1. From problem statement, for isopropanol $P_1^* = 1008$ Torr and for $n$-decane $P_2^* = 48.3$ Torr.

Using the relations $P_i = x_i^g P_{total}$, $a_i = P_i / P_i^*$, and $\gamma_i = \frac{a_i}{x_i}$, the Raoult’s law standard state, the calculated activities and activity coefficients are:

<table>
<thead>
<tr>
<th>$P$(Torr)</th>
<th>$x_i^f$</th>
<th>$x_i^g$</th>
<th>$a_1$</th>
<th>$a_2$</th>
<th>$\gamma_1$</th>
<th>$\gamma_2$</th>
</tr>
</thead>
<tbody>
<tr>
<td>942.6</td>
<td>0.1312</td>
<td>0.0243</td>
<td>0.912</td>
<td>0.474</td>
<td>1.05</td>
<td>3.61</td>
</tr>
<tr>
<td>909.6</td>
<td>0.2040</td>
<td>0.0300</td>
<td>0.875</td>
<td>0.565</td>
<td>1.10</td>
<td>2.77</td>
</tr>
<tr>
<td>883.3</td>
<td>0.2714</td>
<td>0.0342</td>
<td>0.846</td>
<td>0.625</td>
<td>1.16</td>
<td>2.30</td>
</tr>
<tr>
<td>868.4</td>
<td>0.3360</td>
<td>0.0362</td>
<td>0.831</td>
<td>0.651</td>
<td>1.25</td>
<td>1.94</td>
</tr>
<tr>
<td>830.2</td>
<td>0.4425</td>
<td>0.0411</td>
<td>0.790</td>
<td>0.706</td>
<td>1.42</td>
<td>1.60</td>
</tr>
</tbody>
</table>

2. The data is

<table>
<thead>
<tr>
<th>$x_{B_0}$</th>
<th>$P$ (Torr)</th>
<th>$x_{B_0}$</th>
<th>$P$ (Torr)</th>
</tr>
</thead>
<tbody>
<tr>
<td>0.00394</td>
<td>1.52</td>
<td>0.0130</td>
<td>5.43</td>
</tr>
<tr>
<td>0.00420</td>
<td>1.60</td>
<td>0.0236</td>
<td>9.57</td>
</tr>
<tr>
<td>0.00599</td>
<td>2.39</td>
<td>0.0238</td>
<td>9.83</td>
</tr>
<tr>
<td>0.0102</td>
<td>4.27</td>
<td>0.0250</td>
<td>10.27</td>
</tr>
</tbody>
</table>

and plotted
The best fit line in the plot is \( P_{Br_2} \text{(Torr)} = 413 x_{Br_2} - 0.063 \). Therefore, the Henry’s law constant in terms of mole fraction is 413 Torr.

3. Since \( \mu_i = \mu_i^* + RT \ln x_i \) and Gibbs-Duhem states that

\[
d\mu_2 = -\frac{x_i d\mu_1}{x_2} = -\frac{x_i d(\mu_i^* + RT \ln x_i)}{x_2} = -RT \frac{x_i d(\ln x_i)}{x_2} = -RT \frac{x_i d(x_i)}{x_i x_2}.
\]

Since \( x_1 + x_2 = 1 \), \( dx_1 = -dx_2 \) and the above is

\[
d\mu_2 = RT \frac{dx_2}{x_2}. \]

Integrating this gives \( \mu_2 = C + RT \ln x_2 \) with \( C \) a constant of integration. 

\( C \) can be fixed by noting that, since by assumption the relation holds over the entire composition range, we let \( x_2 \to 1 \) whereby \( C = \mu_2^* \), the pure \( 2 \) value.

Therefore, \( \mu_2 = \mu_2^* + RT \ln x_2 \).

4. Using Henry’s law \( x_{N_2}^l = K_{H}^{-1} p_{N_2} \) and noting that

\[
x_{N_2}^l = \frac{n_{N_2}}{n_{N_2} + n_{H_2O}} \approx \frac{n_{N_2}}{n_{H_2O}} \text{ for moles in the liquid},
\]

we have

\[
n_{N_2} = n_{H_2O} K_{H}^{-1} p_{N_2}.
\]

For 1 bar,

\[
n_{N_2} = n_{H_2O} K_{H}^{-1} p_{N_2} = \frac{5.0 \times 10^3 \text{ g}}{18.02 \text{ g/mol}} \times \frac{8.64 \times 10^4 \text{ bar}}{0.8 \text{ bar}} = 2.57 \times 10^{-3} \text{ mol}
\]

For 50 bar,

\[
n_{N_2} = 50 \times 2.57 \times 10^{-3} \text{ mol} = 0.128 \text{ mol}.
\]

b) Using ideal gas, getting additional volume relative to 1 bar volume,

\[
V = \frac{n_{N_2}(p = 50 \text{ bar}) - n_{N_2}(p = 1 \text{ bar})}{p R T} = \frac{(0.128 - 0.00257) \text{ mol} \times 0.08314 \text{ L bar mol}^{-1} \text{ K}^{-1} \times 300 \text{ K}}{1.00 \text{ bar}} = 3.13 \text{ L}
\]

The poor diver probably exploded.
5. It is found that the dissolution of 5.25 g of a substance in 565 g of benzene at 298 K raises the boiling point by 0.625ºC. Note that \( K_f = 5.12 \text{ K kg mol}^{-1} \), \( K_b = 2.53 \text{ K kg mol}^{-1} \), and the density of benzene is 876.6 kg m\(^{-3} \). Calculate the freezing point depression, the ratio of the vapor pressure above the solution to that of the pure solvent, the osmotic pressure, and the molecular weight of the solute. \( P^*_{\text{benzene}} = 103 \text{ Torr} \) at 298 K.

\[
\Delta T_b = K_b m_{\text{solute}}; \quad m_{\text{solute}} = \frac{\Delta T_b}{K_b} = \frac{0.625 \text{ K}}{2.53 \text{ K kg mol}^{-1}} = 0.247 \text{ mol kg}^{-1}
\]

\[
M_{\text{solute}} = \frac{5.25 \text{ g}}{0.247 \text{ mol kg}^{-1} \times 0.565 \text{ kg}} = 37.6 \text{ g mol}^{-1}
\]

\[
\Delta T_f = -K_f m_{\text{solute}} = -5.12 \text{ K kg mol}^{-1} \times 0.247 \text{ mol kg}^{-1} = -1.26 \text{ K}
\]

\[
\frac{P_{\text{benzene}}}{P^*_{\text{benzene}}} = x'_{\text{benzene}} = \frac{n_{\text{benzene}}}{n_{\text{benzene}} + n_{\text{solute}}} \quad (\text{using Raoult's law})
\]

\[
= \frac{565 \text{ g}}{78.11 \text{ g mol}^{-1}} = 0.981
\]

\[
P_{\text{benzene}} = x'_{\text{benzene}} P^*_{\text{benzene}} = 0.981 \times 103 \text{ Torr} = 101.0 \text{ Torr}
\]

\[
\pi = \frac{n_{\text{solute}} RT}{V} = \frac{5.25 \times 10^{-3} \text{ kg}}{37.6 \times 10^{-3} \text{ kg mol}^{-1} \times 8.314 \text{ J mol}^{-1} \text{ K}^{-1} \times 298 \text{ K}} \frac{565 \times 10^{-3} \text{ kg}}{876.6 \text{ kg m}^{-3}} = 5.37 \times 10^5 \text{ Pa} = 5.37 \text{ bar}
\]

where \( V \) is just the volume of the benzene.