

1) The empirical formula of a compound represents the lowest whole number mole ratio of component elements.

2) The molecular formula is the actual ratio of elements found with the naturally-occurring sample. This could also equal the empirical formula if the naturally-occurring formula is also in the lowest whole number mole ratio.

3) Molecular = $C_4H_8N_2O_4$ ← each subscript can be reduced by 2

Empirical = $C_2H_4N_2O_2$

4) Compound pair B is the only set of compounds with the same empirical formula as

5) To determine the empirical formula, convert each atom's provided mass into moles and find the lowest whole number mole ratio.

$$0.8007 \text{ g C} \times 1 \text{ mol C} / 12.01 \text{ g C} = 0.06669 \text{ mol C}$$

$$0.9333 \text{ g H} \times 1 \text{ mol H} / 1.01 \text{ g H} = 0.9241 \text{ mol H}$$

$$0.2016 \text{ g N} \times 1 \text{ mol N} / 14.01 \text{ g N} = 0.0144 \text{ mol N}$$

moles

$$2.130 \text{ g O} \times 1 \text{ mol O} / 16.00 \text{ g O} = 0.1331 \text{ mol O}$$

divide by the smallest # of

$$0.06669 \text{ mol C} / 0.0144 = 4.63 \text{ mol C}$$

$$0.9241 \text{ mol H} / 0.0144 = 64.2 \text{ mol H}$$

$$0.0144 \text{ mol N} / 0.0144 = 1 \text{ mol N}$$

carbamate

$$0.1331 \text{ mol O} / 0.0144 = 9.24 \text{ mol O}$$

$C_4.63H_{64.2}N_1O_{9.24}$ → empirical formula for

6) Determine the empirical formula first, by assuming that 100g of the sample is available, allowing for 24.27% C → 24.27g C, etc... Use the relative masses of each element to determine moles and find the lowest whole number mole ratio. Then compare the mass of the empirical formula to that of the mass of the molecular formula provided. This comparison will allow you to determine the molecular formula.

$$24.27\% \text{ C} \rightarrow 24.27 \text{ g C} \times 1 \text{ mol C} / 12.01 \text{ g C} = 2.021 \text{ mol C}$$

$$4.07\% \text{ H} \rightarrow 4.07 \text{ g H} \times 1 \text{ mol H} / 1.01 \text{ g H} = 4.03 \text{ mol H}$$

$$71.65\% \text{ Cl} \rightarrow 71.65 \text{ g Cl} \times 1 \text{ mol Cl} / 35.45 \text{ g Cl} = 2.021 \text{ mol Cl}$$

$$2.021 \text{ mol C} / 2.021 = 1 \text{ mol C}$$

$$4.03 \text{ mol H} / 2.021 = 2 \text{ mol H}$$

$$2.021 \text{ mol Cl} / 2.021 = 1 \text{ mol Cl}$$

CH_2Cl → empirical formula

$$\text{Empirical mass of } CH_2Cl \rightarrow C: 12.01 \text{ g} +$$

$$H: 2(1.01 \text{ g}) +$$

$$Cl: 35.45 \text{ g}$$

$$= 49.47 \text{ g}$$

To compare empirical & molecular mass:

$$\text{Molecular mass} / \text{empirical mass} = n \quad n(\text{empirical formula}) = \text{molecular}$$

formula

$$98.96 \text{ g} / 49.47 \text{ g} = 2$$

$$2(CH_2Cl) = C_2H_4Cl_2$$

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