Dalton's Law

<u>Dalton's Law of Partial Pressures:</u> the total pressure of a mixture of gases is equal to the <u>Sum</u> of the partial pressures of the component gases.

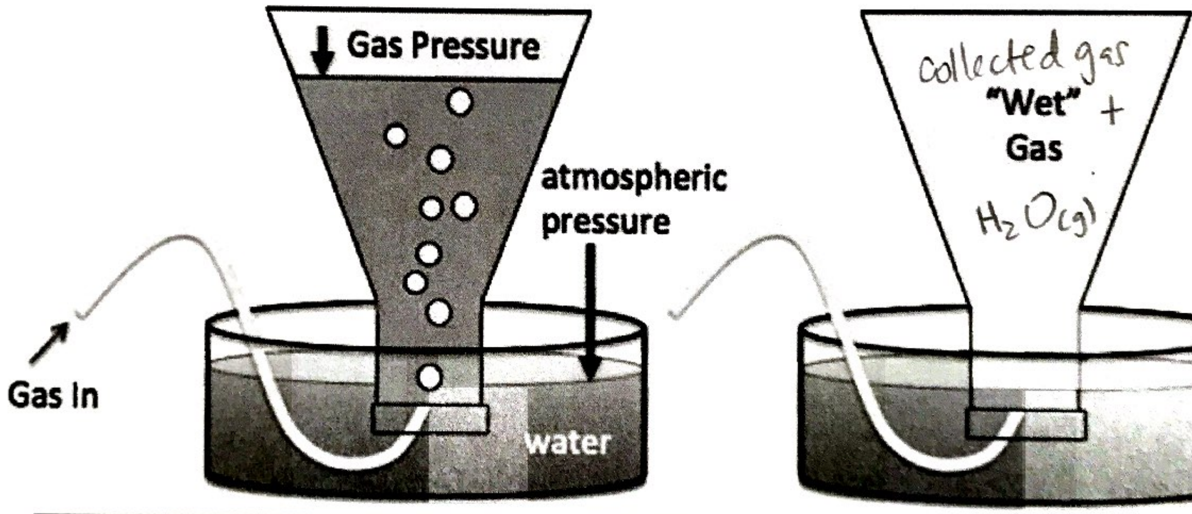
$$P_{total} = P_{gas\,A} + P_{gas\,B} + P_{gas\,C} + \cdots$$

→ The pressure of each gas in a mixture is called the <u>partial</u> pressure of that gas.

<u>Example #1</u>: A sealed container of helium, hydrogen, and nitrogen gases is measured to have a pressure of 288 kPa. If the partial pressures of the helium and the hydrogen are 112 kPa and 83 kPa respectively, what would be the partial pressure of the nitrogen? $P_{tot} = P_{He} + P_{H_2} + P_{N_2}$

$$P_{N_2} = 288 - 112 - 83 = \boxed{93\ kPa}$$

<u>Gas Collection Over Water</u>



- As the gas bubbles in (often from an attached chemical reaction), it is collected in a container.
- As the container fills with gas, it <u>pushes</u> water down.
- At the same time, atmospheric pressure is pushing down on the water outside the container.

- When the water level inside and outside the container are <u>equal</u>, the pressure inside the container equals the atmospheric pressure outside the container.

1. Gases collected in this way are not pure but always mixed with <u>water</u> vapor.
2. Like other gases, water vapor exerts a <u>pressure</u>.
3. To find the "<u>dry</u>" gas pressure, use Dalton's Law to subtract the vapor pressure of water from the atmospheric (or <u>barometric</u>) pressure.

$$P_{atm} = P_{gas} + P_{H_2O}$$

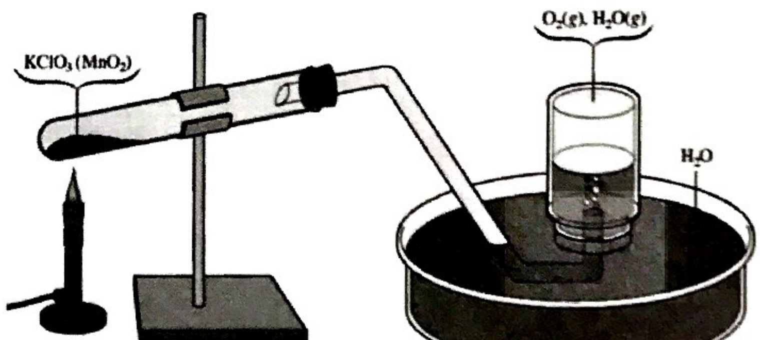
Table 1: Water-Vapor Pressure

Temperature (°C)	Pressure (mmHg)	Temperature (°C)	Pressure (mmHg)	Temperature (°C)	Pressure (mmHg)	Temperature (°C)	Pressure (mmHg)
0.0	4.6	18.5	16.0	23.5	21.7	40.0	55.3
5.0	6.5	19.0	16.5	24.0	22.4	50.0	92.5
10.0	9.2	19.5	17.0	24.5	23.1	60.0	149.4
15.0	12.8	20.0	17.5	25.0	23.8	70.0	233.7
15.5	13.2	20.5	18.1	26.0	25.2	80.0	355.1
16.0	13.6	21.0	18.6	27.0	26.7	90.0	525.8
16.5	14.1	21.5	19.2	28.0	28.3	95.0	633.9
17.0	14.5	22.0	19.8	29.0	30.0	100.0	760.0
17.5	15.0	22.5	20.4	30.0	31.8		
18.0	15.5	23.0	21.1	35.0	42.2		

Now you try!

1. A sample of solid potassium chlorate ($KClO_3$) was heated in a test tube (see the figure to the right) and decomposed by the following reaction:

$$2\ KClO_3(s) \rightarrow 2\ KCl(s) + 3\ O_2(g)$$



The oxygen produced was collected by displacement of water at 22°C at a barometric pressure of 754 torr. The volume of the gas collected was 0.650 L.

a. Calculate the partial pressure of O_2 in the gas collected (the dry pressure of O_2).

$$P_{total} = P_{O_2} + P_{H_2O} \Rightarrow P_{O_2} = 754 - 19.8 = 734.2$$
$$= \boxed{734\ \text{torr}}$$

b. What mass of $KClO_3$ was decomposed in this experiment?

$$n_{O_2} = \frac{PV}{RT} = \frac{(734\ \text{torr})(0.650\ \text{L})}{(62.36\ \frac{\text{L}\cdot\text{torr}}{\text{mol}\cdot\text{K}})(295\ \text{K})} = 0.0259\ \text{mol}\ O_2 \times \frac{2\ \text{mol}\ KClO_3}{3\ \text{mol}\ O_2} \times \frac{122.55\ \text{g}\ KClO_3}{1\ \text{mol}\ KClO_3}$$
$$= \boxed{2.12\ \text{g}\ KClO_3}$$

c. What would the volume of the dry O_2 be at STP?

$$V = \frac{nRT}{P} = \frac{(0.0259\ \text{mol})(0.08206\ \frac{\text{L}\cdot\text{atm}}{\text{mol}\cdot\text{K}})(273.15\ \text{K})}{(1.00\ \text{atm})} = \boxed{0.581\ \text{L}\ O_2}$$

Mixtures of Gases: Mathy Math

Many gas samples are not pure, but instead are mixtures of different gases.

- In certain applications, the gas mixture can be treated as one gas.
    - We can calculate the total number of moles of ALL gases, and use the ideal gas law to calculate other properties of the mixture.

$$P_{total} = P_{gas\,A} + P_{gas\,B} + P_{gas\,C} + \cdots$$
$$= n_A \frac{RT}{V} + n_B \frac{RT}{V} + n_C \frac{RT}{V} + \cdots$$
$$= (n_A + n_B + n_C + \cdots)\frac{RT}{V}$$
$$= (n_{total})\frac{RT}{V}$$

Mole Fractions

Mole fraction (X_A): number of moles of a component in a mixture, divided by the total number of moles in the mixture

$$X_A = \frac{\text{moles A}}{\text{total moles}}$$

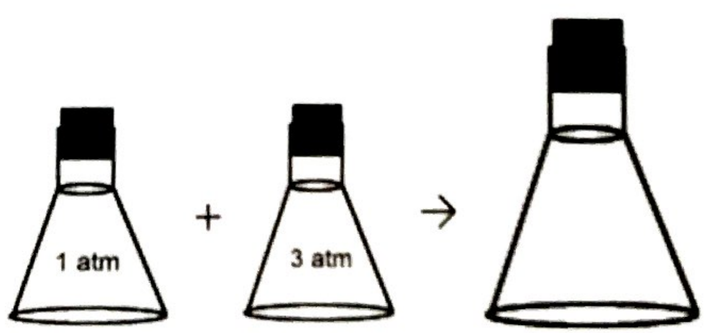
The mole fraction can be used to determine the partial pressure of a component in a gaseous mixture

$$P_A = P_{total} \times X_A$$

Let's Practice!

1. The mole fraction of nitrogen in the air is 0.7808. Calculate the partial pressure of N_2 in the air when the atmospheric pressure is 760. torr.

$$P_{N_2} = (760.\ \text{torr})(0.7808) = \boxed{593\ \text{torr}}$$



Flask A (1.0 L) Flask B (1.0 L) Flask C (4.0 L)

$$\frac{4\ \text{atm}}{4} = 1\ \text{atm}$$

2. Two one-liter flasks (Flask A and Flask B, shown above) are sealed at 25°C. Flask A contains helium gas, and Flask B contains argon gas. If the contents of both flasks are combined into a previously evacuated four-liter flask (Flask C), what would be the total pressure in Flask C at 25°C?

   a. 4.0 atm b. 3.0 atm c. 2.0 atm d. 1.0 atm