Gases Tutorial 1

e.g. Convert 352 Torr to kPa

352 Torr x (1 atm/760 Torr) x (101.325 kPa/1 atm) = 46.9 kPa

e.g. Convert 0.255 atm to bar

0.255 atm x (101325 Pa/atm) x (1 bar / 100000 Pa) = 0.258 bar

e.g. Convert 10^{-7} mbar to Pa

10^{-7} mbar x (1 bar / 1000 mbar) x (100000 Pa / 1 bar) = 10^{-5} Pa

e.g. Which contains the most molecules: 1.00 L $H_{2(g)}$, 1.00 L $O_{2(g)}$ or 1.00 L air (all at 298 K, 1 bar pressure)?

According to the ideal gas law, n = pV/RT. It makes no difference what the gas is - all three samples have the same p, V and T. Thus all have the same number of moles, or molecules.

e.g. Oxygen is commonly sold in 49.0 L containers at 150 bar. What volume in L would this gas occupy if it were allowed to expand to 1.00 bar at the same temperature?

Since n and T are constant, we use Boyle's Law, $p_1V_1 = p_2V_2$ and solve for V_2: $V_2 = p_1V_1 / p_2$ = 150 bar (49.0 L) / 1.00 bar = 7,350 L

e.g. If the gas in the above example is at 25 ^{o}C in the container but cools to 15 ^{o}C when expanded, what is its new volume?

Here, n is constant, but p, V and T are not. Thus we combine Charles' and Boyle's Laws: $p_1V_1 / T_1 = p_2V_2 / T_2$ and solve for V_2:

$V_2 = p_1V_1\ T_2 / p_2\ T_2$ = 150 bar (49.0 L) (15+273.15) K / 1.00 bar (25+273.15) K = 7,100 L (smaller, as expected, because the temperature is lower)

e.g. A compressed gas cylinder at 13.7 MPa and 23^{o}C is in a room where a fire raises the temperature to 950^{o}C. Find the pressure at the new temperature.

According to the laws of Boyle and Charles: $P_1V_1/T_1 = P_2V_2/T_2$
The volume of the cylinder does not change, so $P_1/T_1 = P_2/T_2$, or $P_2 = P_1(T_2/T_1)$ = 13.7 MPa (1223/296) = 56.6 MPa

e.g. Interstellar space is essentially H atoms at 100 K, 1 atom/cm^3. What is the gas pressure in Pa?

If we express p in Pa and V in m^3, the value of R is 8.314 Pa m^3 K^{-1} mol^{-1} = 8.314 J K^{-1} mol^{-1} (Note that 1 J = 1 kg m^2 s^{-2} and that 1 Pa = 1 N/m^2 = 1 kg m s^{-2} m^{-2} = 1 kg s^{-2} m^{-1}. Thus 1 J = 1 Pa m^3.)

p = nRT/V. Note that 1 atom/cm^3 has the same dimensions as n/V. Thus p = (n/V)RT. But n must be in moles and V in m^3.

1 atom/cm^3 x (1 mol / 6.02 x 10^{23} atoms/mol) x (10^6 cm^3 /m^3) = 1.66 x 10^{-18} mol/m^3

p = (n/V)RT = 1.66 x 10^{-18} mol/m^3 (8.314 J/(mol K)) (100 K) = 1.38 x 10^{-18} Pa

e.g. A steel cylinder is filled with 150 mol Argon gas at 25°C and 7.5 MPa. Some of the gas is used, then the pressure is 1.2 MPa at 17°C. What mass of gas remains in the cylinder?

The volume of the cylinder is V = nRT/p = 150 mol(8.314 J K^{-1} mol^{-1})(298 K) / 7.5 x 10^6 Pa = 0.0496 m^3

After some gas is removed, the number of moles remaining is n = pV/RT = 1.2 x 10^6 Pa(0.0496 m^3)/(8.314 J K^{-1} mol^{-1})(290 K) = 24.7 mol

24.7 mol x 39.9 g/mol = 986 g Ar left.

e.g. what volume of $O_{2(g)}$ at 22°C and 763 Torr is required for the combustion of 7.50 L of ethane measured at STP?

$C_2H_{6(g)} + 3.5\ O_{2(g)} \rightarrow 2\ CO_{2(g)} + 3\ H_2O_{(g)}$

7.50 L $C_2H_{6(g)}$ x (1 mol / 22.4 L) x (3.5 mol $O_{2(g)}$ / 1 mol $C_2H_{6(g)}$) = 1.17 mol $O_{2(g)}$ required.

V = nRT/p

= 1.17 mol ((0.082 L atm $K^{-1}mol^{-1}$)(295 K))/(763/760 atm)

= 28.2 L (makes sense, since the number of moles of O_2 in the balanced reaction is 3.5 times that of ethane. Since the molar volumes of the gas are similar (the conditions for measuring the O_2 are close to STP), then the volumes occupied by the gases should be roughly 3.5 : 1.)

e.g. (Problem 2.30 in the text)
Ethylene oxide is produced via the reaction $C_2H_{4(g)} + ½\ O_{2(g)} \rightarrow C_2H_4O_{(g)}$. The yield of this reaction is typically only 65%. Calculate the mass of ethylene oxide produced from 50,000 L of a mixture containing $C_2H_{4(g)}$ and $O_{2(g)}$ in a 1:1 mole ratio.

> The problem gives information about the amounts of both starting materials, so this is a limiting reactant situation. We must calculate the number of moles of each species, construct a table of amounts, and use the results to determine the partial pressure. Use the

ideal gas equation to determine the initial amounts of each gas. The reactor initially contains the gases in 1:1 mole ratio, so

$$p_{\mathrm{i}} = \frac{1.00 \text{ bar}}{2} = 0.500 \text{ bar for each gas}$$

$$n_{\mathrm{i}} = \frac{p_{\mathrm{i}}V}{RT} = \frac{(0.500 \text{ bar})(5.00\times10^{4} \text{ L})}{(0.08314 \text{ L bar mol}^{-1} \text{ K}^{-1})(280+273.15 \text{ K})} = 5.44 \times 10^{2} \text{ mol}$$

Since both gases have the same initial amount, the limiting reactant will be the one with the higher stoichiometric coefficient: C_2H_4 is limiting. Here is the complete amounts table:

Reaction:	**2 C_2H_4 (*g*)**	**+ O_2 (*g*)**	**→ 2 C_2H_4O (*g*)**
Initial amount (10^2 mol)	5.44	5.44	0.00
Change (10^2 mol)	–5.44	(5.44)	+5.44
Final amount (10^2 mol)	0	2.72	5.44

Obtain the mass of product formed from the final amount in the table:

$$m_{C_2H_4O} = nM = 5.44\times10^{2} \text{ mol}\left(\frac{44.05 \text{ g}}{1 \text{ mol}}\right)\left(\frac{10^{-3} \text{ kg}}{1 \text{ g}}\right) = 24.0 \text{ kg}$$

e.g. A hot air balloon is filled to a volume of 4.0 x 10^3 m^3 at 745 Torr and 21°C. The air is then heated to 62°C, causing the balloon to expand to a volume of 4.2 x 10^3 m^3. Find the ratio of the number of moles of air in the heated balloon to the unheated balloon.

$P_1V_1/n_1T_1 = P_2V_2/n_2T_2$

Thus, $n_2/n_1 = P_2V_2T_1/P_1V_1T_2$. Note that the internal pressure of the balloon remains constant, since it is open to the atmosphere at the bottom where the hot gases are blown in. Thus, $P_2 = P_1$, and $n_2/n_1 = V_2T_1/V_1T_2$.

n_2/n_1 = 4.2 x 10^3 m^3 (294 K) / (4.0 x 10^3 m^3 (335 K)) = 0.92