KINETIC THEORY

BEHAVIOR OF GASES

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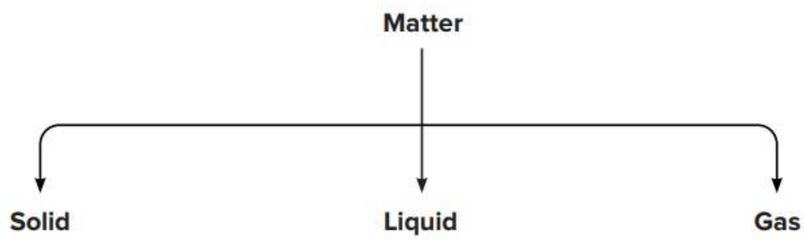
States of Matter

States of matter are different ways that stuff can be. There are three main types of states of matter, and they are:



1. Solid
Solids are things that keep their shape, size, and volume. The tiny particles in solids are stuck together really well, and the only movement they can do is to vibrate in place.

2. Liquid
Unlike solids, these things have a weaker connection between their particles, and they usually don't have a set shape. Liquids can change their shape to match the container they're put in.

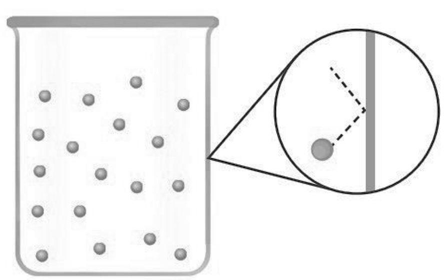
3. Gases
In gases, the tiny bits (atoms or molecules) are spread far apart because they don't stick together much. These gas particles can move freely in any direction.

Characteristics of Gas

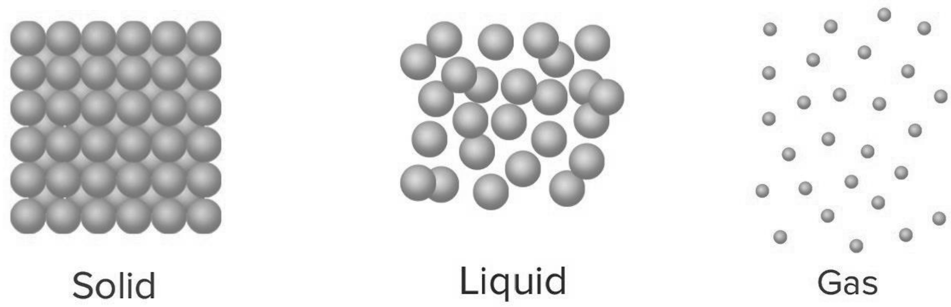
1. Gas doesn't have a fixed shape or size. It can spread out as much as it wants to and fill up all the space it can find.

Volume of the gas = Volume available to the gas = Volume of the container

2. Gas pushes on the container it's in by hitting the walls of the container with its tiny particles. This pushing is what we call pressure.

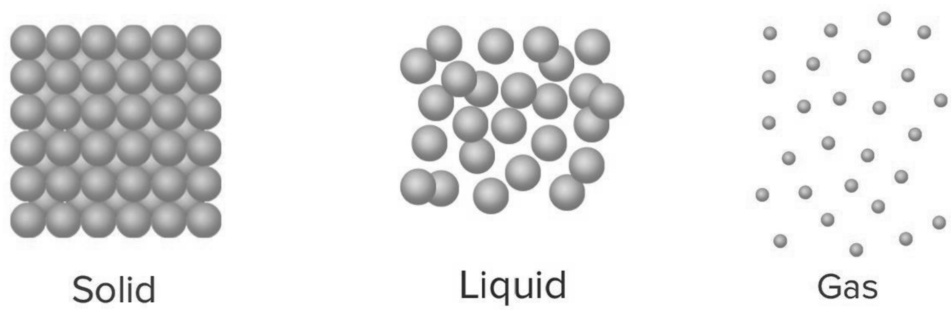


3. Intermolecular space:



Gases have the most empty space between their tiny particles.

4. Density:



Gases are less crowded because there are fewer particles in a given space, which means they have lower density.

5. Kinetic energy:

Gases have the most energy because their tiny particles are always moving around.

Macroscopic Properties of Gases

Pressure

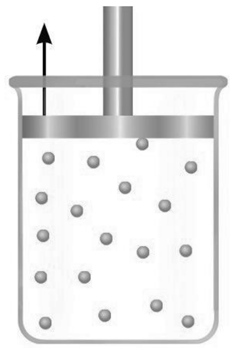
Gas particles are always in motion and they keep bumping into the walls of the container. When they bump into the walls, they transfer some of their push (momentum) to the walls. This push spread out over the wall's area is what we call pressure.

Volume

It means the atoms can move around in the container.

Imagine we have a container filled with gas, and there's a piston on top of it. When we move the piston a bit upwards, the gas fills the extra space that's now available at the top.

Volume of the gas = Volume available to the gas = Volume of the container



Temperature

When the temperature goes up, the gas molecules start moving faster.

Ideal Gas

Assumptions of an ideal gas

To think of a gas as ideal, we make a few assumptions:

- There are lots of tiny molecules in the gas, and they're all moving around randomly.
- The size of these molecules is so small that it doesn't really affect the overall gas volume.
- The forces between the gas molecules (if they exist) are very tiny or can be ignored.
- When the molecules hit each other or the container walls, they bounce back perfectly without losing any energy.

Gas Laws

Boyle's law

When you have a certain amount of ideal gas and you keep the temperature the same, if you push the gas harder (increase the pressure), it will take up less space (the volume gets smaller).

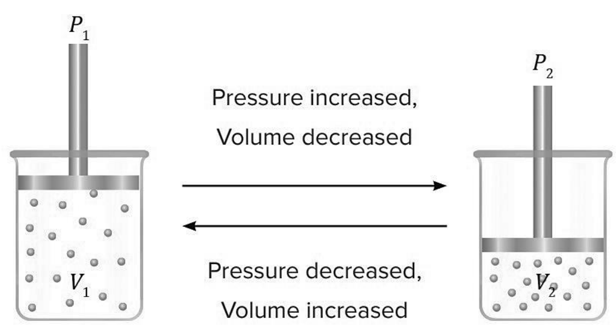
At constant temperature,

$V \propto \frac{1}{\mathrm{P}}$

$\Rightarrow V = \frac{k}{P}$

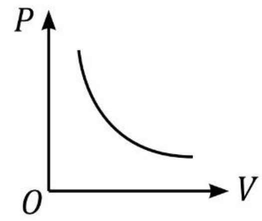
$\Rightarrow PV = k$

Where, k is a constant.



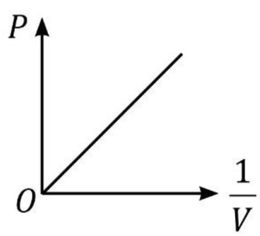
- **Pressure vs volume graph:**

  The result of multiplying the pressure and volume of a certain amount of gas always stays the same. Because of this, the graph that shows their relationship looks like a shape called a hyperbola, which is shown in the figure.

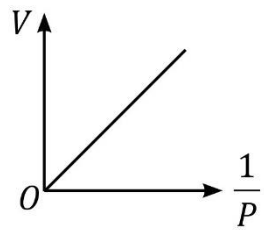


Some important graphs

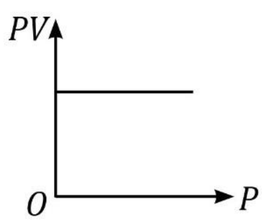
1. P vs $\frac{1}{v}$



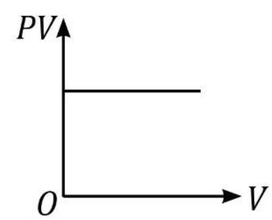
2. V vs $\frac{1}{P}$



3. PV vs P



4. PV vs V



Charles's law

If you keep the pressure the same, then the volume of a certain amount of gas will increase as its temperature goes up. They are directly related.

At constant pressure,

$V \propto T$

$\Rightarrow V = kT$

$\Rightarrow \frac{V}{T} = k$

Where, k is a constant.

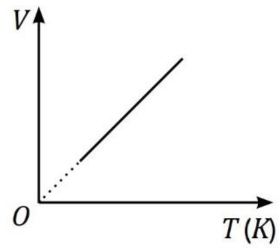
Let us assume that when a gas with initial volume V_1 at temperature T_1 is heated to a temperature T_2, the volume becomes V_2 at a constant pressure.

Therefore,

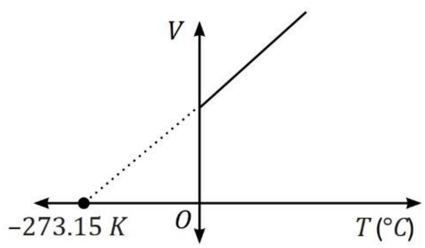
$$\frac{V_1}{T_1} = \frac{V_2}{T_2}$$

Volume vs temperature graph for different scales:

1. Kelvin scale: When it's as cold as it can possibly get, we imagine that the gas has no volume at all.



2. Centigrade scale: At a temperature of -273.15 degrees Celsius, the gas's volume is assumed to be zero.

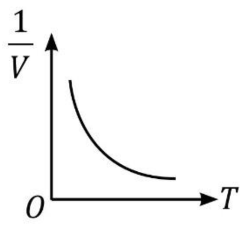


Note:

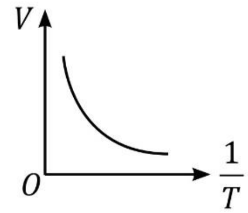
- At absolute zero temperature, $V_{gas} = 0$
- Straight line plots can be extrapolated to $V = 0$ (Practically impossible)
- Kelvin made a temperature scale where the lowest possible temperature is -273.15 degrees Celsius.

Some important graphs

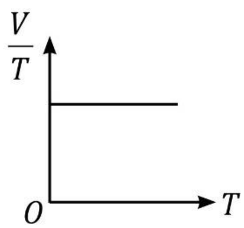
1. $\frac{1}{V}$ vs T



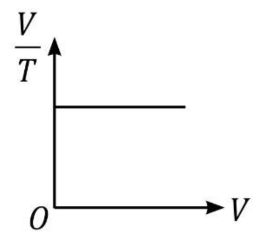
2. V vs $\frac{1}{T}$



3. $\frac{V}{T}$ vs V



4. $\frac{V}{T}$ vs V



Gay-Lussac's law

If you keep the volume the same, the pressure of a certain amount of gas goes up as its temperature gets higher. They're directly related.

At constant volume,

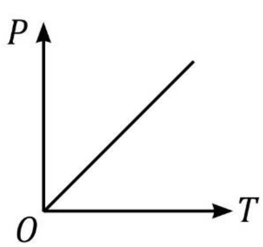
$P \propto T$

$\Rightarrow P = kT$

$\Rightarrow \frac{P}{T} = k$

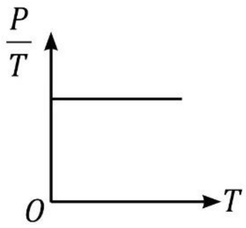
Where is constant.

Pressure vs temperature graph:

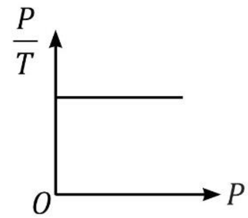


Some important graphs

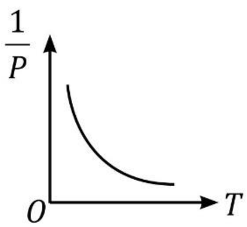
1. $\frac{P}{T}$ vs T



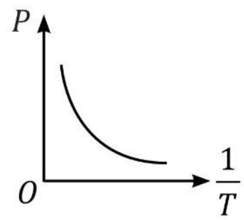
2. $\frac{P}{T}$ vs P



3. $\frac{1}{P}$ vs T



4. P vs $\frac{1}{T}$



Avogadro's law

When you have gases at the same temperature and pressure, if they take up the same amount of space (equal volumes), they also have the same number of molecules inside.

At constant pressure and temperature,

Volume of the gas (V) α Number of molecules in the gas (n)

$\Rightarrow V = kn$

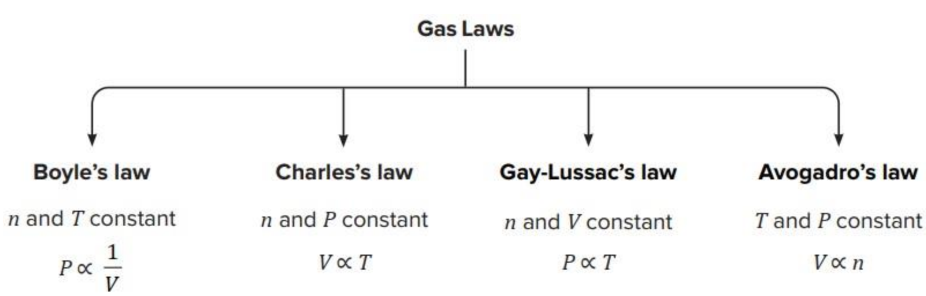
Where, is a constant.

Therefore,

$\frac{V_1}{n_1} = \frac{V_2}{n_2}$

The starting volume and the number of molecules are represented by V_1 and n_1, while the ending volume and number of molecules are shown as V_2 and n_2.

Note:



Ideal Gas Equation

The ideal gas equation helps us connect the pressure (P), volume (V), and temperature (T) of a particular state of an ideal gas.

By combining all four gas laws, we get,

PV α nT

⇒ PV = nRT

Where Universal gas constant, and, Number of moles.

Thus, the universal gas constant, $R = \frac{\mathrm{PV}}{\mathrm{nT}}$

Note: 1 mole gas = 6.023 × 1023 number of gas molecules

Value of universal gas constant (R)

At standard temperature pressure (STP) conditions,

$P = 1\ atm = 101325\ Nm^{-2}$

$n = 1\ mole$

$V = 22.4\ L = 22.4 \times 10^{-3}\ m^3$

$T = 273\ K$

We have,

$$R = \frac{PV}{nT}$$

Substituting all the values, we get the following:

$$R = \frac{101325\ Nm^{-2} \times 22.4 \times 10^{-3}\ m^3}{1\ mol \times 273\ K} = 8.314\ Nm\ mol^{-1}\ K^{-1} = 8.314\ J\, mol^{-1}\, K^{-1}$$

Also, 1 $cal = 4.18\ J$

$$R = 8.314 \times \frac{\frac{1}{4.18}\ cal}{mol\, K} = 1.98\ cal\, mol^{-1}\, K^{-1} \cong 2\ cal\, mol^{-1}\, K^{-1}$$