In the Haber-Bosch process, assuming $\Delta S^o = -389 \text{ J/K}$ and $\Delta H^o = -92 \text{ KJ/mol}$, what would be the temperature at which the reaction starts to become spontaneous? \(2 \text{ marks}\)

What does “spontaneous” mean from a thermodynamics perspective? \(1 \text{ mark}\)

What is the role of a catalyst in any given reaction? If a catalyst is added to a closed system not yet at equilibrium, what will occur? \(2 \text{ marks}\)

\(\Delta (2 \text{ marks})\) At 900 °C, calcium carbonate undergoes thermal decomposition to produce CO\(_2\) according to the below reaction, with $K_{eq} = 0.0108$. What is the concentration of CO\(_2\) at equilibrium?

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

Solids and liquids are not a part of equilibrium.

$$K_{eq} = \frac{[\text{CO}_2]}{[\text{CaO}]}^0$$

If $K_{eq} = 0.0108$,

$$K_{eq} = \frac{[\text{CO}_2]}{[\text{CaO}]}^{eq.}$$

$$0.0108 = [\text{CO}_2]^{eq.}$$

However CO\(_2\) is a gas. So I think it would have a $K_p$, and a partial pressure:
2. Copper reacts with sulfuric acid to produce copper sulfate according to the below reaction.

\[ \text{Cu(s)} + 2\text{H}_2\text{SO}_4(\text{l}) \rightarrow \text{CuSO}_4(\text{aq}) + \text{SO}_2(\text{g}) + 2\text{H}_2\text{O}(\text{l}) \]

a) What is the oxidizing agent? (1 mark)

b) State the oxidation states for each element in the reaction. (2 marks)

3. (3 marks) Hydrogen iodide decomposes according to the following equation:

\[ 2\text{HI}(\text{g}) \rightleftharpoons \text{H}_2(\text{g}) + \text{I}_2(\text{g}) \]

4.0 moles of HI are introduced into a 10.0L container at 500°C. At equilibrium, 70% of the HI remains. What is the value of \( K_{eq} \)?

\[
K_{eq} = \frac{[H_2][I_2]}{[HI]^2}
\]

\[
\begin{align*}
2\text{HI}(\text{g}) & \rightleftharpoons \text{H}_2(\text{g}) + \text{I}_2(\text{g}) \\
\text{I} & 0.40 \text{M} \\
\text{C} & \times \times \times \\
\text{E} & 0.26 \text{M} \\
\end{align*}
\]

\[
M_{\text{HI}}(0.7) = 2.8 \text{ mol Hi} \\
M_{\text{HI}}(0) = 0.26 \text{M} \\
M_{\text{HI}}(0) = 0.26 \text{M} \\
: \ x = 0.06 \text{M} \\

K_{eq} = \frac{[0.06\text{M}][0.06\text{M}]}{(0.26\text{M})^2} = \frac{0.06\times0.06}{(0.26)^2} \\

\text{(Favours reactants)}
\]
Consider the reaction of ethylene (CH$_2$=CH$_2$) with ammonia (NH$_3$) to produce ethylamine (CH$_3$-CH$_2$-NH$_2$). Calculate $\Delta G_{\text{rxn}}$ at 298K and at 1000K. Would the most promising conditions for the reaction be to run it at higher or lower temperatures? Would higher or lower pressures be preferred?

<table>
<thead>
<tr>
<th>Substance</th>
<th>$\Delta G_{\text{f}}^\circ$ Kcal/mole</th>
</tr>
</thead>
<tbody>
<tr>
<td>CH$_2$=CH$_2$</td>
<td>16.28</td>
</tr>
<tr>
<td>CH$_2$=CH$_2$</td>
<td>28.25</td>
</tr>
<tr>
<td>CH$_2$=CH$_2$+NH$_3$</td>
<td>8.91</td>
</tr>
<tr>
<td>CH$_2$=CH$_2$+H$_2$</td>
<td>-40.22</td>
</tr>
<tr>
<td>NH$_3$</td>
<td>-3.86</td>
</tr>
<tr>
<td>H$_2$O</td>
<td>-54.64</td>
</tr>
<tr>
<td>H$_2$</td>
<td>0</td>
</tr>
<tr>
<td>N$_2$</td>
<td>0</td>
</tr>
</tbody>
</table>

Reaction: \[ \text{CH}_2=\text{CH}_2 + \text{NH}_3 \rightarrow \text{CH}_3\text{CH}_2\text{NH}_2 \]

\[ \Delta G_{\text{rxn}} = \sum \Delta G_{\text{f}} - \sum \Delta G_{\text{f}} \]

Case I: 298$^\circ$C

\[ \Delta G_{\text{rxn} I} = (6.91 \text{ Kcal/mole}) - (10.26 \text{ Kcal/mole} + (-3.86 \text{ Kcal/mole})) \]

\[ \Delta G_{\text{rxn} I} = -3.51 \text{ Kcal/mole} \]

Case II: 1000$^\circ$C

\[ \Delta G_{\text{rxn} II} = (60.90 \text{ Kcal/mole}) - (28.25 \text{ Kcal/mole} + 14.85 \text{ Kcal/mole}) \]

\[ \Delta G_{\text{rxn} II} = 17.86 \text{ Kcal/mole} \]

Lower temperatures would be preferred (I $\Rightarrow$ II) as at 298$^\circ$, the reaction is spontaneous and at 1000$^\circ$, the reaction is non-spontaneous. There are 2 moles of gas on the reactants side, and 1 mole of gas on the products side, so higher pressure would push equilibrium more to the products.

**Answer:** Lower (298$^\circ$) Temp., Higher Pressure.
1. b) From a thermodynamic perspective, spontaneous means that no extra energy input is required for a reaction to proceed in a forward direction. The conditions for spontaneous conditions are: $\Delta H < 0$, $\Delta S > 0$, and $\Delta G < 0$.

c) The role of a catalyst is to provide a different mechanism for a reaction, to allow the reaction to proceed more quickly. A catalyst is not consumed in a reaction.

$E_{a1}$ is the reaction with a catalyst, as it lowers the activation energy of a system. When added to a closed system, a catalyst will help reach equilibrium at a faster rate, as catalysts affect chemical kinetics.