1) We have a classic case of the initial rate method to be applied here. Let us start there.
   
a) Going about it the way we did in class, we can look at the ratio of rate$_2$ and rate$_1$ and from that, get one of the reaction orders. Note that rate = k[(CH$_3$)$_3$CCl]$^x$[OH$^-$$]^y$
   \[
   \frac{3.98 \times 10^8}{2.0 \times 10^8} = \frac{k(0.050)^x(0.025)^y}{k(0.025)^x(0.025)^y} = 1.99 = 2^x, \text{ and } x = \text{ order for } [(CH_3)_3CCl] = 1.
   \]
   
   Similarly, we can use rate$_3$ and rate$_1$ and get the other reaction order
   \[
   \frac{2.01 \times 10^8}{2.0 \times 10^8} = \frac{k(0.025)^x(0.051)^y}{k(0.025)^x(0.025)^y} = 1.005 = 2.04^y \text{ and } y = \text{ order for } [OH^-] = 0.
   \]
   
The overall reaction order, therefore, is $1 + 0=1$.

b) Recall we’ve just determined the rate law to be: rate = k[(CH$_3$)$_3$CCl]. Picking any one set of reaction conditions, we can plug in and get the rate constant. Taking experiment 2: \(3.98 \times 10^8 \text{ M/s} = k(0.050 \text{ M}),\) which leads to \(k = 7.96 \times 10^9 \text{ s}^{-1}\)

c) The rate law this mechanism predicts is:
   rate = k$_1$[(CH$_3$)$_3$Cl] (remember, the rate of slower/slowest step controls the rate of the reaction), which is in agreement with the experimental rate law.

2) If the time to rottling at 273 K is four times what it is at 293 K, then the rate constant at 273 K is $1/4^{th}$ of what it is at 293 K. Therefore, we can say that $k_2 = k_1/4$ and \(\ln\left(\frac{k_1}{4} / k_1\right) = \ln\left(\frac{1}{4}\right) = \frac{E_a}{(8.314J/K\cdot mol)} \left[\frac{1}{293} - \frac{1}{273}\right]\) which leads to \(E_a = 46.1 \text{ kJ/mol}\)

3) Let’s take the hint given, and write the balanced equation.
   \[2N_2O_5 \rightarrow 4NO_2 + O_2\] (or similar)
   With the rate constant given (units of s$^{-1}$), we can safely assumethe reaction is first order, with a rate law of: rate = $2.8 \times 10^{-3} \text{ s}^{-1}[N_2O_5]$. (As will be shown below we don’t need this rate law; it is sort of a red-herring.)
   
   The total pressure, $P_{tot}$ in the vessel is simply: $P_{tot} = P_{N_2O_5} + P_{NO_2} + P_{O_2}$. Our stoichiometry tells us that the pressure of NO$_2$ is twice that of N$_2$O$_5$, and four times that of O$_2$. Letting $x = P_{O_2}$, we can say:
   
   \[178 \text{ kPa} = (125 \text{ kPa} – 2x) + 4x + x,\] and solve for $x$. Doing so, we get $x = 17.67$ kPa, which gives us $P_{NO_2} = 4x = 71$ kPa.

4) Let’s do this by process of elimination. We are given an experimental rate law of:
rate = k[NO]^2[H_2]

with which the mechanistically derived rate law needs to agree. The first mechanism
gives us:
\[
rate = k[H_2]^2[NO]^2,
\]

which does not agree with experiment.
The second mechanism gives us the rate law:
\[
Rate = k[H_2][NO],
\]

which also does not agree with experiment.
The third mechanism gives us:
\[
Rate = k[NO]^2[H_2],
\]

which does agree with experiment, so Mechanism III is consistent because it predicts a rate law in agreement with experiment.

5a) Since [E] is increasing with time, E would have to be a product. (Reactants tend to disappear with time.)

b) The average rate of the reaction is faster between points A and B than between points B and C. The why for this is quite simple: the slope of the line between A and B is steeper than between B and C (and this slope represents the average rate of the reaction.)