* Dalton's Law of Partial Pressures

By Shawn P. Shields, Ph. D.

* Recall: The Ideal Gas Law

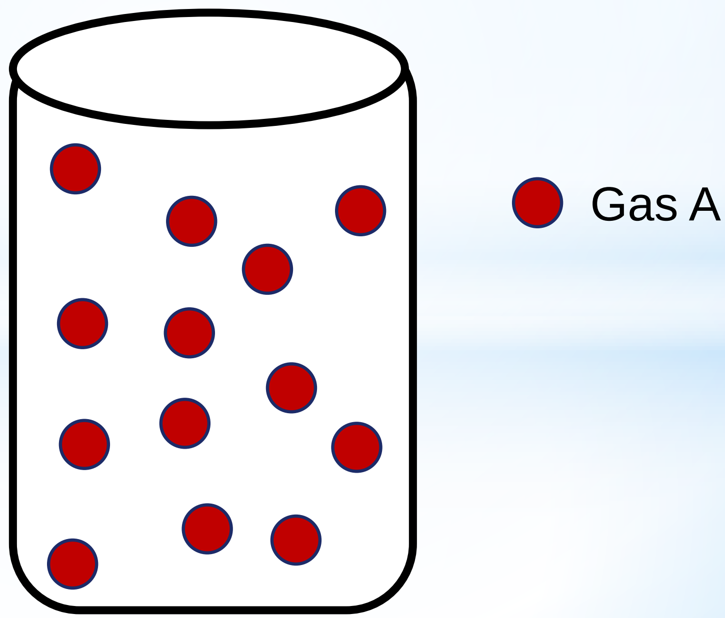
* The equation: $PV = nRT$
* Gases are described in terms of four macroscopic observables:
  * Pressure (P in atm)
  * Temperature (T in Kelvin)
  * Volume (V in L)
  * Moles of gas (n)
  * Recall that R is the gas constant: $R = \frac{0.0826 \text{ L atm}}{\text{mol K}}$

*Dalton's Law of Partial Pressures

Recall that the Ideal Gas Law simply states that

the *behavior of a gas* in terms of its pressure, temperature, and volume depends ONLY on the *number of moles of gas*, NOT on the identity of the gas.

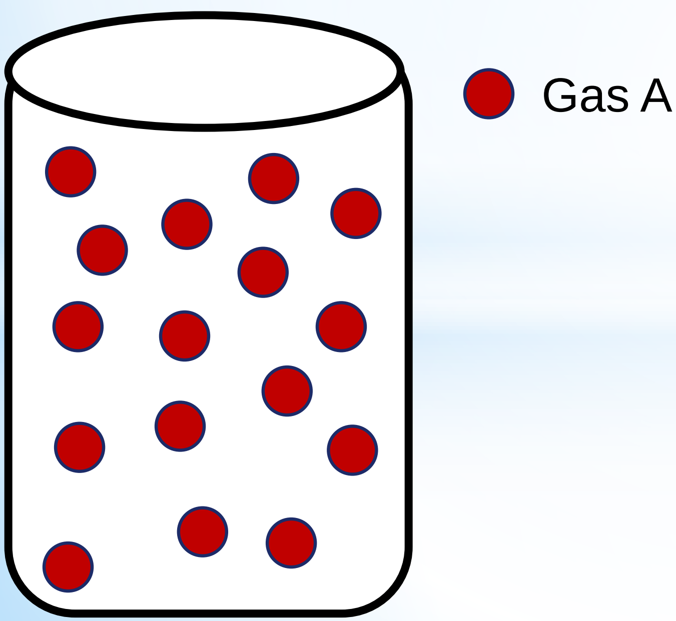
The <u>moles</u> of a gas can be thought of as the <u>number of molecules</u> of gas.

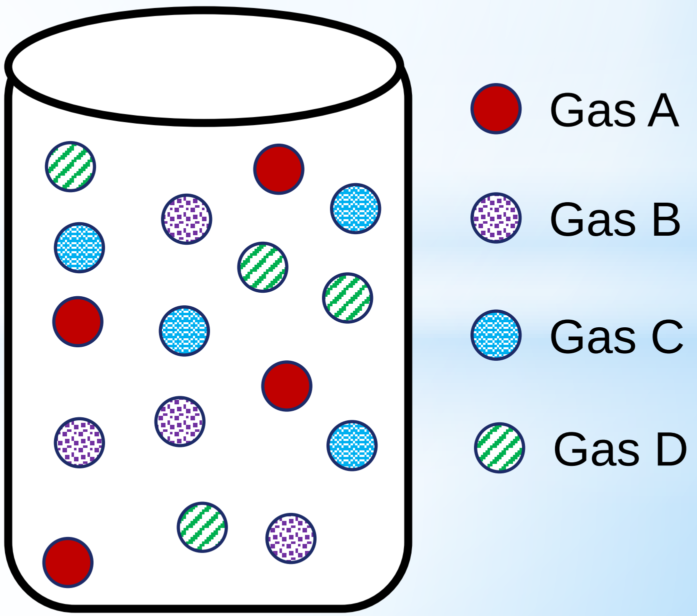


*Dalton's Law of Partial Pressures

Dalton's Law extends this idea to *mixtures of gases*. The P-V-T properties of a mixture of gases will behave just like a pure gas.

The behavior of the gas only depends on the total number of moles of gas, no matter their identity!

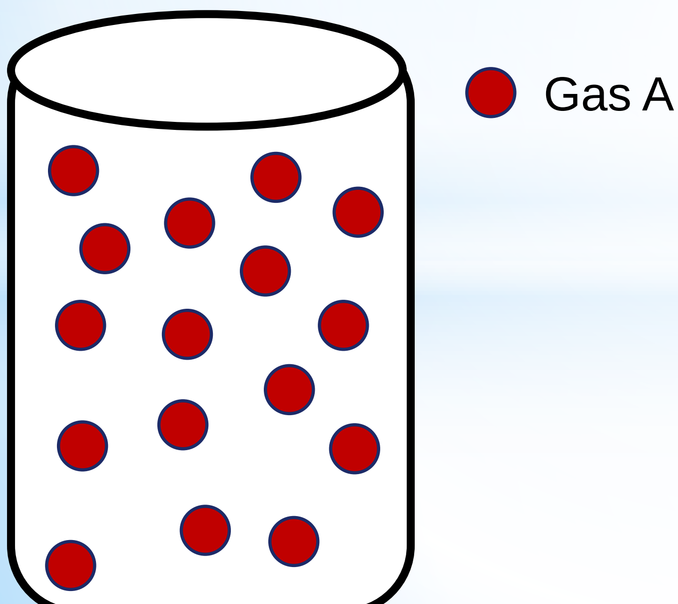


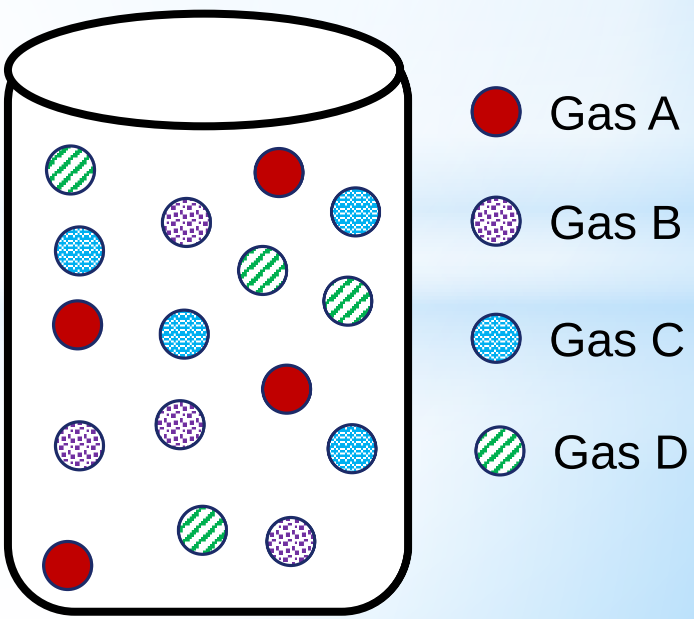


*Dalton's Law of Partial Pressures

Both containers have the same volume and the same number of gas particles (moles of gas).

At the same temperature, both the pure gas and the mixture of gases will exert the same pressure!

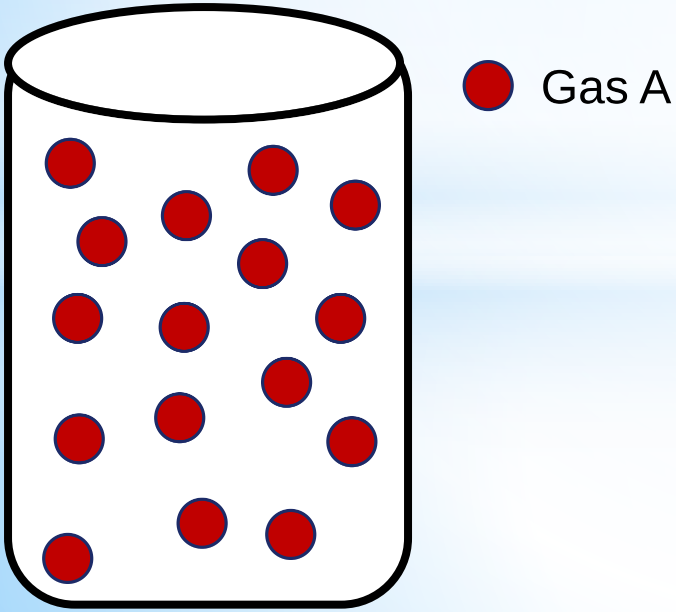




*Dalton's Law of Partial Pressures

We can write an equation to calculate the pressure of gas A in the container as

$$P_A = \frac{n_A RT}{V}$$



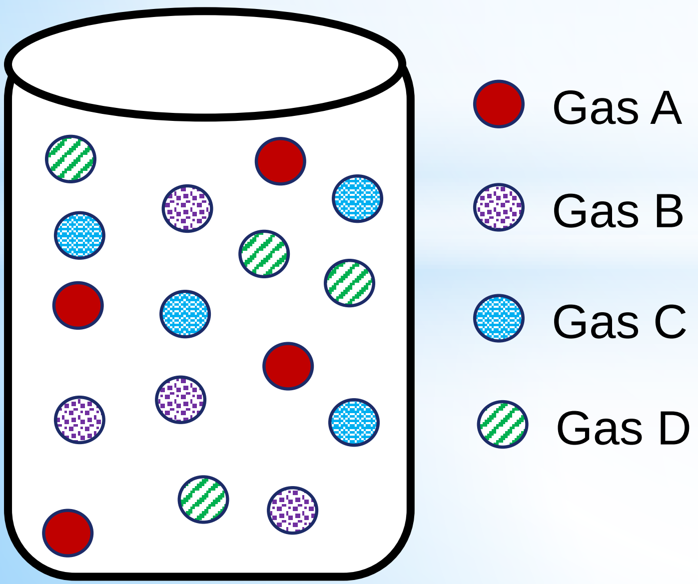
All we've done is identify "gas A" as a subscript in the Ideal Gas Equation.

The pressure due to gas A is just dependent on the moles of gas A, at a given T and V.

*Dalton's Law of Partial Pressures

Now, let's do the same thing for the mixture of gases...
We can write an equation to calculate the total pressure of all four gases in the container as

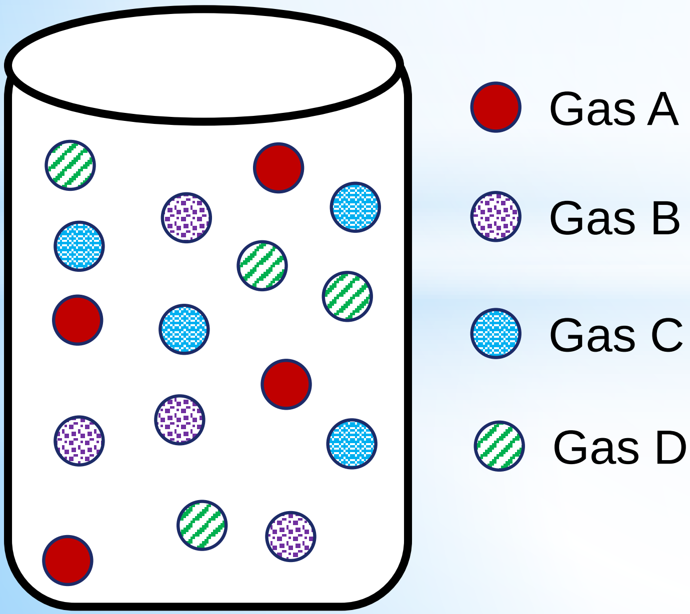
$$P_{Total} = \frac{n_A RT}{V} + \frac{n_B RT}{V} + \frac{n_C RT}{V} + \frac{n_D RT}{V}$$



*Dalton's Law of Partial Pressures

The total pressure due to the gases in the container is

$$P_{Total} = P_A + P_B + P_C + P_D$$



Each gas exerts one-fourth (¼) of the total pressure!

The gases all behave independently of one another.

*The Mole Fraction

Let's just look at the contribution from gas B to the total pressure in the container and write an equation for it.

$$\mathrm{P_B} = \frac{n_B RT}{V}$$

Let's also write the equation for the total pressure another way…

$$\mathrm{P_{Total}} = \frac{n_{Total} RT}{V}$$

*The Mole Fraction

Now, divide the equation for the "partial pressure of gas B" (P_B) by the total pressure

$$\frac{P_B}{P_{Total}} = \frac{\frac{n_B RT}{V}}{\frac{n_{Total} RT}{V}}$$

Cancel out R, T, and V, since each has the same value in both equations…

*The Mole Fraction

Now, divide the equation for the "partial pressure of gas B" (P_B) by the total pressure

$$\frac{P_B}{P_{Total}} = \frac{\frac{n_B \cancel{R}\cancel{T}}{\cancel{V}}}{\frac{n_{Total}\cancel{R}\cancel{T}}{\cancel{V}}}$$

We are left with

$$\frac{P_B}{P_{Total}} = \frac{n_B}{n_{Total}}$$

*The Mole Fraction

Notice that the partial pressure of B is equal to the fraction of gas B in the total moles of gas…

$$\frac{P_B}{P_{Total}} = \frac{n_B}{n_{Total}} = \chi_B$$

We call this the "mole fraction" gas B.

$$\frac{n_B}{n_{Total}} = \chi_B$$

*The Mole Fraction

Now, rearrange the equation to calculate the fraction (or part) of the total pressure that is contributed by gas B…

$$P_B = \chi_B P_{Total}$$

If we multiply the mole fraction of gas B by the total pressure in the container, we get the "partial pressure" of gas B.

In other words, we can calculate the pressure that gas B is contributing all by itself. ☺

* A Simple Example…

The atmosphere is composed of 78% N_2 gas, 21% O_2 gas, and about 1% other gases. If the total pressure of the atmosphere is 1 atm, then what is the partial pressure of N_2?

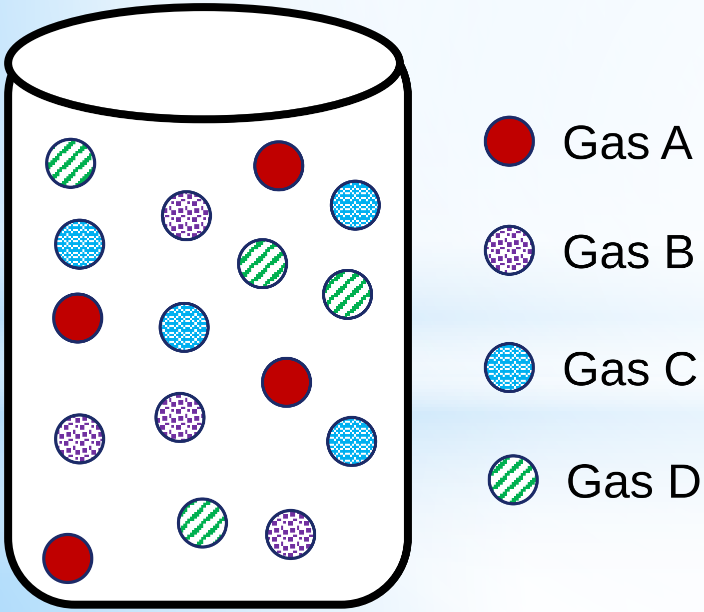
$$P_{N_2} = \chi_{N_2} P_{Total}$$

$$\mathrm{P_{N_2}} = (0.78)(1\,\mathrm{atm}) = 0.78\,\mathrm{atm}$$

What is the partial pressure of O_2?

Summary of Dalton's Law of Partial Pressures

The partial pressure of each gas in the container is only dependent on its mole fraction.



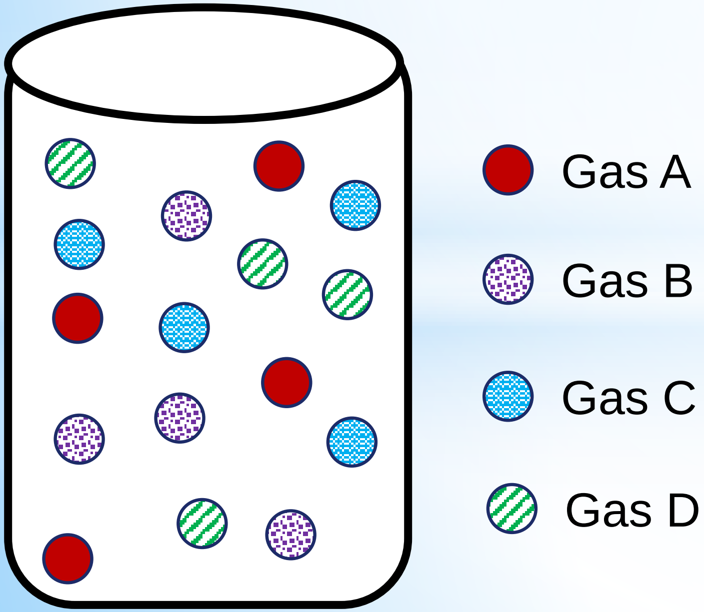
If each gas in the container has a mole fraction of one fourth, then each gas contributes $\frac{1}{4}$ of the total pressure.

The gases do not interact, so the mixture of gases obeys the Ideal Gas Law. ☺

*Summary of Dalton's Law of Partial Pressures

Suppose the total pressure in the container is 6.0 atm. What is the partial pressure of gas D in the container?

(Hint: Count the gas molecules)



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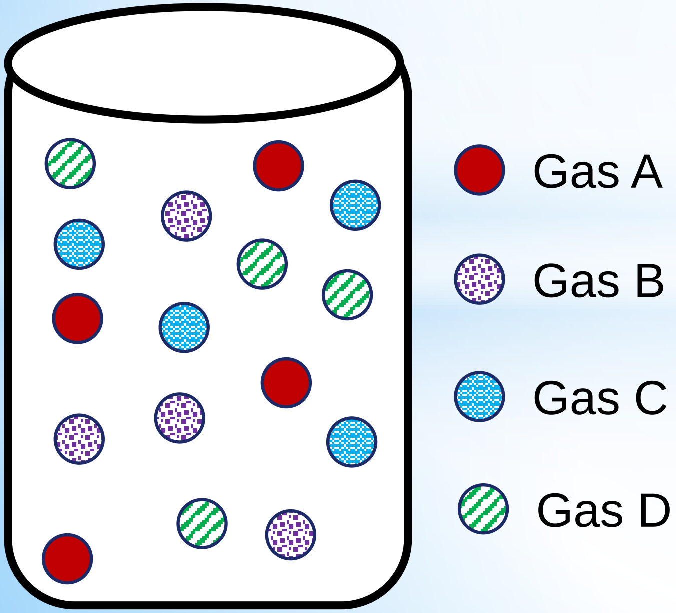


Gas D contributes ¼ of 6.0 atm.

$$P_D = \chi_D P_{Total}$$

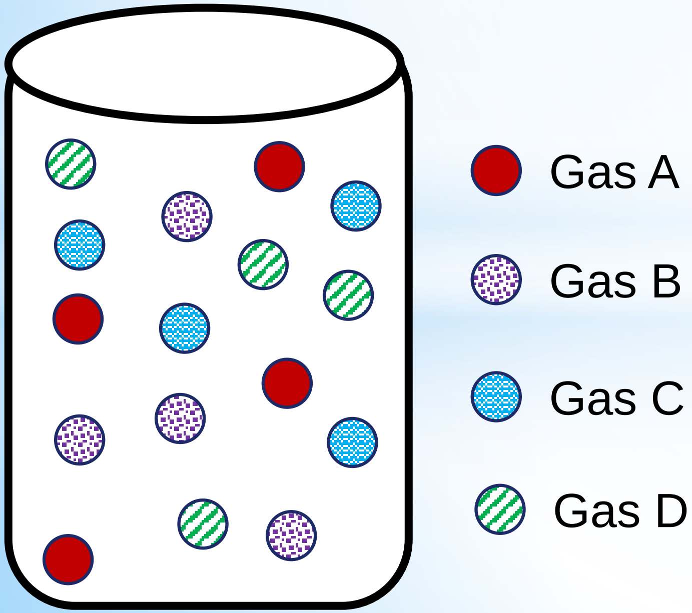
$$P_D = (0.25)(6.0\ atm)$$

$$P_D = 1.5\ atm$$

Could you do this in your head? ☺

*Summary of Dalton's Law of Partial Pressures

Suppose the total pressure in the container is 10 atm. What is the partial pressure of gases B and C in the container? (Hint: Count the gas molecules.)



Summary of Dalton's Law of Partial Pressures

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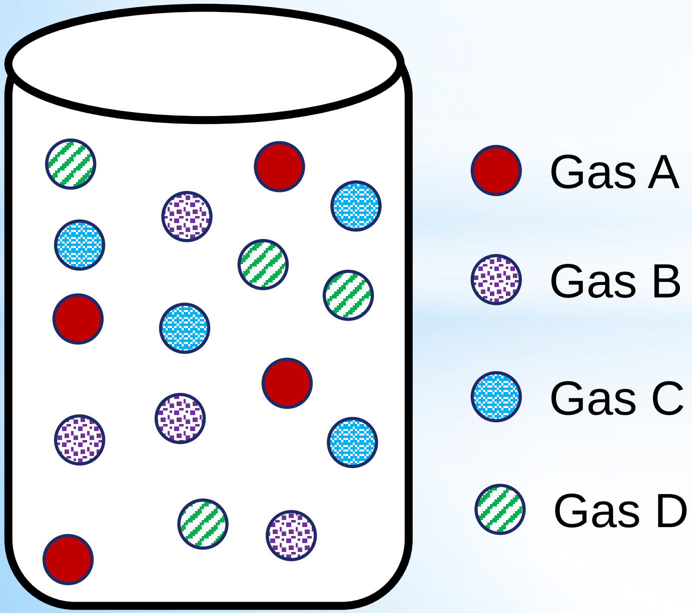

Together, gases B and C contribute ½ of 10 atm.

$$P_{B+C} = (\chi_B + \chi_C)P_{Total}$$

$$P_{B+C} = (0.25 + 0.25)(10\ atm)$$

$$P_{B+C} = 5\ atm$$

☺