A hydrocarbon has a density of 1.02 g/L at 745.3 torr and 25.3 °C. The complete combustion of the hydrocarbon produced carbon dioxide gas and water vapor.

\[ \text{Molecular formula?} \]

\[ pV = nRT \implies \rho m = d \frac{RT}{M} \]

\[ (\text{CH})_2 \implies \text{C}_2\text{H}_2 \]

The product gases exerted a pressure of 8.75 atm when confined in a 0.50 L container at 125 °C.

\[ \frac{p_{\text{tot}} V}{n_{\text{tot}} RT} = \frac{n_{\text{tot}}}{N_{\text{mol}}} \times \frac{0.08206 \text{ L atm}}{\text{mol K}} \times 398 \text{ K} = 0.133 \text{ mol} \]

When the product gases were cooled to 0.0 °C, the pressure fell to 4.00 atm (volume still 0.50 L).

\[ \frac{p_{\text{CO}_2} V}{n_{\text{CO}_2} RT} = \frac{n_{\text{CO}_2}}{N_{\text{mol}}} \times \frac{0.08206 \text{ L atm}}{\text{mol K}} \times 273.2 \text{ K} \]

\[ n_{\text{CO}_2} = 0.0892 \text{ mol} \]
\[ 0.133 \text{ mol CO}_2 + \text{H}_2\text{O} \]
\[ - 0.0892 \text{ mol CO}_2 \]
\[ = 0.0438 \text{ mol H}_2\text{O} \]

\[ 0.0892 \text{ mol CO}_2 \Rightarrow 0.0892 \text{ mol C} \]

\[ 0.0438 \text{ mol H}_2\text{O} \Rightarrow 0.0876 \text{ mol H} \]

\[ \text{CH empirical formula} \]

\[ \text{formula mass} \quad 13.0 \text{ g/mol} \]

\[ \text{mol "formula unit"} \]
\[ P = nRT \]

\[ P = 745.3 \text{ torr} \]
\[ T = 25.3 \, \text{°C} \]
\[ d = 1.02 \text{ g} \]
\[ \frac{1}{2} \]

\[ \frac{745.3 \text{ torr}}{1 \text{ bar}} \left( \frac{1 \text{ atm}}{760 \text{ torr}} \right) (1 \text{ L}) = n \left( \frac{0.08206 \text{ L atm}}{\text{mol K}} \right) (298.15 \text{ K}) \]

\[ n = 0.0400 \text{ mol HCl} \]

\[ M = \frac{1.02 \text{ g HCl}}{0.0400 \text{ mol HCl}} = 25.5 \text{ g/mol} \]

\[ \text{HCl = original hydrocarbon sample} \]

\[ \frac{26 \text{ g}}{1 \text{ mol C-H}} \times \frac{1 \text{ mol C-H}}{13 \text{ g}} = \frac{2 \text{ mol C-H}}{1 \text{ mol C-H}} \]
How much of the error in a gas law calculation results from the non-ideality of real gases?

Molar Volume at STP

Ideal 22.414 L

He 22.434 L

CO₂ 22.260 L

NH₃ 22.079 L

\[
\frac{22.414 - 22.079}{22.414} = 0.015
\]

Reasons for deviations:

1. Particles of real gas are very small but they do have some volume.

   \[ V_{\text{real gas}} > V_{\text{ideal}} \]

2. Particles of real gas do have some attraction for one another.

   \[ V_{\text{real}} < V_{\text{ideal gas}} \]
He, He molecule volume dominates
(no attraction between He, He or Ne, He)

At "high" temp... low ?

attraction dominate
Compressibility Factor \( (Z = PV/nRT) \)

Plot \( Z \) versus \( P \)

At low pressures, attractions are important.
At high pressures, molecule size is more important.
At sufficiently high pressures, all particles will begin to repel one another.
Critical Data

Critical Temperature ($T_c$): highest temperature at which a gas can be liquefied

Critical Pressure ($P_c$): pressure that must be applied to liquefy the gas at its critical temperature

For water, $T_c = 374 \, ^\circ C$. For hydrogen, $T_c = -271 \, ^\circ C$.

Value of $T_c$ tells us something about the strength of the interactions between particles.
Van der Waals’ forces

Natural gas (methane) is a gas at room temperature and atmospheric pressure.

Gasoline is a liquid at room temperature and atmospheric pressure.

Wax is a solid at room temperature and atmospheric pressure.

Dipole-Dipole (polarity)

\[
\begin{align*}
C_2H_6 & \quad \text{non-polar} \quad \text{bp} = -89 \, ^\circ C \quad \boxed{\text{MW} = 30} \\
H_2S & \quad \text{polar} \quad \text{bp} = -61 \, ^\circ C \quad \boxed{\text{MW} = 34}
\end{align*}
\]
Polarization
original vdW force

δ+ δ-
induced dipole

bigger particles have more chance for their electrons to be moved from spherical distribution
### Hydrogen bonding

<table>
<thead>
<tr>
<th>Compound</th>
<th>bp</th>
<th>MW</th>
</tr>
</thead>
<tbody>
<tr>
<td>$\text{H}_2\text{O}(\ell)$</td>
<td>100 °C</td>
<td>18</td>
</tr>
<tr>
<td>$\text{H}_2\text{S}(\ell)$</td>
<td>-61 °C</td>
<td>34</td>
</tr>
<tr>
<td>$\text{H}_2\text{Se}(\ell)$</td>
<td>-42 °C</td>
<td>81</td>
</tr>
<tr>
<td>$\text{H}_2\text{Te}(\ell)$</td>
<td>-2 °C</td>
<td>130</td>
</tr>
</tbody>
</table>

### Ionic forces

<table>
<thead>
<tr>
<th>Compound</th>
<th>mp</th>
</tr>
</thead>
<tbody>
<tr>
<td>$\text{NaCl}$</td>
<td>800 °C</td>
</tr>
</tbody>
</table>
A sample of a gaseous binary compound of boron and chlorine weighing 2.842 g occupies 0.153 L at STP. This sample is decomposed to give solid boron and gaseous chlorine. At STP, the chlorine gas occupies 0.688 L. Determine the molecular formula of the compound.

\[
\text{Bpd} \quad 0.153 \text{ L} \left( \frac{1 \text{ mol}}{22.4 \text{ L}} \right) = 0.00683 \text{ mol Bpd} \quad \text{at STP}
\]
\[
\text{Cl}_2 \quad 0.688 \text{ L} \left( \frac{1 \text{ mol}}{22.4 \text{ L}} \right) = 0.0307 \text{ mol Cl}_2 = 0.0614 \text{ mol Cl atoms}
\]
\[
2.842 \text{ g Bpd} - 2.1795 \text{ g Cl} = 0.6625 \text{ g B}
\]
\[
0.6625 \text{ g B} \left( \frac{1 \text{ mol B}}{10.81 \text{ g B}} \right) = 0.06129 \text{ mol B}
\]  
empirical formula is BCl

\[
\begin{align*}
\text{Bpd} & \quad \frac{2.842 \text{ g}}{0.00683 \text{ mol}} = 4.16 \text{ g Bpd mol}^{-1} \\
4.16 \text{ g Bpd} \left( \frac{1 \text{ mol BCl units}}{46.267 \text{ g BCl units}} \right) & = 0.0899 \text{ mol BCl units}
\end{align*}
\]

Molecular formula \( \text{B}_9 \text{Cl}_{19} \)