10.26 use $PV = nRT$

(a) pressure $= 589.1 \text{ torr} \times \frac{1 \text{ atm}}{760 \text{ torr}} = 0.7751 \text{ atm}$

$T = (312.4 + 273.15)K = 585.6 K$

so $V = \frac{nRT}{P} = \frac{(3.28 \text{ mole})(0.08206 \text{ atm L mol}^{-1} K^{-1})(585.6 K)}{0.7751 \text{ atm}} = 203 L$

(b) $\cancel{V = \frac{nR\frac{1}{2}T}{\frac{1}{2}P} = \frac{nRT}{P} = 203L}$ $P = \frac{nRT}{V} = \frac{nR\frac{1}{2}T}{\frac{1}{2}V} = \frac{nRT}{V} = 0.7751 \text{ atm}$

the pressure remains the same because the "$\frac{1}{2}$"s cancel out.

10.30 convert to moles!

$160. \text{ g } CH_4 \times \frac{1 \text{ mole } CH_4}{16.043 \text{ g } CH_4} = 9.97 \text{ mole } CH_4$

$480. \text{ g } O_2 \times \frac{1 \text{ mole } O_2}{31.9988 \text{ g}} = 15.0 \text{ mole } O_2$

(a) moles of CH_4 & O_2 remaining:

oxygen is the limiting reagent & CH_4 is in excess.

After combustion, no O_2 is left.

$15.0 \text{ mole } O_2 \times \frac{1 \text{ mole } CH_4}{2 \text{ mole } O_2} = 7.50 \text{ mole } CH_4$ react w/ O_2

so $9.97 - 7.50 = 2.47$ mole CH_4 remains.

(b) moles CO_2 & H_2O produced:

CO_2: $15.0 \text{ mole } O_2 \times \frac{1 \text{ mole } CO_2}{2 \text{ mole } O_2} = 7.50 \text{ mole } CO_2$

H_2O: $15.0 \text{ mole } O_2 \times \frac{2 \text{ mole } H_2O}{2 \text{ mole } O_2} = 15.0 \text{ mole } H_2O$