1) a) Let's start with the I.C.F. table (Note \( M_{m(COCl_2)} = 98.916 \) g/mol):

\[
\begin{array}{c|cc|c|c|c}
 & \text{COCl}_2 & \text{CO} & \text{Cl}_2 \\
\hline
\text{I} & 6.72 \times 10^{-2} M & 0 M & 0 M \\
\text{C} & -x & +x & +x \\
\text{F} & 0.0672 - x & x & x \\
\end{array}
\]

\[K_c = 4.6 \times 10^{-3} = \frac{[CO][Cl_2]}{[COCl_2]} = \frac{(x)(x)}{(0.0672 - x)} \]

which leads to: \( x^2 + 4.6 \times 10^{-3} x - 3.0912 \times 10^{-4} = 0 \). Solving for \( x \), we get: \( x = 1.54 \times 10^{-2} M \). Therefore, \([CO] = [Cl_2] = 1.54 \times 10^{-2} M; [COCl_2] = 5.18 \times 10^{-2} M.\)

b) The fraction COCl\(_2\) decomposed is, simply, \( \frac{0.0154}{0.0672} = 0.2292. \)

c) Adding more CO (3.0 g equals 0.107 moles) would have the qualitative effect of shifting the reaction back towards more phosgene, COCl\(_2\).

2) a) Starting with the both the hint (assume all molecules have dimerized and then dissociate) and the I.C.F. method that I proselytize, we get:

\[
\begin{array}{c|cc}
 & \text{(CH}_3\text{CO}_2\text{H})_2 & 2 \text{CH}_3\text{CO}_2\text{H} \\
\hline
\text{I} & 3.75 \times 10^{-3} \text{ atm} & 0 \text{ atm} \\
\text{C} & -x & +2x \\
\text{F} & (3.75 \times 10^{-3} - x) \text{ atm} & 2x \text{ atm} \\
\end{array}
\]

\[K_p = \frac{P(\text{(CH}_3\text{CO}_2\text{H})_2)^2}{P(\text{CH}_3\text{CO}_2\text{H})_2} = \frac{(2x)^2}{(3.75 \times 10^{-3} - x)} = 7.692 \times 10^{-4} \]

(Remember when you reverse an equilibrium, you take the inverse of the equilibrium constant).

Solving the quadratic which results (\( 4x^2 + 7.692 \times 10^{-4}x - 2.885 \times 10^{-6} = 0 \)), we get, as the only sensible answer, \( x = 7.58 \times 10^{-4} \text{ atm.} \) and the resultant pressures are: monomer (\( = \text{CH}_3\text{CO}_2\text{H} \)) = \( 1.52 \times 10^{-3} \text{ atm} \) and dimer (\( = (\text{CH}_3\text{CO}_2\text{H})_2 \)) = \( 2.99 \times 10^{-3} \text{ atm.} \)

(Via an error of omission, the next two parts did not get assigned, nor were you graded on them.)
b) Hydrogen-bonding between the acidic groups, to provide some stabilization in the gas phase. The structure would:

\[
\begin{align*}
\text{H}_3\text{C} & \quad \text{O} \quad \text{H} \quad \text{O} \\
\text{O} & \quad \text{H} \quad \text{O} \quad \text{C} \quad \text{CH}_3
\end{align*}
\]

c) As the temperature decreases, less energy is being input into the system. Since bond breaking is an endothermic process, removing thermal energy from the system would not allow for a great deal of bond breaking. Since the hydrogen bonds can not (or will not) easily break, the amount of dimer should **increase**. Another way to look at it is that forming the hydrogen bonds results in the release of some energy; from a LeChatlier principle argument, this is removing some product (heat) and therefore would shift the equilibrium to the **product side** (dimer) to ameliorate the affect of this change.

3) In order to solve this equilibrium problem, need to determine how much HCl was produced; this we can get from the AgCl precipitated. 3.59 g AgCl (\(M_m = 143.353 \text{ g/mol}\)) gives 0.025 moles AgCl = moles Cl\(^-\) = moles HCl formed so we have **0.025 moles HCl\(_{(g)}\)**.

Now we can use I.C.F

\[
\begin{align*}
\text{LaCl}_3\(_{(s)}\) & + \text{H}_2\text{O}\(_{(g)}\) & \leftrightarrow & \text{LaOCl}\(_{(s)}\) & + 2\text{HCl}\(_{(g)}\) \\
\text{I} & 0.01 \text{ M} & 0.02 \text{ M} & 0 \text{ M} & 0 \text{ M} \\
\text{C} & ---- & -0.0093 \text{ M} & ----- & +0.0185 \\
\text{F} & ---- & \text{0.0107 M} & ----- & \text{0.0185 M}
\end{align*}
\]

(Remember that pure solids do not enter the equilibrium constant expression, so we needn’t worry about their concentrations.

\[
\text{K}_c = \frac{[\text{HCl}]^2}{[\text{H}_2\text{O}]} = \frac{(0.0185)^2}{(0.0107)} = 0.0320 \text{ M}.
\]

4) There are two ways to approach this problem: either using concentrations and moles, and finally substituting into \(PV = nRT\) or using partial pressures and the law of partial pressures. I’ll do both; concentrations first, then pressures.

For both parts, let us remember that \(V = 3.50 \text{ L} \text{ and } K_c = 2.50 \times 10^{-2}\).

\[
\begin{align*}
\text{SbCl}_5\(_{(g)}\) & \leftrightarrow \text{SbCl}_3\(_{(g)}\) + \text{Cl}_2\(_{(g)}\) \\
\text{I} & 4.29 \times 10^{-3} \text{ M} & 0 & 0 \\
\text{C} & -x & +x & +x \\
\text{F} & (4.29 \times 10^{-3} - x) \text{ M} & x \text{ M} & x \text{ M}
\end{align*}
\]

We know that \(K_c = 2.50 \times 10^{-2}\) so we can set up the equilibrium expression:
\[ K_c = \frac{[SbCl_3][Cl_2]}{[SbCl_5]} = \frac{x^2}{4.29 \times 10^{-3} - x} = 2.50 \times 10^{-2} \] which solves out to: \( x = 3.73 \times 10^{-3} M \), which leads to \([SbCl_5] = 5.61 \times 10^{-4} M\), and \([SbCl_3] = [Cl_2] = 3.73 \times 10^{-3} M\), and the number of moles of \( SbCl_5 \) = 1.96 \( \times 10^{-3} \) and the number of moles of \( SbCl_3 = \text{number of moles} = Cl_2 \) = 1.31 \( \times 10^{-2} \), which leads to (finally), total number of moles = 0.02807. Substituting into PV = nRT, we get: \( P = (0.02807)(0.08206)(521)/(3.5) = 0.343 \text{ atm} \).

The \( K_p \) we need is merely \( K_c(RT) \) or 0.0250(0.08206)(521) = 1.07 atm, and the partial pressure of \( SbCl_5 \) is: \( P_5 = 0.183 \text{ atm} \). (I will use \( P_5 \) for pressure of \( SbCl_5 \), \( P_3 \) for pressure of \( SbCl_3 \) and \( P_{Cl} \) for pressure of \( Cl_2 \))

\[
\begin{array}{ccc}
\text{SbCl}_5(\text{g}) & \leftrightarrow & \text{SbCl}_3(\text{g}) + \text{Cl}_2(\text{g}) \\
\text{I} & 0.183 & 0 \\
\text{C} & -x & +x +x \\
\text{F} & (0.183 - x) \text{ atm} & x \text{ atm} & x \text{ atm} \\
\end{array}
\]

Making a similar substitution into the \( K_p \) expression, we get:

\[ K_p = \frac{P_3P_{Cl}}{P_5} = \frac{x^2}{0.183 - x} = 1.07 \text{, which solves out to be:} \] \( P_3 = P_{Cl} = 0.159 \text{ atm and } P_5 = 0.024 \text{ atm} \) which leads to a total pressure, \( P_{\text{tot}} = 0.342 \text{ atm} \). (Not bad).