***Chemistry***

**9: Gases**

**9.3: Stoichiometry of Gaseous Substances, Mixtures, and Reactions**

49. Calculate the density of Freon 12, CF2Cl2, at 30.0 °C and 0.954 atm.

Solution



51. A cylinder of O2(*g*) used in breathing by emphysema patients has a volume of 3.00 L at a pressure of 10.0 atm. If the temperature of the cylinder is 28.0 °C, what mass of oxygen is in the cylinder?

Solution



53. What is the molar mass of a gas if 0.281 g of the gas occupies a volume of 125 mL at a temperature 126 °C and a pressure of 777 torr?

Solution

From the ideal gas law, *PV = nRT*, set  and solve the molar mass.



55. The density of a certain gaseous fluoride of phosphorus is 3.93 g/L at STP. Calculate the molar mass of this fluoride and determine its molecular formula.

Solution





*M*phosphorous = 30.97376 g/mol

*M*fluorine = 18.998403 g/mol



The molecular formula is PF3.

To find this answer you can either use trial and error, or you can realize that since phosphorus is in group 5, it can fill its valence shell by forming three bonds. Fluorine, being in group 7, needs to form only one bond to fill its shell. Thus it makes sense to start with PF3 as a probable formula.

57. A 36.0–L cylinder of a gas used for calibration of blood gas analyzers in medical laboratories contains 350 g CO2, 805 g O2, and 4,880 g N2. What is the pressure in the flask in atmospheres, in torr, and in kilopascals?

Solution

Calculate the moles of each gas present and from that, calculate the pressure from the ideal gas law. Assume 25 °C. The calibration gas contains:







Total moles = 7.953 + 25.157+ 174.202 = 207.312 mol



*P* in torr = 107,000 torr

*P* in kPa = 14,300 kPa

59. A sample of gas isolated from unrefined petroleum contains 90.0% CH4, 8.9% C2H6, and 1.1% C3H8 at a total pressure of 307.2 kPa. What is the partial pressure of each component of this gas? (The percentages given indicate the percent of the total pressure that is due to each component.)

Solution

Since these are percentages of the total pressure, the partial pressure can be calculated as follows:

CH4: 90% of 307.2 kPa = 0.900 × 307.2 = 276 kPa

C2H6: 8.9% of 307.2 kPa = 0.089 × 307.2 = 27 kPa

C3H8: 1.1% of 307.2 kPa = 0.011 × 307.2 = 3.4 kPa

61. Most mixtures of hydrogen gas with oxygen gas are explosive. However, a mixture that contains less than 3.0 % O2 is not. If enough O2 is added to a cylinder of H2 at 33.2 atm to bring the total pressure to 34.5 atm, is the mixture explosive?

Solution

The oxygen increases the pressure within the tank to (34.5 atm – 33.2 atm =) 1.3 atm. The percentage O2 on a mole basis is 100% = 3.77%. The mixture is explosive. However, the percentage is given as a weight percent. Converting to a mass basis increases the percentage of oxygen even more, so the mixture is still explosive.

63. A sample of carbon monoxide was collected over water at a total pressure of 756 torr and a temperature of 18 °C. What is the pressure of the carbon monoxide? (See Table 9.2) for the vapor pressure of water.

Solution

The vapor pressure of water at 18 °C is 15.5 torr. Subtract the vapor pressure of water from the total pressure to find the pressure of the carbon monoxide:

*P*T = *P*gas + *P*water

Rearrangement gives:

*P*T – *P*water = *P*gas

756 torr – 15.5 torr = 740 torr

65. Joseph Priestley first prepared pure oxygen by heating mercuric oxide, HgO:



(a) Outline the steps necessary to answer the following question: What volume of O2 at 23 °C and 0.975 atm is produced by the decomposition of 5.36 g of HgO?

(b) Answer the question.

Solution

(a) Determine the moles of HgO that decompose; using the chemical equation, determine the moles of O2 produced by decomposition of this amount of HgO; and determine the volume of O2 from the moles of O2, temperature, and pressure.

(b)



*PV* = *nRT*

*P* = 0.975 atm

*T* = (23.0 + 273.15) K



67. The chlorofluorocarbon CCl2F2 can be recycled into a different compound by reaction with hydrogen to produce CH2F2(*g*), a compound useful in chemical manufacturing:



(a) Outline the steps necessary to answer the following question: What volume of hydrogen at 225 atm and 35.5 °C would be required to react with 1 ton (1.000  103 kg) of CCl2F2?

(b) Answer the question.

Solution

(a) Determine the molar mass of CCl2F2. From the balanced equation, calculate the moles of H2 needed for the complete reaction. From the ideal gas law, convert moles of H2into volume.

(b) Molar mass of CCl2F2 = 12.011 + 2  18.9984 + 2  35.4527 = 120.913 g/mol





69. Lime, CaO, is produced by heating calcium carbonate, CaCO3; carbon dioxide is the other

product.

(a) Outline the steps necessary to answer the following question: What volume of carbon dioxide at 875 K and 0.966 atm is produced by the decomposition of 1 ton (1.000 × 103 kg) of calcium carbonate?

(b) Answer the question.

Solution

(a) Balance the equation. Determine the grams of CO2 produced and the number of moles. From the ideal gas law, determine the volume of gas.

(b) 

mass CO2 = 

mol CO2 = 



71. Calculate the volume of oxygen required to burn 12.00 L of ethane gas, C2H6, to produce carbon dioxide and water, if the volumes of C2H6 and O2 are measured under the same conditions of temperature and pressure.

Solution



From the balanced equation, we see that 2 mol of C2H6 requires 7 mol of O2 to burn completely. Gay-Lussac’s law states that gases react in simple proportions by volume. As the number of liters is proportional to the number of moles,



*V*(O2) =  = 42.00 L

73. Consider the following questions:

(a) What is the total volume of the CO2(*g*) and H2O(*g*) at 600 °C and 0.888 atm produced by the combustion of 1.00 L of C2H6(*g*) measured at STP?

(b) What is the partial pressure of H2O in the product gases?

Solution

(a) The scheme to solve this problem is:



(b) First, calculate the mol H2O produced:



Second, calculate the pressure of H2O:



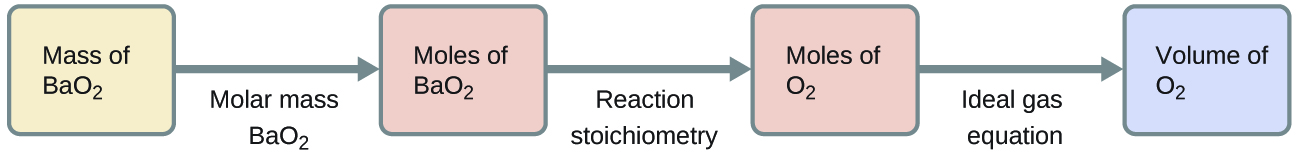
75. What volume of oxygen at 423.0 K and a pressure of 127.4 kPa is produced by the decomposition of 129.7 g of BaO2 to BaO and O2?

Solution

First, we must write a balanced equation to establish the stoichiometry of the reaction:



We are given the mass of BaO2 that decomposes, so the scheme for solving this problem will be:

****

Mass (BaO2) = 137.33 + 2(15.9994) = 169.33 g/mol





77. Ethanol, C2H5OH, is produced industrially from ethylene, C2H4, by the following sequence of reactions:





What volume of ethylene at STP is required to produce 1.000 metric ton (1000 kg) of ethanol if the overall yield of ethanol is 90.1%?

Solution

At 90.1% conversion, a 1.000 × 106 g final yield would require a = 1.1099 × 106 g theoretical yield.

3C2H4 produces 3C2H5OH, giving a 1:1 ratio:



*V*(C2H4) = 22.4 L/mol× 2.409 × 104 mol = 5.40 × 105 L

79. A sample of a compound of xenon and fluorine was confined in a bulb with a pressure of 18 torr. Hydrogen was added to the bulb until the pressure was 72 torr. Passage of an electric spark through the mixture produced Xe and HF. After the HF was removed by reaction with solid KOH, the final pressure of xenon and unreacted hydrogen in the bulb was 36 torr. What is the empirical formula of the xenon fluoride in the original sample? (Note: Xenon fluorides contain only one xenon atom per molecule.)

Solution

The reaction is:



Immediately after the H2 is added (before the reaction):



After the reaction:



And the partial pressure of unreacted H2 is:



The pressure of H2 that reacts is:

48 torr – 24 torr = 24 torr

The number of moles of gas is proportional to the partial pressures. The reaction used 24 torr of XeFx and 24 torr of H2 so:

 = 1 and *x* = 2

This resource file is copyright 2015, Rice University. All Rights Reserved.