(1) A) $C_5H_{12} + 8O_2 \rightarrow 5CO_2 + 6H_2O$

B) $2.5g\ C_5H_{12} \cdot \frac{1\,mol}{72.096} \cdot \frac{5\ CO_2}{1\,mol\ C_5H_{12}} = 0.1734\,mol\ CO_2$

$\left(\frac{785}{760}\right) \cdot V = 0.1734 \cdot 0.0821 \cdot 298$ $\boxed{V = 4.11L}$

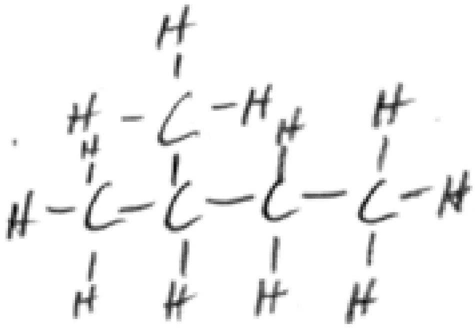
C) $5.00g\ C_5H_{12} \cdot \frac{1\,mol\ C_5H_{12}}{72.096g} = 6.935 \cdot 10^{-2}\,mol\ C_5H_{12}$

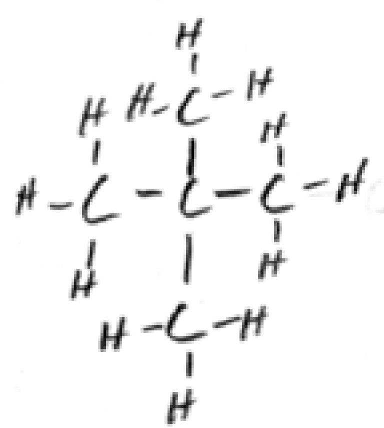
$\frac{243\,kJ}{6.935 \cdot 10^{-2}\,mol} = \boxed{3.50 \cdot 10^{3}\ \frac{kJ}{mol}}$

D) $\frac{R_{unk}}{R_p} = \frac{\sqrt{M_p}}{\sqrt{M_{unk}}}$ $R_{unk} = 2R_p$ $2 = \frac{\sqrt{72.096}}{\sqrt{M_{unk}}}$

$M_{unk} = \frac{72.096}{4} = \boxed{18.0\frac{g}{mol}}$ $4 = \frac{72.096}{M_{unk}}$

E)





(2) A) $\Delta H = [2\Delta H^\circ_{f\,HCl}] - [\Delta H^\circ_{f\,OCl_2} + \Delta H^\circ_{f\,H_2O}]$

$-46\text{ kJ} = [2 \cdot -92.3] - [\Delta H^\circ_f OCl_2 + -242]$

$-46\text{ kJ} = [-184.6] - \Delta H^\circ_{f\,OCl_2} + 242$

$-46\text{ kJ} = 57.4 - \Delta H^\circ_{f\,OCl_2}$

$\Delta H^\circ_{f\,OCl_2} = 1.03 \cdot 10^2 \frac{kJ}{mol}$

B) $\frac{1}{2}O_2 + Cl_2 \rightarrow OCl_2$

C) $\Delta H^\circ_{f\,OCl_2} = [\frac{1}{2}(O{=}O) + (Cl{-}Cl)] - [2(O{-}Cl)]$

$1.03 \cdot 10^2 = [\frac{1}{2}(498) + (243)] - 2(O{-}Cl)$

$1.03 \cdot 10^2 = 492 - 2(O{-}Cl)$

$-389 = -2(O{-}Cl)$

$(O{-}Cl) = 195 \frac{kJ}{mol}$

D) Oxygen has a double bond while Cl_2 has only a single bond. Double bonds are stronger and therefore harder to break because there are more e^- being shared between the atoms.

Question 7

Answer the following questions about thermodynamics.

Substance	Combustion Reaction	Enthalpy of Combustion, ΔH°_{comb}, at 298 K (kJ mol^{-1})
$H_2(g)$	$H_2(g) + \frac{1}{2}O_2(g) \rightarrow H_2O(l)$	–290
$C(s)$	$C(s) + O_2(g) \rightarrow CO_2(g)$	–390
$CH_3OH(l)$		–730

(a) In the empty box in the table above, write a balanced chemical equation for the complete combustion of <u>one</u> mole of $CH_3OH(l)$. Assume products are in their standard states at 298 K. Coefficients do not need to be in whole numbers.

$CH_3OH(l) + \frac{3}{2}O_2(g) \rightarrow CO_2(g) + 2\,H_2O(l)$	One point is earned for the correct products. One point is earned for balancing the equation.

(b) On the basis of your answer to part (a) and the information in the table, determine the enthalpy change for the reaction $C(s) + H_2(g) + H_2O(l) \rightarrow CH_3OH(l)$.

Adding the following three equations,

$C(s) + O_2(g) \rightarrow CO_2(g)$ –390 kJ mol^{-1}

$H_2(g) + \frac{1}{2}O_2(g) \rightarrow H_2O(l)$ –290 kJ mol^{-1}

$CO_2(g) + 2\,H_2O(l) \rightarrow CH_3OH(l) + \frac{3}{2}O_2(g)$ +730 kJ mol^{-1}

yields this equation: $C(s) + H_2(g) + H_2O(l) \rightarrow CH_3OH(l)$ +50 kJ mol^{-1}

One point is earned for the correct equations.

One point is earned for the correct value of ΔH°.

(c) Write the balanced chemical equation that shows the reaction that is used to determine the enthalpy of formation for one mole of $CH_3OH(l)$.

$C(s) + 2\,H_2(g) + \frac{1}{2}O_2(g) \rightarrow CH_3OH(l)$	One point is earned for the correct equation.

Question 7 (continued)

(d) Predict the sign of $\Delta S°$ for the combustion of $H_2(g)$. Explain your reasoning.

$\Delta S°$ for the combustion of $H_2(g)$ is negative. Both reactants are in the gas phase and the product is in the liquid phase. The liquid phase is much more ordered than the gas phase, so the product is more ordered compared to the reactants, meaning that $\Delta S°$ is negative. (Note: There are fewer moles of products than reactants, which also favors a more ordered condition in the products, but the difference in phases is the more important factor.)	One point is earned for the correct sign of $\Delta S°$. One point is earned for a correct explanation.

(e) On the basis of bond energies, explain why the combustion of $H_2(g)$ is exothermic.

The combustion of $H_2(g)$ is exothermic ($\Delta H° < 0$) because more energy is released during the formation of two moles of O–H bonds than is required to break one mole of H–H bonds and one-half of a mole of O–O bonds.	One point is earned for the correct explanation.