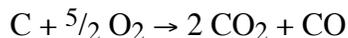


Stoichiometry

1970

A 2.000 gram sample containing graphite (carbon) and an inert substance was burned in oxygen and produced a mixture of carbon dioxide and carbon monoxide in the mole ratio 2.00:1.00. The volume of oxygen used was 747.0 milliliters at 1,092K and 12.00 atmospheres pressure. Calculate the percentage by weight of graphite in the original mixture.

Answer:



$$n = \frac{PV}{RT} = \frac{(12.0\text{atm})(0.7470\text{L})}{(0.08205 \frac{\text{L}\cdot\text{atm}}{\text{mol}\cdot\text{K}})(1092\text{K})} = 0.100 \text{ mol } O_2$$

$$0.100 \text{ mol } O_2 \times \frac{2 \text{ mol } C}{5 \text{ mol } O_2} \times \frac{12.0 \text{ g } C}{1 \text{ mol } C} = 0.480 \text{ g } C$$

$$\frac{0.480 \text{ g } C}{2.000 \text{ g sample}} \times 100 = 24.0\% C$$

1982 B

Water is added to 4.267 grams of UF_6 . The only products are 3.730 grams of a solid containing only uranium, oxygen and fluorine and 0.970 gram of a gas. The gas is 95.0% fluorine, and the remainder is hydrogen.

- From these data, determine the empirical formula of the gas.
- What fraction of the fluorine of the original compound is in the solid and what fraction in the gas after the reaction?
- What is the formula of the solid product?
- Write a balanced equation for the reaction between UF_6 and H_2O . Assume that the empirical formula of the gas is the true formula.

Answer:

- Assume 100 g of compound

$$(95.0 \text{ g F})(1 \text{ mol F}/19.0\text{g}) = 5.0 \text{ mol F}$$

$$(5.0 \text{ g H})(1 \text{ mol H}/1.00\text{g}) = 5.0 \text{ mol H}$$

$$5.0 \text{ mol F} : 5.0 \text{ mol H} = 1 \text{ F} : 1 \text{ H}, = \text{HF}$$

- $4.267 \text{ g } UF_6 \times \frac{1 \text{ mol } UF_6}{352.0 \text{ g } UF_6} \times \frac{6 \text{ mol F}}{1 \text{ mol } UF_6} =$

$$= 0.07273 \text{ mol F in original compound}$$

$$0.970 \text{ g HF} \times \frac{1 \text{ mol HF}}{20.0 \text{ g HF}} \times \frac{1 \text{ mol F}}{1 \text{ mol HF}} = 0.0485 \text{ mol F}$$

$$\frac{0.0485 \text{ mol}}{0.07273 \text{ mol}} \times 100 = 66.68\% \text{ F in gas}$$

$$(100.0 - 66.68)\% = 33.32\% \text{ F in solid}$$

- $4.267 \text{ g } UF_6 \times \frac{1 \text{ mol } UF_6}{352.0 \text{ g } UF_6} \times \frac{1 \text{ mol U}}{1 \text{ mol } UF_6} =$

$$= 0.01212 \text{ mol U}$$

$$(0.07273 \text{ mol F in original compound}) - (0.0485 \text{ mol F in gas}) = 0.02433 \text{ mol F in solid}$$

$$(4.267 + X)\text{g} = (3.730 + 0.970)\text{g}; X = 0.433 \text{ g } H_2O$$

$$0.433 \text{ g } H_2O \times \frac{1 \text{ mol } H_2O}{18.0 \text{ g } H_2O} \times \frac{1 \text{ mol O}}{1 \text{ mol } H_2O} =$$

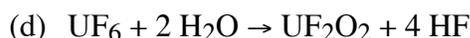
$$= 0.02406 \text{ mol O}$$

$$0.01212 \text{ mol U}/0.01212 \text{ mol} = 1$$

$$0.02433 \text{ mol F}/0.01212 \text{ mol} = 2.007$$

$$0.02406 \text{ mol O}/0.01212 \text{ mol} = 1.985$$

$$= UF_2O_2$$



1986 B

Three volatile compounds X, Y, and Z each contain element Q. The percent by weight of element Q in each compound was determined. Some of the data obtained are given below.

Compound	Percent by weight of Element Q	Molecular Weight
X	64.8%	?
Y	73.0%	104.
Z	59.3%	64.0

- The vapor density of compound X at 27°C and 750. mm Hg was determined to be 3.53 grams per litre. Calculate the molecular weight of compound X.
- Determine the mass of element Q contained in 1.00 mole of each of the three compounds.
- Calculate the most probable value of the atomic weight of element Q.
- Compound Z contains carbon, hydrogen, and element Q. When 1.00 gram of compound Z is oxidized and all of the carbon and hydrogen are converted to oxides, 1.37 grams of CO_2 and 0.281 gram of water are produced. Determine the most probable molecular formula of compound Z.

Answer:

Stoichiometry

$$(a) \text{ mol.wt.} = \frac{gRT}{PV} = \frac{(3.53g) \left(0.0821 \frac{\text{L atm}}{\text{mol K}}\right) (300\text{K})}{\left(\frac{750}{760} \text{ atm}\right) (1.00\text{L})}$$

$$= 88.1 \text{ g/mol}$$

	X	Y	Z
	88.1 g/mol	104	64.0
% Q	64.8	73.0	59.3
g Q	57.1	75.9	38.0

- (c) ratio 1.5 2 1
- masses must be integral multiples of atomic weight
- therefore, 3 4 2
- which gives an atomic weight of Q = 19

$$(d) 1.37\text{gCO}_2 \times \frac{1 \text{ mol CO}_2}{44.0\text{gCO}_2} \times \frac{1 \text{ mol C}}{1 \text{ mol CO}_2} = 0.0311\text{mol}$$

$$0.281\text{gH}_2\text{O} \times \frac{1 \text{ mol H}_2\text{O}}{18.0\text{gH}_2\text{O}} \times \frac{2 \text{ mol H}}{1 \text{ mol H}_2\text{O}} = 0.0312\text{molH}$$

$$1.00 \text{ g Z is } 59.3\% \text{ Q} = 0.593 \text{ g Q}$$

$$0.0593 \text{ g Q} \times \frac{1 \text{ mol}}{19.0 \text{ g}} = 0.0312 \text{ mol Q}$$

therefore, the empirical formula = CHQ, the smallest whole number ratio of moles.

formula wt. of CHQ = 32.0, if mol. wt. Z = 64 then the formula of Z = (CHQ)₂ or C₂H₂Q₂

1991 B

The molecular formula of a hydrocarbon is to be determined by analyzing its combustion products and investigating its colligative properties.

- (a) The hydrocarbon burns completely, producing 7.2 grams of water and 7.2 liters of CO₂ at standard conditions. What is the empirical formula of the hydrocarbon?
- (b) Calculate the mass in grams of O₂ required for the complete combustion of the sample of the hydrocarbon described in (a).
- (c) The hydrocarbon dissolves readily in CHCl₃. The freezing point of a solution prepared by mixing 100. grams of CHCl₃ and 0.600 gram of the hydrocarbon is -64.0°C. The molal freezing-point depression constant of CHCl₃ is 4.68°C/molal and its normal freezing point is -63.5°C. Calculate the molecular weight of the hydrocarbon.
- (d) What is the molecular formula of the hydrocarbon?

Answer:

- (a)

Stoichiometry

$$7.2 \text{ g H}_2\text{O} \times \frac{1 \text{ mol H}_2\text{O}}{18.0 \text{ g H}_2\text{O}} \times \frac{2 \text{ mol H}}{1 \text{ mol H}_2\text{O}} = 0.80 \text{ mol H}$$

$$7.2 \text{ L CO}_2 \times \frac{1 \text{ mol CO}_2}{22.4 \text{ L CO}_2} \times \frac{1 \text{ mol C}}{1 \text{ mol CO}_2} = 0.32 \text{ mol C}$$

(or $PV=nRT$ could be used to solve for n)

$$\frac{0.80 \text{ mol H}}{0.32 \text{ mol C}} = \frac{2.5 \text{ H}}{1 \text{ C}} = \frac{5 \text{ H}}{2 \text{ C}}$$

$$= \text{C}_2\text{H}_5$$

(b) 0.40 mol oxygen in water + 0.64 mol oxygen in $\text{CO}_2 = 1.04 \text{ mol O} = 0.52 \text{ mol O}_2 = 16.64 \text{ g} = 17 \text{ g}$ of oxygen gas (alternative approach for mol O_2 from balanced equation)

(c) $-63.5^\circ\text{C} - (-64.0^\circ\text{C}) = 0.5^\circ\text{C}$

$$0.5^\circ\text{C} \times \frac{1 \text{ molal}}{4.68^\circ\text{C}} = 0.107 \text{ molal} = \frac{0.107 \text{ mol solute}}{1.0 \text{ kg solvent}}$$

$$\frac{0.107 \text{ mol HC}}{1.0 \text{ kg CHCl}_3} \times 0.100 \text{ kg CHCl}_3 = 0.0107 \text{ mol}$$

$$\text{mol.wt.} = \frac{0.60 \text{ g HC}}{0.0107 \text{ mol}} = 56.2 \frac{\text{g}}{\text{mol}}$$

OR

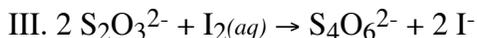
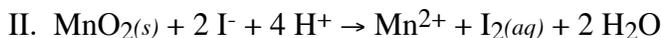
solve for mol. wt. using

$$\Delta T = \frac{K \left(\frac{\text{g}}{\text{mol.wt.}} \right)}{\text{kg solvent}}$$

(d) $\text{C}_2\text{H}_5 = 29 \text{ g mol}^{-1}$

$$56.2/29 = 1.93 = 2, \therefore (\text{C}_2\text{H}_5)_2 = \text{C}_4\text{H}_{10}$$

1993 B



The amount of oxygen, O_2 , dissolved in water can be determined by titration. First, MnSO_4 and NaOH are added to a sample of water to convert all of the dissolved O_2 to MnO_2 , as shown in equation I above. Then H_2SO_4 and KI are added and the reaction represented by equation II proceeds. Finally, the I_2 that is formed is titrated with standard sodium thiosulfate, $\text{Na}_2\text{S}_2\text{O}_3$, according to equation III.

(a) According to the equation above, how many moles of $\text{S}_2\text{O}_3^{2-}$ are required for analyzing 1.00 mole of O_2 dissolved in water?

(b) A student found that a 50.0-milliliter sample of water required 4.86 milliliters of 0.0112-molar $\text{Na}_2\text{S}_2\text{O}_3$ to reach the equivalence point. Calculate the number of moles of O_2 dissolved in this sample.

(c) How would the results in (b) be affected if some I_2 were lost before the $\text{S}_2\text{O}_3^{2-}$ was added? Explain.

Stoichiometry

- (d) What volume of dry O₂ measured at 25°C and 1.00 atmosphere of pressure would have to be dissolved in 1.00 liter of pure water in order to prepare a solution of the same concentration as that obtained in (b)?
- (e) Name an appropriate indicator for the reaction shown in equation III and describe the change you would observe at the end point of the titration.

Answer:

- $$\frac{1 \text{ mol O}_2}{1 \text{ mol O}_2} \times \frac{2 \text{ mol MnO}_2}{1 \text{ mol O}_2} \times \frac{1 \text{ mol I}_2}{1 \text{ mol MnO}_2} \times$$
- (a) $\frac{2 \text{ mol S}_2\text{O}_3^{2-}}{1 \text{ mol I}_2} = 4 \text{ mol S}_2\text{O}_3^{2-}$
- (b) $\text{mol S}_2\text{O}_3^{2-} = (0.00486 \text{ L})(0.0112 \text{ M}) = 5.44 \times 10^{-5} \text{ mol S}_2\text{O}_3^{2-}$
- $$\frac{5.44 \times 10^{-5} \text{ mol S}_2\text{O}_3^{2-}}{4 \text{ mol S}_2\text{O}_3^{2-}} \times \frac{1 \text{ mol O}_2}{4 \text{ mol S}_2\text{O}_3^{2-}} =$$
- $$1.36 \times 10^{-5} \text{ mol O}_2$$
- (c) less I₂ means less thiosulfate ion required thus indicating a lower amount of dissolved oxygen.
- (d) molarity of solution in (b) = $1.36 \times 10^{-5} \text{ mol O}_2 / 0.050 \text{ L} = 2.72 \times 10^{-4} \text{ M}$
- $$V = \frac{nRT}{P} = \frac{(2.72 \times 10^{-4} \text{ M})(0.08205 \frac{\text{L atm}}{\text{mol K}})(298 \text{ K})}{1.00 \text{ atm}}$$
- $$= 6.65 \times 10^{-5} \text{ L or } 6.65 \text{ mL O}_2$$
- (e) starch indicator
color disappears or blue disappears
[color Δ alone is not sufficient for 2nd pt.; any other color w/starch is not sufficient for 2nd pt.]

1995 B

A sample of dolomitic limestone containing only CaCO₃ and MgCO₃ was analyzed.

- (a) When a 0.2800 gram sample of this limestone was decomposed by heating, 75.0 milliliters of CO₂ at 750 mm Hg and 20°C were evolved. How many grams of CO₂ were produced.
- (b) Write equations for the decomposition of both carbonates described above.
- (c) It was also determined that the initial sample contained 0.0448 gram of calcium. What percent

of the limestone by mass was CaCO₃?

- (d) How many grams of the magnesium-containing product were present in the sample in (a) after it had been heated?

Answer:

- (a) $n = \frac{PV}{RT} = \frac{(750 \text{ mm Hg})(75.0 \text{ mL})}{\left(62400 \frac{\text{mL mm Hg}}{\text{mol K}}\right)(293 \text{ K})}$
- $$= 3.08 \times 10^{-3} \text{ mol}$$
- $$3.08 \times 10^{-3} \text{ mol} \times \frac{44.0 \text{ g CO}_2}{1 \text{ mol}} = 0.135 \text{ g CO}_2$$
- (b) CaCO₃ → CaO + CO₂
MgCO₃ → MgO + CO₂
- (c) $0.0448 \text{ g Ca} \times \frac{1 \text{ mol Ca}}{40.08 \text{ g Ca}} \times \frac{1 \text{ mol CaCO}_3}{1 \text{ mol Ca}}$
- $$\times \frac{100.1 \text{ g CaCO}_3}{1 \text{ mol CaCO}_3} = 0.112 \text{ g CaCO}_3$$
- $$\frac{0.112 \text{ g CaCO}_3}{0.2800 \text{ g sample}} = 40.0\% \text{ CaCO}_3$$
- (d) 60.0% of 0.2800 g sample = 0.168 g of MgCO₃
- $$0.168 \text{ g MgCO}_3 \times \frac{1 \text{ mol MgCO}_3}{84.3 \text{ g MgCO}_3} \times$$
- $$\frac{1 \text{ mol MgO}}{1 \text{ mol MgCO}_3} \times \frac{40.3 \text{ g MgO}}{1 \text{ mol MgO}} = 0.0803 \text{ g MgO}$$

1998 B

An unknown compound contains only the three elements C, H, and O. A pure sample of the compound is analyzed and found to be 65.60 percent C and 9.44 percent H by mass.

- (a) Determine the empirical formula of the compound.
- (b) A solution of 1.570 grams of the compound in 16.08 grams of camphor is observed to freeze at a temperature 15.2 Celsius degrees below the normal freezing point of pure camphor. Determine the molar mass and apparent molecular formula of the compound. (The molal freezing-point depression constant, *K_f*, for camphor is 40.0 kg·K·mol⁻¹.)
- (c) When 1.570 grams of the compound is vaporized at 300°C and 1.00 atmosphere, the gas occupies a volume of 577 milliliters. What is the molar mass of the compound based on this

Stoichiometry

result?

- (d) Briefly describe what occurs in solution that accounts for the difference between the results obtained in parts (b) and (c).

Answer

$$(a) \quad \frac{65.60}{12.011} = 5.462 \quad \frac{9.44}{1.0079} = 9.37 \quad \frac{24.96}{16.00} = 1.56$$

$$\frac{5.462}{1.56} = 3.50 \quad \frac{9.37}{1.56} = 6.00 \quad \frac{1.56}{1.56} = 1.00$$

$$(C_{3.5}H_6O_1)_2 = C_7H_{12}O_2$$

$$(b) \quad \Delta T_{FP} = k_f m = k_f \frac{\text{mol solute}}{\text{kg solvent}} = k_f \frac{\frac{\text{g solute}}{\text{molar mass}}}{\text{kg solvent}}$$

$$\text{molar mass} = \frac{(\text{g solute})(k_f)}{(\text{kg solvent})(\Delta T_{FP})}$$

$$= \frac{(1.570 \text{ g})(40.0 \frac{\text{kg K}}{\text{mol}})}{(0.01608 \text{ kg})(15.2^\circ\text{C})} = 257 \text{ g/mol}$$

$$C_7H_{12}O_2 = 128 \text{ g/mol} ; 257/128 = 2$$

$$\therefore, (C_7H_{12}O_2)_2 = C_{14}H_{24}O_4$$

$$(c) \quad PV = \frac{\text{grams}}{\text{molar mass}} RT$$

$$\text{molar mass} = \frac{\text{grams} \times R \times T}{P \times V}$$

$$\frac{(1.570 \text{ g}) \times \left(0.08205 \frac{\text{L} \cdot \text{atm}}{\text{mol} \cdot \text{K}}\right) \times 573 \text{K}}{1.00 \text{ atm} \times 0.577 \text{ L}} = 128 \text{ g/mol}$$

- (d) since the apparent molar mass from freezing point change is twice that determined by the vapor, the molecules of the compound must associate in camphor to form dimers, **OR** the compound must dissociate in the gas phase.