Supplemental Questions
for
Gases

1. Is N_2 (boiling point 77.4 K) or Cl_2 (boiling point 238.6 K) more likely to behave as an ideal gas at STP? Why?

N_2 is more likely to behave as an ideal gas at STP (273.15 K and 1 atm) than Cl_2 because under these conditions N_2 is well above its boiling point, but Cl_2 is not. Gases behave ideally when they are at low pressures and at high temperatures relative to their boiling points.

2. Sodium azide, NaN_3, decomposes explosively when heated according to the balanced chemical reaction given below. If 0.377 g NaN_3 are allowed to decompose in a closed 10.00-L flask that originally contained 1.00 atm of air at 24.2 °C, what will be the total pressure of gas in the flask when the reaction is over? The molar mass of NaN_3 is 65.01 g/mole.

$$2\ NaN_3\ (s) \rightarrow 2\ Na\ (s) + 3\ N_2\ (g)$$

Find moles of N_2 formed.

$$0.377\ g\ NaN_3\left(\frac{1\ mole\ NaN_3}{65.01\ g\ NaN_3}\right)\left(\frac{3\ mole\ N_2}{2\ mole\ NaN_3}\right) = 8.69_{87}\ x\ 10^{-3}\ mole\ N_2$$

Find pressure in the flask due to N_2.

$$T = 273.15 + 24.2 = 297.3_5\ K$$

$$pV = nRT$$
$$p = \frac{nRT}{V} = \frac{\left(8.69_{87} x 10^{-3}\ moleN_2\right)\left(0.082057\ L\cdot atm\cdot K^{-1}\cdot mole^{-1}\right)\left(297.3_5\ K\right)}{10.00\ L}$$
$$p = 0.0212\ atm$$

The total pressure is, according to Dalton's Law of Partial Pressure, the sum of the partial pressure of all the gases present. The Na does not affect the pressure because it is a solid.

$$p_{total} = p_{N_2} + p_{atmosphere} = 0.0212\ atm + 1.00\ atm = 1.02\ atm$$

The final pressure in the flask is 1.02 atm.

3. To a 2.50-L flask containing 3.00 g NiO is added gaseous ClF_3 until the pressure of the gas is 250.0 mm Hg at 20.0 °C. The flask is heated and the following reaction takes place.

$$6\ NiO\ (s) + 4\ ClF_3\ (g) \rightarrow 6\ NiF_2\ (s) + 2\ Cl_2\ (g) + 3\ O_2\ (g)$$

What is the pressure in the flask, in mm Hg, when the reaction is finished and the temperature of the flask has returned to 20.0 °C? The molar mass of NiO is 74.6928 g/mole.

Calculate moles of each reactant (volume of NiF_2 is deemed to make an insignificant contribution to the volume).

$$3.00\ g\ NiO\left(\frac{1\ mole}{74.6928\ g}\right) = 0.0402\ mole\ NiO$$

$$n = \frac{pV}{RT} = \frac{\left(\frac{250.0}{760.0}\ atm\right)(2.50\ L)}{\left(0.082057\frac{L\ atm}{K\ mole}\right)(293.2\ K)} = 0.0342\ mole\ ClF_3$$

Need to calculate limiting reagent. If all of the NiO reacts,

$$0.0402\ mole\ NiO\left(\frac{2\ mole\ Cl_2}{6\ mole\ NiO}\right) = 0.0134\ mole\ Cl_2$$

$$0.0342\ mole\ ClF_3\left(\frac{2\ mole\ Cl_2}{4\ mole\ ClF_3}\right) = 0.0171\ mole\ Cl_2$$

NiO is the limiting reagent.

Calculate moles of O_2 formed and moles of ClF_3 that reacted.

$$0.0402\ mole\ NiO\left(\frac{3\ mole\ O_2}{6\ mole\ NiO}\right) = 0.0201\ mole\ O_2$$

$$0.0402\ mole\ NiO\left(\frac{4\ mole\ ClF_3}{6\ mole\ NiO}\right) = 0.0268\ mole\ ClF_3$$

There are 0.0074 mole ClF_3 remaining when the reaction is over (0.0342 mole ClF_3 initially, 0.0268 mole ClF_3 reacted, leaving 0.0074 mole ClF_3).

To find total pressure at the end use Dalton's Law.

$$p_{total} = p_{ClF_3} + p_{O_2} + p_{Cl_2}$$

Substitute using the Ideal Gas Law to get

$$p_{total} = \left(n_{ClF_3} + n_{O_2} + n_{Cl_2}\right)\frac{RT}{V}$$

$$p_{total} = \left(0.0074 + 0.0201 + 0.0134\right) mole \frac{\left(0.082057 \frac{L \cdot atm}{K \cdot mole}\right)\left(293.2\ K\right)}{2.50\ L}$$

$$p_{total} = 0.395\ atm \left(\frac{760\ mmHg}{1\ atm}\right) = 300.\ mm\ Hg$$

The final pressure at 20.0 °C is 300. mm Hg.

4. A 1.00-g sample of $CaCO_3$ is placed into an empty 5.00-L vessel and is heated to 1000.0 K. Some of the $CaCO_3$ decomposes to CaO and CO_2. If the pressure in the flask due to CO_2 is 4.00×10^{-2} atm at 1000.0 K, what percentage of the $CaCO_3$ reacted? The molar mass of $CaCO_3$ is 100.09 g/mole.

First find moles of CO_2 formed.

$$PV = nRT$$

$$n = \frac{PV}{RT} = \frac{(4.00\ x\ 10^{-2}\ atm)(5.00\ L)}{(0.082057 \frac{L \cdot atm}{K \cdot mole})(1000.0\ K)} = 2.44\ x\ 10^{-3}\ mole\ CO_2$$

Now from balanced chemical equation determine the number of grams of $CaCO_3$ that reacted using stoichiometry.

$CaCO_3$ (s) → CaO (s) + CO_2 (g)

$$2.44\ x\ 10^{-3}\ mole\ CO_2 \left(\frac{1\ mole\ CaCO_3}{1\ mole\ CO_2}\right)\left(\frac{100.09\ g}{1\ mole\ CaCO_3}\right) = 0.244\ g\ CaCO_3$$

The percentage that reacted is then

$$\frac{0.244\ g\ reacted}{1.00\ g\ total} x100 = 24.4\%\ reacted$$

So 24.4% of the $CaCO_3$ reacted.

5a. Nitroglycerin, $C_3H_5(NO_3)_3$, decomposes upon heating or when shocked according to the following balanced chemical reaction.

$$4\ C_3H_5(NO_3)_3\ (l) \rightarrow 6\ N_2\ (g) + O_2\ (g) + 12\ CO_2\ (g) + 10\ H_2O\ (g)$$

If 1.03 g of nitroglycerin is decomposed completely in an empty 1.00-L sealed vessel, what will be the total pressure in the vessel if the temperature is 100.0 °C? You are given the molar mass of $C_3H_5(NO_3)_3$ is 227.08 g/mole.

When the reaction is over, the vessel will be filled with N_2, O_2, CO_2 and H_2O in the gaseous state. The total pressure is given by Dalton's Law of Partial Pressure.

$$P_{total} = P_{N_2} + P_{O_2} + P_{CO_2} + P_{H_2O}$$

Assuming that all the gases are ideal, we can rewrite this in terms of moles of each gas present (this just simplifies the math).

$$P_{total} = \left(n_{N_2} + n_{O_2} + n_{CO_2} + n_{H_2O}\right)\frac{RT}{V}$$

Find moles of nitroglycerin, and then by stoichiometry from balanced chemical equation, find moles of each gas.

$$1.03\ g\ nitroglycerin\left(\frac{1\ mole}{227.08\ g}\right) = 4.53_6\ x10^{-3}\ mole\ nitroglycerin$$

$$4.53_6\ x10^{-3}\ mole\ nitroglycerin\left(\frac{6\ mole\ N_2}{4\ mole\ nitroglycerin}\right) = 6.80_4 x10^{-3}\ mole\ N_2$$

$$4.53_6\ x10^{-3}\ mole\ nitroglycerin\left(\frac{1\ mole\ O_2}{4\ mole\ nitroglycerin}\right) = 1.13_{25} x10^{-3}\ mole\ O_2$$

$$4.53_6\ x10^{-3}\ mole\ nitroglycerin\left(\frac{12\ mole\ CO_2}{4\ mole\ nitroglycerin}\right) = 1.36_{08} x10^{-2}\ mole\ CO_2$$

$$4.53_6\ x10^{-3}\ mole\ nitroglycerin\left(\frac{10\ mole\ H_2O}{4\ mole\ nitroglycerin}\right) = 1.13_4 x10^{-2}\ mole\ H_2O$$

Find the total pressure.

$$P_{total} = \left(6.80_4 x10^{-3} + 1.13_{25} x10^{-3} + 136_{08} x10^{-2} + 1.13_4 x10^{-2}\ mole\right) \times \frac{(0.082057\ L \cdot atm \cdot K^{-1} \cdot mole^{-1})(373.2\ K)}{1.00\ L}$$

$P_{total} = 1.01\ atm$

The total pressure in the flask is 1.01 atm.

Note that the vessel was empty except for 1.03 g of nitroglycerin, and this was enough to raise the pressure from 0 to about 1 atm!

b. Give the two reasons why the pressure in this vessel will fall as the temperature is lowered.

The Ideal Gas Law shows P and T are directly correlated. As T falls so does P.

As the temperature falls, the water condenses. The number of moles of gas decreases so the pressure in the vessel falls (again Ideal Gas Law tells us that T and P are directly correlated). We can also see this relationship in the equation we derived above.