2. (a) $? \text{ moles } N_2 = 24.5 \text{ g } N_2 \left(\frac{1 \text{ mole } N_2}{28.0 \text{ g } N_2}\right) = 0.875 \text{ moles } N_2$

$PV = nRT$
$P(5.00 \text{ L}) = (0.875 \text{ moles})(0.0821 \frac{\text{L}\cdot\text{atm}}{\text{K}\cdot\text{mol}})(298 \text{ K})$
$P_{N_2} = 4.28 \text{ atm}$
$? \text{ moles } O_2 = 28.0 \text{ g } O_2 \left(\frac{1 \text{ mole } O_2}{32.0 \text{ g } O_2}\right) = 0.875 \text{ moles } O_2$

$PV = nRT$
$P(5.00 \text{ L}) = (0.875 \text{ moles})(0.0821 \frac{\text{L}\cdot\text{atm}}{\text{K}\cdot\text{mol}})(298 \text{ K})$
$P_{O_2} = 4.28 \text{ atm}$
$P_{Total} = P_{N_2} + P_{O_2} = 4.28 \text{ atm} + 4.28 \text{ atm}$

$\boxed{P_{Total} = 9.56 \text{ atm}}$

(b) (i) $X_{N_2} = \frac{\#\text{ moles } N_2}{\text{Total }\#\text{ of moles}} = \frac{0.875 \text{ moles } N_2}{0.875 \text{ moles } N_2 + 0.875 \text{ moles } O_2} = \boxed{0.5 = X_{N_2}}$

(ii) $PV = nRT$
$P(5.00 \text{ L}) = (0.875 \text{ moles})(0.0821 \frac{\text{L}\cdot\text{atm}}{\text{K}\cdot\text{mol}})(280 \text{ K})$

$\boxed{P_{N_2} = 4.02 \text{ atm}}$

(c) Both gases have the same average kinetic energy since they are at the same temperature. However, since the nitrogen molecule has less mass than the oxygen molecule, the nitrogen molecule will be moving faster at this temperature and, thus, diffuse faster out of the cylinder. Therefore, the ratio $\frac{\text{moles of } N_2(g)}{\text{moles of } O_2(g)}$ would decrease.

(d) $2 \text{ NO}(g) + O_2(g) \rightarrow 2 \text{ NO}_2(g)$

(e) According to the balanced equation the mole ratio in the reaction 2:1:2 [$NO(g)$:$O_2(g)$:$NO_2(g)$]. This means that all of the 0.176 moles of $NO(g)$ will react with half of the $O_2(g)$ [leaving 0.088 moles $O_2(g)$] to form 0.176 moles of $NO_2(g)$. Therefore after the reaction is complete there will be no $NO(g)$, 0.088 moles of $O_2(g)$, and 0.176 moles of $NO_2(g)$. So, the total number of moles in the cylinder will be 0.176 moles + 0.088 moles = 0.264 moles = n_{total}. According to the ideal gas law all molecules are equivalent so n_{total} can be used in the ideal gas equation to calculate total pressure.

$PV = nRT$
$P(5.00 \text{ L}) = (0.264 \text{ moles})(0.0821 \frac{\text{L}\cdot\text{atm}}{\text{K}\cdot\text{mol}})(298 \text{ K})$

$\boxed{P_{Total} = 1.29 \text{ atm}}$

3. (a) $C_5H_{12} + 8O_2 \rightarrow 5CO_2 + 6H_2O$

(b) $? \text{ moles } CO_2 = 2.50 \text{ g } C_5H_{12}\left(\frac{1 \text{ mole } C_5H_{12}}{72.15 \text{ g } C_5H_{12}}\right)\left(\frac{5 \text{ moles } CO_2}{1 \text{ mole } C_5H_{12}}\right) = 0.173 \text{ moles } CO_2 = n$

$PV = nRT$

$$\left(\frac{785 \text{ mm Hg}}{760 \ \text{mm Hg}/\text{atm}}\right)V = (0.173 \text{ moles})\left(0.0821 \ \frac{L \cdot atm}{K \cdot mol}\right)(298 \text{ K})$$

$V = 4.10$ L of CO_2

(c) $? \text{ kJ} = 1 \text{ mole } C_5H_{12}\left(\frac{72.15 \text{ g } C_5H_{12}}{1 \text{ mole } C_5H_{12}}\right)\left(\frac{243 \text{ kJ}}{5.00 \text{ g } C_5H_{12}}\right)$

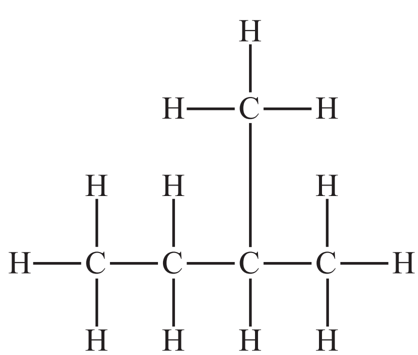
$\Delta H = 3.50 \times 10^3$ kJ

(d) $$\frac{\text{Rate}_{\text{unkwnown}}}{\text{Rate}_{\text{pentane}}} = \sqrt{\frac{\text{Molar Mass}_{\text{pentane}}}{\text{Molar Mass}_{\text{unknown}}}}$$

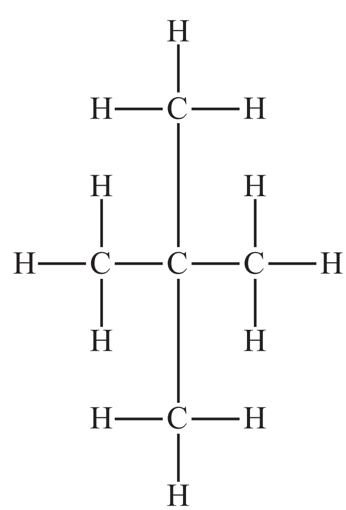
$$2 = \sqrt{\frac{72.15 \text{ g pentane}}{\text{Molar Mass}_{\text{unknown}}}}$$

$\text{Molar Mass}_{\text{unknown}} = 18.04$ g

(e)



isopentane, or
2-methyl butane



neopentane, or
2,2-dimethyl propane