

SOLUTIONS TO TOPIC C HOMEWORK

1) For parts a, b and c, the combined gas law is the most convenient way to solve.

$$\frac{P_1 V_1}{T_1} = \frac{P_2 V_2}{T_2}$$

a) The volume is constant, so we can eliminate it from both sides of the combined gas law, giving us $P_1/T_1 = P_2/T_2$. Then we substitute the pressures and initial temperature, and solve. Be sure to use Kelvin temperatures!

$$T_1 = 18 + 273 = 291 \text{ K}$$

$$T_2 = -61 + 273 = 212 \text{ K}$$

(Since the temperatures are only given to the nearest whole number, there is no need to use 273.15 as the “conversion factor” between Celsius and Kelvin.)

$$\frac{739 \text{ torr}}{291 \text{ K}} = \frac{P_2}{212 \text{ K}}$$

$$\begin{aligned} P_2 &= 538.38 \text{ torr} \\ &= \mathbf{538 \text{ torr}} \text{ (to 3 sig figs)} \end{aligned}$$

b) Here the temperature is constant, so we can eliminate the temperature from both sides of the combined gas law, giving us $P_1 V_1 = P_2 V_2$. Then we substitute the initial pressure and the volumes. Note that the initial pressure here is the final pressure from part a. Keep a couple of extra digits to guard against rounding errors.

$$538.38 \text{ torr} \times 455 \text{ mL} = P_2 \times 623 \text{ mL}$$

$$\begin{aligned} P &= 393.20 \text{ torr} \\ &= \mathbf{393 \text{ torr}} \text{ (to 3 sig figs)} \end{aligned}$$

c) Here the pressure is constant, so we eliminate the pressure from both sides of the combined gas law, giving us $V_1/T_1 = V_2/T_2$. Then we substitute the temperatures and the initial volume. The initial volume here is 623 mL (the final volume from part b) and the initial temperature is 212 K (the final temperature from part a, since the temperature did not change in part b).

$$T_2 = 83 + 273 = 356 \text{ K}$$

$$\frac{623 \text{ mL}}{212 \text{ K}} = \frac{V_2}{356 \text{ K}}$$

$$\begin{aligned} V &= 1046.2 \text{ mL} \\ &= \mathbf{1.05 \times 10^3 \text{ mL}} \text{ (to 3 sig figs)} \end{aligned}$$

d) To calculate the mass of the oxygen, we must first calculate the number of moles using the ideal gas law ($PV = nRT$). We can use any of the conditions in the previous sections of this problem (since changing P , V or T does not change the number of moles), but it is safest to use the initial conditions, just in case we made an arithmetic error in the earlier steps.

When using the ideal gas law, we must be sure that the units of P , V and T match the units of R . There are two values of R that are commonly used in the ideal gas law:

$$0.08206 \text{ L}\cdot\text{atm}/\text{mol}\cdot\text{K} \quad \text{and} \quad 62.36 \text{ L}\cdot\text{torr}/\text{mol}\cdot\text{K}$$

In this problem, the pressure is given in torr, so the second value of R is more convenient. Both values require the volume to be in liters ($455 \text{ mL} = 0.455 \text{ L}$) and the temperature to be in kelvins ($18^\circ\text{C} = 291 \text{ K}$).

$$(739 \text{ torr})(0.455 \text{ L}) = n(62.36 \text{ L}\cdot\text{torr}/\text{mol}\cdot\text{K})(291 \text{ K})$$

$$n = 0.018529 \text{ mol}$$

Finally, we convert from moles to grams. Pure oxygen is a diatomic gas (O_2), so one mole of oxygen weighs 32.00 grams.

$$0.018529 \text{ mol} \times \frac{32.00 \text{ g}}{1 \text{ mol}} = \mathbf{0.593 \text{ g}} \text{ (to 3 sig figs)}$$

2) We need to do two things before we can get the molecular formula:

- Use the percent composition to determine the empirical formula
- Use the density information to determine the molecular weight

Once we know these, we can combine them to get the molecular formula. Let's work on the empirical formula first. You learned how to do this in topic A...

100 g of the compound contains 88.8 g of carbon and 11.2 g of hydrogen.

$$88.8 \text{ g C} \times \frac{1 \text{ mol C}}{12.01 \text{ g C}} = 7.3938 \text{ mol C}$$

$$11.2 \text{ g H} \times \frac{1 \text{ mol H}}{1.008 \text{ g H}} = 11.111 \text{ mol H}$$

Divide both of these by the smaller value:

$$\frac{7.3938 \text{ mol C}}{7.3938 \text{ mol C}} = 1 \text{ (C)} \qquad \frac{11.111 \text{ mol H}}{7.3938 \text{ mol C}} = 1.50 \text{ (H)}$$

To convert these into whole numbers, multiply both of them by 2, giving 2 carbon atoms and 3 hydrogen atoms. The empirical formula of the compound is therefore C_2H_3 .

Next, we calculate the molar mass. You can memorize the formula in the book that relates density to mass, but it is easier on your memory to just use the ideal gas law to calculate the number of moles, then divide grams by moles. The density of the gas is 2.32 g/L, which means

that one liter of the gas weighs 2.32 grams. How many moles are there in one liter of a gas at 150°C (423 K) and 565 torr?

$$565 \text{ torr} \times 1 \text{ L} = n \times 62.36 \text{ L} \cdot \text{torr/mol} \cdot \text{K} \times 423 \text{ K}$$

$$n = 0.021419 \text{ mol}$$

We now know that 0.021419 moles of this gas weigh 2.32 grams, so the molar mass must be:

$$\frac{2.32 \text{ g}}{0.021419 \text{ mol}} = 108 \text{ g/mol}$$

Now we can determine the molecular formula. If the compound were C_2H_3 , the molar mass would be 27 g/mol. The actual molar mass must be a multiple of this. In fact, 108 g/mol is four times 27 g/mol:

$$108 \div 27 = 4$$

Therefore, the molecular formula must be four times the empirical formula, or C_8H_{12} .

3) To calculate the mass of a gas, we need to first determine the number of moles. The trick here is working out the volume. One milliliter is the same as one cubic centimeter, so we can calculate the volume of the room (in mL) by finding the volume of the room in cm^3 . A lab room is a rectangular solid, so its volume equals length x width x height. Converting these measurements to centimeters, we get:

$$\text{width} = 8.9 \text{ m} = 890 \text{ cm} \quad \text{length} = 7.5 \text{ m} = 750 \text{ cm} \quad \text{height} = 4.6 \text{ m} = 460 \text{ cm}$$

$$\text{volume} = 890 \text{ cm} \times 750 \text{ cm} \times 460 \text{ cm} = 3.0705 \times 10^8 \text{ cm}^3$$

$$= 3.0705 \times 10^8 \text{ mL}$$

The ideal gas law requires that the volume be in liters and the temperature be in kelvins. We can leave the pressure in atmospheres if we use the value $R = 0.08206 \text{ L} \cdot \text{atm/mol} \cdot \text{K}$.

$$V = (3.0705 \times 10^8 \text{ mL}) \times \frac{1 \text{ L}}{1000 \text{ mL}} = 3.0705 \times 10^5 \text{ L}$$

$$T = 20 + 273 = 293 \text{ K}$$

$$1 \text{ atm} \times (3.0705 \times 10^5 \text{ L}) = n \times 0.08206 \text{ L} \cdot \text{atm/mol} \cdot \text{K} \times 293 \text{ K}$$

$$n = 1.2771 \times 10^4 \text{ mol}$$

Now we can use the percentages of the various gases to calculate the mass of each gas.

$$\left(1.2771 \times 10^4 \text{ mol air}\right) \times \frac{0.770 \text{ mol N}_2}{1 \text{ mol air}} \times \frac{28.02 \text{ g N}_2}{1 \text{ mol N}_2} = 2.7553 \times 10^5 \text{ g N}_2$$

$$\left(1.2771 \times 10^4 \text{ mol air}\right) \times \frac{0.207 \text{ mol O}_2}{1 \text{ mol air}} \times \frac{32.00 \text{ g O}_2}{1 \text{ mol O}_2} = 8.4592 \times 10^4 \text{ g O}_2$$

$$\left(1.2771 \times 10^4 \text{ mol air}\right) \times \frac{0.014 \text{ mol H}_2\text{O}}{1 \text{ mol air}} \times \frac{18.016 \text{ g H}_2\text{O}}{1 \text{ mol H}_2\text{O}} = 3.2210 \times 10^3 \text{ g H}_2\text{O}$$

$$\left(1.2771 \times 10^4 \text{ mol air}\right) \times \frac{0.009 \text{ mol Ar}}{1 \text{ mol air}} \times \frac{39.95 \text{ g Ar}}{1 \text{ mol Ar}} = 4.5917 \times 10^3 \text{ g Ar}$$

The total mass is the sum of these, which works out to $3.679 \times 10^5 \text{ g}$. This is a very rough approximation, because we have not subtracted the volume of the lab benches and other furniture, so our answer is probably only valid to one significant figure. The air in S-201 weighs roughly $4 \times 10^5 \text{ grams}$ (400 kilograms). Reporting two significant figures would be reasonable also, if somewhat overoptimistic...

4) First, let's work out the number of moles of CrCl_2 .

$$50.0 \text{ mL} \times \frac{1 \text{ L}}{1000 \text{ mL}} \times \frac{0.123 \text{ mol CrCl}_2}{1 \text{ L}} = 0.00615 \text{ mol CrCl}_2$$

Next, we can calculate the number of moles of O_2 that will react with the CrCl_2 .

$$0.00615 \text{ mol CrCl}_2 \times \frac{1 \text{ mol O}_2}{4 \text{ mol CrCl}_2} = 0.0015375 \text{ mol O}_2$$

Now let us work out the number of moles of O_2 that were originally present. Before adding the CrCl_2 solution, we had a volume of 500.0 L (0.5000 L), a temperature of 18°C (291 K), and an oxygen pressure of 156 torr. Using the ideal gas law gives us:

$$156 \text{ torr} \times 0.5000 \text{ L} = n \times 62.36 \text{ L} \cdot \text{torr/mol} \cdot \text{K} \times 291 \text{ K}$$

$$n = 0.0042983 \text{ moles of O}_2$$

Note that this value is valid to three significant figures, because the Kelvin temperature is valid to three sig figs (not two).

The reaction consumed 0.0015375 moles of O_2 , so we are left with:

$$0.0042983 \text{ mol} - 0.0015375 \text{ mol} = 0.0027608 \text{ moles of O}_2$$

Finally, we calculate the pressure of O_2 . The CrCl_2 solution occupies 50.0 mL, so the air in the container occupies $500.0 \text{ mL} - 50.0 \text{ mL} = 450.0 \text{ mL}$ (0.4500 L). Using the ideal gas law gives us:

$$P \times 0.4500 \text{ L} = 0.0027608 \text{ mol} \times 62.36 \text{ L} \cdot \text{torr/mol} \cdot \text{K} \times 291 \text{ K}$$

$$P_{\text{O}_2} = \mathbf{111 \text{ torr}}$$

5) a) This is a stoichiometry problem disguised as a gas law problem. We must calculate the initial moles of each reactant (NH_3 and HCl), then do stoichiometry to determine the moles of product.

To calculate the moles of each reactant, use $PV = nRT$. Remember your units: the volumes must be converted to liters, the temperatures must be converted to kelvins, and we must use the correct R value for pressures in atm.

$$\text{For } \text{NH}_3 : 0.875 \text{ atm} \times 0.635 \text{ L} = n \times 0.08206 \text{ L} \cdot \text{atm/mol} \cdot \text{K} \times 300 \text{ K}$$

$$n = 0.022570 \text{ mol } \text{NH}_3$$

$$\text{For } \text{HCl}: 1.613 \text{ atm} \times 0.429 \text{ L} = n \times 0.08206 \text{ L} \cdot \text{atm/mol} \cdot \text{K} \times 300 \text{ K}$$

$$n = 0.028109 \text{ mol } \text{HCl}$$

Next, we do stoichiometry. The reaction is 1:1, so NH_3 must be the limiting reactant. Here is an ICE table for this reaction:

	NH_3	+	HCl	\rightarrow	NH_4Cl
Initial moles	0.022570		0.028109		0
Change	-0.022570		-0.022570		+0.022570
Ending moles	0		0.005539		0.022570

The molar mass of NH_4Cl is 53.492 g/mol, so the mass of ammonium chloride that will be formed is:

$$0.022570 \text{ mol} \times \frac{53.492 \text{ g}}{1 \text{ mol}} = \mathbf{1.21 \text{ g}}$$

b) The total pressure in the apparatus will be the sum of the pressures of HCl and Ne . (There is no NH_3 left after the reaction, and NH_4Cl is a solid and does not exert any pressure.) We can calculate the pressure of HCl using the ideal gas law. After the valve is opened, the remaining HCl is free to occupy the entire apparatus, so we must use the total volume:

$$635 \text{ mL} + 429 \text{ mL} = 1064 \text{ mL}$$

Using the ideal gas law gives us...

$$P_{\text{HCl}} \times 1.064 \text{ L} = 0.005539 \text{ mol} \times 0.08206 \text{ L} \cdot \text{atm/mol} \cdot \text{K} \times 300 \text{ K}$$

$$P_{\text{HCl}} = 0.128 \text{ atm}$$

The neon did not react, but it too will now occupy the entire container. Since its effective volume increases from 635 mL to 1064 mL, its pressure will decrease. We can calculate the new pressure of Ne using the combined gas law. The temperature is constant, so we eliminate it, giving $P_1V_1 = P_2V_2$.

$$0.893 \text{ atm} \times 635 \text{ mL} = P_2 \times 1064 \text{ mL}$$

$$P_2 = 0.533 \text{ atm} (= P_{\text{Ne}})$$

The total pressure is $0.128 \text{ atm} + 0.533 \text{ atm} = \mathbf{0.661 \text{ atm}}$.

6) a) When a gas is collected over water, the gas sample always contains water vapor. The partial pressure of water equals its vapor pressure, which is 20.5 torr at this temperature. The total pressure of the gases in our sample is 751 torr, so the partial pressure of the acetylene in this sample must be:

$$751 \text{ torr} - 20.5 \text{ torr} = 730.5 \text{ torr}$$

The acetylene occupies the entire volume (261 mL = 0.261 L) and is at a temperature of $22.5^\circ\text{C} = 295.65 \text{ K}$ (using 273.15 to translate Celsius into Kelvin here, since the temperature is given to the nearest tenth of a degree). The ideal gas law now gives us:

$$730.5 \text{ torr} \times 0.261 \text{ L} = n \times 62.36 \text{ L} \cdot \text{torr/mol} \cdot \text{K} \times 295.65 \text{ K}$$

$$n = 0.010341 \text{ mol}$$

$$= \mathbf{0.0103 \text{ moles of } C_2H_2} \quad (3 \text{ sig figs})$$

b) We can use the number of moles of acetylene and the balanced equation to calculate the number of moles of CaC_2 , then convert moles into grams. The molar mass of CaC_2 is 64.10 g/mol. Be sure to keep a couple of extra digits in your moles of C_2H_2 .

$$0.010341 \text{ mol } C_2H_2 \times \frac{1 \text{ mol } CaC_2}{1 \text{ mol } C_2H_2} \times \frac{64.10 \text{ g } CaC_2}{1 \text{ mol } CaC_2} = \mathbf{0.663 \text{ g } CaC_2}$$

7) a) For a single object, the kinetic energy is given by $KE = \frac{1}{2}mv^2$, where m is the mass in kilograms, v is the velocity in meters per second, and KE is the kinetic energy in joules. Substituting in the given values gives us:

$$KE = \frac{1}{2}(625.0 \text{ kg})(18.0 \text{ m/s})^2$$

$$= \mathbf{1.01 \times 10^5 \text{ J}} \quad (\text{or } 101 \text{ kJ})$$

The units work out to $\text{kg} \cdot \text{m}^2/\text{s}^2$; you must remember that this combination is a joule.

b) For a single molecule, we work out the mass starting with the molar mass. The molar mass of CO_2 is 44.01 g/mol, so we have:

$$\frac{44.01 \text{ g}}{1 \text{ mol}} \times \frac{1 \text{ mol}}{6.022 \times 10^{23} \text{ molecules}} \times \frac{1 \text{ kg}}{1000 \text{ g}} \times 1 \text{ molecule} = 7.3082 \times 10^{-26} \text{ kg}$$

$$KE = \frac{1}{2}(7.3082 \times 10^{-26} \text{ kg})(419 \text{ m/s})^2 = \mathbf{6.42 \times 10^{-21} \text{ J}}$$

c) This is similar to part b. The mass of one N_2 molecule (molar mass = 28.02 g/mol) is 4.6529×10^{-26} kg, and its kinetic energy at 511 m/s is **6.07×10^{-21} J**.

d) To convert to a per mole basis, multiply by Avogadro's number.

$$\text{For part b: } 6.42 \times 10^{-21} \text{ J} \times \frac{1 \text{ kJ}}{1000 \text{ J}} \times \frac{6.022 \times 10^{23} \text{ molecules}}{1 \text{ mol}} = \mathbf{3.86 \text{ kJ/mol}}$$

$$\text{For part c: } 6.07 \times 10^{-21} \text{ J} \times \frac{1 \text{ kJ}}{1000 \text{ J}} \times \frac{6.022 \times 10^{23} \text{ molecules}}{1 \text{ mol}} = \mathbf{3.66 \text{ kJ/mol}}$$

8) The easiest way to do this is to do everything “per mole”: express the mass in kg/mol and the kinetic energy in J/mol. We can then use our kinetic energy formula, $\text{KE} = \frac{1}{2}mv^2$, to solve for the velocity.

$$\begin{aligned} \text{KE} &= 3.86 \text{ kJ/mol} = 3860 \text{ J/mol} \\ m &= 4.003 \text{ g/mol} = 0.004003 \text{ kg/mol} \end{aligned}$$

$$3860 \text{ J/mol} = \frac{1}{2}(0.004003 \text{ kg/mol}) \cdot v^2$$

$$v^2 = \frac{2(3860 \text{ J/mol})}{0.004003 \text{ kg/mol}} = 1.9286 \times 10^6 \text{ J/kg}$$

$$v = \sqrt{1.9286 \times 10^6 \text{ J/kg}} = 1.39 \times 10^3 \sqrt{\text{J/kg}}$$

To make sense of the unit, remember that a joule is a $\text{kg} \cdot \text{m}^2/\text{s}^2$, so a J/kg is the same as a m^2/s^2 . Taking the square root gives us m/s, as we expect, so our answer is **1.39×10^3 m/s**.

You can also do this problem based on the energy and mass of a single atom. The mass of a single helium atom is 6.6473×10^{-27} kg (see problem 7 for details on how to calculate the mass of a single atom or molecule from the molar mass), and 3.86 kJ/mol converts to...

$$\frac{3.86 \text{ kJ}}{1 \text{ mol}} \times \frac{1000 \text{ J}}{1 \text{ kJ}} \times \frac{1 \text{ mol}}{6.022 \times 10^{23} \text{ atoms}} \times 1 \text{ atom} = 6.4098 \times 10^{-21} \text{ J}$$

Substituting these values into $\text{KE} = \frac{1}{2}mv^2$ gives us:

$$6.4098 \times 10^{-21} \text{ J} = \frac{1}{2}(6.6473 \times 10^{-27} \text{ kg}) \cdot v^2$$

$$v^2 = 1.9286 \times 10^6 \text{ J/kg}$$

$$v = 1.39 \times 10^3 \text{ m/s}$$

We get the same answer (as we must, since it's the same helium atom!).

9) The average velocity of a gas molecule is...

$$u_{\text{ave}} = \sqrt{\frac{8RT}{\pi M}}$$

where T is the Kelvin temperature (as always!!), M is the molar mass in kg/mol, and R is the gas constant in the form 8.314 J/mol·K.

$$T = 20 + 273 = 293 \text{ K}$$

$$M = 28.02 \text{ g/mol} \times \frac{1 \text{ kg}}{1000 \text{ g}} = 0.02802 \text{ kg/mol}$$

Substituting these values into the equation gives us:

$$\begin{aligned} u_{\text{ave}} &= \sqrt{\frac{8 \times 8.314 \text{ J/mol} \cdot \text{K} \times 293 \text{ K}}{3.14159 \times 0.02802 \text{ kg/mol}}} && \text{(the mol and K units cancel)} \\ &= \sqrt{221386 \text{ J/kg}} \end{aligned}$$

To make sense of the unit, we must recall that a joule is a $\text{kg} \cdot \text{m}^2/\text{sec}^2$. (You can reason this out from the kinetic energy formula $\text{KE} = \frac{1}{2}mv^2$, remembering that m must be in kg and v in m/sec.) Substituting $\text{kg} \cdot \text{m}^2/\text{sec}^2$ in place of J in the square root gives us...

$$\begin{aligned} u_{\text{ave}} &= \sqrt{221386 (\text{kg} \cdot \text{m}^2/\text{sec}^2)/\text{kg}} \\ &= \sqrt{221386 \text{ m}^2/\text{sec}^2} && \text{(the kg units cancel)} \\ &= 471 \text{ m/sec} \end{aligned}$$

To compare this with the speed of sound, we must convert the speed of sound from miles/hr to m/sec. This requires converting miles into meters **and** converting hours into seconds. Here is one possible setup:

$$\frac{768 \text{ miles}}{\text{hr}} \times \frac{1.6093 \text{ km}}{1 \text{ mile}} \times \frac{1000 \text{ m}}{1 \text{ km}} \times \frac{1 \text{ hr}}{60 \text{ min}} \times \frac{1 \text{ min}}{60 \text{ sec}} = \frac{343 \text{ m}}{\text{sec}}$$

The speed of sound is 343 m/sec and the average speed of a nitrogen molecule is 471 m/sec, so **the average speed of nitrogen molecules is the larger of the two.**

10) a) As you reduce the size of the container, the number of molecules per cm^3 increases. This in turn increases the number of times per second that a molecule collides with the wall of the container. Pressure is the sum of the force from these collisions, so the pressure increases.

b) As you reduce the temperature, the kinetic energies of the molecules decrease. Kinetic energy is proportional to mass and to velocity; since the mass of each molecule does not change, the molecular velocities decrease as you decrease the temperature. This has two effects: the molecules hit the walls of the container less often, and each collision has less force. Both of these effects reduce the pressure on the walls of the container, with the second being more important.

c) As you increase the temperature, the fraction of molecules that have very high kinetic energies increases. In order to react, a molecule must have a kinetic energy that is above some minimum value (called the activation energy); only molecules that have more KE than this can react. (Molecules that have less energy simply bounce off one another without reacting.) Increasing the temperature increases the fraction of molecules that have KE's above this minimum value, so a greater number of molecules can react when they collide.

11) a) The rms velocity depends on both the molar mass and the temperature, so it is easiest to simply calculate the rms velocity for each gas.

For N_2 at $25^\circ C$, $T = 298 K$ and $M = 0.02802 \text{ kg/mol}$

$$\begin{aligned}u_{\text{rms}} &= \sqrt{\frac{3 \times 8.314 \text{ J/mol} \cdot \text{K} \times 298 \text{ K}}{0.02802 \text{ kg/mol}}} \\ &= 515 \text{ m/sec}\end{aligned}$$

(See the solution to problem 9 for details about how the units cancel.)

For O_2 at $75^\circ C$, $T = 348 K$ and $M = 0.03200 \text{ kg/mol}$, which gives $u_{\text{rms}} = 521 \text{ m/sec}$

Therefore, the rms velocity of **oxygen at $75^\circ C$** is larger.

b) The kinetic energy distribution depends only on temperature. In this case, the two gases are at the same temperature, so **they have the same most probable kinetic energies**.

c) If the temperature is constant, the velocity distribution is inversely dependent on the molar mass; the smaller the molar mass, the faster the molecules. In this case, hydrogen has a lower molar mass than helium (2.016 g/mol versus 4.003 g/mol), so the average velocity of **hydrogen** is larger.

d) This is similar to part c. The gas with the smaller molar mass will have a larger fraction of "fast molecules" and a smaller fraction of "slow molecules", no matter where we choose the cutoff between "fast" and "slow". CH_4 has a smaller molar mass than NH_3 does (16 g/mol versus 17 g/mol). Therefore, the fraction of **CH_4 molecules** with velocities greater than 500 m/sec is larger.

e) To work this out, we must recognize that the kinetic energy distribution depends only on temperature, not on molar mass. Therefore, the helium and the neon will have **exactly the same fraction of atoms** with kinetic energies in any range we select.

12) This problem is not trivial, because both the molar mass and the temperature are different for each sample of gas. We can work this out by comparing the most probable velocity for each gas sample.

For neon at $0^\circ C$, $T = 273 K$ and $M = 0.02018 \text{ kg/mol}$

$$u_{\text{mp}} = \sqrt{\frac{2 \times 8.314 \text{ J/mol} \cdot \text{K} \times 273 \text{ K}}{0.02018 \text{ kg/mol}}}$$

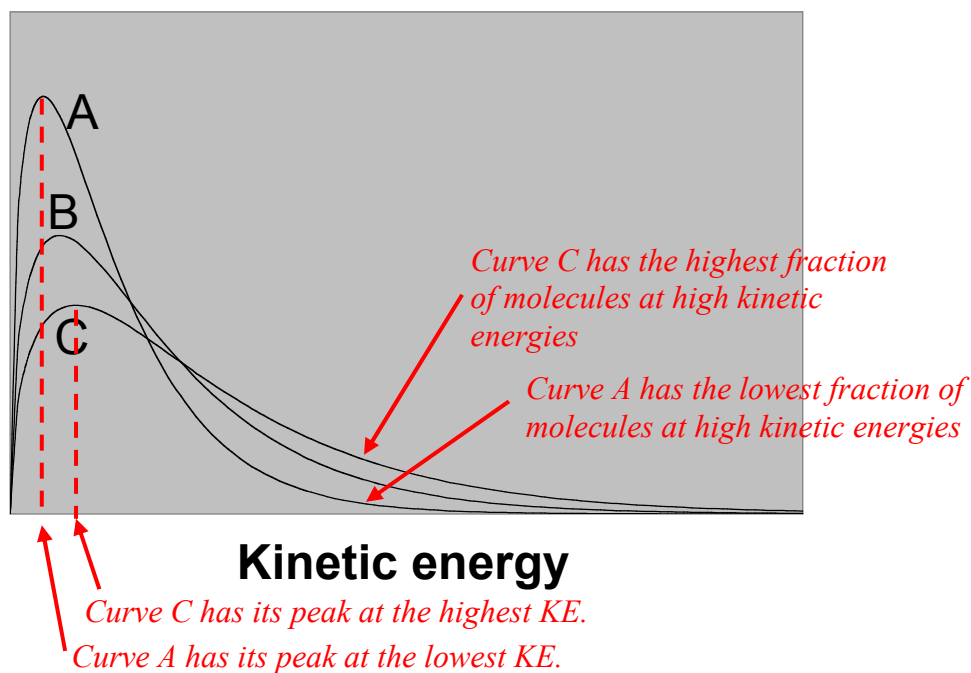
$$= 474 \text{ m/sec}$$

For carbon dioxide at 25°C, $T = 298 \text{ K}$ and $M = 0.04401 \text{ kg/mol}$. Substituting these into the formula gives $u_{\text{mp}} = 336 \text{ m/sec}$.

For hydrogen at 50°C, $T = 323 \text{ K}$ and $M = 0.002016 \text{ kg/mol}$. Substituting these into the formula gives $u_{\text{mp}} = 1632 \text{ m/sec}$.

The most probable velocity is the velocity that corresponds to the peak of the curve (the point that has the largest fraction of molecules). From the graph, we see that curve A has the lowest u_{mp} (its peak is farthest to the left) and curve C has the highest u_{mp} . Therefore, **curve A corresponds to the CO_2 at 25°C, curve B corresponds to the Ne at 0°C, and curve C corresponds to the H_2 at 50°C.**

13) The kinetic energy distribution of a gas depends only on its temperature, not on the molar mass. As you raise the temperature of a gas, the KE distribution shifts toward larger kinetic energies; the peak moves to the right and becomes smaller, and the number of molecules with high KE's increases. Looking at the curves, we see that curve A has its peak at the lowest KE and has the fewest molecules at very high KE's, so curve A corresponds to hydrogen at 50 K. Curve C its peak at the highest KE and has the most molecules at very high KE's, so curve C corresponds to neon at 150 K. Curve B has intermediate values and corresponds to oxygen at 100 K.



14) We will need to use Graham's Law of effusion to calculate the molar mass of the unknown compound.

$$\frac{\text{rate 1}}{\text{rate 2}} = \sqrt{\frac{M_2}{M_1}} \quad \text{or} \quad \frac{\text{rate 1}}{\text{rate 2}} = \frac{\sqrt{M_2}}{\sqrt{M_1}}$$

For the unknown compound, the effusion rate is $12.1 \text{ mL} \div 10.0 \text{ min} = 1.21 \text{ mL/min}$.

For the CO_2 , $M = 44.01 \text{ g/mol}$ and the effusion rate is $11.9 \text{ mL} \div 5.0 \text{ min} = 2.38 \text{ mL/min}$.

Let the unknown compound be “compound 2” and the CO_2 be “compound 1” (so our unknown molar mass is in the numerator of the fraction).

$$\frac{2.38 \text{ mL/min}}{1.21 \text{ mL/min}} = \frac{\sqrt{M_2}}{\sqrt{44.01 \text{ g/mol}}}$$

$$1.9669 = \frac{\sqrt{M_2}}{\sqrt{44.01 \text{ g/mol}}}$$

Square both sides to give M_2 .

$$1.9669^2 = \frac{M_2}{44.01 \text{ g/mol}}$$

$$\begin{aligned} M_2 &= 3.869 \times 44.01 \text{ g/mol} \\ &= 170.3 \text{ g/mol} \end{aligned}$$

The atomic weight of nickel is 58.69 g/mol , so the carbon monoxide molecules that are bonded to Ni must weigh a total of $170.3 \text{ g/mol} - 58.69 \text{ g/mol} = 111.6 \text{ g/mol}$. The molar mass of CO is 28.01 g/mol , so...

$$111.6 \text{ g/mol} \div 28.01 \text{ g/mol} = 3.98 \approx 4$$

Therefore, the unknown compound contains four molecules of CO for each atom of Ni, making its molecular formula **Ni(CO)₄**.

15) a) Using the ideal gas law gives us:

$$150 \text{ atm} \times 253 \text{ L} = n \times 0.08206 \text{ L} \cdot \text{atm/mol} \cdot \text{K} \times 523 \text{ K}$$

$$\begin{aligned} n &= 884.26 \text{ mol} \\ &= \mathbf{884 \text{ moles of O}_2} \text{ (to three sig figs)} \end{aligned}$$

b) The van der Waals equation for gases is

$$\left[P + a \left(\frac{n}{V} \right)^2 \right] (V - nb) = nRT$$

For oxygen, $a = 1.36 \text{ atm}\cdot\text{L}^2/\text{mol}^2$ and $b = 0.0318 \text{ L/mol}$. We already have $n = 884.26 \text{ mol}$, $V = 253 \text{ L}$, and $T = 523 \text{ K}$. Substituting these values into the equation gives us:

$$\left[P + 1.36 \text{ atm}\cdot\text{L}^2/\text{mol}^2 \left(\frac{884.26 \text{ mol}}{253 \text{ L}} \right)^2 \right] (253 \text{ L} - 884.26 \text{ mol} \times 0.0318 \text{ L/mol}) = 884.26 \text{ mol} \times 0.08206 \text{ L}\cdot\text{atm}/\text{mol}\cdot\text{K} \times 523 \text{ K}$$

Doing all of the arithmetic and canceling units gives us:

$$(P + 16.6134 \text{ atm})(224.881 \text{ L}) = 37950.1 \text{ L}\cdot\text{atm}$$

$$P + 16.6134 \text{ atm} = \frac{37950.1 \text{ L}\cdot\text{atm}}{224.881 \text{ L}}$$

$$P + 16.6134 \text{ atm} = 168.756 \text{ atm}$$

$$\begin{aligned} P &= 152.143 \text{ atm} \\ &= \mathbf{152 \text{ atm}} \quad (\text{to 3 sig figs}) \end{aligned}$$

This is slightly above the safety limit for the reactor (150 atm), so the engineer should decrease the maximum number of moles of O_2 accordingly.