**Chemistry 534**

**Pretest 1.2** **Show all work. Use loose leaf. Solutions in blue**

1. **If a closed vessel (constant volume)'s temperature goes from $0^0$C to 546 K, what will happen to its pressure? It was originally under 20.00 kPa.**

$$\frac{P_1}{T_1} = \frac{P_2}{T_2}$$

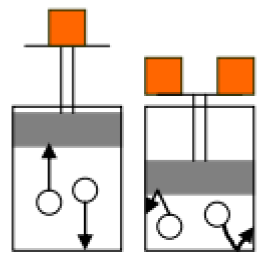
**(constant volume and moles)**

$$\frac{20.0}{0 + 273} = \frac{P_2}{546}$$

**$P_2$ = 40.0 kPa**

2. Use diagrams of molecules in cylinders (with pistons) to show that if pressure is halved, volume doubles.

In the original cylinder, the molecules are crowded: four collisions result. With half the pressure, the molecules have just as much kinetic energy, but they are covering more room in the same instant because there are less collisions between the molecules and between the walls of the container.



3. Consider the first step in the industrial production of nitric acid:

$4\ NH_{3(g)} + 5\ O_{2(g)} \rightarrow 4\ NO_{(g)} + 6\ H_2O_{(g)}$

a) What is the most oxygen (in liters at STP) that could react with 68.0 g of $NH_{3(g)}$?

68.0g/(17.0g/mole) = 4.00 moles of $NH_{3(g)}$.

The ratio of $NH_{3(g)}$ to $O_{2(g)}$ is 4 to 5, so to completely react with 4.00 moles of $NH_{3(g)}$, 5.00 moles of oxygen will be needed

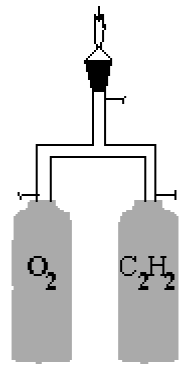
At STP, 5.00 moles of oxygen occupy 5.00 *22.4 = 112 L

b) In burning ammonia, the STP equivalent of 224 L of oxygen were consumed. How many grams of NO must be contended with at STP?

224 L/22.4 L = 10.0 moles of oxygen at STP.

The ratio of oxygen to NO is 5 to 4, so 8.00 moles of NO will be created.

8.00 moles (14.0+16.0 g/mole) = 240. g or 2.40 X $10^2$ g of NO




4. Acetylene reacts with oxygen according to the following equation.

$$\mathbf{2\ C_2H_{2(g)} + 5\ O_{2(g)} \rightarrow 4\ CO_{2(g)} + 2\ H_2O_{(g)}}$$

In an *oxyacetylene blow torch*, steel cylinders containing acetylene and oxygen are connected through hoses that join together, and then connect to a nozzle. The mixture is combusted and produces a flame capable of cutting through metal.

When a steel cylinder of oxygen with a volume of 14.5 L was used to supply oxygen to an oxyacetylene torch, the pressure in the oxygen cylinder changed from $2.080 \times 10^3$ kPa to $2.010 \times 10^3$ kPa. The temperature of both cylinders was 22.0°C at the times of both pressure readings.

What mass of acetylene was combusted from the other cylinder?

First find $n_1$ for oxygen using PV = $n$RT:

$n_1 = P_1V/RT$ = moles oxygen originally present.

$n_2 = P_2V/RT$ = final moles oxygen present.

Amount of oxygen consumed = $n_1 - n_2 = P_1V/RT - P_2V/RT = V/RT\ (P_1 - P_2)$ =14.5/(8.31*[22+273])(2080-2010) = 0.414 moles $O_2$

The ratio of $C_2H_2$ to $O_2$ is 2 to 5, so the amount of acetylene that reacted with 0.414 moles $O_2$ is (2/5)* 0.414 = 0.1656 moles of $C_2H_2$. The mass of 0.1656 moles of $C_2H_2$.= 0.1656 *(2 * 12.011+2 *1.00797) g/mole = 4.31 g.

5. Use $\frac{P_1V_1}{n_1T_1} = \frac{P_2V_2}{n_2T_2}$ to come up with 4 linear relationships between two variables.

Mention the constants in each case.

- P vs T at constant n and constant volume
- P vs n at constant T and V
- V vs n at constant P and T
- V vs T at constant n and constant pressure

Only P vs V and T vs n are *not* linear.

6. What is the density of argon at –50.0° C and 200.0 kPa?

PV = nRT

n/V = P/RT = 200.0/[(8.31*(-50.0 + 273)) = 0.108 moles/L = 0.108 moles*39.948(g/mole)/L = 4.31 g/L.

7. A student wants to triple the pressure of an ideal gas, while decreasing the volume by a factor of 0.80 and increasing the temperature from 200.0 K to 250.0 K. If there were 2.0 moles of gas originally in the gas tank, should he remove gas? Add gas? Explain.

$P_2 = 3P_1$; $V_2 = 0.8V_1$;

Substitute into:

$$\frac{P_1V_1}{n_1T_1} = \frac{P_2V_2}{n_2T_2}$$

$n_2$ = moles. So he should add 3.84-2.00 = 1.84 more moles of gas to meet those conditions. **With sig figs, the answer is rounded to 1.8 more moles needed**.

8. Using the idea that each Pa of pressure is exactly 1 $N/m^2$ and that

pressure in Pa = (9.8 N/kg)**m*ass /area, find the minimum pressure of a tire , knowing that four tires have to support a 2001 kg van ? The area of contact for each tire is 205 $cm^2$. Express your answer in psi's, knowing that exactly one psi =6.89475729 kPa.

**Each of the 4 tires will support 2001kg/4 = 500.25 kg. This is a weight of 500.25 kg*9.8 N/kg = 4902.450 N.**

**Each tire must support a minimum pressure of 4902.450 N/(205 $cm^2$* $m^2/100^2 cm^2$) = 239143.9024 Pa. If you are bedazzled and baffled by this step, it's because you're forgetting that although there are 100 cm in 1m, there are $100^2$ $cm^2$ in 1 $m^2$.**

**239143.9024 Pa = 239.143 kPa**

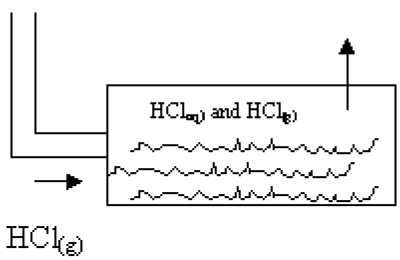
**239.143 kPa (1psi/6.895 kPa) = 35 psi with SF**

9. Hydrogen chloride (HCl) can form from the following reaction:

$2\ NaCl + H_2SO_4 \rightarrow Na_2SO_4 + 2\ HCl_{(g)}$

As the gas forms, it is cooled to 20.0 °C, and it first passes through a 0.20 L tank of water, and whatever does not dissolve escapes into a lab that holds 125 000 L of air at 101.3 kPa.

Will there be enough HCl to kill someone if 400.0 g of $H_2SO_4$react?




Lethal dose of HCl is 0.0018 g /L of air.

At 20.0 °C, the solubility of HCl in water is 720 g/L.

400.0 g/ ([**98.0791**] g/mole) = 4.078moles of $H_2SO_4$

Ratio of $H_2SO_4$ to HCl is 1:2 so , 4.078.*2 = 8.1566 moles of HCl will be released

8.1566 moles of HCl will be released(36.5 g/mole) = 297.6940 g of HCl

But some will get stuck in water:

0.20L *720g/L= 144 g get stuck, leaving us with 297.6940 g – 144 = 153.6940 g HCl

Concentration = 153.6940 g HCl/ 125000L = 0.0012g/L < 0.0018 g/L ; not lethal, but close enough to cause health problems

10 If 2.0 L of HBr gas are mixed with 6.0 L of $NH_3$ gas, what is the most $NH_4Br$ gas that can form, if all gases are measured under the same conditions? Show why.

$HBr + NH_3 \rightarrow NH_4Br$

2.0 L. By Avogadro's principle, there are an equal number of molecules in 2.0 L of each reactant. They react with a 1:1 ratio, so they will produce 2.0 L of $NH_4Br$ product. There will be 4.0 L of unreacted $NH_3$.

**11.**

Given: $2\ NH_{3(g)} + 4\ O_{3(g)} \rightarrow NH_4NO_3 + 4\ O_{2(g)} + H_2O_{(l)}$

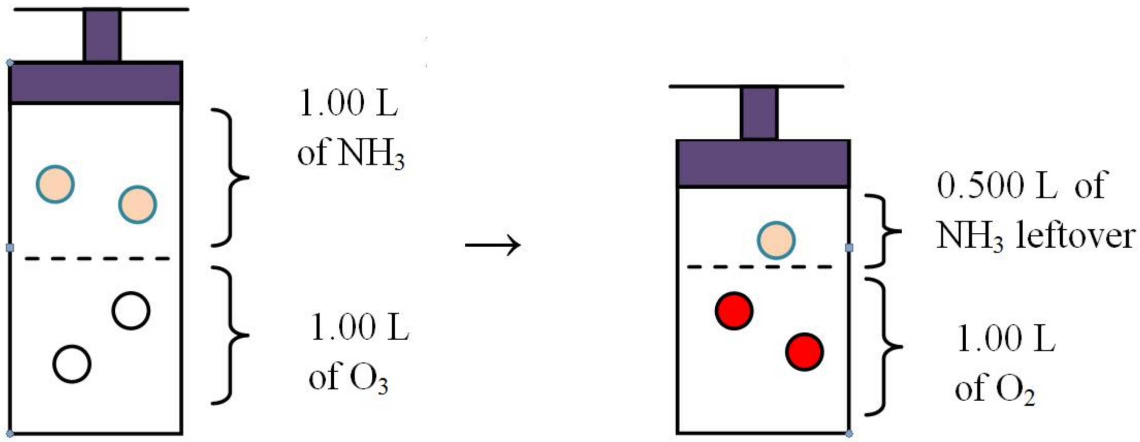
If all three gases' volumes are measured at the same pressure and temperature, and if we mix 1.00 L of ammonia ($NH_{3(g)}$) with 1.00 L of ozone($O_{3(g)}$) what is the greatest volume of oxygen($O_{2(g)}$) that will be produced? Will any of the reactants be in excess?

Answer

$NH_3$ and $O_3$ react in a 1: 2 ratio(2:4=1:2). That means that 1.00 L of ammonia ($NH_{3(g)}$) cannot all react with 1.00 L of ozone because with need twice as much. But the entire 1.00 L of ozone could react with 0.500 L of $NH_{3(g)}$. Since 0.500 L of $NH_{3(g)}$ will produce twice as much oxygen(ratio is (2:4) then 2*0.500 L = 1.00 L of oxygen is the most that can be made, and there will be 1.00 – 0.500 L = 0.500L of $NH_{3(g)}$ left over. The volumes can be treated like moles because of Avogadro's Law

$2\ NH_{3(g)} + 4\ O_{3(g)} \rightarrow NH_4NO_3 + 4\ O_{2(g)} + H_2O_{(l)}$

Is like $1\ NH_{3(g)} + 2\ O_{3(g)} \rightarrow 0.5NH_4NO_3 + 2\ O_{2(g)} + 0.5H_2O_{(l)}$




12 Click here for extra problem and solution.