

Empirical and Molecular Formula Problems

Name KEY

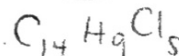
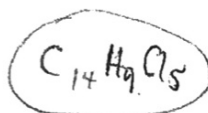
1. The insecticide DDT has the following composition by mass: 47.5 % C, 2.54 % H, and 50.0 % Cl. Determine the empirical formula of DDT.

Multiply all by 2 - not better, then 3  
Need to be closer than 0.1, if possible - lab data  
Give good lab data on test

$$\frac{47.5 \text{ g C}}{1} \bigg| \frac{1 \text{ mole C}}{12.01 \text{ g C}} = \frac{3.96 \text{ mol C}}{1.41} = 2.81 \times 5 = 14.1$$

$$\frac{2.54 \text{ g H}}{1} \bigg| \frac{1 \text{ mole H}}{1.01 \text{ g H}} = \frac{2.51 \text{ mol H}}{1.41} = 1.78 \times 5 = 8.90$$

$$\frac{50.0 \text{ g Cl}}{1} \bigg| \frac{1 \text{ mole Cl}}{35.45 \text{ g Cl}} = \frac{1.41 \text{ mol Cl}}{1.41} = 1 \times 5 = 5$$

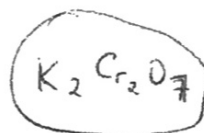


2. A compound gave on analysis the following percent composition: K = 26.57 %, Cr = 35.36 %, O = 38.07 %. Derive the empirical formula of the compound.

$$\frac{26.57 \text{ g K}}{1} \bigg| \frac{1 \text{ mole K}}{39.10 \text{ g K}} = \frac{0.6795 \text{ mol K}}{0.6795} = 1 \times 2 = 2$$

$$\frac{35.36 \text{ g Cr}}{1} \bigg| \frac{1 \text{ mole Cr}}{52.00 \text{ g Cr}} = \frac{0.6800 \text{ mol Cr}}{0.6795} = 1.001 \times 2 = 2$$

$$\frac{38.07 \text{ g O}}{1} \bigg| \frac{1 \text{ mole O}}{16.00 \text{ g O}} = \frac{2.379 \text{ mol O}}{0.6795} = 3.501 \times 2 = 7$$

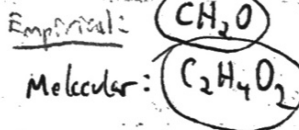


3. A compound has the following percent composition: C = 40.0 %, H = 6.67 %, O = 53.3 %. Its molar mass is 60.0 g/mole. Derive its empirical and molecular formulas.

$$\frac{40.0 \text{ g C}}{1} \bigg| \frac{1 \text{ mole C}}{12.01 \text{ g C}} = \frac{3.33 \text{ mol C}}{3.33} = 1 \text{ mol C}$$

$$\frac{6.67 \text{ g H}}{1} \bigg| \frac{1 \text{ mole H}}{1.01 \text{ g H}} = \frac{6.60 \text{ mol H}}{3.33} = 1.98 \text{ mol H}$$

$$\frac{53.3 \text{ g O}}{1} \bigg| \frac{1 \text{ mole O}}{16.00 \text{ g O}} = \frac{3.33 \text{ mol O}}{3.33} = 1 \text{ mol O}$$



$$1C \times 12.01 = 12.01 \text{ g}$$

$$2H \times 1.01 = 2.02 \text{ g}$$

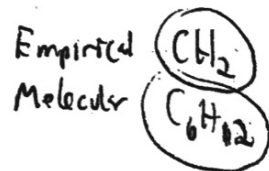
$$1O \times 16.00 = 16.00 \text{ g}$$

$$30.03 \text{ g/mol} \times 2 = 60.06 \text{ g/mol}$$

4. Determine the empirical formula and the molecular formula of a hydrocarbon which has a molar mass of 84.0 grams and contains 85.7 % carbon. (A hydrocarbon is a compound that contains only carbon and hydrogen.)

$$\frac{85.7 \text{ g C}}{1} \bigg| \frac{1 \text{ mole C}}{12.01 \text{ g C}} = \frac{7.14 \text{ mol C}}{7.14} = 1 \text{ mol C}$$

$$\frac{14.3 \text{ g H}}{1} \bigg| \frac{1 \text{ mole H}}{1.01 \text{ g H}} = \frac{14.2 \text{ mol H}}{7.14} = 1.99 \text{ mol H} \approx 2 \text{ mol H}$$



$$1C \times 12.01 = 12.01 \text{ g}$$

$$2H \times 1.01 = 2.02 \text{ g}$$

$$14.03 \text{ g} \times \sim 6 = 84.0 \text{ g}$$

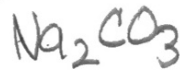
Multiply empirical formula by 6

5. A sample of a pure compound contains 2.04 g of sodium,  $2.65 \times 10^{22}$  atoms of carbon, and 0.132 mole of oxygen atoms. Find the empirical formula. (Think in terms of moles.)

$$\frac{2.04 \text{ g Na}}{1} \times \frac{1 \text{ mol}}{22.99 \text{ g}} = \frac{.0887 \text{ mol Na}}{.0440} = 2.02$$

$$\frac{2.65 \times 10^{22} \text{ atoms C}}{1} \times \frac{1 \text{ mol C}}{6.02 \times 10^{23} \text{ atoms}} = \frac{.0440 \text{ mol C}}{.0440} = 1$$

$$\frac{.132 \text{ mol O}}{.0440} = 3$$



6. A 2.500 g sample of uranium was heated in the air. The resulting oxide weighed 2.949 g. Determine the empirical formula of the compound.

$$2.949 \text{ g} - 2.500 = .449 \text{ g Oxygen}$$

$$\frac{.449 \text{ g O}}{1} \times \frac{1 \text{ mol O}}{16.00 \text{ g}} = \frac{.0281 \text{ mol O}}{.01050} = 2.68 \times 3 = 8.04$$

$$\frac{2.500 \text{ g U}}{1} \times \frac{1 \text{ mol U}}{238.03 \text{ g}} = \frac{.01050}{.01050} = 1 \times 3 = 3$$



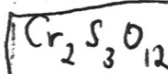
7. Determine the simplest formula of a compound that has the following composition:

Cr = 26.52 %, S = 24.52 %, O = 48.96 %

$$\frac{26.52 \text{ g Cr}}{1} \times \frac{1 \text{ mol Cr}}{52.00 \text{ g Cr}} = \frac{0.5100 \text{ mol Cr}}{0.5100} = 1 \text{ mol Cr} \times 2 = 2 \text{ mol Cr}$$

$$\frac{24.52 \text{ g S}}{1} \times \frac{1 \text{ mol S}}{32.07 \text{ g S}} = \frac{0.7646 \text{ mol S}}{0.5100} = 1.499 \text{ mol S} \times 2 = \sim 3 \text{ mol S}$$

$$\frac{48.96 \text{ g O}}{1} \times \frac{1 \text{ mol O}}{16.00 \text{ g O}} = \frac{3.060 \text{ mol O}}{0.5100} = 6 \text{ mol O} \times 2 = 12 \text{ mol O}$$

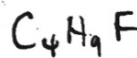


8. A compound contains 63.1 % carbon, 11.92 % hydrogen, and 24.97 % fluorine. Derive its empirical formula.

$$\frac{63.1 \text{ g C}}{1} \times \frac{1 \text{ mol C}}{12.01 \text{ g C}} = \frac{5.25 \text{ mol C}}{1.314} = \sim 4 \text{ mol C}$$

$$\frac{11.92 \text{ g H}}{1} \times \frac{1 \text{ mol H}}{1.01 \text{ g H}} = \frac{11.8 \text{ mol H}}{1.314} = \sim 9 \text{ mol H}$$

$$\frac{24.97 \text{ g F}}{1} \times \frac{1 \text{ mol F}}{19.00 \text{ g F}} = \frac{1.314 \text{ mol F}}{1.314} = 1 \text{ mol F}$$



9. A compound with a molar mass of about 175 grams/mole consists of 40.0 % carbon, 6.7 % hydrogen, and 53.3 % oxygen. What is the empirical and molecular formula of the compound?

$$\frac{40.0 \text{ g C}}{1} \times \frac{1 \text{ mole C}}{12.01 \text{ g C}} = \frac{3.33 \text{ mol C}}{3.33} = 1 \text{ mol C}$$

$$\frac{6.7 \text{ g H}}{1} \times \frac{1 \text{ mole H}}{1.01 \text{ g H}} = \frac{6.6 \text{ mol H}}{3.33} = 2.0 \text{ mol H}$$

$$\frac{53.3 \text{ g O}}{1} \times \frac{1 \text{ mol O}}{16.00 \text{ g O}} = \frac{3.33 \text{ mol O}}{3.33} = 1 \text{ mol O}$$



$$\text{C} = 12.01 \times 1 = 12.01 \text{ g}$$

$$\text{H} = 1.01 \times 2 = 2.02 \text{ g}$$

$$\text{O} = 16.00 \times 1 = 16.00 \text{ g}$$

$$\sim 6 \times \frac{20.03 \text{ g/mol}}{175}$$

