Announcements

1.) For this week: October 22-26: Group 1 is doing the Acid/Base lab and Group 2 is doing the Strong Acids & Redox tutorial.

2.) You are not responsible for any of the historical tidbits I throw your way. They are for your amusement only.

3.) Tentative Date of X-mas exam: Monday December 17, 2007 9 am – 12:00 pm.

More problems to consider

Example: How much volume is occupied by one mole of gas at standard temperature and pressure (abbreviated STP, and defined as 0^oC and 1 atm)?

$$PV = nRT$$

$$\therefore \quad V = \frac{nRT}{P}$$

$$\therefore \quad V = \frac{(1\,\text{mol})(8.314\,\text{kPa L mol}^{-1}\,\text{K}^{-1})(273\,\text{K})}{101.3\,\text{kPa}}$$

$$\boxed{V = 22.4\,\text{L}}$$

Note: this result is independent of the gas identity! A mole of any gas molecule will occupy 22.4 L at STP. This is known as the molar volume of a gas

Example *where n is not constant but no chemical reactions are occurring:*
A 25.0 L cylinder of N_2 gas has an internal pressure of 5.0 x 10^3 kPa at 20°C. A valve is opened and the gas is released until the internal pressure becomes 2.0 x 10^3 kPa. How many moles of nitrogen gas were released during this constant temperature process?

To solve this problem, calculate the number of moles in the cylinder before and after the valve was opened, and then find that difference.

$$PV = nRT$$

$$\text{n before} = \frac{PV}{RT} = \frac{(5 \times 10^3 \, \text{kPa})(25 \, \text{L})}{(8.314 \, \text{kPa L mol}^{-1} \, \text{K}^{-1})(293 \, \text{K})} = 51.3 \text{ mol}$$

$$\text{n after} = \frac{PV}{RT} = \frac{(2 \times 10^3 \, \text{kPa})(25 \, \text{L})}{(8.314 \, \text{kPa L mol}^{-1} \, \text{K}^{-1})(293 \, \text{K})} = 20.5 \text{ mol}$$

$$\therefore \quad \Delta n = \text{n before} - \text{n after} \quad = 51.3 - 20.5 = \boxed{30.8 \text{ mol}}$$

Example: A 250 mL flask, open to the atmosphere, contains 0.0110 mol of air at 0 °C *(the initial state)*. On heating, part of the air escapes.
How much air remains in the flask at 100 °C at 100 °C *(the final state)*?

First realize that P and V are constant. Only T and n change.

The initial state:

$$n_1 T_1 = \frac{PV}{R}$$

The final state:

$$n_2 T_2 = \frac{PV}{R}$$

$\therefore \quad n_1 T_1 = n_2 T_2$ We know everything except for n_2

$\therefore \quad (0.0110 \, \text{mol})(273 \, \text{K}) = n_2 (373 \, \text{K})$

$\therefore \quad \boxed{n_2 = 0.00805 \, \text{mol}}$

Note:

Approximately 100 - (0.008/0.011) x 100 = 27% of the molecules escape when the temperature is raised by 100 degrees. This makes the hot air less dense (fewer molecules in the same volume, which is why a hot air balloon rises!)

Example: A sample of acetone (nail polish remover) is placed in a 3.00 L flask and vaporized by heating to 95 °C and 1.02 atm. Under these conditions, the vapour filling the flask weighs 5.87 g, and it is assumed there is no air in the flask.

a) What is the density of the gas vapour?

$$d = \frac{5.87g}{3.00L} = 1.96\,g/L$$

b) Calculate the molar mass of acetone.

We now know that 1 L contains 1.96 g of gas. How many moles is this?

$$PV = nRT$$

1.02 atm = 103.3 kPa; T = 95 °C = 368 K

$$\therefore \quad n = \frac{PV}{RT} \quad = \frac{(103.3\,\text{kPa})(1\,\text{L})}{(8.314\,\text{kPa}\,\text{L}\,\text{mol}^{-1}\,\text{K}^{-1})(368\,\text{K})}$$

$$\therefore \quad n = 0.0337\,\text{mol in 1L of gas}$$

This is 1.96 g, so the molar mass = 1.96 g/0.0337 mol

$$= 58\ \text{g mol}^{-1}$$

Note: if the question had not asked for the density, you could have found the number of mol in 3 L, and then divided 5.87 g by that number of mol to get the same answer.

Example: The minor planet of Xebon has an atmosphere composed solely of Xe. If the pressure at the surface were 1.01 kPa and the temperature 20 °C, what would be the density of the atmosphere at the surface in g L^{-1}?

$M(Xe) = 131.1$ g mol^{-1}

This can be solved using the density equation or as here, from scratch.

First find the number of moles in 1 L of gas:

$$PV = nRT$$

$$\therefore n = \frac{PV}{RT} = \frac{(1.01\,\text{kPa})(1\,\text{L})}{(8.314\,\text{kPa L mol}^{-1}\,\text{K}^{-1})(293\,\text{K})} = 4.15 \times 10^{-4}\,\text{mol}$$

$$(4.15 \times 10^{-4}\,\text{mol}) \times 131.3\,\text{g mol}^{-1} = 0.054\,\text{g}$$

$$\therefore \boxed{d = 0.054 \text{ g/L}}$$

Example: At what temperature will the density of CO_2 be 2.00 g/L at 1.00 atm pressure?

Use either the derived equation or from scratch using the Ideal Gas Equation and some logic

$M(CO_2)$ = 44 g/mol.   Therefore, the number of moles in 1 L:

$$= n = \frac{2.00\,\text{g}}{44.0\,\text{g mol}^{-1}} = 0.0455\,\text{mol}$$

Plug into PV = nRT and solve for the temperature T

$$T = \frac{PV}{nR}$$

$$\therefore T = \frac{(101.3\,\text{kPa})(1\,\text{L})}{(0.0455\,\text{mol})(8.314\,\text{kPa L mol}^{-1}\,\text{K}^{-1})} \qquad = \boxed{268\,\text{K}}$$

An alternate approach is to realize that because the density is 2.00 g/L and M = 44 g/mol, there are therefore 22 L per mol = molar volume of the gas.

$$\therefore \text{T} = \frac{(101.3\,\text{kPa})(22\,\text{L})}{(1\,\text{mol})(8.314\,\text{kPa}\,\text{L}\,\text{mol}^{-1}\,\text{K}^{-1})} = 268\,\text{K}$$

There are many ways of approaching gas problems through the Ideal Gas Equation. Pick your favorite.

Partial Pressures

Each gas in a mixture is unaffected by other gases
(that is, each gas acts as if it were in a vacuum)

Therefore, the ideal gas applies to each single gas in a mixture as well as
their total, and the **total gas pressure** is a **sum of the individual gas pressures**

that is: $$P_{TOT} = P_1 + P_2 + \cdots + P_n = \sum_{i=1}^{n} P_i$$ = **Dalton's Law of Partial Pressures**

The **individual** pressure P_i of gas I in a **mixture of gases** is termed the **partial pressure of gas i**

For a mixture of gases, the number of moles of gas i is $$n_i = \frac{P_i V}{RT}$$

and the total number of moles of gas is $$n_{TOT} = \frac{P_{TOT} V}{RT}$$

The **mole fraction** of gas i is X_i

$$X_i = \frac{n_i}{n_{TOT}} = \frac{P_i V}{RT} \div \frac{P_{TOT} V}{RT} = \frac{P_i V}{RT} \cdot \frac{RT}{P_{TOT} V} = \frac{P_i}{P_{TOT}}$$

$$\therefore \quad \boxed{\frac{n_i}{n_{TOT}} = \frac{P_i}{P_{TOT}} = X_i}$$ where X_i = mole fraction of i in a mixture

Therefore, for any gas in a mixture its mole fraction (n_i/n_{TOT}) is the same as its pressure fraction (P_i/P_{TOT})

Note: $$\sum_i X_i = 1$$

Also since $$P_i = \frac{n_i}{n_{TOT}} \cdot P_{TOT}$$

$$\therefore \quad \boxed{P_i = X_i P_{TOT}}$$ i.e., the partial pressure of any gas i in a mixture is equal to its mole fraction x the total pressure

Who was Dalton?

• John Dalton was an English chemist born in 1766, and died in 1844

• He is credited with developing modern atomic theory at a time when the idea of atoms was in doubt.

• He also studied color blindness which is known as Daltonism in his honor

• The validity of his experimental work was recently reproduced using equipment fashioned from his notebooks here at Western by Professor Mel Usselman in the Chemistry department.