Supplemental Activities

Module: States of Matter

Section: Gas Laws & Gas Mixtures – Key

Gas Laws

Activity 1

The purpose of this activity is to practice your understanding of Gas Laws.

1. Test yourself: how many of the gas law equations can you write down?

   Answers here will vary. Here are some of the valid answers:

$$\boxed{\begin{array}{l} PV = k \\ (or\ P_1V_1 = P_2V_2) \\ \\ \dfrac{V}{T} = k \\ (or\ V_1T_2 = V_2T_1) \\ \\ \dfrac{V}{n} = k \\ (or\ V_1n_2 = V_2n_1) \\ \\ \dfrac{PV}{T} = k \\ \left(or\ \dfrac{P_1V_1}{T_1} = \dfrac{P_2V_2}{T_2}\right) \\ \\ PV = nRT \end{array}}$$

2. Torr, mmHg, atm, bar and Pa are all units of gas pressure, which is the ratio of the combined force of the gas particle impacts and the surface area of the gas container.

3. Robert Boyle studied the relationship between pressure and volume of a fixed amount of gas at a constant temperature. Explain Boyle's law with words and an equation.

   Here is the equation for Boyle's law:

$$\boxed{\begin{array}{l} PV = k \\ (or\ P_1V_1 = P_2V_2) \end{array}}$$

   where P is the pressure of the gas, V is the volume of the gas and k is a constant.

Boyle's law states that for a fixed amount of gas at a constant temperature, the pressure and volume of that gas are inversely proportional.

4. Jacques Charles studied the relationship between temperature and volume of a fixed amount of gas at a constant pressure. Explain Charles' law with words and an equation.

   Here is the equation for Boyle's law:

$$\boxed{\begin{array}{l} \dfrac{V}{T} = k \\ (or\ V_1T_2 = V_2T_1) \end{array}}$$

   where V is the volume of the gas, T is the absolute temperature of the gas and k is a constant.

   Charles' law states that for a fixed amount of gas at a constant temperature, the temperature and volume of that gas are directly proportional.

5. Lord Kelvin developed an absolute temperature scale called the Kelvin scale. It defines an absolute zero point at which substances have their minimum value of thermal energy and at which pure substances exist as perfect crystals/lattices.

6. The combined gas law combines which two gas laws? Explain the combined gas law with words and an equation.

   The combined gas law combines Charles' law and Boyle's law.

   Here is the equation for the combined gas law:

$$\boxed{\begin{array}{l} \dfrac{PV}{T} = k \\ \left(or\ \dfrac{P_1V_1}{T_1} = \dfrac{P_2V_2}{T_2}\right) \end{array}}$$

   where P is the pressure of the gas, V is the volume of the gas, T is the absolute temperature of the gas and k is a constant.

   The combined gas law states that the ratio of the product of the pressure and volume of a gas with its absolute temperature is a constant.

7. Boyle's law, Charles' law and the combined gas law can be written as two state laws as well as one state laws. What is meant when we discuss the "state" of a gas sample?

   The state of a gas sample is simply the set of variables that define the conditions of the gas at a particular moment. So the state of the gas may include the measurements for the pressure, volume, temperature and amount of the gas.

8. Assuming a constant molar quantity of gas, how could you produce the following effects?

a. Decrease pressure
You could decrease temperature or increase volume

b. Decrease volume
You could decrease temperature or increase external pressure

c. Increase pressure
You could increase temperature or decrease volume

d. Increase volume
You could increase temperature or decrease external pressure

Activity 2

The purpose of this activity is to practice your mastery of the quantitative nature of gas laws.

1. A gas occupies 11.2 liters at 0.860 atm. What is the pressure when the volume is 15.0 liters? Assume that the temperature and amount of gas remain the same.

   Here we recognize that only values for volume and pressure are given and asked for in the problem. The temperature and amount of gas remain constant. Therefore, we can use Boyle's Law to solve this problem.

$$\boxed{\begin{aligned} &P_1 = 0.860atm \\ &V_1 = 11.2L \\ &V_2 = 15.0L \\ &P_2 = ? \\ &P_1V_1 = P_2V_2 \\ &(0.860atm)(11.2L) = (P_2)(15.0L) \\ &P_2 = \frac{(0.860atm)(11.2L)}{(15.0L)} \\ &P_2 = 0.642atm \end{aligned}}$$

   We observe that as the volume of the gas increased, the pressure of the gas decreased. This observation makes sense according to Boyle's Law, which describes the inversely proportional relationship between volume and pressure of a fixed amount of gas at a constant temperature.

2. You're working on your chemistry homework with a friend. Your friend considers a problem describing a gas that is cooled from 120˚C to 40˚C. Your friend assumes that the volume must have decreased by a factor of three from 1.50L down to 0.50L. Is your friend correct? Why or why not?

   No, your friend is not correct. It is important to remember that calculations with the gas laws must be done in the absolute temperature scale. On the absolute temperature scale a decrease from 120˚C to 40˚C corresponds to a decrease from 393K to 313K. The volume of a gas is directly proportional to the absolute temperature (Kelvin) so the volume would've actually decreased from 1.50L only down to about 1.19L.

3. A gas occupies 60.0 mL at 33˚C. What *change* in volume does this gas experience if it is cooled down to 5.00˚C? Assume that the pressure and amount of gas remain constant.

   Here we recognize that only values for volume and temperature are given and asked for in the problem. The pressure and amount of gas remain constant. Therefore, we can use Charles' Law to solve this problem.

$$
\begin{aligned}
&V_1 = 0.600L \\
&T_1 = 33.0^\circ C = 306.15K \\
&T_2 = 5.00^\circ C = 278.15K \\
&V_2 = ? \\
&\Delta V = ? \\
&V_1 T_2 = V_2 T_1 \\
&(0.600L)(278.15K) = (V_2)(306.15K) \\
&V_2 = \frac{(0.600L)(278.15K)}{(306.15K)} \\
&V_2 = 0.545L \\
&\Delta V = V_2 - V_1 = 0.545L - 0.600L = -0.0549L
\end{aligned}
$$

   We observe that as the temperature of the gas decreased, the volume of the gas decreased. This observation makes sense according to Charles' Law, which describes the directly proportional relationship between volume and temperature of a fixed amount of gas at a constant pressure.

4. A 3250 mL gas sample at 24.5˚C has a pressure of 1825 mmHg. You change the temperature of the gas. The new volume is 4250 mL and the new pressure is 1.50 atm. What is the new temperature?

   Here we recognize that values for volume, pressure temperature are given and asked for in the problem. Therefore, we must use the Combined Gas Law to solve this problem. Note though that the two pressure values are in different units. We *must* convert one of them so that their units match in order to solve the problem. Let's convert to atm.

$$\boxed{\begin{array}{l}
P_1 = 1825 mmHg \times \frac{1 atm}{760 mmHg} = 2.401 atm \\
V_1 = 3250 mL \\
T_1 = 24.5 ^\circ C = 297.65 K \\
P_2 = 1.50 atm \\
V_2 = 4250 mL \\
T_2 = ? \\
\frac{P_1V_1}{T_1} = \frac{P_2V_2}{T_2} \\
\frac{(2.401 atm)(3250 mL)}{(297.65 K)} = \frac{(1.50 atm)(4250 mL)}{T_2} \\
T_2 = \frac{(1.50 atm)(4250 mL)(297.65 K)}{(2.401 atm)(3250 mL)} \\
T_2 = 243.14 K = -30.0 ^\circ C
\end{array}}$$

Activity 3

The purpose of this activity is to investigate Avogadro's law and start developing a small particle model of gases.

1. Amedeo Avogadro studied the relationship between volume and amount/moles of a gas at a constant temperature and at a constant pressure. Explain Avogadro's law with words and an equation.

   Here is the equation for Avogadro's law:

   $$\boxed{\begin{array}{l} \frac{V}{n} = k \\ (or\ V_1 n_2 = V_2 n_1) \end{array}}$$

   where V is the volume of the gas, n is the moles of the gas and k is a constant.

   Avogadro's law states that for a gas at a constant temperature and pressure, the volume of that gas is directly proportional to the amount of gas present.

2. How does a small particle simulator illustrate the microscopic behavior of gases? In your explanation, describe how are gases represented and the observance of the gas laws.

   A small particle simulator illustrates the microscopic behavior of gases by representing the gas particles as small spheres that do not interact with one another other than colliding elastically. The gas laws are obeyed in the small particle simulator. For example, if the pressure and number of particles are held constant while the temperature of the sample is cooled, the volume of the sample lowers. In this way, the small particle simulator illustrates Charles' law.

3. Imagine that you are an argon atom inside a sealed balloon filled with only argon gas. Describe your behavior and how your behavior affects the macroscopic measurements of pressure and volume. Describe what happens macroscopically and microscopically to the pressure and volume if the balloon begins to leak.

   As an argon atom, you bounce around the volume of the balloon. Your impacts with the sides of the balloon, in combination with the impacts of your fellow argon atoms, create an internal pressure that is equal to the external pressure. Thanks to your collisions with the sides of the balloon, the balloon maintains a particular volume, which has been dictated by the pressure outside the balloon and the number of atoms inside the balloon. Unfortunately, when your nice balloon springs a leak, some of your fellow argon atoms begin to leave the balloon. There are fewer atoms to make the collision necessary to maintain an internal pressure that is equal to the external pressure. The balloon begins to shrink and by decreasing its volume, the surface area inside the balloon is also decreasing. Therefore, even with fewer atoms, you and your remaining compatriot atoms can collide with the walls with enough frequency to maintain equilibrium between the external and internal pressure.

Ideal Gas Law

Activity 1

The purpose of this activity is to practice your understanding of the Ideal Gas Law

1. How is the ideal gas law different than the other gas laws previously discussed?

   The ideal gas law is a one state law. The ideal gas law only considers one set of conditions for a gas sample. The other laws are either one or two states laws (they are most often used as two state laws).

2. Explain the ideal gas law with words and an equation.

   Here is the equation for the ideal gas law:

   $$\boxed{PV = nRT}$$

   where P is the pressure of the gas, V is the volume of the gas, n is the moles of the gas, R is the universal gas constant and T is temperature.

   If we rearrange the ideal gas law for R we see that the ratio of pressure times volume over moles times temperature gives a constant value for any ideal gas.

   $$\boxed{R = \frac{PV}{nT}}$$

3. Write down different values for *R* that you could use when pressure is in atm and volume is in liters, when pressure is in torr and volume is in liters, when pressure is in Pa and volume is in m^3, and when you need to calculate joules.

   R is a constant. You can manipulate the units to best address the problem at hand. Or you can convert other values in the problem to match the value of *R* that you want to use. This problem here is simply highlighting the fact that *R* can have many different units:

   - Pressure in atm, volume in L → $R = 0.08206$ Latm $mol^{-1}K^{-1}$
   - Pressure in mmHg, volume in L → $R = 62.36$ Ltorr $mol^{-1}K^{-1}$
   - Pressure in kPa, volume in cm^3 → $R = 8.314$ m^3Pa $mol^{-1}K^{-1}$
   - For joules → $R = 8.314$ J$mol^{-1}K^{-1}$

4. What do we assume about ideal gases?

   Ideal gases are infinitely small, hard spheres that do not interact with each other. They are essentially "blind" to other gas molecules and will bounce off of each other just as they would bounce of a wall.

5. You are scuba diving in a large fish tank. While you are at the bottom of the tank, you release a balloon full of air and watch it as it rises to the surface. What do you notice about the volume of the balloon?

   The pressure at the bottom of the tank is greater than the pressure at the top of the tank, so as it rises, you notice that its volume increases.

ACTIVITY 2

The purpose of this activity is to practice your understanding of the quantitative side of the Ideal Gas Law.

1. Give the different pressure and temperature conditions for both STP and SATP for gases.

   STP: 1 atm and 0˚C (273.15 K)   SATP: 1 bar and 25˚C (298.15 K)

2. What does the term "molar volume" mean? What is the value for molar volume of an ideal gas at STP?

   Molar volume is the volume occupied by one mole. For ideal gases at STP, the molar volume value is 22.4 L.

3. Calculate the moles of gas present in a 910 mL sample at 38˚C and 650 Torr.

$$PV = nRT$$
$$P = 650 Torr$$
$$V = 910 mL = 0.910 L$$
$$R = 62.36 \tfrac{L \cdot Torr}{mol \cdot K}$$
$$T = 38^\circ C = 311.15 K$$
$$n = ?$$
$$n = \frac{PV}{RT} = \frac{(650 Torr)(0.910 L)}{(62.36 \tfrac{L \cdot Torr}{mol \cdot K})(311.15 K)}$$
$$n = 3.05 \times 10^{-2} mol$$

4. How many atoms of argon gas are in a 137 mL container if the pressure in the container is $8.80 \text{ X } 10^{-5}$ mmHg and the temperature is 794 K?

$$PV = nRT$$
$$P = 8.80 \times 10^{-5} mmHg$$
$$V = 137 mL = 0.137 L$$
$$R = 62.36 \tfrac{L \cdot mmHg}{mol \cdot K}$$
$$T = 794 K$$
$$n = ?$$
$$PV = nRT$$
$$n = \frac{PV}{RT} = \frac{(8.80 \times 10^{-5} mmHg)(0.137 L)}{(62.36 \tfrac{L \cdot mmHg}{mol \cdot K})(794 K)}$$
$$n = 2.43 \times 10^{-10} mol$$
$$2.43 \times 10^{-10} mol\ Ar \times \frac{6.022 \times 10^{23} atoms\ Ar}{1 mol\ Ar} = 1.47 \times 10^{14} atoms\ Ar$$

5. What is the temperature of .75 moles of argon in a 18 L container with a pressure of 790 Torr?

$$
\boxed{\begin{aligned}
&PV = nRT\\
&P = 790Torr \times \frac{1atm}{760torr} = 1.04atm\\
&V = 18L\\
&R = 0.082\,\tfrac{L\cdot atm}{mol\cdot K}\\
&T = ?\\
&n = 0.75mol\\
&PV = nRT\\
&T = \frac{PV}{nR} = \frac{1.04atm * 18L}{0.75mol \times 0.082\frac{L\cdot atm}{mol\cdot K}} = 304.4K
\end{aligned}}
$$

6. Is it possible for 1 mole of air in an adult's lungs to be at STP? Explain and prove by use the ideal gas law.

   No it's not possible, you would die. 1 mole of an ideal gas at STP has a volume of 22.4 L. This would be way too much for your lungs, which hold about 6L.
   PV = nRT
   At STP: 1 atm and 273K
   $V = nRT/P$
   $V = \frac{(1\ mole)\left(0.082\frac{Latm}{Kmol}\right)(273K)}{1\ atm}$
   $V = 22.386\ L$

Gas Density

Activity 1

The purpose of this activity is to demonstrate a thorough understanding of the difference between number density and mass density

1. The ratio of the mass of a substance and the volume that the mass occupies is what is considered to be the density of that substance.

2. Mass density is the ratio of mass and volume while number density is the ratio of moles or molecules and volume.

3. As some given pressure and temperature, how could two 500 mL closed containers filled only with ideal gas(es) have the same number density but different mass density? Provide a theoretical example.

   Two containers with the same number density implies that they each contain the same number of particles per unit volume. The implication is proven true when we consider the ideal gas law: the temperature, volume and pressure are identical between the two containers and so the number of moles in each container is identical as well ($PV = nRT$ where R is a constant). However, if one container were filled with a heavier gas, it would have a greater mass density than the container with a lighter gas even though they have the same number of moles per unit volume. For example, argon gas is about four times heavier than neon gas. If one 500 mL container held 2 moles of argon and the other held 2 moles of neon, the argon would have a mass density of about four times that of the helium.

4. Consider two balloons each filled with the same amount of helium gas. Both balloons are at SATP. Then, you take one balloon and carefully place it into an open container of liquid nitrogen (77K). The other balloon remains unchanged. Draw both balloons in their initial and final states and include values for number and mass density in each drawing.

   If both balloons have the same amount of gas and are both at standard ambient temperature and pressure, then they must have the same volume initially. Therefore, their number densities will be identical initially. Furthermore, because the gases in both balloons are the same, they will have identical mass densities initially. In the final state, the first balloon will have the same values for number and mass density because it is unchanged. The second balloon however experiences a temperature and volume decrease (Charles' law). Let's calculate the number and mass densities for each balloon in each state:

$$\underline{Initial\ Number\ Density}$$

$$Balloon\ One\ and\ Balloon\ Two:$$

$$PV = nRT$$

$$\frac{n}{V} = \frac{P}{RT} = numberdensity$$

$$P = 1.0 bar$$

$$R = 0.08314\ \tfrac{L \cdot bar}{mol \cdot K}$$

$$T = 298.15K$$

$$\frac{n}{V} = \frac{P}{RT} = \frac{(1.0 bar)}{(0.08314\ \frac{L \cdot bar}{mol \cdot K})(298.15K)} = 0.0403\ \tfrac{mol}{L}$$

$$\frac{n}{V} \approx 4.0 \times 10^{-2}\ \tfrac{mol}{L}$$

Notice that we did not need to calculate volume or number of moles to determine the number density of the gas. Knowing that mass and moles are related through the molar mass of the sample, we can simply multiply the number density by the molar mass of the gas to calculate mass density:

$$\underline{Initial\ Mass\ Density}$$
$$Balloon\ One\ and\ Balloon\ Two:$$
$$PV = nRT$$
$$\frac{n}{V} = \frac{P}{RT}$$
$$mass = n \times molarmass$$
$$massdensity = \frac{n}{V} \times molarmass$$
$$molarmass = 4.0\,\tfrac{g}{mol}$$
$$massdensity = 0.0403\,\tfrac{mol}{L} \times 4.0\,\tfrac{g}{mol} = 0.161\,\tfrac{g}{L}$$
$$massdensity \approx 1.6 \times 10^{-1}\,\tfrac{g}{L}$$

In the final state, balloon one will have the same values for its number density and its mass density because nothing was changed. The second balloon (in the liquid nitrogen) is where we will find new values:

$$\underline{Final\ Number\ Density}$$
$$Balloon\ One:$$
$$\frac{n}{V} \approx 4.0 \times 10^{-2}\,\tfrac{mol}{L}$$

$$Balloon\ Two:$$
$$PV = nRT$$
$$\frac{n}{V} = \frac{P}{RT} = numberdensity$$
$$P = 1.0bar$$
$$R = 0.08314\,\tfrac{L\cdot bar}{mol\cdot K}$$
$$T = 77K$$

$$\frac{n}{V} = \frac{P}{RT} = \frac{(1.0bar)}{(0.08314\,\tfrac{L\cdot bar}{mol\cdot K})(77K)} = 0.156\,\tfrac{mol}{L}$$
$$\frac{n}{V} \approx 1.6 \times 10^{-1}\,\tfrac{mol}{L}$$

$\underline{Final\ Mass\ Density}$

$Balloon\ One:$

$massdensity \approx 1.6 \times 10^{-1}\,\frac{g}{L}$

$Balloon\ Two:$

$PV = nRT$

$\frac{n}{V} = \frac{P}{RT}$

$mass = n \times molarmass$

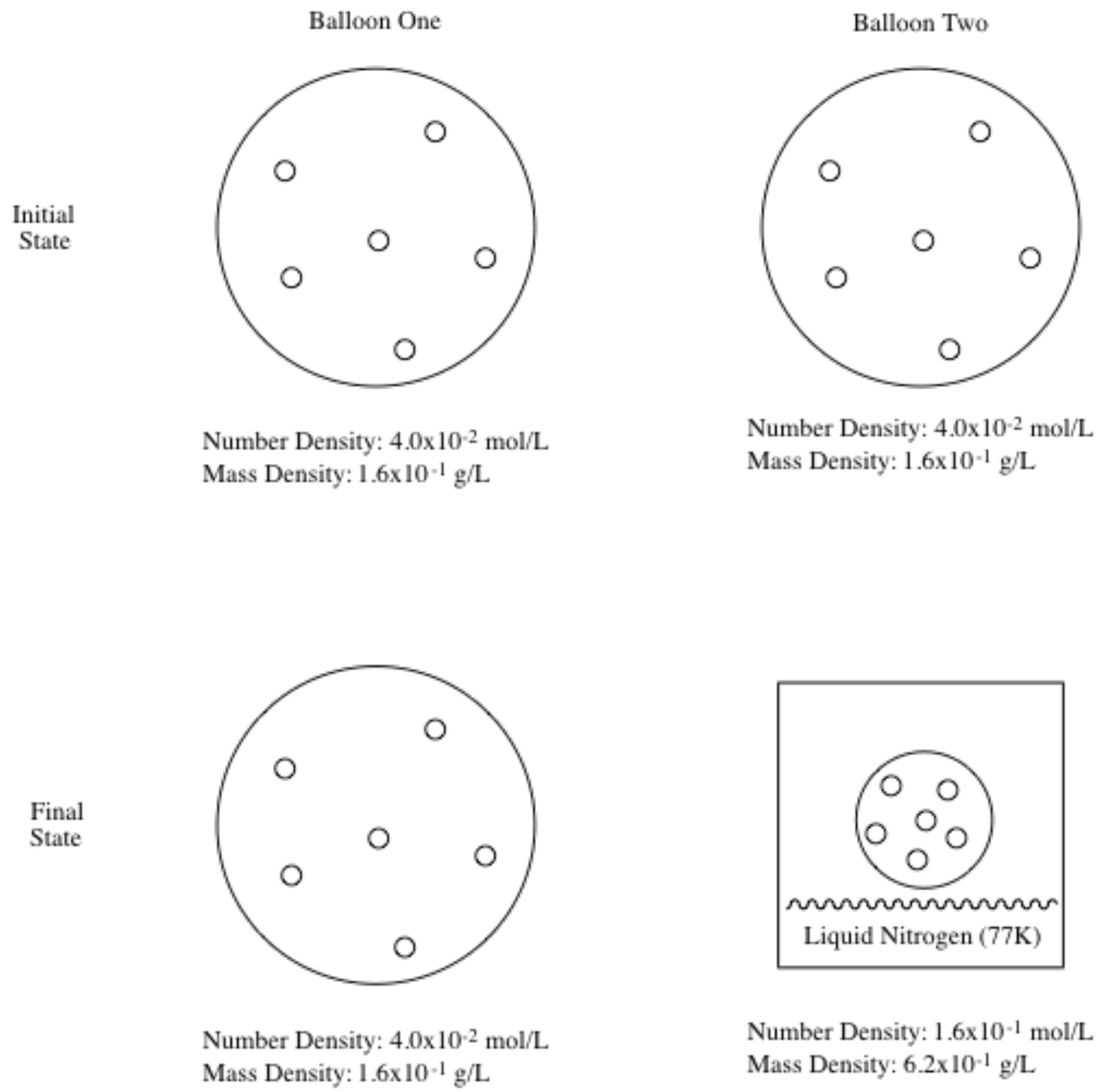
$massdensity = \frac{n}{V} \times molarmass$

$molarmass = 4.0\,\frac{g}{mol}$

$massdensity = 0.156\,\frac{mol}{L} \times 4.0\,\frac{g}{mol} = 0.625\,\frac{g}{L}$

$massdensity \approx 6.2 \times 10^{-1}\,\frac{g}{L}$

Due to the decrease in temperature, balloon two decreased in volume. However, this balloon still contained the same amount of gas as in its initial state. Therefore, we observe an increase in both its number density and its mass density. Notice that because the liquid nitrogen container is open, the pressure remains the same as in the initial state.



5. An adult's lungs can hold about 6L. What mass of air can an adult hold at a pressure of 102 kPa? Normal body temperature is 37 °C and air is about 20% oxygen and 80% nitrogen. (101,325 Pa = 1 atm)

Air is 20% O2 and 80% N2, oxygen and nitrogen are diatomic molecules
Oxygen: 32 g/mol
Nitrogen: 28 g/mol
Oxygen contributes 6.4 g/mol of air and nitrogen contributes 22.4 g/mol of air. So, air has a total of 28.8 g/mol.
(1 atm / 101,325 Pa)(102000 Pa) = 1.007 atm
37 °C + 273 = 310 K

$$\boxed{\begin{array}{l} PV = nRT \\ n = \dfrac{PV}{RT} = \dfrac{1.007 atm \times 6L}{0.082 \dfrac{Latm}{molk} \times 310K} = 0.238 mol \\ 0.238 mol \times \dfrac{28.8g}{mol} = 6.85 grams\_of\_air \end{array}}$$

ACTIVITY 2

The purpose of this activity is practice your understanding using gas data to compute molar mass of gas

1. Use your knowledge of the ideal gas law to write two equations to solve for molar mass of an ideal gas – one where mass is a variable and the other where density is a variable.

$$\boxed{\begin{array}{l} molarmass = \dfrac{mass \times RT}{PV} \\ \\ molarmass = \dfrac{\rho RT}{P} \end{array}}$$

2. Given that a 2.16 gram sample of gas occupies 0.43 L at a pressure of 1.2 atm at a temperature of 298 K. Calculate the molar mass of the gas.

$$\boxed{\begin{array}{l} molarmass = \dfrac{mass \times RT}{PV} \\ mass = 2.16g \\ R = 0.08206 \frac{L \cdot atm}{mol \cdot K} \\ T = 298K \\ P = 1.2atm \\ V = 0.43L \\ molarmass = \dfrac{(2.16g)(0.08206 \frac{L \cdot atm}{mol \cdot K})(298K)}{(1.2atm)(0.43L)} \\ molarmass = 102.4 \frac{g}{mol} \end{array}}$$

3. A gas has a molar mass of 100.0 g/mol. At 25˚C, 1.40 moles of the gas exerts a pressure of 380 torr. What is the density of the gas under these conditions?

$$
\boxed{
\begin{aligned}
&PV = nRT \\
&n = \frac{PV}{RT} \\
&n = \frac{mass}{molarmass} \\
&denisty = \rho = \frac{mass}{volume} \\
&mass = denisty \times volume = \rho V \\
&n = \frac{\rho V}{molarmass} \\
&\frac{\rho V}{molarmass} = \frac{PV}{RT} \\
&molarmass \times PV = \rho VRT \\
&\rho = \frac{molarmass \times P}{RT} \\
&molarmass = 100.0 \tfrac{g}{mol} \\
&P = 380 torr \\
&R = 62.36 \tfrac{L \cdot torr}{mol \cdot K} \\
&T = 25^\circ C = 298.15K \\
&\rho = \frac{(100.0 \tfrac{g}{mol})(380 torr)}{(62.36 \tfrac{L \cdot torr}{mol \cdot K})(298.15K)} \\
&\rho = 2.04 \tfrac{g}{L}
\end{aligned}
}
$$

4. The unknown gas sample at 137°C and 0.989 bar, has a density of 0.698 g/L. Calculate the molar mass of this unknown gas.

$$\boxed{\begin{aligned}
&molarmass = \frac{\rho RT}{P} \\
&\rho = 0.698\,\tfrac{g}{L} \\
&R = 0.08314\,\tfrac{L\cdot bar}{mol\cdot K} \\
&T = 137^{\circ}C = 410.15K \\
&P = 0.989 bar \\
&molarmass = \frac{(0.698\,\tfrac{g}{L})(0.08314\,\tfrac{L\cdot bar}{mol\cdot K})(410.15K)}{(0.989 bar)} \\
&molarmass = 24.1\,\tfrac{g}{mol}
\end{aligned}}$$

5. A gas exerts a pressure of 1.12 atm in a 4 L container at º19C. You know the density of the gas is 1.5 g/L. What is the molecule?

$PV = nRT$

$(1.12\ atm)(4L) = n\,(0.082\,\frac{Latm}{K\ mol})(19 + 273) = .187$

$(0.187 g/4L)(x\ g/mol) = 1.5\ g/L$

$x = 32 g/mol$

Probably O_2

Gas Mixtures

Activity 1

The purpose of this activity is for you check your understanding of gas mixtures.

1. Explain the relationship between the partial pressure of a gas in a mixture and the total pressure of the mixture with words and an equation.

Here is the equation for this relationship:

$$\boxed{P_i = X_i P_{Total}}$$

where P_i is the partial pressure of a gas in a mixture, X_i is mole fraction of a gas in a mixture, and P_{Total} is the total pressure of the gas mixture.

The relationship states that the partial pressure of a gas in a mixture is directly proportional to its concentration in the mixture (mole fraction).

2. Explain Dalton's Law of partial pressure with words and an equation.

   Here is the equation for this relationship:

$$\boxed{\begin{aligned} &P_{Total} = \sum_{i=1}^{n} P_i \\ &OR \\ &P_{Total} = P_1 + P_2 + \ldots + P_n \end{aligned}}$$

   where P_i is the partial pressure of a gas in a mixture, n represents the number of individual gases in the mixture and P_{Total} is the total pressure of the gas mixture.

   Dalton's Law of partial pressure states that the total pressure of a gas mixture is the sum of all the partial pressures of each gas within the mixture.

3. A mole fraction is considered a measure of concentration. Explain why this is true.

   Mole fraction is the ratio of the moles of an individual substance to the total number of moles. This provides us with a value that expresses the relative presence of that substance in the mixture. In this way, mole fraction is a measure of concentration.

4. True or False? It is impossible to actually measure the individual pressures of a particular gas within a mixture of gases. Explain your answer.

   True. A pressure sensor cannot differentiate between different types of molecules that collide with it inside a container. All collisions are simply felt as collisions. So we need Dalton's Law of partial pressures to help us calculate partial pressures.

Activity 2

1. In the troposhere of Titan, Saturn's largest moon, the atmospheric pressure is about 1.5 bar. At this point in the atmosphere there is approximately 4.9% methane (the vast majority of the atmosphere is nitrogen) by number. Calculate the partial pressure of methane in the stratosphere of Titan.

   The problem states the atmosphere is 4.9% methane "by number." We can interpret "by number" to mean 4.9% of the molecules (and therefore moles) at this point in Titan's atmosphere are methane. This makes the calculation of the mole fraction of methane fairly straightforward. If we have say 100 mole sample of Titan's atmosphere, then 4.9 moles of this mixture are methane. This makes the mole fraction of methane 0.049.

$$\boxed{\begin{aligned} &P_{CH_4} = X_{CH_4} P_{total} \\ &X_{CH_4} = \frac{n_{CH_4}}{n_{total}} = \frac{4.9 mol}{100 mol} = 0.049 \\ &P_{CH_4} = (0.049)(1.5 bar) \\ &P_{CH_4} = 0.0735 bar \end{aligned}}$$

2. You are the Titan expert of your lab group and you decide to recreate the atmosphere in the stratosphere of Titan in a container with a fixed volume of 7.00L. You dutifully remind your lab partners that the stratosphere of Titan is, by number, 98.4% nitrogen gas, 1.40% methane gas and the rest is hydrogen gas. Finally, you remind them to make sure all calculations are for a total pressure of 1.50 bar when the mixture is at a temperature –179˚C in order to best mimic the conditions on Titan. Luckily, you already created a freezer that maintains a perfect temperature of –179˚C into which you will place your model Titan atmosphere. Draw a small particle model of what is going on in the container when the mixture is at a final pressure of 1.5 bar and final temperature of –179˚C. Determine the pressure of each gas in the container and compute how many grams of each gas you will place into this container (after first evacuating it of course).

First we need to know how many total moles of gas will be present with the given "Titan"-like conditions of 1.5 bar and –179˚C:

$$\boxed{\begin{aligned}
&Total\ Moles\ in\ Mixture:\\
&PV = nRT\\
&P = 1.5bar\\
&V = 7.00L\\
&R = 0.08314\tfrac{L\cdot bar}{mol\cdot K}\\
&T = -179^{\circ}C = 94.15K\\
&PV = nRT\\
&n_{total} = \frac{PV}{RT}\\
&n_{total} = \frac{(1.5bar)(7.00L)}{(0.08314\frac{L\cdot bar}{mol\cdot K})(94.15K)} = 1.341402mol
\end{aligned}}$$

So our mini-Titan stratosphere contains 1.341402 moles of gas total in 7.00L at –179˚C to achieve a pressure of 1.5 bar.

Now, we know the percentage breakdown of the three gases present in this mixture. We also know the percentages are "by number" meaning these percentages pertain to the molar composition and not the mass composition. We can determine the number of moles of each type of gas:

$$\boxed{\begin{aligned}
&Moles\ of\ each\ gas:\\
&n_x = (percentage)\times n_{total}\\
&\\
&n_{N_2} = 0.984\times 1.341402mol = 1.319939mol\\
&n_{CH_4} = 0.014\times 1.341402mol = 0.01877963mol\\
&n_{H_2} = 0.002\times 1.341402mol = 0.002682804mol
\end{aligned}}$$

The partial pressures of each gas could be calculated in two ways. We could go through the trouble of calculating the mole fractions of each gas and then calculate partial pressure. OR, we could recognize that the percentages by number are equivalent to the mole fraction of each gas! So we can simply use the percentages given in the original problem to calculate partial pressure:

$$\boxed{\begin{aligned}
&P_{total} = 1.5bar \\
&P_{N_2} = (percentage_{N_2}) \times P_{total} \\
&P_{N_2} = (0.984)(1.5bar) \\
&P_{N_2} = 1.476bar \\
\\
&P_{CH_4} = (percentage_{CH_4}) \times P_{total} \\
&P_{CH_4} = (0.014)(1.5bar) \\
&P_{CH_4} = 0.021bar \\
\\
&P_{H_2} = (percentage_{H_2}) \times P_{total} \\
&P_{H_2} = (0.002)(1.5bar) \\
&P_{H_2} = 0.003bar
\end{aligned}}$$

Finally, we can calculate the grams needed to be added to the container by multiplying the number of moles by their molar masses:

$$\boxed{\begin{aligned}
&Mass\ of\ each\ gas: \\
&mass = n \times molarmass \\
&mass_{N_2} = 1.319939mol\ N_2 \times \frac{28g\ N_2}{1mol\ N_2} = 36.95831g\ N_2 \\
&mass_{CH_4} = 0.01877963mol\ CH_4 \times \frac{16g\ CH_4}{1mol\ CH_4} = 0.3004740g\ CH_4 \\
&mass_{H_2} = 0.002682804mol \times \frac{2g\ H_2}{1mol\ H_2} = 0.005365608g\ H_2 \\
&mass_{N_2} \approx 37.0g\ N_2 \\
&mass_{CH_4} \approx 3.00 \times 10^{-1}g\ CH_4 \\
&mass_{H_2} \approx 5.37 \times 10^{-3}g\ H_2
\end{aligned}}$$

Here's a small particle representation of the 7.00L mini-Titan atmosphere at –179˚C and 1.5 bar pressure.