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!

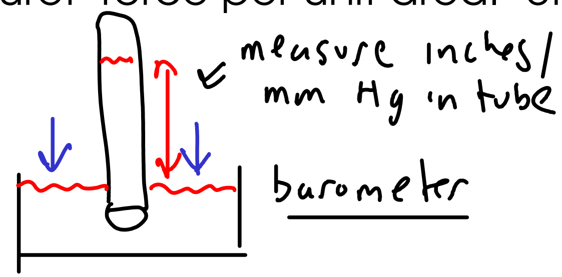
  - Gases show a much larger change of volume on heating or cooling than the other phases.

  - 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

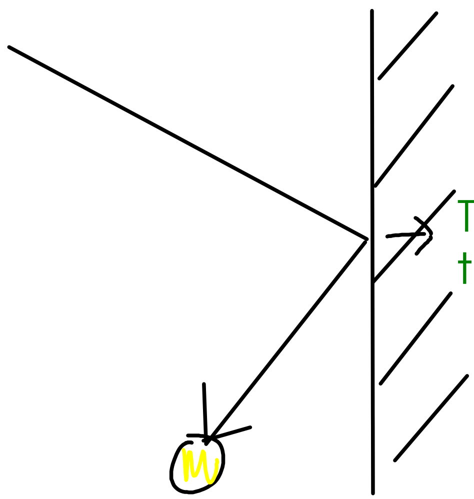
- is a way to explain the behavior of gases.
- views the properties of gases as arising from them being molecules in motion.

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



760 mm Hg = 1 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!

201 - 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.

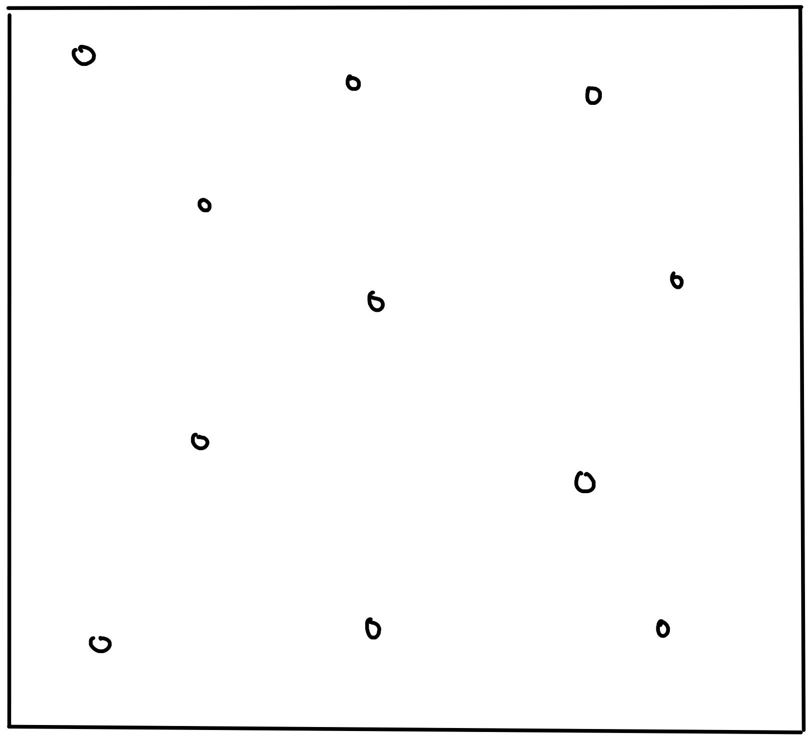
- KELVIN: metric absolute temperature scale.

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

Quick comparison of temperature scales!

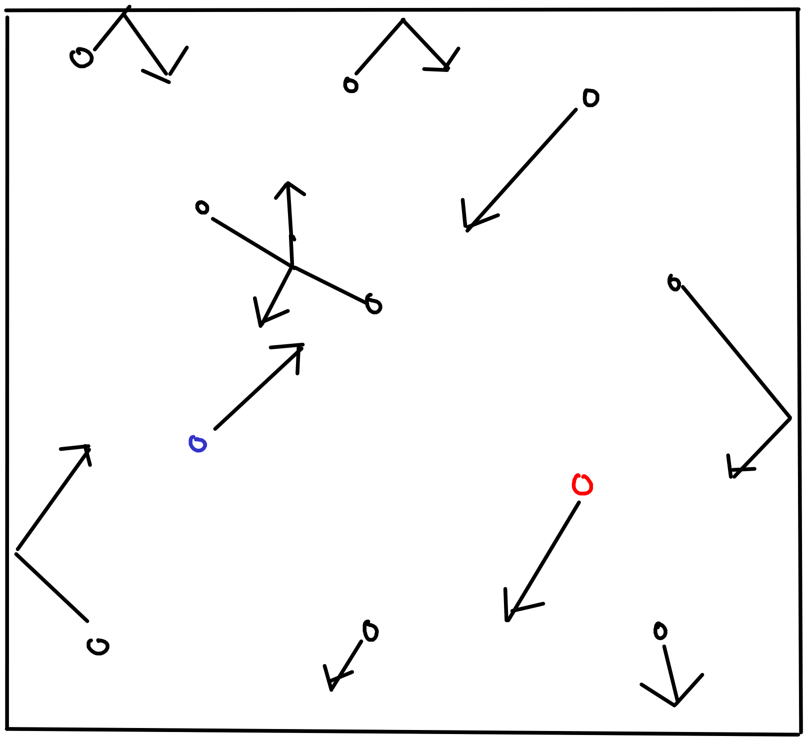
212	100	373	Water boils
77	25	298	Room temperature
32	0	273	Water freezes
-460	-273	0	Absolute zero!
°F	°C	K	

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!

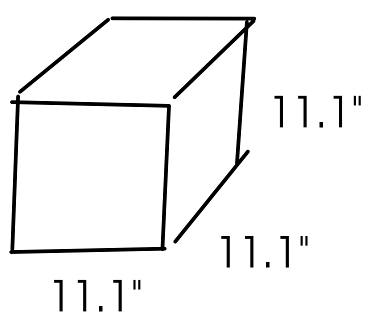
Avogadro's law:

- 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:

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

... 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

A balloon is taken from a room where the temperature is 27.0 C to a freezer where the temperature is -5.0 C. If the balloon has a volume of 3.5 L in the 27.0 C room, what is the volume of the balloon in the freezer. Assume pressure is constant.

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

$V_1 = 3.5 L$

$T_1 = 27.0°C = 300.2 K$

$V_2 = ?$

$T_2 = -5.0°C = 268.2 K$

$$\frac{3.5 L}{300.2 K} = \frac{V_2}{268.2 K}$$

$$\boxed{3.1 L = V_2}$$

2.25 L of nitrogen gas is trapped in a piston at 25.0 C and 1.00 atm pressure. If the piston is pushed in so that the gas's volume is 1.00 L while the temperature increases to 31.0 C, what is the pressure of the gas in the piston?

$P_1 = 1.00 atm$ $V_1 = 2.25 L$

$T_1 = 25.0°C = 298.2 K$

$P_2 = ?$ $V_2 = 1.00 L$

$T_2 = 31.0°C = 304.2 K$

$$\frac{P_1V_1}{T_1} = \frac{P_2V_2}{T_2}$$

$$\frac{(1.00 atm)(2.25 L)}{298.2 K} = \frac{P_2(1.00 L)}{304.2 K}$$

$$\boxed{2.30 atm = P_2}$$

Calculate the mass of 22650 L of oxygen gas at 25.0 C and 1.18 atm pressure.

O_2

O_2: 32.00 g O_2 = mol O_2

* Volume of a 10'x10'x8' room

We'd like to use PV=nRT ... (ideal gas equation)

1) Use PV=nRT to find the moles of oxygen gas.
2) Convert moles oxygen to mass using formula weight.

$PV = nRT$

$n = \frac{PV}{RT}$

$P = 1.18$ atm

$V = 22650$ L

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

$T = 25.0°C = 298.2$ K

① $n_{O_2} = \frac{(1.18\,atm)(22650\,L)}{\left(0.08206\frac{L\cdot atm}{mol\cdot K}\right)(298.2\,K)} = 1092.222357$ mol O_2

② 1092.222357 mol $O_2 \times \frac{32.00\,g\,O_2}{mol\,O_2} = \boxed{35\bar{0}00\,g\,O_2}$

35.0 kg O_2

~77 lb O_2