GasesKey

Wednesday, October 19, 2016 7:34 AM

Gas Laws and Chemical Reactions

1. 1 gram of CH_4 is added to a 1L flask and pressurized to 4 atm. What temperature is the flask at?

$\frac{1\,g\,CH_4}{} \left| \frac{mol}{16.05\,g} \right. = 0.0623\ mol$

$PV = nRT$

$T = \frac{PV}{nR} = \frac{(4\,atm)(1L)}{(0.0623\,mol)(0.08206\,\frac{L\,atm}{mol\,K})}$

$T = 782.4\ K$

2. 10 moles of an ideal gas is added to gas cylinder at 300 K. 15% of the gas is released from the cylinder while holding the pressure and volume constant. What is the new temperature of the flask?

$\rightarrow 10(0.15) = 1.5$ moles lost

Strategy: Identify constants and variables. Derive a gas law that relates the variables. Determine the initial and final variables. Solve for T_2.

$PV = nRT$ $\quad nT = \frac{PV}{R}$ $\quad$ so... $n_1T_1 = n_2T_2$

$(10\,mol)(300K) = 8.5\,mol\,T_2$

$T_2 = 352.9\ K$

3. Inside a house where the room temperature is **25 °C**, a child is handed a **2 L** birthday balloon containing helium – this, of course, makes little Bobby really happy! When Bobby walks outside to the frigid Siberian winter day, the balloon **loses 10% of its volume** – Bobby cries. Stupid gas laws made a kid cry on his birthday. What is the temperature outside? Assume that the pressure is the same inside and outside. Report your answer in °C.

298.15K

$2(0.1) = 0.2\,L$ lost

$1.8\,L$ left

$PV = nRT$

$\frac{V}{\cancel{P}} = \frac{\cancel{nR}T}{P}$

$\frac{V_1}{T_1} = \frac{V_2}{T_2}$

$\frac{2\,L}{298.15K} = \frac{1.8\,L}{T_2}$

$T_2 = 268.3\ K = \boxed{-4.8\,°C}$

4. 100 grams of a **noble gas** is added to a 10 L flask at 300 K. The pressure of this flask is 2.94 atm. What is this gas? Hint: the only way to identify a gas is by determining the molecular weight.

g / mol $\qquad$ 100 g $\qquad$ mol

Strategy: What do you need to calculate a MW? Do you know any of these values? What else do you need to know? Use the information given to calculate moles. Calculate the MW.

$PV = nRT$

$n = \frac{PV}{RT} = \frac{(10\,L)(2.94\,atm)}{(0.08206\,\frac{L\,atm}{mol\cdot K})(300K)} = 1.194\ mol$

$MW = \frac{100g}{1.194\,mol} = 83.73\ g/mol$

Krypton

5. 4 liters of N_2O_4 (g) decomposes to nitrogen and oxygen gas. If this decomposition occurs at **STP** (so constant temperature and pressure!), determine the **total volume** of gas that is produced.

Strategy: Start with a balanced reaction. Calculate moles of the reactant. Determine the moles of each product that is formed. Calculate volume for each product. Calculate total volume.

$N_2O_4(g) \rightarrow N_2(g) + 2O_2(g)$

$n = \frac{PV}{RT} = \frac{(1\,atm)(4\,L)}{(0.08206)(298.5\,K)} = 0.16\,mol$

$\frac{0.163\,mol\,N_2O_4 \mid 1\,N_2}{\mid 1\,N_2O_4} = 0.163\,mol\,N_2$

$\frac{0.163\,mol\,N_2O_4 \mid 2\,O_2}{\mid 1\,N_2O_4} = 0.327\,mol\,O_2$

0.49 mol of gas

$V = \frac{(0.49)R(298)}{1\,atm} = 12\,L$

using shortcut: (constant T + P)

$\frac{4\,L\,N_2O_4 \mid 1\,N_2}{\mid 1\,N_2O_4} = 4\,L\,N_2$

$\frac{4\,L\,N_2O_4 \mid 2\,O_2}{\mid 1\,N_2O_4} = 8\,L\,O_2$

Total 12 L

6. 5 grams of solid phosphorus trichloride is added to a 4 L reaction flask that contains chlorine gas at STP. Solid phosphorus pentachloride is produced.
   a. Calculate the mass of product that is formed.

Strategy: Start with a balanced reaction. Calculate moles of each reactant. Calculate moles of product – remember to think about limiting reactants. Convert moles of product to mass.

$PCl_3(s) + Cl_2(g) \rightarrow PCl_5(s)$

PCl_3 is LR

$\frac{5\,g\,PCl_3 \mid mol \mid 1\,PCl_5}{\mid 137.32\,g \mid 1\,PCl_3} = \frac{0.036\,mol\,PCl_5 \mid 208.22\,g}{\mid mol} = \boxed{7.495\,g\,PCl_5}$

Cl_2: $n = \frac{PV}{RT} = \frac{(1)(4)}{(.08206)(298.5)} = \frac{0.163\,mol\,Cl_2 \mid 1\,PCl_5}{\mid 1\,Cl_2} = 0.163\,mol$

   b. Assuming that the volume and temperature do not change, what is the pressure in the flask after the reaction?

Strategy: Only gases contribute to pressure, so find the moles of the gas left over when the reaction is complete. Convert to pressure.

start: 0.163 mol Cl_2

used: $\frac{0.036\,mol\,PCl_5 \mid 1\,mol\,Cl_2}{\mid 1\,mol\,PCl_5} = 0.036\,mol\,Cl_2$ used

left over: $0.163 - 0.036 = 0.127\,mol$

$P = \frac{nRT}{V} = \frac{(0.127)(.08206)(298.15\,K)}{4\,L}$

$P = 0.78\,atm$

$C_5H_{12}(s) + 8O_2(g) \rightarrow 5CO_2(g) + 6H_2O(g)$

7. 1 gram of C_5H_{12} is combusted in a 2.5 L reaction flask at 400 K.
   a. How many moles of O_2 are needed to react with C_5H_{12}?

$$\frac{1\,g\,C_5H_{12}}{} \left| \frac{mol}{72.17\,g} \right| \frac{8\,O_2}{1\,C_5H_{12}} = 0.1108\text{ mol } O_2$$

   b. What pressure of O_2 is needed to react with all of the C_5H_{12}? Remember this is in a 2.5 L flask at 400 K.

$$P = \frac{nRT}{V} = \frac{(0.1108)(0.08206)(400K)}{2.5L} = 1.455\text{ atm}$$

   c. Assuming that all of the reactants are consumed:
      i. What is the partial pressure of CO_2 in the flask after the reaction? *Strategy: remember that partial pressure is a fancy way of saying "pressure of CO_2." So if you can determine the moles of CO_2 produced, you can calculate the pressure from CO_2.*

$$\frac{0.1108\text{ mol } O_2}{} \left| \frac{5CO_2}{8O_2} \right. = 0.06925\text{ mol } CO_2$$

$$P = \frac{nRT}{V} = \frac{(0.06925)(0.08206)(400K)}{2.5L}$$

      ii. What is the partial pressure of H_2O in the flask after the reaction? $P = 0.91$ atm

$$\frac{0.1108\text{ mol } O_2}{} \left| \frac{6\,H_2O}{8\,O_2} \right. = 0.0831\text{ mol } H_2O$$

$$P = \frac{(0.0831)(0.08206)(400K)}{2.5L} = 1.091\text{ atm}$$

      iii. What is the partial pressure of O_2 in the flask after the reaction?

$\emptyset \rightarrow$ all O_2 is used

      iv. What is the total pressure in the flask?

$$P_{Tot} = P_{O_2} + P_{CO_2} + P_{H_2O} = 0 + 0.91 + 1.091 = \boxed{2\text{ atm}}$$

8. .8 grams of glucose ($C_6H_{12}O_6$) is combusted in a 2.6 L reaction chamber at pressurized to 3 atm with oxygen at 400 K. Determine the **total pressure** in the flask after the reaction is complete.

$$C_6H_{12}O_6\,(s) + 6O_2\,(g) \rightarrow 6CO_2\,(g) + 6H_2O\,(g)$$

$$0.8\,g\ C_6H_{12}O_6 \times \frac{mol}{180.18\,g} \times \frac{6\ CO_2}{1\ C_6H_{12}O_6} = \boxed{0.0266\ mol\ CO_2}$$

$C_6H_{12}O_6 = L.R.$

$$n_{O_2} = \frac{PV}{RT} = \frac{(3atm)(2.6L)}{(0.08206)(400K)} = 0.2376\ mol\ O_2 \times \frac{6\ CO_2}{6\ O_2} = 0.2376\ mol\ CO_2$$

$$H_2O:\quad 0.0266\ mol\ CO_2 \times \frac{6\ H_2O}{6\ CO_2} = \boxed{0.0266\ mol\ H_2O}$$

O_2: start with 0.2376 mol

used: $0.0266\ mol\ CO_2 \times \frac{6\ O_2}{6\ CO_2} = 0.0266\ mol$

left: $0.2376 - 0.0266 = \boxed{0.211\ mol\ O_2}$

$$P_{O_2} = \frac{nRT}{V} = \frac{(0.211)(.08206)(400K)}{2.6} = 2.6\ atm$$

$$P_{CO_2} = \frac{(0.0266)(0.08206)(400K)}{2.6} = 0.336\ atm$$

$$P_{H_2O} = \frac{(0.0266)(0.08206)(400)}{2.6} = 0.336\ atm$$

$\boxed{P_{tot} = 3.27\ atm}$