GASES

- Gases differ from the other two phases of matter in many ways:

  - They have very low viscosity (resistance to flow), so they flow from one place to another very easily.

  - They will take the volume of their container. In other words, gas volumes are variable.

  - They are the least dense of all three phases.

  - Most gases are transparent, and many are invisible.

thermal expansion!

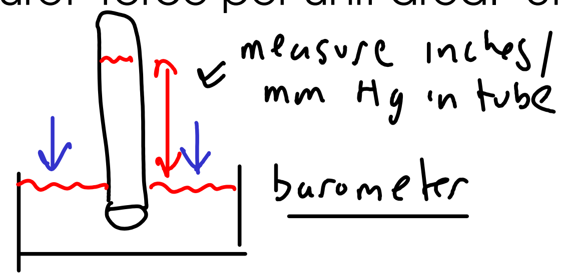
  - Gases show a much larger change of volume on heating or cooling than the other phases.

  - Gases react to changes in temperature and pressure in a very similar way. This reaction often <u>does not depend</u> on what the gas is actually made of.

KINETIC THEORY

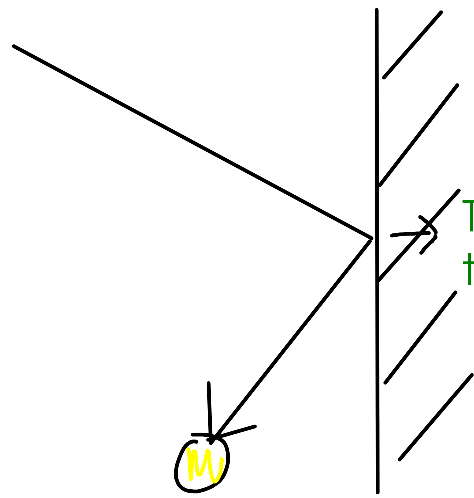
- is a way to explain the behavior of gases.
- views the properties of gases as arising from them being molecules in motion.

- Pressure: force per unit area. Units: Pascal, bar, mm Hg, in Hg, atm, etc.



$$\boxed{760\ \text{mm Hg} = 1\ \text{atm}}$$

- According to kinetic theory, pressure is caused by collisions of gas molecules with each other and the walls of the container the gas is in.



This collision exerts a force on the wall!

- Temperature:

  - a measure of the average kinetic energy of the molecules of the gas

$$E_K = \frac{1}{2} m v^2$$

velocity

mass

  - The faster the gas molecules move, the higher the temperature!

  - The temperature scales used when working with gases are ABSOLUTE scales.

  - ABSOLUTE: scales which have no values less than zero.

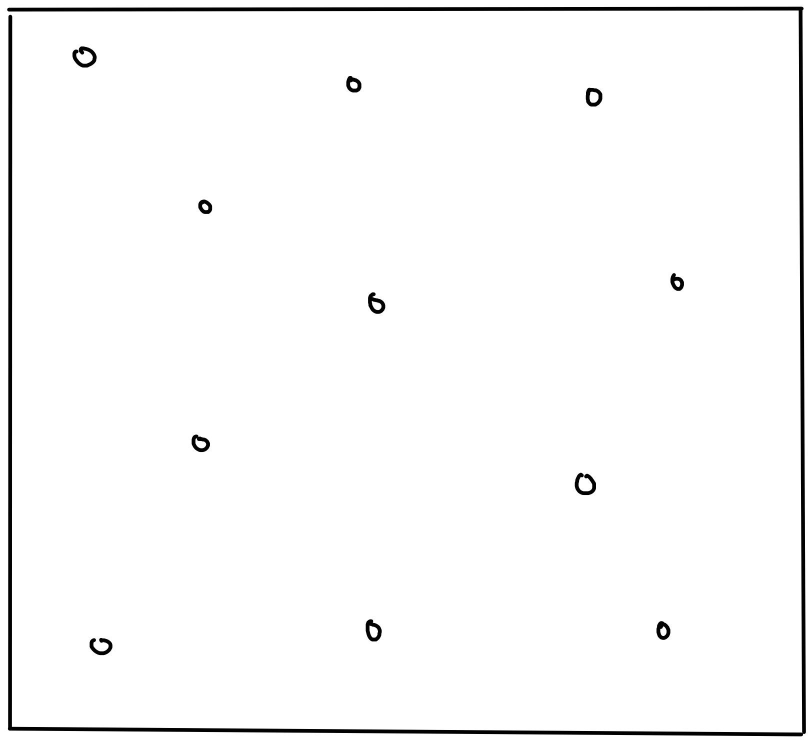
  - KELVIN: metric absolute temperature scale.

$$\boxed{K = 273.15 + {}^\circ C}$$

Quick comparison of temperature scales!

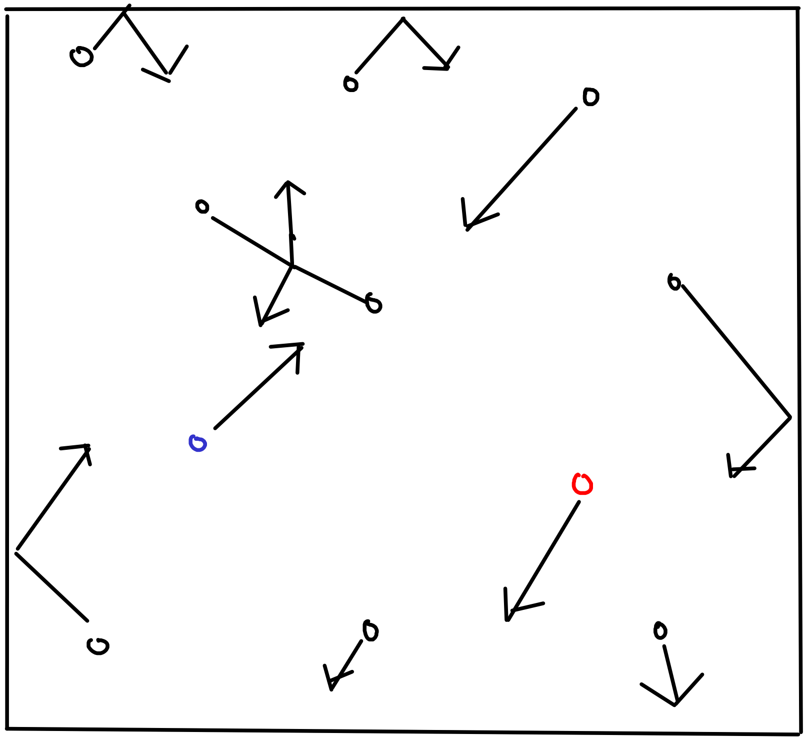
°F	°C	K	
212	100	373	Water boils
77	25	298	Room temperature
32	0	273	Water freezes
-460	-273	0	Absolute zero!

THE KINETIC PICTURE OF GASES



1 Gas molecules are small compared to the space between the gas molecules!

LOW DENSITY!



2 Gas molecules are constantly in motion. They move in straight lines in random directions and with various speeds.

3 Attractive and repulsive forces between gas molecules are so small that they can be neglected except in a collision.

- Each gas molecule behaves independently of the others.

4 Collisions between gas molecules and each other or the walls are ELASTIC.

5 The average kinetic energy of gas molecules is proportional to the absolute temperature.

How does this picture explain the properties of gases?

- Gases expanding to fill their container? Agrees with kinetic picture, since gas molecules are independent
- Thermal expansion of gas at constant pressure? Agrees, because the container has to EXPAND to keep the pressure (from collisions) constant when the gas molecules move faster.
- Pressure increases with temperature at constant volume: Agrees, because the number and force of collisions increases with molecular speed.

GAS LAWS

- were derived by experiment long before kinetic theory, but agree with the kinetic picture!

Boyle's Law:

$$PV = \text{constant}$$

True at constant temperature

$$P_1 V_1 = \text{constant} \qquad P_2 V_2 = \text{constant}$$

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

True at constant temperature

Charles's Law:

$$\frac{V}{T} = \text{constant}$$

True at constant pressure, and using ABSOLUTE temperature

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

True at constant pressure, and using ABSOLUTE temperature

Combined gas law: $$\frac{PV}{T} = \text{constant}$$

Must use ABSOLUTE temperature units!

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

Must use ABSOLUTE temperature units!

amount (moles) of gas must be constant!

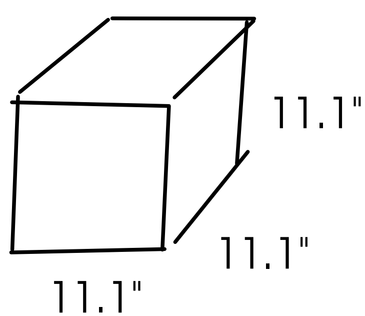
Avogadro's law:

- a mole of any gas at the same conditions has the same volume.

1 mol gas molecules @ 0°C and 1 atm "STP"

Standard Temperature and Pressure

volume = 22.4 L



= 22.4 L

Ideal gas law:

$$\frac{PV}{T} = \text{constant}$$

... but this constant actually depends on the amount of gas!

$$= n \times \text{"}R\text{"}$$

The ideal gas constant.

$$0.08206 \ \frac{L \cdot atm}{mol \cdot K}$$

... combining this together ...

$$\frac{PV}{T} = nR$$

$$\boxed{PV = nRT}$$

P = pressure atm

V = volume L

T = ABSOLUTE temperature K

R = ideal gas constant

n = number of moles of gas molecules

CHEMICAL CALCULATIONS WITH THE GAS LAWS

$FW_{NaHCO_3} = 84.007\ g/mol$

$$H_2SO_4(aq) + 2NaHCO_3(s) \rightarrow 2H_2O(l) + 2CO_2(g) + Na_2SO_4(aq)$$

Given 25.0 g of sodium bicarbonate and sufficient sulfuric acid, what volume of carbon dioxide gas would be produced at 25.0 C and 0.950 atm pressure?

1 - Convert mass sodium bicarbonate to moles (formula weight of sodium bicarbonate)
2 - Convert moles sodium bicarbonate to moles carbon dioxide using chemical equation
3 - Convert moles carbon dioxide to volume using the ideal gas equation

$84.007\ g\ NaHCO_3 = mol\ NaHCO_3$ | $2\ mol\ NaHCO_3 = 2\ mol\ CO_2$

$$25.0\ g\ NaHCO_3 \times \underset{①}{\frac{mol\ NaHCO_3}{84.007\ g\ NaHCO_3}} \times \underset{②}{\frac{2\ mol\ CO_2}{2\ mol\ NaHCO_3}} = 0.297594\ mol\ CO_2$$

$PV = nRT$

$\downarrow$

$V = \frac{nRT}{P}$

$n = 0.297594\ mol\ CO_2$ $\quad P = 0.950\ atm$

$R = 0.08206\ \frac{L \cdot atm}{mol \cdot K}$ $\quad V = ???$

$T = 25.0°C = 298.2\ K$

$$③\ V = \frac{(0.297594\ mol\ CO_2)(0.08206\ \frac{L \cdot atm}{mol \cdot K})(298.2\ K)}{(0.950\ atm)} = \boxed{7.67\ L\ CO_2}$$

What volume would the gas in the last example problem have at STP?

STP: 0°C, 1 atm

STP: "Standard Temperature and Pressure" (0 C and 1 atm)

$P_1 = 0.950$ atm

$V_1 = 7.67$ L

$T_1 = 298.2$ K

$P_2 = 1$ atm

$V_2 = ???$

$T_2 = 273.2$ K

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

... we can use the combined gas law here if we wish.

$$\frac{P_1 V_1 T_2}{T_1 P_2} = V_2 = \frac{(0.950 \text{ atm})(7.67 \text{ L})(273.2 \text{ K})}{(298.2 \text{ K})(1 \text{ atm})} = \boxed{6.68 \text{ L } CO_2 \text{ at STP}}$$

Alternate solution: Since we knew the number of moles, we could also use the ideal gas equation to solve this problem. You'll get the same answer doing it that way.

$$\boxed{FW_{NH_4NO_3} = 80.0434\ g/mol}$$

$$2\,NH_4NO_3(s) \longrightarrow 2\,N_2(g) + O_2(g) + 4\,H_2O(g)$$

At 300^{0}C, ammonium nitrate violently decomposes to produce nitrogen gas, oxygen gas, and water vapor. What is the total volume of gas that would be produced at 1.00 atm by the decomposition of 15.0 grams of ammonium nitrate?

To simplify this problem, calculate the TOTAL MOLES OF GAS directly!

1 - Convert mass ammonium nitrate to moles ammonium nitrate using formula weight
2 - Convert moles ammonium nitrate to TOTAL moles gas using chemical equation
3 - Convert moles gas to volume using ideal gas equation

80.0434 g NH_4NO_3 = mol NH_4NO_3 | 2 mol NH_4NO_3 = 7 mol gas

$$15.0\ g\ NH_4NO_3 \times \frac{mol\ NH_4NO_3}{80.0434\ g\ NH_4NO_3} \times \frac{7\ mol\ gas}{2\ mol\ NH_4NO_3} = 0.655894\ mol\ gas$$

$$V = \frac{nRT}{P}$$

$n = 0.655894$ mol gas, $R = 0.08206\ \frac{L \cdot atm}{mol \cdot K}$, $V = ???$

$T = 300.°C = 573\ K$, $P = 1.00$ atm

$$V = \frac{(0.655894\ mol\ gas)\left(0.08206\ \frac{L\cdot atm}{mol\cdot K}\right)(573\ K)}{(1.00\ atm)} = \boxed{30.8\ L\ gas}$$

... What if this gas was trapped in a 250 mL container? (This is a constant VOLUME problem)

$$P = \frac{nRT}{V} = \frac{(0.655894\ mol\ gas)\left(0.08206\ \frac{L\cdot atm}{mol\cdot K}\right)(573\ K)}{(0.250\ L)}$$

$$\doteq 123\ atm$$

REAL GASES

- The empirical gas laws (including the ideal gas equation) do not always apply.

    - The gas laws don't apply in situations where the assumptions made by kinetic theory are not valid.

        - When would it be FALSE that the space between gas molecules is much larger than the molecules themselves?

            - at high pressure, molecules would be much closer together!

        - When would it be FALSE that attractive and repulsive forces would be negligible?

            - at high pressure, attractions and repulsions should be stronger!

            - at low temperature, attractions and repulsions have a more significant affect on the paths of molecules

-The gas laws are highly inaccurate near the point where a gas changes to liquid!

- In general, the lower the pressure and the higher the temperature, the more IDEAL a gas behaves.

van der Waals equation

- an attempt to modify PV = nRT to account for several facts.
  - gas molecules actually have SIZE (they take up space)
  - attractive and repulsive forces

$$PV = nRT$$ Ideal gas equation

$$\Big(\underbrace{P + \frac{n^2a}{V^2}}\Big)\Big(\underbrace{V - nb}\Big) = nRT$$ van der Waals equation

$V - nb$: attempts to account for molecular size

$\frac{n^2a}{V^2}$: attempts to account for attractive / repulsive forces

* "a" and "b" are experimentally determined parameters that are different for each gas. p 211

He: $a = 0{,}0346$, $b = 0{,}0238$ tiny, no special attractive forces

H_2O: $a = 5.537$, $b = 0.03049$ small, but strong attractions between moleculres

CH_3CH_2OH: $a = 12.56$ $b = 0{,}08710$ larger, and strong attractions between molecules