Upon combustion, a 0.8233 g sample of a compound containing only carbon, hydrogen, and oxygen produced 2.445 g CO_2 and 0.6003 g H_2O. Find the empirical formula.

Due to the limitations of CCLE, **enter only the subscript of oxygen.**

Upon combustion, a 0.8233-g sample of a compound containing only carbon, hydrogen, and oxygen produced 2.445 g CO_2 and 0.6003 g H_2O. Find the empirical formula of the compound.

Given 0.8233 g sample, 2.445 g CO_2, 0.6003 g H_2O

Find empirical formula

$$2.445\ \cancel{g\ CO_2} \times \frac{1\ mol\ CO_2}{44.01\ \cancel{g\ CO_2}} = 0.05556\ mol\ CO_2$$

$$0.6003\ \cancel{g\ H_2O} \times \frac{1\ mol\ H_2O}{18.01\ \cancel{g\ H_2O}} = 0.03331\ mol\ H_2O$$

$$0.05556\ \cancel{mol\ CO_2} \times \frac{1\ mol\ C}{1\ \cancel{mol\ CO_2}} = 0.05556\ mol\ C$$

$$0.03331\ \cancel{mol\ H_2O} \times \frac{2\ mol\ H}{1\ \cancel{mol\ H_2O}} = 0.06662\ mol\ H$$

$$\text{Mass C} = 0.05556\ \cancel{mol\ C} \times \frac{12.01\ g\ C}{\cancel{mol\ C}} = 0.6673\ g\ C$$

$$\text{Mass H} = 0.06662\ \cancel{mol\ H} \times \frac{1.008\ g\ H}{\cancel{mol\ H}} = 0.06715\ g\ H$$

$$\text{Mass O} = 0.8233\ g - (0.6673\ g + 0.06715\ g) = 0.0889\ g$$

$$\text{Mol O} = 0.0889\ \cancel{g\ O} \times \frac{mol\ O}{16.00\ \cancel{g\ O}} = 0.00556\ mol\ O$$

$$C_{0.05556}H_{0.06662}O_{0.00556}$$

$$C_{\frac{0.05556}{0.00556}}H_{\frac{0.06662}{0.00556}}O_{\frac{0.00556}{0.00556}} \rightarrow C_{10}H_{12}O_1$$

The subscripts are whole numbers; no additional multiplication is needed.

The correct empirical formula is $C_{10}H_{12}O$.

Find the mass (in grams) of hydrogen contained in 1.00 gallon of water. Assume the density of water is 1.00 g/mL. Use any data/information from the textbook as needed. Report your answer to the correct number of significant figures. Enter your number using standard format (e.g. 103.2) instead of scientific notation (e.g. 1.032 x 10^2).

Given 1.00 gal H_2O

$d_{H_2O} = 1.00\ \text{g/mL}$

Find g H

Conceptual Plan

$$\boxed{\text{gal } H_2O} \xrightarrow{\frac{3.785\ \text{L}}{1\ \text{gal}}} \boxed{\text{L } H_2O} \xrightarrow{\frac{1000\ \text{mL}}{1\ \text{L}}} \boxed{\text{mL } H_2O} \xrightarrow{\frac{1.00\ \text{g}}{1\ \text{mL}}} \boxed{\text{g } H_2O}$$

$$\boxed{\text{g } H_2O} \xrightarrow{\frac{1\ \text{mol } H_2O}{18.02\ \text{g } H_2O}} \boxed{\text{mol } H_2O} \xrightarrow{\frac{2\ \text{mol H}}{1\ \text{mol } H_2O}} \boxed{\text{mol H}} \xrightarrow{\frac{1.008\ \text{g H}}{1\ \text{mol H}}} \boxed{\text{g H}}$$

Relationships Used

3.785 L = 1 gal (Table 1.3)

1000 mL = 1 L

1.00 g H_2O = 1 mL H_2O (density of H_2O)

Molar mass $H_2O = 2(1.008) + 16.00 = 18.02$ g/mol

2 mol H : 1 mol H_2O

1.008 g H = 1 mol H

Solution

$$1.00\ \cancel{\text{gal } H_2O} \times \frac{3.785\ \cancel{\text{L}}}{\cancel{1\ \text{gal}}} \times \frac{1000\ \cancel{\text{mL}}}{1\ \cancel{\text{L}}} \times \frac{1.0\ \text{g}}{\cancel{\text{mL}}} = 3.785 \times 10^3\ \text{g } H_2O$$

$$3.785 \times 10^3\ \cancel{\text{g } H_2O} \times \frac{1\ \cancel{\text{mol } H_2O}}{18.02\ \cancel{\text{g } H_2O}} \times \frac{2\ \cancel{\text{mol H}}}{1\ \cancel{\text{mol } H_2O}} \times \frac{1.008\ \text{g H}}{1\ \cancel{\text{mol H}}} = 4.23 \times 10^2\ \text{g H}$$

Titanium metal can be formed by the following reaction:

$TiO_2(s) + 2\,C(s) \longrightarrow Ti(s) + 2\,CO(g)$

When 28.6 kg of C is allowed to react with 88.2 kg of TiO_2, 42.8 kg of Ti is produced. Calculate the percentage yield. Report your answer as a percentage to 1 decimal place. Do not enter the "%" sign, just the numerical value.

Given 28.6 kg C, 88.2 kg TiO_2, 42.8 kg Ti produced

Find limiting reactant, theoretical yield, % yield

Conceptual Plan

$$\text{kg C} \xrightarrow{\frac{1000\text{ g}}{1\text{ kg}}} \text{g C} \xrightarrow{\frac{1\text{ mol C}}{12.01\text{ g C}}} \text{mol C} \xrightarrow{\frac{1\text{ mol Ti}}{2\text{ mol C}}} \text{mol Ti}$$

$$\text{kg TiO}_2 \xrightarrow{\frac{1000\text{ g}}{1\text{ kg}}} \text{g TiO}_2 \xrightarrow{\frac{1\text{ mol TiO}_2}{79.87\text{ g TiO}_2}} \text{mol TiO}_2 \xrightarrow{\frac{1\text{ mol Ti}}{1\text{ mol TiO}_2}} \text{mol Ti}$$

Smallest amount determines limiting reactant.

$$\text{mol Ti} \xrightarrow{\frac{47.87\text{ g Ti}}{\text{mol Ti}}} \text{g Ti} \xrightarrow{\frac{1\text{ kg}}{1000\text{ g}}} \text{kg Ti}$$

Relationships Used

1000 g = 1 kg
molar mass of C = 12.01 g/mol
molar mass of TiO_2 = 79.87 g/mol
1 mol TiO_2 : 1 mol Ti
2 mol C : 1 mol Ti
molar mass of Ti = 47.87 g/mol

Solution

$$28.6\,\cancel{\text{kg C}} \times \frac{1000\,\cancel{\text{g}}}{1\,\cancel{\text{kg}}} \times \frac{1\,\cancel{\text{mol C}}}{12.01\,\cancel{\text{g C}}} \times \frac{1\text{ mol Ti}}{2\,\cancel{\text{mol C}}} = 1.1\underline{9}07 \times 10^3\text{ mol Ti}$$

Limiting reactant

Least amount of product

$$\boxed{88.2\,\cancel{\text{kg TiO}_2}} \times \frac{1000\,\cancel{\text{g}}}{1\,\cancel{\text{kg}}} \times \frac{1\,\cancel{\text{mol TiO}_2}}{79.87\,\cancel{\text{g TiO}_2}} \times \frac{1\,\cancel{\text{mol Ti}}}{1\,\cancel{\text{mol TiO}_2}} = \boxed{1.1\underline{0}43 \times 10^3\text{ mol Ti}}$$

$$1.1\underline{0}43 \times 10^3\,\cancel{\text{mol Ti}} \times \frac{47.87\,\cancel{\text{g Ti}}}{1\,\cancel{\text{mol Ti}}} \times \frac{1\text{ kg}}{1000\,\cancel{\text{g}}} = 52.9\text{ kg Ti}$$

Since TiO_2 makes the least amount of product, it is the limiting reactant and 52.9 kg Ti is the theoretical yield.

$$\%\text{ yield} = \frac{\text{actual yield}}{\text{theoretical yield}} \times 100\%$$

$$= \frac{42.8\text{ g}}{52.9\text{ g}} \times 100\% = 80.9\%$$

The red light emitted by a barcode scanner has a frequency of $4.62 \times 10^{14}\ s^{-1}$. Determine its wavelength in units of nanometers. Report your answer to the correct number of significant figures. Enter your number using standard format (e.g. 103.2) instead of scientific notation (e.g. 1.032×10^2).

$$\nu = \frac{c}{\lambda}$$

$$\lambda = \frac{c}{\nu} = \frac{3.00 \times 10^8\ \text{m}/\cancel{\text{s}}}{4.62 \times 10^{14}/\cancel{\text{s}}}$$

$$= 6.49 \times 10^{-7}\ \text{m}$$

$$6.49 \times 10^{-7}\ \cancel{\text{m}} \times \frac{1\ \text{nm}}{10^{-9}\ \cancel{\text{m}}} = 649\ \text{nm}$$

Calculate the speed of an electron having a wavelength of 2.74×10^{-10} m. Report your answer in units of meters per second to the correct number of significant figures. Use any data/information from the textbook as needed. Enter your number using standard format (e.g. 103.2) instead of scientific notation (e.g. 1.032×10^2).

$v = h / m\lambda = 2650000$ m/s