas turbines, process heaters, reactors, and boilers.
The burner combusets fuel and generates products of combustion and heat.
A firebox contains the flame envelope.
The fire heats water in the tubes to boiling.
The steam rises to a steam drum that separates the liquid and vapor phases, returning water to the tubes and passing steam.
The steam may be further heated in a superheater.
Superheaters raise the temperature of the steam above the boiling point, using either radiant and/or convective heat transfer mechanisms.

- Most large boilers have water in the tubes and fire outside
  — called water-tube boilers.
- Fire-tube boilers put the fire and hot gases in tubes surrounded by water.
- Reactors such as cracking furnaces and reforming furnaces are more extreme versions of process heaters.
- Here, the process fluid undergoes chemical transformations to a different substance. For example, in an ethylene cracking furnace, liquid or gas feedstock transforms to ethylene (C2H4), an intermediate in the production of polyethylene and other plastics.
- There are many specialized types of reactors using combustion as the heat source.
Consider the combustion of methane (CH4) and air. CH4 is the major component of natural gas. The combustion of CH4 produces carbon dioxide (CO2) and water vapor (H2O).

\[
\text{fuel + oxidizer} \rightarrow \text{CO}_2 + \text{H}_2\text{O} + \text{Other species}
\]

**Stoichiometric Equation:**

\[
\text{CH}_4 + 2\text{O}_2 \rightarrow \text{CO}_2 + 2\text{H}_2\text{O}
\]

- In a stoichiometric equation, the subscripted numbers define the proportions of elements in a molecule:
  - the methane molecule comprises four hydrogen atoms for every carbon atom,
  - and that an oxygen molecule comprises two oxygen atoms
- A mole comprises 6.02 x 10^{23} molecules \(\Rightarrow\) *Avogadro’s number*
A global combustion reaction using CH$_4$ as the fuel can be written as:

\[ \text{CH}_4 + (x\text{O}_2 + y\text{N}_2) \rightarrow \text{CO}_2, \text{CO}_2, \text{H}_2, \text{H}_2\text{O}, \text{N}_2, \text{NO}_x, \text{O}_2, \text{Trace species} \]

**CONSERVATION OF MASS**

- Stoichiometric equations are always balanced equations
- Combustion reactions conserve mass; that is, the mass of the reactants must equal the mass of the products
- Stoichiometrically, 1 mole of CH$_4$ reacts with 2 moles of O$_2$ to produce 1 CO$_2$ and 2 moles of H$_2$O
- mass is conserved, but not in moles

**TABLE 2.2  Molar Ratios for Some Combustion Reactions and Products**

<table>
<thead>
<tr>
<th>Reaction</th>
<th>Moles Reactants</th>
<th>Moles Products</th>
</tr>
</thead>
<tbody>
<tr>
<td>H$_2$ + 0.5 O$_2$ → H$_2$O</td>
<td>1.5</td>
<td>1.0</td>
</tr>
<tr>
<td>CO + 0.5 O$_2$ → CO$_2$</td>
<td>1.5</td>
<td>1.0</td>
</tr>
<tr>
<td>CH$_4$ + 2 O$_2$ → CO$_2$ + 2 H$_2$O</td>
<td>3.0</td>
<td>3.0</td>
</tr>
<tr>
<td>C$_2$H$_4$ + 3 O$_2$ → 2 CO$_2$ + H$_2$O</td>
<td>2.5</td>
<td>3.0</td>
</tr>
<tr>
<td>C$_2$H$_4$ + 5 O$_2$ → 3 CO$_2$ + 4 H$_2$O</td>
<td>6.0</td>
<td>7.0</td>
</tr>
<tr>
<td>C$_2$H$_4$ + 6.5 O$_2$ → 4 CO$_2$ + 5 H$_2$O</td>
<td>8.5</td>
<td>9.0</td>
</tr>
</tbody>
</table>

*Note: Most combustion reactions do not conserve moles.*
The ideal gas law applies for typical combustion reactions and relates the pressure, volume, and number of moles.

\[ PV = nRT \]

where:
- \( P \) = Pressure of the gas, psia
- \( V \) = Volume of the gas, ft\(^3\)
- \( n \) = Number of moles
- \( R \) = Gas constant = 10.73 psia-ft/lbmol \(^\circ\)R
- \( T \) = Absolute temperature, \(^\circ\)R

Another useful form of the ideal gas law:

\[ PM = \rho RT \]

where:
- \( \rho \) = Density of the gas, lb/ft\(^3\)
- \( M \) = Molecular weight of the gas, lb/lbmol

**TABLE 2.3 Molecular Weights and Stoichiometric Coefficients for Common Gaseous Fuels**

<table>
<thead>
<tr>
<th>Common Name</th>
<th>Formula</th>
<th>Molecular Weight</th>
<th>( O_2 ) (moles)</th>
<th>( CO_2 ) (moles)</th>
<th>( H_2O ) (moles)</th>
</tr>
</thead>
<tbody>
<tr>
<td>Hydrogen</td>
<td>( H_2 )</td>
<td>2.02</td>
<td>0.5</td>
<td>0.0</td>
<td>1.0</td>
</tr>
<tr>
<td>Carbon monoxide</td>
<td>( CO )</td>
<td>28.01</td>
<td>0.5</td>
<td>1.0</td>
<td>0.0</td>
</tr>
<tr>
<td>Methane</td>
<td>( CH_4 )</td>
<td>16.05</td>
<td>2.0</td>
<td>1.0</td>
<td>2.0</td>
</tr>
<tr>
<td>Ethane</td>
<td>( C_2H_6 )</td>
<td>30.08</td>
<td>5.0</td>
<td>2.0</td>
<td>3.0</td>
</tr>
<tr>
<td>Ethene, ethylene</td>
<td>( C_2H_4 )</td>
<td>28.06</td>
<td>4.0</td>
<td>2.0</td>
<td>2.0</td>
</tr>
<tr>
<td>Acetylene, ethyne</td>
<td>( C_2H_2 )</td>
<td>26.04</td>
<td>3.0</td>
<td>2.0</td>
<td>1.0</td>
</tr>
<tr>
<td>Propane</td>
<td>( C_3H_8 )</td>
<td>44.11</td>
<td>7.0</td>
<td>3.0</td>
<td>4.0</td>
</tr>
<tr>
<td>Propene, propylene</td>
<td>( C_3H_6 )</td>
<td>42.09</td>
<td>6.0</td>
<td>3.0</td>
<td>3.0</td>
</tr>
<tr>
<td>Butane</td>
<td>( C_4H_{10} )</td>
<td>58.14</td>
<td>7.0</td>
<td>4.0</td>
<td>5.0</td>
</tr>
<tr>
<td>Butene, butylene</td>
<td>( C_5H_{12} )</td>
<td>56.12</td>
<td>8.0</td>
<td>4.0</td>
<td>4.0</td>
</tr>
<tr>
<td>Generic hydrocarbon</td>
<td>( C_{x}H_{y} )</td>
<td>12.01 ( x + 1.01 y ) x + y/2</td>
<td>x</td>
<td>y/2</td>
<td></td>
</tr>
</tbody>
</table>
STOICHIOMETRIC RATIO AND EXCESS AIR

- The stoichiometric coefficient for oxygen → the theoretical oxygen required for combustion.
- The theoretical air comprising this amount of oxygen, it is necessary to define a mole of air as: 0.21 lbmol O2 + 0.79 lbmol N2.
- Therefore:

\[ CH_4 + 2O_2 \rightarrow CO_2 + 2H_2O \]

- In industrial practice, perfect mixing cannot be achieved. It is actually more cost-effective to ensure complete combustion with the addition of excess air.
- Excess air is that amount beyond theoretical added to ensure complete combustion of the fuel.
- To account for excess air, the above Equation is modified with $\varepsilon$, the fraction of excess air:

\[
CH_4 + 2(1 + \varepsilon)\left[O_2 + \frac{79}{21} N_2\right] \rightarrow \\
CO_2 + 2H_2O + 2\varepsilon O_2 + 2(1 + \varepsilon)\left(\frac{79}{21}\right) N_2
\]

- Two important chemical features of complete combustion → no carbon monoxide (CO) and no some unreacted oxygen appear in the combustion products.
- To account for any hydrocarbon fuel, the Equation is modified by $x$ (the H/C molar ratio).
- The following Equation gives a generic equation for hydrocarbons with air:

\[
CH_x + (1 + \varepsilon)(1 + x/4)\left[O_2 + \frac{79}{21} N_2\right] \rightarrow \\
CO_2 + x/2 H_2O + \varepsilon(1 + x/4) O_2 \\
+ (1 + x/4)(1 + \varepsilon)\left(\frac{79}{21}\right) N_2
\]
Volumes of flue gas species to excess air for a given fuel H/C:
- oxygen-to-fuel ratio (O/F):
  \[ \frac{O}{F} = (1 + \frac{e}{4})(1 + e) \]
- air-to-fuel ratio (A/F):
  \[ \frac{A}{F} = (100/21)(O/F) \]
- Total wet product (TWP):
  \[ \text{TWP} = \frac{A}{F} + \frac{e}{4} \]
- Total dry product (TDP):
  \[ \text{TDP} = \frac{A}{F} - \frac{e}{4} \]

In situ analyzers measure the flue gas species in the actual hot wet environment ➞ wet
extractive analyzers remove the flue gas, condense the water, and measure the concentration of the flue gas species in the dry gas ➞ dry

Volume fraction (f):

\[
\begin{align*}
  f_{O_2,\text{wet}} &= \frac{\varepsilon(1 + e/4)}{\text{TWP}} \\
  f_{O_2,\text{dry}} &= \frac{\varepsilon(1 + e/4)}{\text{TDP}} \\
  f_{CO_2,\text{wet}} &= \frac{1}{\text{TWP}} \\
  f_{CO_2,\text{dry}} &= \frac{1}{\text{TDP}} \\
  f_{H_2,\text{wet}} &= \frac{121(1 + e)(1 + e/4)}{21 \text{TWP}} \\
  f_{H_2,\text{dry}} &= \frac{121(1 + e)(1 + e/4)}{21 \text{TDP}} \\
  f_{O_2,\text{wet}} &= \frac{\varepsilon}{2 \text{TWP}}
\end{align*}
\]

Because of the strong relationship between oxygen and excess air, the excess oxygen can be used as a measure of excess air:

\[
\begin{align*}
  f_{O_2,\text{wet}} &= \frac{0.21 \epsilon}{K_{\text{wet}} + \epsilon} \\
  f_{O_2,\text{dry}} &= \frac{0.21 \epsilon}{K_{\text{dry}} + \epsilon} \\
  K_{\text{wet}} &= \frac{4 + 1.21x}{4 + x} \\
  K_{\text{dry}} &= \frac{4 + 0.79x}{4 + x} \\
  \epsilon &= K_{\text{wet}} \left( \frac{f_{O_2,\text{wet}}}{0.21 - f_{O_2,\text{wet}}} \right) \\
  \epsilon &= K_{\text{dry}} \left( \frac{f_{O_2,\text{dry}}}{0.21 - f_{O_2,\text{dry}}} \right)
\end{align*}
\]
Heat of Combustion

- One measure of the chemical energy of a fuel is the **heat of combustion**
- Heat of combustion is reported as either **net heating value** (lower heating value, LHV) or **gross heating value** (higher heating value, HHV).
- When methane burns, it produces two products: CO₂ and H₂O. The CO₂ is a gas, however, H₂O is either a liquid or a vapor, depending on how much heat is extracted from the process.
- If so much heat is extracted that the H₂O **condenses**, then the combustion yields its **HHV**.
- If **water is released from the stack as a vapor**, then combustion yields the **LHV**. The LHV excludes the heat of vaporization.
- The process industry usually uses the LHV. Boiler and turbine calculations usually use the HHV must be consistent.

Heating Value

- **Determination of the heating value** of a fuel involves two arbitrary but conventional standard states for the water formed in the reaction:
  - All the water formed is a liquid (gross heating value, frequently called higher heating value [HHV])
  - All the water formed is a gas (net heating value, frequently called lower heating value [LHV])
- The heating value is normally calculated at 60°F and 1 atm (15.6°C and 1.01 atm), standard conditions for the gas industry, and, thus at equilibrium, the water would be partially liquid and partially vapor.
- Heating values for custody transfer are determined either by direct measurement, in which bomb calorimetry is used, or by computation of the value on the basis of gas analysis.

\[
H_v^{(dry)} = \sum_{j=1}^{s} x_j H_v^{d_j} \\
H_v^{(sat)} = (1 - x_v) \sum_{j=1}^{s} x_j H_v^{d_j}
\]
Example of Heating Value Calculation

- **Problem:** A natural gas has a volumetric analysis of 95% CH₄, 3% C₂H₆, and 2% CO₂. For conditions of 14.7 psia and 77°F, calculate (a) the higher heating value of the fuel, Btu/ft³ of gas; and (b) the lower heating value of the fuel, Btu/ft³ of gas.

- **Solution:**
  - **Stoichiometric equation:**
    \[
    \begin{align*}
    0.95 \text{CH}_4 + 0.03 \text{C}_2\text{H}_6 + 0.02 \text{CO}_2 &\rightarrow b\text{CO}_2 + c\text{H}_2\text{O} + d\text{N}_2 \\
    \text{Energy balance:} & \\
    b = 1.03 \\
    c = 1.99 \\
    a = 2.005 \\
    d = 3.76a = (3.76)(2.005) = 7.539
    \end{align*}
    \]

- **Energy balance:**
  \[
  Q = \sum_{i=\text{prod}} N_i[\overline{h}_f^0 + \Delta\overline{h}]_i = \sum_{j=\text{react}} N_j[\overline{h}_f^0 + \Delta\overline{h}]_j
  \]

  or
  \[
  Q = 1.03[\overline{h}_f^0 + \Delta\overline{h}]_{\text{CO}_2} + 1.99[\overline{h}_f^0 + \Delta\overline{h}]_{\text{H}_2\text{O}} \\
  + (3.76)(2.005)[\overline{h}_f^0 + \Delta\overline{h}]_{\text{N}_2} - 0.95[\overline{h}_f^0 + \Delta\overline{h}]_{\text{CH}_4} \\
  - 0.03[\overline{h}_f^0 + \Delta\overline{h}]_{\text{C}_2\text{H}_6} - 0.02[\overline{h}_f^0 + \Delta\overline{h}]_{\text{CO}_2} \\
  - 2.005[\overline{h}_f^0 + \Delta\overline{h}]_{\text{O}_2} - (2.005)(3.76)[\overline{h}_f^0 + \Delta\overline{h}]_{\text{N}_2}
  \]

  - **Higher heating value assumes water in the products is a liquid**

  \[
  Q = (1.03)(-94,054) + 1.99(-68,317) - 0.95(-17,889) \\
  - 0.03(-20,236) - 0.02(-94,054) = -213,344
  \]

  \[
  \overline{HHV} = 213,340 \text{ cal/gmole} \\
  = (213,340 \text{ cal/gmole})(1.8001 \text{ Btu/llbmole/cal/gmole}) \\
  = 384,033 \text{ Btu/llbmole}
  \]
**Density of Fuel:**

\[
\bar{\rho} = \frac{P}{R T} = \frac{(14.7 \text{ lb/in}^2)(144 \text{ in}^2 / \text{ft}^2)}{(1.545 \text{ ft} \cdot \text{lb} / \text{lbmole} \cdot \text{°R})(537°\text{R})} \\
= 0.00255 \text{ lbmole/ft}^3
\]

- Higher heating value, water as liquid in product:
  \[HHV = (384,033 \text{ Btu/lbmole})(0.00255 \text{ lbmole/ft}^3)\]
  a. \(HHV = 980 \text{ Btu/ft}^3\)

- Lower heating value, water as vapor in product:
  \[LHV = HHV - 1.99 \text{ h}_f \{68°F\} \]
  \[LHV = +384,033 \text{ Btu/lbmole fuel} \]
  \[- \left(1.99 \text{ lbmole H}_2\text{O} \right) \left(1.054 \text{ Btu/lbm H}_2\text{O} \right) \left(18 \frac{\text{ lbm}}{\text{lbmole H}_2\text{O}} \right)\]
  b. \(LHV = 346,279 \text{ Btu/lbmole}\)

---

**Adiabatic Flame Temperature**

- **Heat** (Q) \(=====>\) **Temperature** (T)\\
- **Heat is energy** in transit. When a body absorbs heat, it stores it as another form of energy, increasing the body’s **temperature** and expanding it.
- The same amount of heat absorbed in different materials will yield different temperature increases and expansions \(\Rightarrow\) depends on **heat capacity** (Cp)
- For a given mass of fuel, m, the total energy is:

\[Q = m\Delta H = mC_p\Delta T\]
Example 1:

- **Problem**: If 1 lb CH₄ combusts in 15% excess air, what is the maximum possible flame temperature.
- **Solution**:
  - From stoichiometric equation:
    
    \[
    \text{CH}_4 + (1+\varepsilon)(1+x/4)[O_2 + 79/21 N_2] \rightarrow \\
    \text{CO}_2 + x/2 \text{H}_2\text{O} + \varepsilon(1+x/4) \text{O}_2 \\
    +(1+x/4)(1+\varepsilon)(79/21) \text{N}_2 \\
    \]
    
    \[
    \text{CH}_4 + 1.15 \times 2(O_2 + 79/21 N_2) \rightarrow \\
    \text{CO}_2 + 2 \text{H}_2\text{O} + 0.30 \text{O}_2 + 8.65 \text{N}_2 \\
    \]
  
  - Use the basis of 1 lb CH₄ and ratio all other components by the molecular weight of CH₄ to obtain the following:

<table>
<thead>
<tr>
<th>IN</th>
<th>OUT</th>
</tr>
</thead>
<tbody>
<tr>
<td>CH₄</td>
<td>1.00 lb CO₂</td>
</tr>
<tr>
<td>1*(44.01/16.05)</td>
<td>2.74 lb H₂O</td>
</tr>
<tr>
<td>2*(18.02/16.05)</td>
<td>2.24 lb</td>
</tr>
<tr>
<td>O₂: 1.15<em>2</em>(32.00/16.05)</td>
<td>5.59 lb O₂</td>
</tr>
<tr>
<td>N₂: 1.15<em>2</em>79/21*(28.02/16.05)</td>
<td>15.11 lb N₂</td>
</tr>
<tr>
<td>20.7 lb</td>
<td>20.7</td>
</tr>
</tbody>
</table>
• The highest possible flame temperature presumes no loss from the flame whatsoever. This is known as the adiabatic flame temperature.

• total energy from combustion of 1 lb of fuel:
  
  \[ 1 \text{ lb} \times 22,000 \text{ Btu/lb} = 22,000 \text{ Btu} \]

  \[ \Delta T = \Delta H / mC_p \]

• Because there are several species in the flue gas, the contribution of each species must be used for \( mC_p \).

  \[ mC_p = m_{\text{CO}_2} C_{\text{pCO}_2} + m_{\text{H}_2\text{O}} C_{\text{pH}_2\text{O}} + m_{\text{O}_2} C_{\text{pO}_2} + m_{\text{N}_2} C_{\text{pN}_2} \]

• If \( C_p = 0.30 \text{ BTU/lb.}^{\circ}\text{F} \), then the adiabatic flame temperature becomes:

  \[ \Delta T = 22,000 \frac{\text{Btu}}{\text{lb}} \left( \frac{\text{lb}^{\circ}\text{F}}{0.30 \text{ Btu} \left( \frac{1}{20.7 \text{ lb}} \right)} \right) = 3543^{\circ}\text{F} \]

• For air and fuel at 60\(^{\circ}\text{F}\), the adiabatic fuel temperature becomes:

  \[ \text{AFT} = 3543 + 60 \Rightarrow 3603^{\circ}\text{F} \]

• Note that the actual flame temperature will be much cooler than this, because heat will transfer from the flame to the surroundings via convection and radiation.

• at high temperatures, \( \text{CO}_2 \rightarrow \text{CO} + 0.5 \text{O}_2 \) and \( \text{H}_2\text{O} \rightarrow \text{H}_2 + 0.5 \text{O}_2 \), reducing the adiabatic flame temperature.
### TABLE 2.1
Adiabatic Flame Temperatures

<table>
<thead>
<tr>
<th>Fuel</th>
<th>Air</th>
<th>O₂</th>
</tr>
</thead>
<tbody>
<tr>
<td></td>
<td>⁰F</td>
<td>K</td>
</tr>
<tr>
<td>H₂</td>
<td>3807</td>
<td>2370</td>
</tr>
<tr>
<td>CH₄</td>
<td>3542</td>
<td>2223</td>
</tr>
<tr>
<td>C₂H₂</td>
<td>4104</td>
<td>2535</td>
</tr>
<tr>
<td>C₂H₄</td>
<td>3790</td>
<td>2361</td>
</tr>
<tr>
<td>C₃H₆</td>
<td>3607</td>
<td>2259</td>
</tr>
<tr>
<td>C₄H₈</td>
<td>4725</td>
<td>2334</td>
</tr>
<tr>
<td>C₅H₁₀</td>
<td>3610</td>
<td>2261</td>
</tr>
<tr>
<td>C₆H₁₂</td>
<td>3583</td>
<td>2246</td>
</tr>
<tr>
<td>CO</td>
<td>3826</td>
<td>2381</td>
</tr>
</tbody>
</table>

---

**Graphs:**

- **Graph 1:**
  - Title: Adiabatic Flame Temperature (⁰F, K)
  - X-axis: H₂ in CH₄ (Vol. %)
  - Y-axis: Adiabatic Flame Temperature (⁰F, K)
  - Data points for different air conditions and oxygen concentrations.

- **Graph 2:**
  - Title: Adiabatic Flame Temperature (⁰F, K)
  - X-axis: O₂/CH₄ Stoichiometry
  - Y-axis: Adiabatic Flame Temperature (⁰F, K)
  - Data points for different oxygen concentrations.

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16
NEXT ON “Sub-stoichiometric Combustion; equilibrium thermodynamics; Combustion Kinetics; Flame Properties”