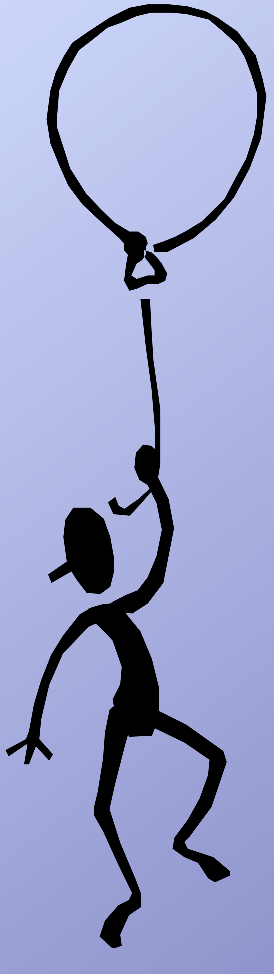
Explain what is happening here….

- https://youtu.be/JsoE4F2Pb20

Ideal Gases & Ideal Gas Law



Ideal Gases vs. Real Gases

- Particles have no volume
- Particles are not attracted or repelled to each other
- Ideal gases can never be liquefied or solidified

- Particles have volume
- Particles experience intermolecular forces
- Real gases can be liquefied and solidified

Deviations from Ideal Behavior

Likely to behave nearly ideally	*Likely not to behave ideally*
Gases at high temperature and low pressure	Gases at low temperature and high pressure
Small non-polar gas molecules	Large, polar gas molecules

Variables that Describe a Gas

- Pressure (atm, mmHg, kPa)
- Volume (L)
- Temperature (°C or K)
- Number of moles (mol)

The Ideal Gas Law

- Allows us to solve for a property of an ideal gas when properties are constant!

PV=nRT

P=pressure

V=volume **(Liters!!)**

T=temperature **(Kelvin!!)**

n=number of moles

$$R = 8.31 \frac{L \times kPa}{K \times mol} \quad \text{or} \quad 0.08206 \frac{L \times atm}{K \times mol}$$

Lets Practice

Determine the volume occupied by 0.582 mol of a gas at 15°C if the pressure is 81.1 kPa.

- Identify what you know

  n=0.582 mol

  T=15°C + 273 = 288 K

  P=81.1 kPa.

  R= 8.31 L x kPa / K x mol

- V=?

PV=nRT

81.1 Kpa x V=(0.582 mol) (8.31 L x kPa / K x mol) (288K)

81.1 KPa x V=1392.89 L x kPa

V=17.2 L

Two rules!

- Temp= Kelvin, Volume =Liters, and pressure= atm or kPa
- If you are given grams convert into moles.

$$36 \text{ grams of } H_2O \times \frac{1 \text{ mole } H_2O}{18 \text{ g } H_2O} = 2 \text{ mol } H_2O$$

Dalton's Law of Partial Pressures

- the total pressure of a mixture of gases is equal to the sum of the partial pressures of all the gases present. (at constant Temp & Volume)

Formula:

$$P_{total} = P_1 + P_2 + P_3 + P_4 \ldots$$

Let’s Practice

1) What is the total pressure of a gas mixture if it contains CO_2 at 40.8 torr and O_2 at 1009.9 torr & H2 at 791.4 torr?

PT=P1 +P2+P3

PT= 40.8 + 1009.9 + 791.4 = 1841.2 torr

2) What is the pressure of Ne gas if the total pressure of the gas is 100.6 atm, and the mixture contains 40.4 atm of He and 22.6 atm of HCl?

PT=P1 +P2+P3

100.6 atm = P_{Ne} + 40.4 atm + 22.6 atm = 37.6 atm

Water Displacement

- Gases produced in the laboratory are often collected over water. Thus, the gas is not pure but is always mixed with water vapor.
- The collection process makes it so the total pressure inside the bottle would be the same as the atmospheric pressure.

$$P_{atm}=P_{gas} + P_{H2O}$$

Lets practice – (Hw check: pg 367 #1)

3) Some hydrogen gas is collected over water at 20.0 °C. The levels of water inside and outside the gas-collection bottle are the same. The partial pressure of hydrogen is 742.5 torr (mmHg).

What is the barometric pressure at the time the gas is collected?

Formula: $P_{atm} = P_{gas} + P_{H_2O}$ (to find the P use table A-8 pg. 859)

$P_{atm} = 742.5 \text{ torr} + 17.5 \text{ torr} = 760.0 \text{ torr}$

Lets Practice Continued….

4. Some CO_2 gas is collected over water at 15.0 °C. The levels of water inside and outside the gas-collection bottle are the same. The barometric pressure at the time the gas is collected is 780 mmHg. What is the partial pressure of CO_2?

Formula: Patm=Pgas + P_{H2O}

780 mmHg = P_{CO2} + 12.8 mmHg= 767.2mmHg

Dalton's Law

<u>Example #1:</u> What is the partial pressure of oxygen at 101.3kPa of total pressure if the partial pressure of nitrogen is 79.10 kPa and carbon dioxide is 0.040kPa?

Solution:

$P_{total} = P_o + P_{N_2} + P_{CO_2}$

$P_o = P_{total} – (P_{N_2} + P_{CO_2})$

Po = 101.30kPa – (79.10kPa + 0.040 kPa)

=22.16 kPa

Dalton Example #2

What is the partial pressure of each gas if you have a mixture of 2.0 moles of O_2 and 2.0 moles of CO_2 that exert at total pressure of 700 torr?

$$O_2 = \frac{2.0 \text{ mol } O_2}{4.0 \text{ total mol}} \times 700 \text{ torr} = 350 \text{ torr}$$

$$CO_2 = \frac{2.0 \text{ mol } CO_2}{4.0 \text{ total mol}} \times 700 \text{ torr} = 350 \text{ torr}$$

When to use the equations

- **If the conditions of the gas** change
  - **Then use the** combined gas law

- **If the conditions of the gas are** fixed
  - **Then use the** ideal gas law