

CHEM131 HOMEWORK #3 KEY

3-38. Caffeine, (Check various graphics in your Lab Manual!)  $C_8H_{10}N_4O_2$ : Molar Mass =  $8 \times 12.01 + 10 \times 1.008 + 4 \times 14.01 + 2 \times 16.00 = 194.2$  g/mole. Mass % C =  $8 \times 12.01 \times 100 / 194.2 = 49.47\%$  C

Sucrose,  $C_{12}H_{22}O_{11}$ : Molar Mass =  $12 \times 12.01 + 22 \times 1.008 + 11 \times 16.00 = 342.3$  g/mole. Mass % C =  $12 \times 12.01 \times 100 / 342.3 = 42.10\%$  C

Ethanol,  $C_2H_6O$ : Molar Mass =  $2 \times 12.01 + 6 \times 1.008 + 16.00 = 46.07$  g/mole. Mass % C =  $24.02 / 46.07 = 52.14\%$  C (Note that in each case, the units of g/mole in the numerator and the denominator cancel.)

3-40. If you begin with 100 g of cyanocobalamin (abbreviated B12 here) your sample contains 4.34 g Co.  
 $4.34 \text{ g Co} / 58.93 \text{ g/mole Co} = 0.07365 \text{ mole Co} = 0.07365 \text{ mole B12}$  (from assumption of 1 Co/B12 molecule).  $100 \text{ g B12} / 0.07365 \text{ mole B12} = 1360 \text{ g/mole B12}$ .

3-46. Beginning with 100.0 g adrenalin, you have 56.79 g C, 6.56 g H, 28.37 g O, and 8.28 g N. Finding the # of moles of each, we have  $56.79 / 12.01 = 4.728$  mol C,  $6.56 / 1.008 = 6.51$  mol H,  $28.37 / 16.00 = 1.773$  mol O,  $8.28 / 14.01 = 0.591$  mol N. Divide the # of moles of each atom by the smallest, 0.591, gives  $4.728 / 0.591 = 8.00$  C:N,  $6.51 / 0.591 = 11.0$  H:N,  $1.773 / 0.591 = 3.00$  O:N and the empirical formula is  $C_8H_{11}O_3N$ . The next step would be to find the complete formula, which I did by looking up adrenalin in a text, only to find that the formula for adrenalin is  $C_9H_{13}O_3N$ . Here is a case of Zumdahl cooking the books. In a real determination of an empirical formula you would expect measuring uncertainties to give you ratios that were close to integers, but not exactly integer values. And he started with the wrong formula to boot!

3-47. This is an involved problem that is best solved by planning your strategy first. From earlier examples, such as problem 3-40, we know how to get the empirical formula if we know the percentages of C, H, O, and N. From the combustion experiment we can find out how much C and how much H was contained in 0.157 g of X. From the ammonia experiment we can find out how much N was contained in 0.103 g of X. (It would have been convenient if both experiments had started with the same mass, but we can adjust for that.)

$$(213 \text{ mg CO}_2)(12.01 \text{ g C} / 44.01 \text{ g CO}_2) = 58.1 \text{ mg C}; (58.1 \text{ mg} / 157 \text{ mg})100 = 37.0\% \text{ C}$$

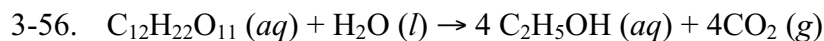
$$(31.0 \text{ mg H}_2\text{O})(2 \times 1.008 \text{ g H} / 18.02 \text{ g H}_2\text{O}) = 3.47 \text{ mg H}; (3.47 \text{ mg} / 157 \text{ mg})100 = 2.21\% \text{ H}$$

$$(23.0 \text{ mg NH}_3)(14.01 \text{ g N} / 17.03 \text{ g NH}_3) = 18.9 \text{ mg N}; (18.9 \text{ mg} / 103 \text{ mg})100 = 18.3\% \text{ N}$$

$$100\% = 37.0\% + 2.2\% + 18.3\% + \% \text{ oxygen}; \% \text{ O} = 42.5\%$$

Now consider the canonical 100 g sample (enough of these meager milligrams) and find the number of moles of each atom in 100 g. Units will be the same as in problem 3-40.  $37.0/12.01 = 3.08$  mol C;  $2.21/1.008 = 2.19$  mol H;  $18.3/14.01 = 1.31$  mol N;  $42.5/16.00 = 2.66$  mol O. Dividing each by the smallest gives C:H:N:O, but not in integer ratios yet.

$3.08/1.31 = 2.35$ ,  $2.19/1.31 = 1.67$ ,  $1.31/1.31 = 1.00$ ,  $2.66/1.31 = 2.03$ . Multiplying 2.35:1.67:1.00:2.03 by 3 gives 7.05:5.01:3:6.09 and the empirical formula is  $C_7H_5N_3O_6$ .



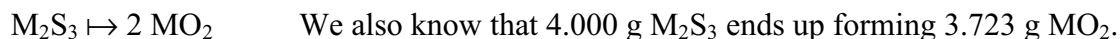
3-57.  $1000 \text{ g Al}/26.98 \text{ g Al/mol} = 37.06 \text{ mol Al}$ . Since 3 mol of ammonium perchlorate react with 3 mol aluminum, we need 37.06 mol  $NH_4ClO_4$ , which has a molecular mass of 117.49 g/mol.  $37.06 \times 117.49 = 4355 \text{ g } NH_4ClO_4 = 4.355 \text{ kg}$ .

3-64. Starting with a balanced equation is ALWAYS a good idea:  $2Al + 3Br_2 \rightarrow 2AlBr_3$ . 6.0 g Al is 0.222 mol Al and for 100% yield you would get 0.222 mole  $AlBr_3$ , weighing  $0.222 \times 266.68 = 59 \text{ g } AlBr_3$ .  $50.3 \times 100/59 = 85\%$  actual yield.

3-89.  $X_2Z$ : 40% X and 60% Z by mass; and from the chemical formula we know molX/molZ =  $2 = (40.0/A_x)/(60.0 A_z)$  where A is the molar mass of X or Z, respectively. Solving  $2 = (40.0 A_z)/(60.0 A_x)$  we get  $A_z = 3 A_x$ .

For  $XZ_2$ , its molar mass is  $A_x + 2 A_z = A_x + 2 (3 A_x) = 7 A_x$ . Thus,  $\%X = A_x/7A_x \times 100 = 1/7 \times 100 = 14.3 \%$ .  $\%Z = 100.0 \% - 14.3 \% = 85.7 \% Z$ .

3-99. Always start with the balanced chemical equation...in this case the best we can do is to balance element "M"



So – from the chemical equation we know that 2 mol  $MO_2 = 1 \text{ mol } M_2S_3$ .

Let A = molar mass of M –

this means  $2(4.000 \text{ g } M_2S_3)/(2A+3(32.07 \text{ g/mol S})) = 1(3.723 \text{ g } MO_2)/(A+2(16.00 \text{ g/mol O}))$   
or  $8.000 A + 256.0 = 7.446 A + 358.2$ ,  $0.554 A = 102.2$ , or  $A = 184 \text{ g/mol}$ , atomic mass = 184 amu.

H.W.4 Ch.4: 13, 16, 19, 25, 29, 30, 35, 86\*