Combustion Quiz

3.80 grams of a compound consisting of C, H, and O was completely combusted in oxygen to give 0.0666 moles of carbon dioxide and 0.167 moles of water. How many grams of oxygen must have been in the original compound before it was burned? What is the empirical formula of this compound?

$$C_xH_yO_z + O_2 \rightarrow CO_2 + H_2O$$

3.80g ; .0666 moles ; 0.167 moles

$$\#gC = .0666 \text{ moles } CO_2 \times \frac{1 \text{ molesC}}{1 \text{ molesCO}_2} \times \frac{12.01g}{1\text{moleC}} = .800gC$$

$$\#gH = 0.167 \text{ moles } H_2O \times \frac{2 \text{ moleH}}{1\text{moleH}_2O} \times \frac{1.008}{1\text{moleH}} = .337gH$$

$$\#gO = 3.80g - (.800g + .337g) = \boxed{2.66g\ O \text{ in original}}$$

$$\#\text{moles C} = .800gC \times \frac{1\text{mole}}{12.01g} = \frac{.067 \text{ moles}}{.067} = 1 \times 2 = 2$$

$$\#\text{moles H} = .337g \times \frac{1\text{mole}}{1.008g} = \frac{.334 \text{ moles}}{.067} = 5 \times 2 = 10$$

$$\#\text{moles O} = 2.66g \times \frac{1\text{mole}}{16.00g} = \frac{.166 \text{ moles}}{.067} = 2.5 \times 2 = 5$$

$$\boxed{C_2H_{10}O_5}$$

Combustion Quiz

3.80 grams of a compound consisting of C, H, and O was completely combusted in oxygen to give 0.0666 moles of carbon dioxide and 0.167 moles of water. How many grams of oxygen must have been in the original compound before it was burned? What is the empirical formula of this compound?

Combust This

1. The general form of a combustion reaction is:

$$CHO + O_2 \rightarrow CO_2 + H_2O$$
$$CH$$

2. Using the Law of Conservation of Mass, you can determine the mass of the reactants in a combustion reaction by stoichiometric means using the products.

Problem 1991 B

The molecular formula of a hydrocarbon is to be determined by analysis of its combustion products. The hydrocarbon burns completely, producing 7.2 grams of water and 7.2 liters of CO_2 at standard conditions. What is the empirical formula of the hydrocarbon?

$$CH + O_2 \rightarrow \underset{7.2L}{CO_2} + \underset{7.2g}{H_2O}$$

$$\# grams\ C = 7.2L\ CO_2 \times \frac{1 mole\ CO_2}{22.4\ L\ CO_2} \times \frac{1\ mole\ C}{1\ mole\ CO_2} \times \frac{12.01g}{1 mole\ C} = 3.86g$$

$$\# grams\ H = 7.2g H_2O \times \frac{1 mole\ H_2O}{18.02\ g\ H_2O} \times \frac{2\ mole\ H}{1 mole\ H_2O} \times \frac{1.008g}{1 mole\ H} = .806g$$

$$\# moles\ C = 3.86g \times \frac{1 mole}{12.01\ g} = \frac{.321\ moles\ C}{.321} = 1 \times 2 = 2$$

$$\# moles\ H = .806g \times \frac{1 mole}{1.008} = \frac{.800\ mole\ H}{.321} = 2.5 \times 2 = 5$$

$$\boxed{C_2H_5}$$

Problem 1982

A combustion device was used to determine the empirical formula of a compound containing only carbon, hydrogen and oxygen. A 0.6349 g sample of the unknown produced 1.603 g of CO_2 and ~~0.810~~ 0.2810 g of H_2O. Determine the empirical formula of the compound.

$$\underset{.6349}{CHO} + O_2 \rightarrow \underset{1.603g}{CO_2} + \underset{.810g}{H_2O}$$

$$\# g\ C = 1.603g\ CO_2 \times \frac{1 mole\ CO_2}{44.01 g CO_2} \times \frac{1 mole C}{1 mole CO_2} \times \frac{12.01g}{1 mole C} = .437g$$

$$\# g H = .2810g\ H_2O \times \frac{1 mole\ H_2O}{18.02\ g H_2O} \times \frac{2 mole H}{1 mole H_2O} \times \frac{1.008g}{1 mole H} = .0314g$$

$$\# g O = .6349 - (.437 + .0314) = .1665g$$

$$\# moles\ C = .437g \times \frac{1 mole}{12.01 g} = \frac{.036\ mole C}{.010} \sim 3.5 \times 2 = 7$$

$$\# moles\ H = .0314g \times \frac{1 mole}{1.008g} = \frac{.031}{.010}\ mole H \sim 3. \times 2 = 6$$

$$\# moles\ O = .1665g \times \frac{1 mole}{16.00g} = \frac{.010}{.010}\ mole O = 1 \times 2 = 2$$

$$\boxed{C_7H_6O_2}$$