Question 3

3. Answer the following questions that relate to the analysis of chemical compounds.

(a) A compound containing the elements C, H, N, and O is analyzed. When a 1.2359 g sample is burned in excess oxygen, 2.241 g of $CO_2(g)$ is formed. The combustion analysis also showed that the sample contained 0.0648 g of H.

(i) Determine the mass, in grams, of C in the 1.2359 g sample of the compound.

$2.241\text{ g }CO_2(g) \times \frac{1\text{ mol }CO_2}{44.01\text{ g }CO_2} \times \frac{1\text{ mol C}}{1\text{ mol }CO_2} \times \frac{12.011\text{ g C}}{1\text{ mol C}}$ $= 0.6116\text{ g C}$	One point is earned for the correct answer.

(ii) When the compound is analyzed for N content only, the mass percent of N is found to be 28.84 percent. Determine the mass, in grams, of N in the original 1.2359 g sample of the compound.

$1.2359\text{ g sample} \times 0.2884 = 0.3564\text{ g N}$	One point is earned for the correct answer.

(iii) Determine the mass, in grams, of O in the original 1.2359 g sample of the compound.

Because the compound contains only C, H, N, and O, mass of O = g sample − (g H + g C + g N) $= 1.2359 - (0.0648 + 0.6116 + 0.3564) = 0.2031\text{ g}$	One point is earned for the answer consistent with the answers in parts (a)(i) and (a)(ii).

(iv) Determine the empirical formula of the compound.

Converting all masses to moles, $0.6116\text{ g C} \times \frac{1\text{ mol C}}{12.011\text{ g C}} = 0.05092\text{ mol C}$ $0.0648\text{ g H} \times \frac{1\text{ mol H}}{1.0079\text{ g H}} = 0.06429\text{ mol H}$ $0.3564\text{ g N} \times \frac{1\text{ mol N}}{14.007\text{ g N}} = 0.02544\text{ mol N}$ $0.2031\text{ g O} \times \frac{1\text{ mol O}}{16.00\text{ g O}} = 0.01269\text{ mol O}$	One point is earned for all masses converted to moles. Note: Moles of C may be shown in part (a)(i).

Divide all mole quantities by the smallest number of moles: $0.05092 \text{ mol} \div 0.01269 \text{ mol} = 4.013$ $0.06429 \text{ mol} \div 0.01269 \text{ mol} = 5.066$ $0.02544 \text{ mol} \div 0.01269 \text{ mol} = 2.005$ $0.01269 \text{ mol} \div 0.01269 \text{ mol} = 1.000$ $\Rightarrow$ Empirical formula is $C_4H_5N_2O$	One point is earned for dividing by the smallest number of moles. One point is earned for the empirical formula consistent with the ratio of moles calculated.

(b) A different compound, which has the empirical formula CH_2Br, has a vapor density of 6.00 g L^{-1} at 375 K and 0.983 atm. Using these data, determine the following.

(i) The molar mass of the compound

$PV = nRT \Rightarrow \frac{PV}{RT} = n$ $\frac{(0.983 \text{ atm})(1.00 \text{ L})}{(0.0821 \text{ L atm mol}^{-1}\text{K}^{-1})(375 \text{ K})} = 0.0319 \text{ mol}$ molar mass of gas (M) $= \frac{6.00 \text{ g}}{0.0319 \text{ mol}} = 188 \text{ g mol}^{-1}$ *OR* $M = \frac{DRT}{P} = \frac{6.00 \text{ g L}^{-1} \times 0.0821 \text{ L atm mol}^{-1}\text{ K}^{-1} \times 375 \text{ K}}{0.983 \text{ atm}}$ $= 188 \text{ g mol}^{-1}$	One point is earned for applying the gas law to calculate n. One point is earned for calculating the molar mass. *OR* Two points are earned for calculating the molar mass using $M = \frac{DRT}{P}$

(ii) The molecular formula of the compound

Each CH_2Br unit has mass of $12.011 + 2(1.0079) + 79.90 = 93.9$ g, and $\frac{188 \text{ g}}{93.9 \text{ g}} = 2.00$, so there must be two CH_2Br units per molecule. Therefore, the molecular formula of the compound is $C_2H_4Br_2$.	One point is earned for the molecular formula that is consistent with the molar mass calculated in part (b)(i).