1. You have four identical 1.00 L **unbreakable** container**s** filled with gases. The molar masses of the gases are given in curly brackets {}. Assume all gases are ideal.

| | | | |
|---|---|---|---|
| **Flask A:** | 1.0 mol He | {4.00 g/mol} | 100 K |
| **Flask B:** | 0.40 mol CO | {28.01 g/mol} | 400 K |
| **Flask C:** | 1.0 atm $Cl_2$ | {70.91 g/mol} | 298 K |
| **Flask D:** | 1.0 atm $NO_2$ | {46.01 g/mol} | 273 K |

(a) Which flask has the greatest pressure?

Flasks C and D are both at 1 atm pressure. Determine the pressure of flasks A and B:

$$P_A = \frac{nRT}{V} = \frac{(1.0\ \text{mol})\left(0.08206\,\frac{\text{L·atm}}{\text{mol·K}}\right)(100\ \text{K})}{1.0\ \text{L}} = 8.2\ \text{atm} \qquad P_B = \frac{nRT}{V} = \frac{(0.40\ \text{mol})\left(0.08206\,\frac{\text{L·atm}}{\text{mol·K}}\right)(400\ \text{K})}{1.0\ \text{L}} = 13\ \text{atm}$$

**Answer: Flask B**

(b) Which flask is at STP?

Recall that STP is T = 0 °C = 273 K and P = 1.0 atm.

**Answer: Flask D**

(c) In which flask would the contents take up the smallest volume if brought to STP?

Recall that at STP, the volume of 1 mol of any ideal gas is 22.4 L.
Therefore, the volume is directly proportional to the number of moles of each gas ($V_{STP}$ α n).
Flask A has 1.0 mol of gas and flask B has 0.40 mol of gas.
Determine the number of moles in flasks C and D:

$$n_C = \frac{PV}{RT} = \frac{(1.0\ \text{atm})(1.0\ \text{L})}{\left(0.08206\,\frac{\text{L·atm}}{\text{mol·K}}\right)(298\ \text{K})} = 0.041\ \text{mol}\ Cl_2 \qquad n_D = \frac{PV}{RT} = \frac{(1.0\ \text{atm})(1.0\ \text{L})}{\left(0.08206\,\frac{\text{L·atm}}{\text{mol·K}}\right)(273\ \text{K})} = 0.045\ \text{mol}\ NO_2$$

**Answer: Flask C**

(d) In which flask will diffusion of the gas be fastest?

$$\text{rate}_{\text{diff}} \propto \sqrt{T/_M} \propto T/_M$$

| | *Flask A* | *Flask B* | *Flask C* | *Flask D* |
|---|---|---|---|---|
| T/M | 25 | 14 | 4.2 | 5.9 |

**Answer: Flask A**

2. A gas tank has a volume of 32.0 L, a temperature of 27.0 °C, a pressure of 1.10×10⁵ Torr, and contains 748 g of an unknown gas. What is the identity of the gas?

Determine the number of moles of unknown gas (X) using the ideal gas law:

$$n_X = \frac{PV}{RT} = \frac{\left(2125\ \text{psi} \times \frac{1\ \text{atm}}{14.7\ \text{psi}}\right)(32\ \text{L})}{\left(0.08206\,\frac{\text{L}\cdot\text{atm}}{\text{mol}\cdot\text{K}}\right)(27 + 273.15\ \text{K})} = 18_{7.8}\ \text{mol X}$$

We can determine the molar mass of gas X ($MM_X$)now:

$$MM_X = \frac{\text{mass}_X}{n_X} = \frac{748\ \text{g X}}{18_{7.8}\ \text{mol X}} = 4.0\,\frac{\text{g}}{\text{mol}}$$

The gas is helium (He).

3. An average person consumes 31 g of $O_2$ per hour through the following balanced chemical equation (cellular respiration):

$$6\ O_2\ (g) + C_6H_{12}O_6\ (aq) \rightarrow 36\ ATP\ (aq) + 6\ CO_2\ (g) + 6\ H_2O\ (l)$$

What volume of $O_2$ (at STP) is consumed in 30.0 minutes?

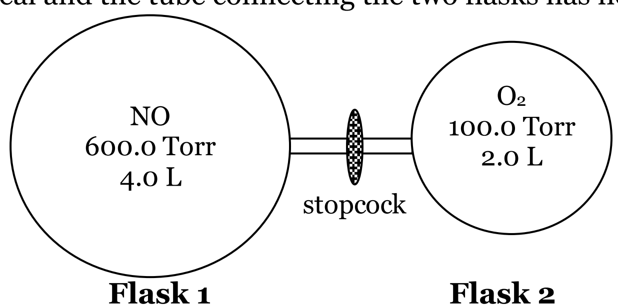
Determine number of moles of $O_2$ required:

$$30\text{ min} \times \frac{1\text{ hr}}{60\text{ min}} \times \frac{31\text{ g O}_2}{1\text{ hr}} \times \frac{1\text{ mol O}_2}{32.00\text{ g O}_2} = 0.48_4\text{ mol O}_2$$

Determine volume of this much $O_2$ using the ideal gas law:

$$V_{O_2} = \frac{n_{O_2}RT}{P} = \frac{(0.48_4\text{ mol O}_2)\left(0.08206\ \frac{\text{L}\cdot\text{atm}}{\text{mol}\cdot\text{K}}\right)(273.15\text{ K})}{1.00\text{ atm}} = 11\text{ L}$$

4. Consider the following arrangement of two flasks at 400.0 K, connected by a stopcock. Assume that the gases are ideal and the tube connecting the two flasks has negligible volume.




(a) Assuming <u>no chemical reaction</u> between NO and $O_2$, calculate the partial pressures of NO and $O_2$ if the stopcock were opened. Assume no temperature changes.

Use pressure-volume gas relationship to find new partial pressures: $P_1V_1 = P_2V_2$

$$P_{O_2} = \frac{(100.0\text{ Torr})(2.0\text{ L})}{6.0\text{ L}} = 33\text{ Torr}$$

$$P_{NO} = \frac{(600.0\text{ Torr})(4.0\text{ L})}{6.0\text{ L}} = 4.0 \times 10^2\text{ Torr}$$

(b) Now assume that when the stopcock is opened, the NO and $O_2$ react to form $NO_2$. Assume that there is no temperature changes.

Calculate the partial pressures of NO and $O_2$ after the reaction is complete.

$$2NO\ (g) + O_2\ (g) \rightarrow 2NO_2\ (g)$$

First, determine the number of moles of each reactant through the ideal gas law:

$$n_{NO} = \frac{PV}{RT} = \frac{\left(600.0\text{ Torr} \times \frac{1\text{ atm}}{760\text{ Torr}}\right)(4.0\text{ L})}{\left(0.08206\ \frac{\text{L}\cdot\text{atm}}{\text{mol}\cdot\text{K}}\right)(400.0\text{ K})} = 0.096_2\text{ mol NO}$$

$$n_{O_2} = \frac{PV}{RT} = \frac{\left(100.0\text{ Torr} \times \frac{1\text{ atm}}{760\text{ Torr}}\right)(2.0\text{ L})}{\left(0.08206\ \frac{\text{L}\cdot\text{atm}}{\text{mol}\cdot\text{K}}\right)(400.0\text{ K})} = 0.0080_2\text{ mol O}_2$$

Second, determine that the limiting reactant is $O_2$:

| | |
|---|---|
| $0.096_2\text{ mol NO} \times \frac{1\text{ mol O}_2}{2\text{ mol NO}} = 0.048_1\text{ mol O}_2$ <br> → We have less $O_2$ than we need. | $0.096_2\text{ mol NO} \times \frac{2\text{ mol NO}_2}{2\text{ mol NO}} = 0.096_2\text{ mol NO}_2$ <br> $0.0080_2\text{ mol O}_2 \times \frac{2\text{ mol NO}_2}{1\text{ mol O}_2} = 0.016_0\text{ mol NO}_2$ <br> → $O_2$ produces less $NO_2$ product. |

Since $O_2$ is limiting, none will be left after the reaction. However, some NO gas will be left over:

$$0.0080_2 \text{ mol } O_2 \times \frac{2 \text{ mol NO}}{1 \text{ mol } O_2} = 0.016_0 \text{ mol NO reacted}$$

$$n_{\text{NO,left over}} = 0.096_2 \text{ mol} - 0.016_0 \text{ mol} = 0.080_2 \text{ mol NO left over}$$

Third, find the partial pressures of NO gas leftover:

$$\mathbf{P_{NO}} = \frac{\text{nRT}}{\text{V}} = \frac{(0.080_2 \text{ mol})\left(0.08206 \frac{\text{L} \cdot \text{atm}}{\text{mol} \cdot \text{K}}\right)(400.0 \text{ K})}{(2.0 \text{ L} + 4.0 \text{ L})} = 0.43_9 \text{ atm} \times \frac{760 \text{ Torr}}{1 \text{ atm}} = \mathbf{330\ Torr}$$