# 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!

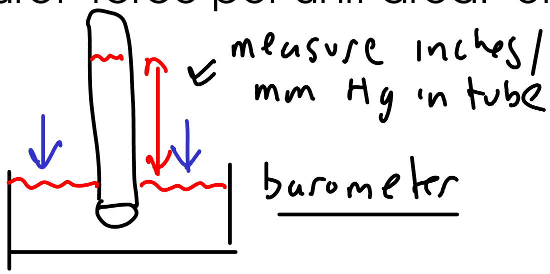
  - Gases show a much larger change of volume on heating or cooling than the other phases.

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

# KINETIC THEORY

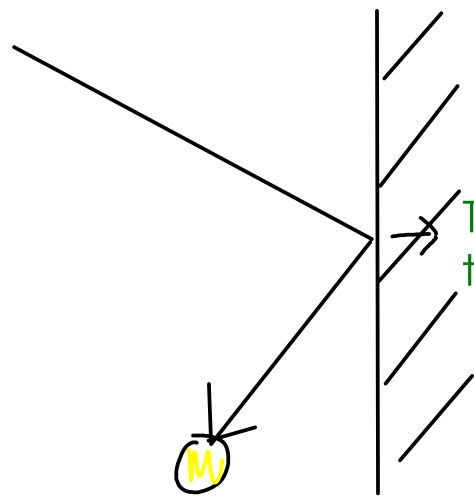
- is a way to explain the behavior of gases.
- views the properties of gases as arising from them being molecules in motion.

- 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!

136- 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.

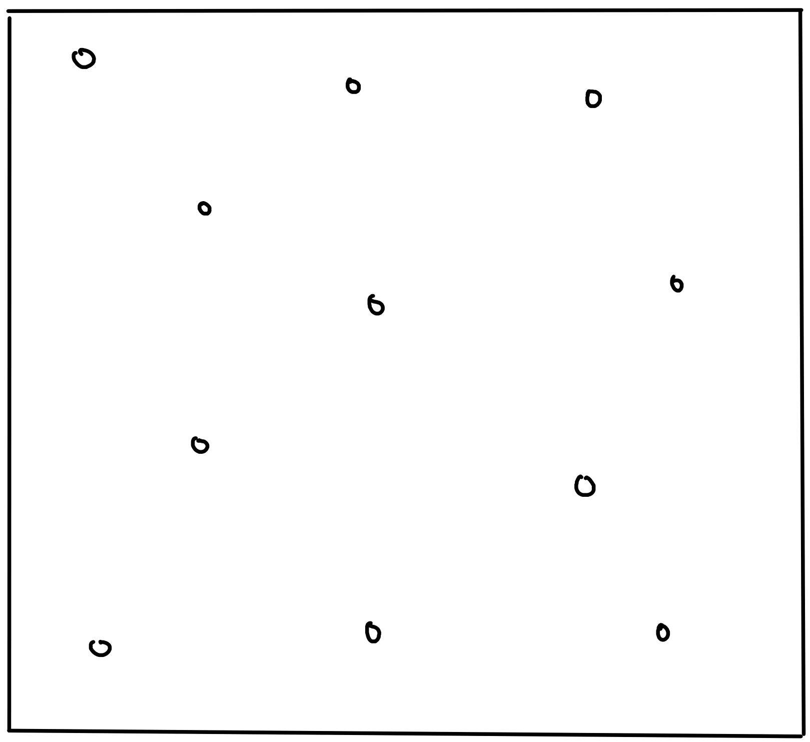
- KELVIN: metric absolute temperature scale.

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

Quick comparison of temperature scales!

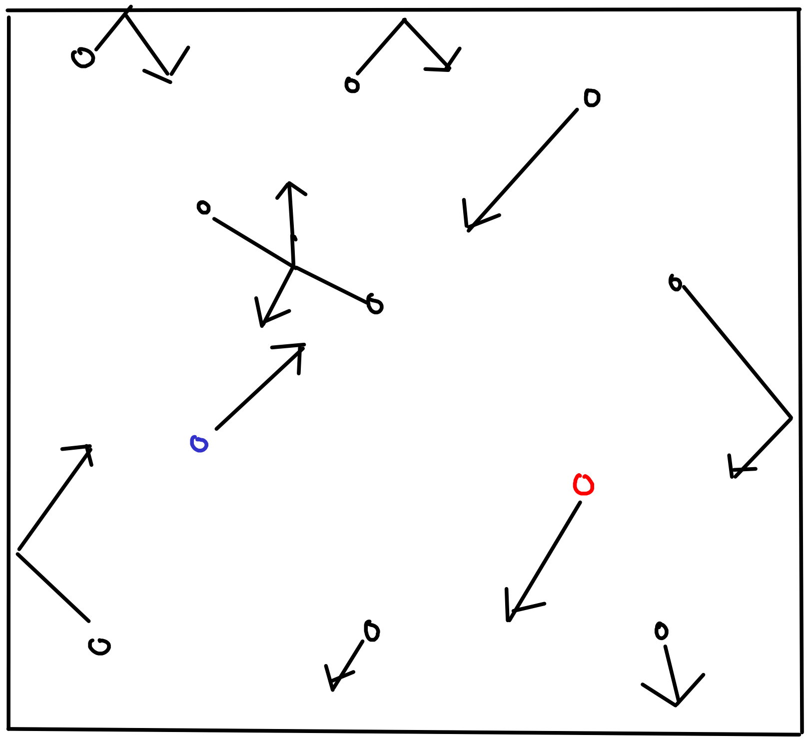
| °F | °C | K | |
|---|---|---|---|
| 212 | 100 | 373 | Water boils |
| 77 | 25 | 298 | Room temperature |
| 32 | 0 | 273 | Water freezes |
| -460 | -273 | 0 | Absolute zero! |

# 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

$$P_1V_1 = \text{constant} \qquad P_2V_2 = \text{constant}$$

$$\boxed{P_1V_1 = P_2V_2}$$

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!

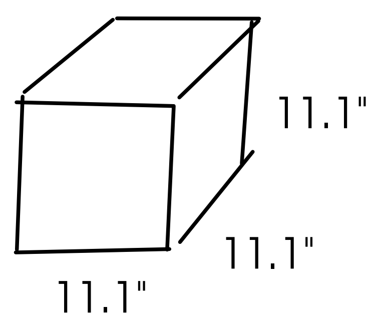
Avogadro's law:

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

1 mol gas molecules @ 0°C and 1 atm

volume = 22.4 L

"STP"
Standard Temperature and Pressure




= 22.4 L

Ideal gas law:

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

... 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 these 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

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

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

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 25.0 grams of sodium bicarbonate to moles using FORMULA WEIGHT.
2 - Convert moles sodium bicarbonate to moles carbon dioxide using chemical equation.
3 - Convert moles carbon dioxide to VOLUME using IDEAL GAS EQUATION.

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

$$25.0\ g\ NaHCO_3 \times \underbrace{\frac{mol\ NaHCO_3}{84.007\ g\ NaHCO_3}}_{1} \times \underbrace{\frac{2\ mol\ CO_2}{2\ mol\ NaHCO_3}}_{2} = 0.297594281\ mol\ CO_2$$

3 $PV = nRT$

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

$n = 0.297594281\ mol\ CO_2$ $\quad T = 25.0°C = 298.2\ K$

$R = 0.08206\ \dfrac{L \cdot atm}{mol \cdot K}$ $\quad P = 0.950\ atm$

$$V = \frac{(0.297594281\ mol\ CO_2)\left(0.08206\ \frac{L \cdot atm}{mol \cdot K}\right)(298.2\ K)}{(0.950\ atm)} = \boxed{7.67\ L\ CO_2}$$

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

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

Let's use the COMBINED GAS LAW to solve this one.

$$\frac{P_1V_1}{T_1} = \frac{P_2V_2}{T_2} \quad ; \quad \frac{P_1V_1T_2}{T_1P_2} = V_2$$

$P_1 = 0.950$ atm

$V_1 = 7.67$ L

$T_1 = 298.2$ K

$P_2 = 1$ atm

$T_2 = 273.2$ K

$V_2 = ?$

$$V_2 = \frac{(0.950 \text{ atm})(7.67 \text{ L})(273.2 \text{ K})}{(298.2 \text{ K})(1 \text{ atm})} = \boxed{6.67 \text{ L } CO_2 \text{ @ STP}}$$

Alternate solution: Since we know the number of moles of gas (from the previous problem), we could just use the ideal gas equation with the STP temperature and pressure. We'll get the same answer if we do it this way. (Try it!)

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

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

At 300°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, we'll calculate the TOTAL MOLES OF GAS instead of looking at each product individually.

1 - Convert 15.0 grams amonium nitrate to moles using FORMULA WEIGHT.
2 - Convert moles ammonoium nitrate to TOTAL MOLES OF GAS using chemical equation.
3 - Convert TOTAL MOLES OF GAS to volume using the ideal gas equation.

$80.0434\ g\ NH_4NO_3 = mol\ NH_4NO_3$ | $2\ mol\ NH_4NO_3 = 7\ mol\ gas$ $(2+1+4)$

$$15.0\ g\ NH_4NO_3 \times \underset{1}{\frac{mol\ NH_4NO_3}{80.0434\ g\ NH_4NO_3}} \times \underset{2}{\frac{7\ mol\ gas}{2\ mol\ NH_4NO_3}} = 0.655894 1774\ mol\ gas$$

3 $V = \frac{nRT}{P}$ | $n = 0.6558941774\ mol\ gas$ $T = 300°C = 573\ K$

$R = 0.08206\ \frac{L \cdot atm}{mol \cdot K}$ $P = 1.00\ atm$

$$V = \frac{(0.6558941774\ mol\ gas)\left(0.08206\ \frac{L\cdot atm}{mol\cdot K}\right)(573\ K)}{(1.00\ atm)} = \boxed{30.8\ L\ gas}$$ *

* More than 2000x the size of the original solid!

# 145 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

fast (high T) slow (low T)

-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.