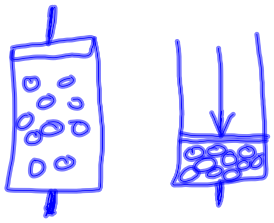
The Kinetic Molecular Theory of Matter

- a theory that is based on the idea that particles of matter are always in constant motion

- it can be used to explain the properties of solids, liquids, and gases in terms of energy of particles and forces that act between them



The Kinetic Molecular Theory of Gases

* **Ideal Gas** - an imaginary gas that perfectly fits all the assumptions of the kinetic molecular theory

1) Gases consist of large numbers of tiny particles that are far apart relative to their size

   - most of the volume occupied by a gas is empty space

2) Collisions between gas particles and between particles and container walls are elastic collisions

   - a collision where there is no net loss of kinetic energy

3) Gas particles are in continuous, rapid, and random motion. They therefore contain kinetic energy which is the energy of motion

4) There are no forces of attraction or repulsion between gas particles

5) The average kinetic energy of gas particles depends on the temperature of the gas

   - the higher the temp. the more kinetic energy (they move faster)

$\boxed{KE = 1/2\ mv^2}$ m = mass v = speed velocity

T= ↑ V=↑ KE↑

T=↓ V=↓ KE↓

Physical Properties of Gases

1) Expansion - gases don't have definite shape and volume. They will expand to whatever size container they are put in

2) Fluidity - because there is no attractive force between gas particles, they slide past one another easily. Because liquids and gases flow they are known as fluids

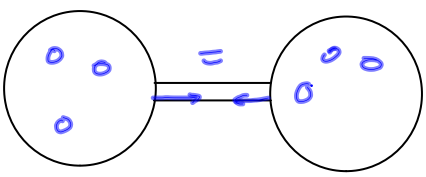
3) Low density - the density of a substance in the gaseous state is about 1/1000 the density of the same substance in the solid or liquid form

4) Compressibility - gases can be compressed. Initially the particles are far apart, but when a gas is compressed they come close together

5) Diffusion and Effusion - Gases spread out and mix with one another even if they are not stirred

   Effusion: process by which gas particles pass through a tiny opening

   Diffusion: the spontaneous mixing of the particles of 2 substances caused by their random motion High concentration to low concentration

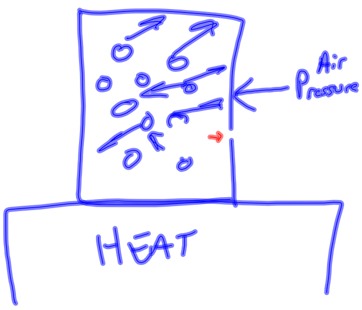


Real Gases

- A gas that does not behave completely according to the assumptions of the kinetic molecular theory

   - real gases occupy space and exert an attractive force on one another

Ideal Gases Do NOT Exist



Pressure

mass gravity
$F = mg$

Pressure (P) - the force per unit area on a surface

pressure = force/area $P = \frac{F}{A}$

Unit for force is the Newton (N)

the force that will increase the speed of one kilogram mass by one meter per second each second it is applied

$F = mg$

$F = \frac{kg\,m}{s^2} = \underline{\underline{1N}}$

Gravity = 9.8 m/s^2

How much force does a 50 kg person exert on the floor?

$P = \frac{F}{A}$ $F = mg$

$F = mg \longrightarrow F = (50.kg)(9.8\,m/s^2)$

$F = 490\,N$

How much pressure does this person exert if they occupy 3.5 m^2 of area on the floor?

$P = \frac{F}{A} \longrightarrow P = \frac{490N}{3.5m^2}$

$P = 140\,\frac{N}{m^2} \rightarrow \underline{\underline{Pa}}$

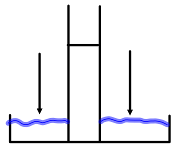
$KE = \frac{1}{2}mv^2$

$F = mg$

$P = \frac{F}{A}$

Measuring Pressure

Barometer - a device used to measure atmospheric pressure



Units of Pressure

mm of Hg = millimeters of mercury (used with mercury barometer)

torr = named after Torricelli who invented the barometer

1 torr = 1 mm Hg

atmospheres (atm) = average atmospheric pressure at sea level and 0^o C

1 atm = 760 mm Hg
= 760 torr
= 1.01325 x 10^5 Pa
= 101.325 KPa

Pascal (Pa) = SI unit of pressure, named after Blaise Pascal

- the pressure exerted by a force of one newton (N) acting on an area of 1 m^2

- Units = N/m^2 $\rightarrow$ Pa

The average atmospheric pressure in Denver CO is 0.830 atm. Express this pressure in a) mm Hg and b) Pa

$\frac{0.830\,\cancel{atm} \mid 760\,mm\,Hg}{\mid 1\,\cancel{atm}} = \boxed{631\,mm\,Hg}$

$\frac{0.830\,\cancel{atm} \mid 1.01325 \times 10^5\,Pa}{\mid 1\,\cancel{atm}} = 84099. \quad \boxed{84100\,Pa}$

Convert 570. torr to atm and Pa

$\frac{570.\,\cancel{torr} \mid 1\,atm}{\mid 760\,\cancel{torr}} = 0.750\,atm$

$\frac{570.\,\cancel{torr} \mid 1\,\cancel{atm} \mid 1.01325 \times 10^5\,Pa}{\mid 760\,\cancel{torr} \mid 1\,\cancel{atm}}$

$\frac{570.\,\cancel{torr} \mid 1.01325 \times 10^5\,Pa}{\mid 760\,torr}$

Convert a pressure of 1.75 atm to mm Hg

$\frac{1.75\,\cancel{atm} \mid 760\,mm\,Hg}{\mid 1\,\cancel{atm}}$

The Gas Laws

- simple mathematical relationships between volume, temperature, pressure, and the amount of a gas

Boyle's Law

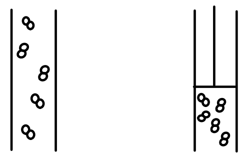
- Gas law that relates volume and pressure of a gas (temp remains constant)

- Boyle realized that if you double the pressure of a gas, you decreased the volume by 1/2. If you tripled the pressure of a gas, you decreased the volume to 1/3 its orignial. If you decreased the pressure of a gas by 1/2 you doubled the volume of the gas

- Pressure and volume are inversely related

P↑ V↓ P↓ V↑

- pressure is caused by gas particles running into the side of their container. If volume is decreased, the gas particles will hit the container more causing the pressure to go up



$P \cdot V = K$

$P \cdot V = K$

* $\boxed{P_1 V_1 = P_2 V_2}$ *

P = pressure

V = volume

$KE = \frac{1}{2}mv^2$

$F = mg$

$P = \frac{F}{A}$

$P_1V_1 = P_2V_2$

A sample of oxygen has a volume of 150. mL when its pressure is 0.947 atm. What will the volume of the gas be at a pressure of 0.987 atm, if the temperature remains constant.

$V_1 = 150.\ mL$

$P_1 = 0.947\ atm$

$V_2 = ?\ mL$

$P_2 = 0.987\ atm$

$P_1V_1 = P_2V_2$

$\frac{P_1V_1}{P_2} = V_2$

$\frac{(0.947\ \cancel{atm})(150.\ mL)}{(0.987\ \cancel{atm})} = V_2$

$\boxed{V_2 = 144\ mL}$

Balloon has volume of 500 mL at a pressure of 1 atm. When the pressure is 0.5 atm, what is the volume? constant temp.

$V_1 = 500\ mL$

$P_1 = 1\ atm$

$P_2 = 0.5\ atm$

$V_2 = ?\ mL$

$\frac{P_1V_1}{P_2} = V_2$

$\frac{(1\ atm)(500\ mL)}{0.5\ atm} = V_2$

$\boxed{V_2 = 1000\ mL}$

A gas has a pressure of 1.26 atm and occupies a volume of 7.40 L. If the gas is compressed to a volume of 2.93 L what is the final pressure in torr?

$P_1 = 958\ torr$

$V_1 = 7.40\ L$

$P_2 = ?\ torr$

$V_2 = 2.93\ L$

$\frac{P_1V_1}{V_1} = P_2$

$1.26\ \cancel{atm} \times \frac{760\ torr}{1\ \cancel{atm}}$

$\frac{(958\ torr)(7.40\ L)}{(2.93\ L)} = P_2$

$\boxed{P_2 = 2420\ torr}$

Charles' Law

- Gas law that relates volume and temperature of a gas when the temperature is measure in Kelvin (pressure remains constant)

- Charles that gases expand when heated. He also saw that gases change in their volume by 1/273 for each degree Celsius (heating or cooling)

- If a gas is cooled to -273 °C it would have 0 volume. This is impossible. This lead to the development of the Kelvin temperature scale

Kelvin scale

all motion stops

- based on start temperature of -273.15 °C (absolute zero)
- -273.15 °C = absolute zero and is considered 0 on Kelvin scale
- $\boxed{K = 273.15 + {}^{\circ}C}$
- The units for Kelvin is Kelvin (K) NOT °K

- Temperature (in Kelvin) and volume are directly related

T↑ V↑ T↓ V↓

$$\boxed{\frac{V_1}{T_1} = \frac{V_2}{T_2}}$$

A sample of neon gas occupies a volume of 752 mL at 25°C. What volume will the gas occupy at 50°C at constant pressure

$V_1 = 752\ mL$

$T_1 = 298\ K$

$V_2 = ?\ mL$

$T_2 = 323\ K$

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

$V_2 = \frac{V_1T_2}{T_1}$

$V_2 = \frac{(752\ mL)(323\ \cancel{K})}{(298\ \cancel{K})}$

$\boxed{V_2 = 815\ mL}$

A balloon filled with helium has a volume of 2.75 L at 20 °C. The volume of the balloon decreases to 2.46 L after it is placed outside. What is the outside temperature in Kelvin?

$V_1 = 2.75\ L$

$T_1 = 293\ K$

$V_2 = 2.46\ L$

$T_2 = ?\ K$

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

$\frac{V_2T_1}{V_1} = T_2$

$\frac{(2.46\ \cancel{L})(293\ K)}{2.75\ \cancel{L}} = T_2$

$\boxed{T_2 = 262\ K}$

Gay-Lussac's Law

- Gas law that relates temperature (in Kelvin) and pressure of a gas (volume remains constant)

- Pressure and temperature are directly related

P↑ T↑ P↓ T↓

- pressure is caused by gas particles running into the side of their container. If temperature is increased the gas particles will move faster and collide more with the walls of the container more causing a higher pressure

$$\boxed{\frac{P_1}{T_1} = \frac{P_2}{T_2}}$$

The pressure in an aerosol can is 3.00 atm at 25°C. What would the gas pressure be at 52 °C? volume is constant

$\frac{3.00 \text{ atm}}{298 \text{ K}} = \frac{\text{atm}}{325 \text{K}}$

$= 3.27 \text{ atm}$

A sample of helium is at 1.20 atm at 22 °C. At what temperature will the helium reach a pressure of 2.00 atm?

$\frac{1.20 \text{atm}}{295 \text{K}} = \frac{2.00 \text{atm}}{T_2}$

$T_2 = 492 \text{K}$

Combined Gas Law

- A gas law that expresses the relationship between pressure, volume, and temperature
- A gas sample often goes through changes in volume, pressure, and temperature all at the same time
- Boyle's law, Charle's law, and Gay-Lussac's law can be combined into a single expression.

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

STP = standard temperature and pressure

T = 0 °C P = 1 atm
273.15 K 101.325 kPa

A balloon has a volume of 50.0 L at 25°C and 1.08 atm. What volume will it have at 0.855 atm and 10. °C?

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

$V_1 = 50.0 \text{L}$ $V_2 = ?\text{L}$

$T_1 = 298 \text{K}$ $T_2 = 283 \text{K}$

$P_1 = 1.08 \text{atm}$ $P_2 = 0.855 \text{atm}$

$\frac{(283 \cancel{K})(1.08 \cancel{atm})(50.0 \text{L})}{(0.855 \cancel{atm})(298 \cancel{K})} = \boxed{V_2 = 60.0 \text{L}}$

A 700. mL gas sample at STP is compressed to a volume of 200. mL and the temperature is increased to 30.0 °C. What is the new pressure of the gas in Pa?

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

$V_1 = 700. \text{mL}$ $V_2 = 200. \text{mL}$

$P_1 = 1.01325 \times 10^5 \text{Pa}$ $P_2 = ? \text{Pa}$

$T_1 = 273.15 \text{K}$ $T_2 = 303.2 \text{K}$

$P_2 = \frac{(303.2 \cancel{K})(1.01325 \times 10^5 \text{Pa})(700. \cancel{mL})}{(200. \cancel{mL})(273.15 \cancel{K})}$

$P_2 = 393652$

$\boxed{P_2 = 394000 \text{ Pa}}$

$P_2 = 3.89 \text{atm}$

$3.89 \text{atm} \times \frac{1.01325 \times 10^5 \text{Pa}}{1 \text{atm}} =$

Dalton's Law of Partial Pressures

the total pressure of a mixture of gases is equal to the sum of the partial pressures of the component gases

- Discovered that in the absence of a chemical reaction, the pressure of a gas mixture is the sum of the individual pressures of each gas alone

ex. A 1.0 L container has O_2 gas in it at a pressure of .12 atm

Another 1.0 L container has N_2 gas in it at a pressure of .12 atm

The gas samples are combined in a 1.0 L container and the pressure of the gas mixture is .24 atm

- The pressure of each gas in a mixture is the partial pressure of that gas

$$\boxed{P_T = P_1 + P_2 + P_3 + P_4 + ...}$$

P_T = 5 atm, 4 atm, 8 atm, 7 atm

$P_T = 24 \text{ atm}$

Gases Collected by Water Displacement

- Gases in the lab are often collected over water because the gas will displace the water which is more dense. The collected gas is mixed with some water vapor so a *water-vapor pressure* is exerted as well as the pressure exerted by the gas that is collected

According to Dalton's law of partial pressures

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

- to find P of the water at various temperatures, you can look it up on table A-8 in the book (pg 899)
- The atmospheric pressure will be given to you

Oxygen gas is collected over water at 20°C and a pressure of 731.0 torr. What is the partial pressure of the oxygen gas that is collected?

$P_{atm} = P_{gas} + P_{H_2O}$ 1 mm Hg = 1 torr

$731.0 \text{ torr} = P_{gas} + 17.5 \text{ torr}$

$P_{gas} = 731.0 \text{ torr} - 17.5 \text{ torr}$

$P_{gas} = 713.5 \text{ torr}$

Hydrogen is collected over water at 25.0 °C. What is the partial pressure of the hydrogen gas if the barometric pressure is 750.0 mm Hg

$750.0 \text{ mm Hg} = P_{gas} + 23.8 \text{ mm Hg}$

$P_{gas} = 726.2 \text{ mm Hg}$

$P_1 \cdot V_1 = \underline{26}$

$\underbrace{P_2 \cdot V_2}_{\text{inverse relationship}} = \underline{26}$

$P_1 V_1 = P_2 V_2$ Boyles law

$\underbrace{\frac{V_1}{T_1}}_{\text{Direct}} = \underline{26}$ $\frac{V_2}{T_2} = \underline{26}$

$\frac{V_1}{T_1} = \frac{V_2}{T_2}$ charles law

$P_1 T_1 = 14$ $P_2 T_2 = 29$

$\frac{P_1}{T_1} = 26$ $\frac{P_2}{T_2} = 26$

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