10.30 (c) Total pressure of gases after reaction at 125°C.

$$n_{total} = n_{O_2} + n_{CH_4} + n_{CO_2} + n_{H_2O} = 0\ \text{mole} + 2.47\ \text{mole} + 7.50\ \text{mole} + 15.0\ \text{mole}$$
$$= 24.97\ \text{mole}$$

Given a final volume of 1.00 L and temp of $273 + 125 = 398\text{K}$

$$P = \frac{n_{tot} RT}{V} = \frac{(24.97\ \text{mole})(0.08206\ \text{atm L mol}^{-1}\text{K}^{-1})(398\text{K})}{1.00\text{L}}$$

$$= 816\ \text{atm}$$

10.31 convert to moles!

$$n_{total} = n_{CO_2} + n_{H_2O} + n_{CH_4} = 2.00\text{g}\ CO_2 \times \frac{1\ \text{mole}\ CO_2}{44.010\text{g}\ CO_2} + 2.00\text{g}\ H_2O \times \frac{1\ \text{mole}\ H_2O}{18.015\text{g}}$$

$$+ 2.00\text{g}\ CH_4 \times \frac{1\ \text{mol}\ CH_4}{16.043\text{g}\ CH_4} = 0.28113\ \text{mole}$$

then use ideal gas equation to establish temperature:

$$T = \frac{PV}{n_{tot} R} = \frac{(10.0\ \text{atm})(1.00\text{L})}{(0.28113\ \text{mol})(0.08206\ \text{atm L mol}^{-1}\text{K}^{-1})} = 433\text{K}$$

10.35 convert to moles! (use $PV = nRT$) $\quad \frac{1}{n} = \frac{RT}{PV}$

$$\frac{1}{n} = \frac{(0.08206\ \text{atm L mol}^{-1}\text{K}^{-1})(273\text{K})}{(1.00\ \text{atm})(1.50\ \text{L})} = 0.15\ \text{moles}^{-1}$$

$$m.w. = 2.14\text{g} \times 0.15\ \text{mole}^{-1} = 32\ \text{g/mole}$$ so the diatomic gas is O_2

10.36 find the molecular weight as in #35:

$$\frac{1}{n} = \frac{(0.08206\ \text{atm L mol}^{-1}\text{K}^{-1})(273\text{K})}{(1.00\ \text{atm})(17.7\ \text{L})} = 1.27\ \text{moles}^{-1}$$

$$m.w. = 34.84\ \text{g} \times 1.27\ \text{mole}^{-1} = 44.4\ \text{g/mole}$$