1) We need to find the percentages of all the elements present (since there were two different analyses), using all the data we are given. From these data, we can find the empirical formula, and using the last bit of data we are given, the \( M_m \) and the molecular formula. Let’s start with the C, H data.

\[
\text{? g C} = 0.2766 \text{g CO}_2 \left( \frac{12.011 \text{g C}}{44.01 \text{g CO}_2} \right) = 0.07549 \text{ g C} \quad \text{and} \quad \% \text{C} = \frac{0.07549 \text{g}}{0.1023 \text{g}} (100) = 73.79%.
\]

\[
\text{? g H} = 0.0991 \text{ g H}_2\text{O} \left( \frac{2.016 \text{g H}}{18.02 \text{g H}_2\text{O}} \right) = 0.0111 \text{ g H} \quad \text{and} \quad \% \text{H} = \frac{0.0111 \text{g}}{0.1023 \text{g}} (100) = 10.84%.
\]

Now we can use the N data to get the percent N, and from that, finally, the percent O.

\[
\text{? moles N}_2 = \frac{(1 \text{ atm})(0.0276 \text{L})}{(0.08206 \text{L atm/K mol})(273.15 \text{K})} = 1.23 \times 10^{-3} \text{ moles which equals 0.0345 g N, and a } \% \text{ N} = \frac{0.0345 \text{g}}{0.4831 \text{g}} (100) = 7.14%. \text{ These data leave 8.23% O in our molecule.}
\]

Using our normal set up, we can obtain:

<table>
<thead>
<tr>
<th>Element</th>
<th>Mass (out of 100)</th>
<th>Moles</th>
<th>Simplest ratio</th>
</tr>
</thead>
<tbody>
<tr>
<td>C</td>
<td>73.79</td>
<td>6.144</td>
<td>12</td>
</tr>
<tr>
<td>H</td>
<td>10.84</td>
<td>10.754</td>
<td>21</td>
</tr>
<tr>
<td>N</td>
<td>7.14</td>
<td>0.510</td>
<td>1</td>
</tr>
<tr>
<td>O</td>
<td>8.23</td>
<td>0.514</td>
<td>1</td>
</tr>
</tbody>
</table>

These data give us an **empirical formula of C\(_{12}\)H\(_{21}\)NO**, and an empirical formula mass of 195 g/mol. Now on to the molar mass part of the problem.

We are told the density at certain conditions, We can use \( P M_m = dRT \) to get the \( M_m \).

\[
M_m = \frac{(4.02 \text{ g }/ \text{ L})(0.08206 \text{ L atm/K mol})(400 \text{ K})}{256/760} = 391.73 \text{ g/mol which is just about twice the empirical mass, so the **molecular formula is C}\(_{24}\)H\(_{42}\)N\(_2\)O\(_2\)**}
\]

2) We can not use \( P M_m = dRT \), since we do not have a pressure. (In fact, we are asked to find the total pressure.) We need to determine the amount of each gas present from its own \( M_m \).
a) We can obtain the $M_m$ of both $CH_4$ and $C_2H_6$: $M_m(CH_4) = 16.043 \text{ g/mol}$; $M_m(C_2H_6) = 30.07 \text{ g/mol}$. Now, if we let $x = \text{ moles of } CH_4$, and $y = \text{ moles of } C_2H_6$, we can say:

\[
0.915 = \frac{x}{x + y} \quad \text{(that is, the mole fraction of } CH_4), \quad \text{and} \quad 0.085 = \frac{y}{x + y} \quad \text{(that is, the mole fraction of } C_2H_6). \quad \text{We can also say:} \quad 16.043x + 30.07y = 15.50 \quad \text{(that is, the amount of mass each contributes summed must be the total mass.)}
\]

Solving one of the first two equations for one of the variables, we get:

\[
y = \frac{0.085x}{0.915} = 9.29 \times 10^{-2} x. \quad \text{We can now substitute for } y \text{ in the second equation to get:} \quad 16.043x + 30.07(0.0929x) = 15.50, \quad \text{or} \quad 18.836x = 15.50, \quad \text{yielding } x = 0.823 \text{ moles, } y = 0.076 \text{ moles.}
\]

b) We now have a total number of moles, which is $0.823 + 0.076 = 0.899 \text{ moles}$. We can now plug this into $PV = nRT$, and get the pressure:

\[
P = \frac{(0.899 \text{ moles})(0.08206L \cdot \text{atm} / \text{K} \cdot \text{mol})(293.15K)}{15.00L} = 1.44 \text{ atm.}
\]

c) Our individual partial pressures are merely the mole fraction of the substance times the total pressure, so $P_{\text{methane}} = (0.915)(1.44 \text{ atm}) = 1.32 \text{ atm}$; $P_{\text{ethane}} = 0.122 \text{ atm}$. 

3) The balanced equation is either:

\[
2 \text{HCl} + Na_2CO_3 \rightarrow H_2O + CO_2 + 2 \text{NaCl}
\]

or

\[
H_2SO_4 + Na_2CO_3 \rightarrow H_2O + CO_2 + Na_2SO_4.
\]

We are given the volume of CO$_2$ produced, the temperature and pressure, and the concentration and volume of the acid used. From all these, we can make the necessary determination. Using $PV = nRT$, we can get the number of moles of CO$_2$, which is: $n = \left(\frac{722 \text{ mmHg}}{760 \text{ mmHg} / \text{atm}}\right)(0.125L) = 5.0 \times 10^{-3} \text{ moles CO}_2$. From the volume and molarity of the acid given, we can get the number of moles of acid used: moles acid = $(0.2003 \text{ M})(0.04980 \text{ L}) = 1.02 \times 10^{-2} \text{ moles acid}$. Since the number of moles of acid = 2(moles CO$_2$), the only stoichiometric equation which fits is that for HCl, so we have hydrochloric acid, HCl.

4) Our two equations we need to use are:

\[
BaO(s) + CO_2(g) \rightarrow BaCO_3(s)
\]
and

$$\text{CaO}(s) + \text{CO}_2(g) \rightarrow \text{CaCO}_3(s)$$

We are also given a volume, temperature, and pressure for our starting and ending \text{CO}_2(g), so we can get the number of moles consumed. Then, solving two equations in two unknowns (just like the last problem set) we can get our answer. Here we go.

The number of moles, \(n_i\), of \text{CO}_2 to start is:

$$n_i = \frac{(750 \text{torr} / 760 \text{torr} / \text{atm}) \times (1.50 \text{L})}{(0.08206 \text{L} \cdot \text{atm} / \text{K} \cdot \text{mol})(303.15 \text{K})} = 0.0595 \text{ moles}$$

the number of moles, \(n_f\), of \text{CO}_2 at the end is:

$$n_f = \frac{(230 \text{torr} / 760 \text{torr} / \text{atm}) \times (1.50 \text{L})}{(0.08206 \text{L} \cdot \text{atm} / \text{K} \cdot \text{mol})(303.15 \text{K})} = 0.0182$$

so we now know that

$$(0.0595 - 0.0182) = 0.0413 \text{ moles}$$

of \text{CO}_2 was consumed in the reaction.

Let \(x = \text{mass of BaO}\), and \((5.14 - x) = \text{mass of CaO}\). Then we can say:

$$0.0413 \text{ moles} = \frac{x \text{g}}{153.33} + \frac{(5.14 - x) \text{g}}{56.08}$$

which is the same as saying:

$$0.0413 \text{ moles} = 6.521 \times 10^{-3} x + 0.09165 - 1.783 \times 10^{-2} x.$$ This equation reduces to:

$$-0.0504 = -0.0113 x,$$

and

$$x = 4.46 \text{ g} = \text{mass of BaO}; (5.14 - 4.46) = 0.68 \text{ g} = \text{mass of CaO}.$$ Our percentages are, therefore, 86.8\% BaO, 13.2\% CaO.

5) There are two equations which are our overriding concern: \(P \text{M}_m = dRT\), (which relates density, pressure, temperature, and molar mass), and \(\frac{P_1V_1}{T_1} = \frac{P_2V_2}{T_2}\) (which relates temperature, pressure, and volume.) In the first case, we know from the second equation that as temperature goes up (at constant volume), so does pressure; since \(M_m\) and \(R\) remain constant (for a given gas), the only way this equation remains an equality is if density remains the same too. (That is, we know \(P = cT\), and with \(M_m\) and \(R\) constants, \(d\) must also be constant for the value of \(c\) to be constant.) Now, in the second case, when pressure is constant (atmospheric), and volume can vary, as the same amount of gas (assumed) occupies a greater volume, density, being mass/volume, must decrease.