1. Details of Module and its structure

Module Detail	
Subject Name	Physics
Course Name	Physics 02(Physics Part 2,Class XI)
Module Name/Title	Unit 9, Module 1, Ideal Gas Chapter 13, Kinetic theory of gases
Module Id	keph_201301_econtent
Pre-requisites	Pressure, Volume, temperature and internal energy of a gas
Objectives	After going through this module, the learners will be able to: • Understand the concept of an ideal gas • Know the ideal gas equation and its applications • Appreciate the Gas laws Avogadro's law, Boyle's law, Charles' law, Gay-Lussac's law. Dalton's law of partial pressures. • Rationalise deviations from ideal gas behaviour of real gases
Keywords	Ideal gas equation, Avogadro's law, Boyle's law, Charles' law, Gay-Lussac's law. Dalton's law of partial pressures.

2. Development Team

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TABLE OF CONTENTS:

1. UNIT SYLLABUS

Unit 9: Behaviour of perfect gases and kinetic theory

Chapter13: Kinetic Theory:

Equation of state of a perfect gas, work done in compressing a gas. Kinetic theory of gases - assumptions, concept of pressure. Kinetic interpretation of temperature; rms speed of gas molecules; degrees of freedom, law of equi-partition of energy (statement only) and application to specific heat capacities of gases; concept of mean free path, Avogadro's number.

2. MODULE-WISE DISTRIBUTION OF UNIT SYLLABUS 4 Modules

Module 1	• **Microscopic and macroscopic view of interaction of heat with matter** • **Kinetic theory** • **Equation of state for a perfect gas** • **PV= nRT** • **Statement of gas laws – Boyles law, Charles’ law, pressure law, Dalton law of partial pressure**
Module 2	• **Kinetic theory of gases** • **Assumptions made regarding molecules in a gas** • **Concepts of pressure**

	• **Derivation for pressure exerted by a gas using the kinetic theory model**
Module 3	• **Kinetic energy of gas molecules** • **rms speed of gas molecules** • **Kinetic interpretation of temperature** • **Derive gas laws from kinetic theory**
Module 4	• **Degrees of freedom** • **Law of equipartition of energy** • **Specific heat capacities of gases depends upon the atomicity of its molecules** • **monoatomic molecule** • **diatomic molecules** • **polyatomic molecules**

MODULE 1

3. WORDS YOU MUST KNOW

Equation of state: an equation showing the relationship between the values of the pressure, volume, and temperature of a quantity of a particular substance.

Gauge Pressure: the amount by which the pressure measured in a fluid exceeds that of the atmosphere.

Internal energy: the energy associated with the random, disordered motion of molecules. It is separated in scale from the macroscopic ordered energy associated with moving objects; it refers to the invisible microscopic energy on the atomic and molecular scale.

Microscopic variables: Microscopic variables deal with the state of each molecule in terms of the mass, position, velocity etc. of the each molecule.

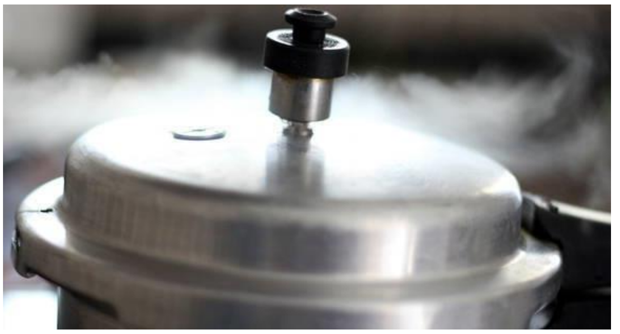
Macroscopic variables: Macroscopic variables are the physical quantities which describe the state of a system as a whole e.g. Pressure, Volume, Temperature etc. They can be easily measured with laboratory instruments.

Partial pressure: the pressure that would be exerted by one of the gases in a mixture if it occupied the same volume on its own.

4. INTRODUCTION

Every day we hear the pressure cooker whistling in the kitchen. As the temperature of the food kept inside the pressure cooker increases, steam is produced. The steam pressure increases with temperature. When this pressure becomes quite high, the whistle blows which reduces the pressure inside.

This example from our daily life shows that the pressure and temperature of the vapour produced inside are directly co-related.



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In this module, we will consider the internal energy of a gas system.

When we consider a gas of fixed mass it is characterised by its variables i.e. pressure, volume and temperature. These are its four parameters.

5. MICROSCOPIC AND MACROSCOPIC VIEW OF INTERACTION OF HEAT WITH MATTER

A gas is characterised by its variables i.e. **pressure, volume and temperature**. These are called **macroscopic variables** which can be measured using instruments **like manometer, thermometer etc.**

In the kinetic theory of gases, these macroscopic variables like pressure, temperature, volume etc. of the gas are related to the sum of the individual behaviour of the molecules.

For example motion of the molecules, the mass of the molecules, the number of the molecules of the gas etc.

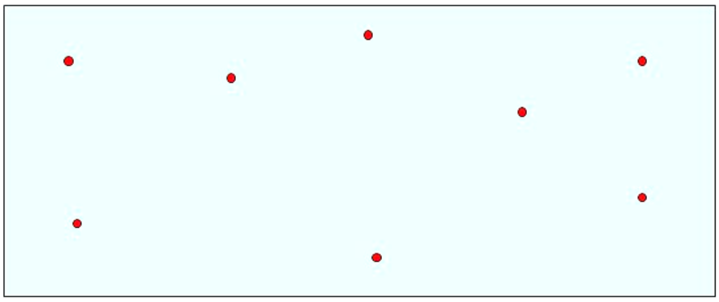
These properties cannot be measured directly from outside and they are a characteristic of molecules of the gas. They are referred to as the microscopic variables of the gas.

Day-to-day life examples show that a change in any one of the variables i.e. pressure, temperature, and volume is accompanied by a change in the other. Many laws have been formulated to show the variation of one of these variables w.r.t the other. But these laws are not obeyed strictly by the ordinary gases. There are deviations from the formulated laws, so we consider the existence of a model gas in which all these laws are strictly obeyed. We call this model gas, the ideal gas.

6. IDEAL GAS

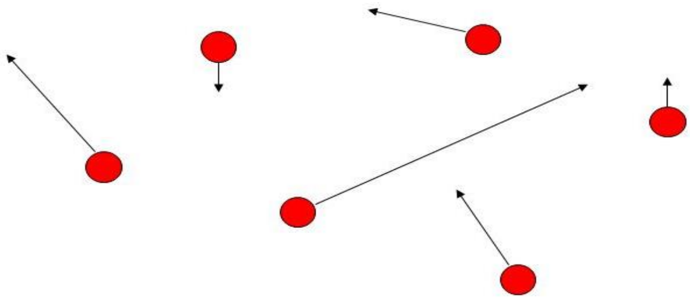
To make things simpler, we assume a model gas which is based on certain assumptions

- **The volume of the individual gas molecules is very small compared to the total volume occupied by the gas.**



- **The collisions between the molecules and the walls of the container and between the particles themselves are elastic. This means that there is no kinetic energy lost in the collisions**.

- **The particles in a gas move with a range of speeds, randomly.**



- **There are no forces of attraction between the molecules of the gas.**

- **This in turn means that the potential energy of the gas molecules is zero. All the internal energy of the gas molecules is in the form of the kinetic energy.**

7. IDEAL GAS EQUATION

The ideal gas follows the following equation which is known as the ideal gas equation.

$$PV = nRT$$

Where
P = pressure, V= volume,
T = Temperature,
n = number of moles and
R = Universal Gas constant whose value is 8.314J/(mol·K) or 0.0821 L atm / (mol·K)

All the gas laws can be arrived from this ideal gas equation which gives the dependence between the variables of an ideal gas.

8. GAS LAWS

Macroscopic rules followed by gases at temperatures higher than temperature of liquefaction. The rules or laws were formulated empirically or experimentally by many researchers. The extrapolations were a result of theoretical calculations. these together help us understand the behaviour of gases.

(i) Boyle's law:

Robert Boyle formulated a law in the 17th century which related the pressure and volume of a gas. According to this law,

At constant temperature, the pressure of a fixed amount of ideal gas is inversely proportional to it volume.

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

$$PV = \text{constant} \qquad \text{(at constant temperature)}$$

This means that an increase in pressure will be accompanied by a decrease in volume and vice versa at a fixed temperature.

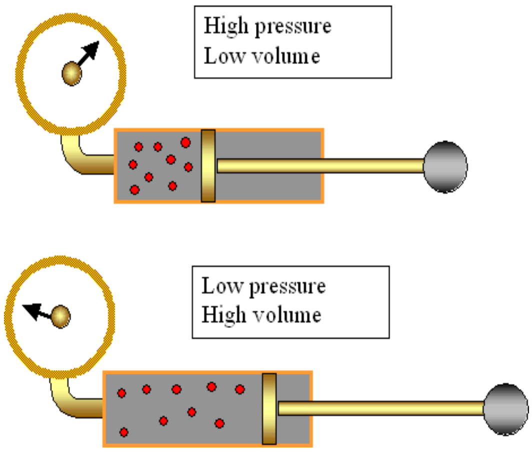


Figure shows the expansion of a gas in a cylinder when the piston is pulled out. The pressure decrease as the volume increases

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EXAMPLE

A scuba diver is 10 m below the surface of a lake. He releases an air bubble with a volume of 10 cm^3. The bubble rises to the surface. What is the volume of the bubble right before it breaks the surface? Take density of water = 1000kg/m^3 and g = 10m/s^2. Consider the temperature of water to be same at the surface and at a depth of 10m.

SOLUTION:

Pressure at 10m below the surface of a lake = hρg + Atmospheric pressure

$$h\rho g = 10 \times 1000 \times 10 = 10^5 \text{Pa}$$

Atmospheric pressure = 1.013 x 10^5 Pa = 10^5 Pa (approx.)

So the pressure 10m below the surface of the lake = 2 x 10^5 Pa

Pressure just below the surface of the lake = atmospheric pressure = 10^5Pa

Since temperature is same

$$P_1V_1 = P_2 V_2$$

$$(2\times 10^5) \times 10 = (10^5) \times V_2$$

$$V_2 = 20\text{cm}^3$$

Note:

As the pressure at the surface of the lake is half its value at a depth of 10m, the volume of the air bubble doubles as it comes to the surface.

(ii) Charles' law

According to Charles's law:

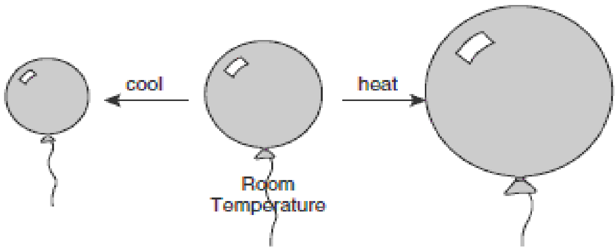
If the pressure of an ideal gas is held constant, the volume of the gas is directly proportional to the temperature (in Kelvin) of the gas.

$$V \propto T$$

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

(Here V = volume and T = Temperature)

This means that an increase in temperature will be accompanied by an increase in volume of a gas and vice versa at a constant pressure. This is pictorially represented in the figure below.



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EXAMPLE

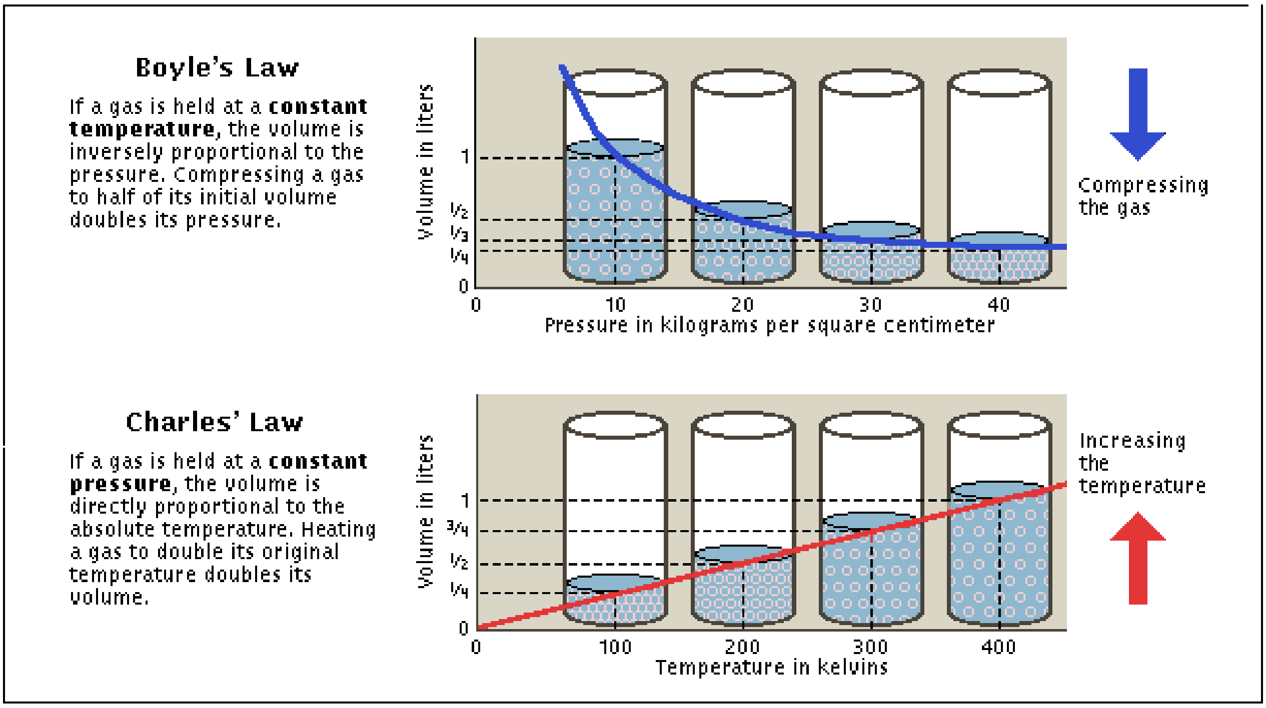
Variation of volume of gas with its pressure, when temperature is held constant

S.No	Pressure (kg. wt./cm^2)	Volume in (litres)
1	10	1
2	20	½
3	30	1/3
4	40	¼

Variation of volume of gas with its temperature, when pressure is held constant

S.No	Temperature (Kelvin)	Volume (litres)
1	100	¼
2	200	½
3	300	¾
4	400	1

A pictorial graph for the above data is made below.



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Point to ponder: What happens to the volume of a gas when its temperature becomes 0K?

EXAMPLE

A gas is kept at constant pressure. If its temperature is changed from 20°C to 100°C, by what factor does the volume change?

SOLUTION

Since the pressure is constant

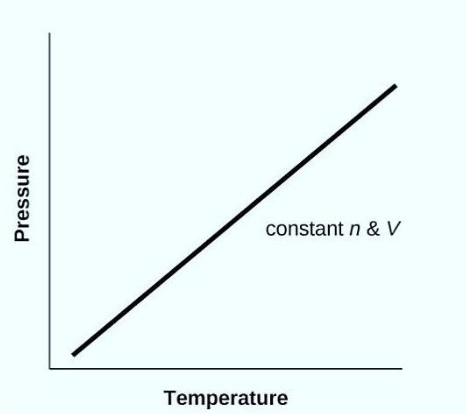
$$\frac{V_1}{T_1} = \frac{V_2}{T_2}$$

Initial temperature $T_1 = 20^oC = 293K$

Final temperature $T_2 = 100^oC = 373K$

The factor by which volume changes is given by

$$\frac{V_2}{V_1} = \frac{T_2}{T_1} = \frac{373}{293} = 1.27$$



Point to be noted: Although the ratio of final and initial temperature in degree Celsius is 5:1, this ratio in kelvin is only 1.27. Gas laws use temperature in kelvin.

So one must remember to convert the temperature into kelvin

(iii) Gay-Lussac's law

According to this law:

If the volume of a certain mass of an ideal gas is fixed, its pressure varies directly as the temperature of the gas (in kelvin).

$P \propto T$

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

This means that an increase in temperature of the gas will be accompanied by an increase its pressure and vice versa at constant volume of the gas.

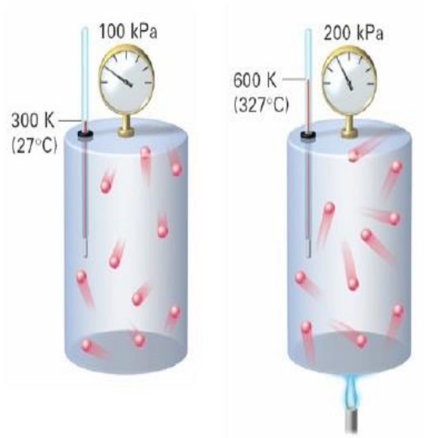


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Figure shows that the pressure in the cylinder doubles (100kPa to 200kPa) when the temperature (in kelvin) increases to twice its original value (300K to 600K).

Point to ponder: What happens to the pressure of the gas if the temperature becomes 0K?

EXAMPLE

An automobile tyre is filled to a gauge pressure of 150 kPa when its temperature is 27°C. After the car has been driven at high speed,the tyre temperature increases to 47°C. Take one atm = 100kPa

(a) Assuming that the volume of the tyre does not change and that air behaves as an ideal gas, find the gauge pressure of the air in the tyre.
(b) Calculate the gauge pressure if the volume of the tyre expands by 5%.

SOLUTION:

(a) If the volume of the tyre does not change, the increase in pressure will depend directly on the increase in temperature.

Initial temperature $T_1 = 27^oC = 300K$

Final temperature $T_2 = 47^oC = 320K$

Initial gauge pressure = 150kPa

Total pressure inside the tyre P_1 = 150kPa + 100kPa = 250kPa

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

Hence, we have P_2 = (250 x 320) / 300 = 266.7kPa

So the gauge pressure in the tyre will become 166.7kPa

(b) If the volume of the tyre increases by 5%

$$V_2/V_1 = 105/100$$

We have to use the equation $\frac{PV}{T}$ = constant

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

So P_2 = (250 x 320x100)/ (300 x105)

= 266.7 x 100/105 = 254kPa

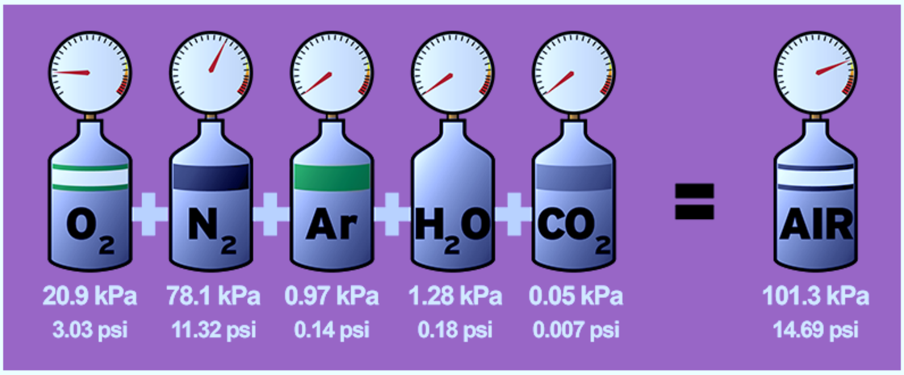
(iv) Dalton's law of partial pressures

According to this law:

In a mixture of non-reacting gases, the total pressure exerted is equal to the sum of the partial pressures exerted by each of the individual gases forming the mixtures.

$$P_{total} = P_1 + P_2 + P_3 + P_4 \text{------------} + P_n$$

Where P_{total} is the total pressure exerted by the gas while P_1, P_2, P_3 --------P_n are the partial pressures



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The figure above shows that the total pressure of 101.3 kPa exerted by the atmosphere is equal to the sum of the individual pressure of the component gases which make up the air.

(Note that the pressure due to Nitrogen gas is maximum and Carbon dioxide gas is minimum**)**

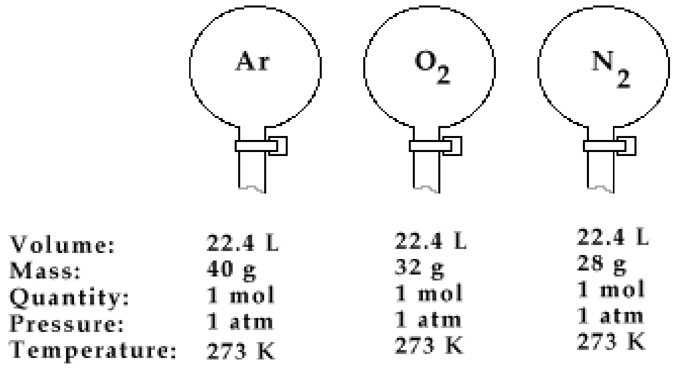
(v) Avogadro's law

According to this law**:**

Equal volume of all gases, at the same temperature and pressure, have equal number of molecules.

An interesting consequence of this law is that if the volume of a gas is 22.4 litres, the temperature 273K and pressure is equal to 1 atmospheric pressure, the number of molecules of the gas is equal to **6.023×10^{23}.** This is called the **Avogadro's number** to honour the great scientist.

The quantity of the gas is called one mole.



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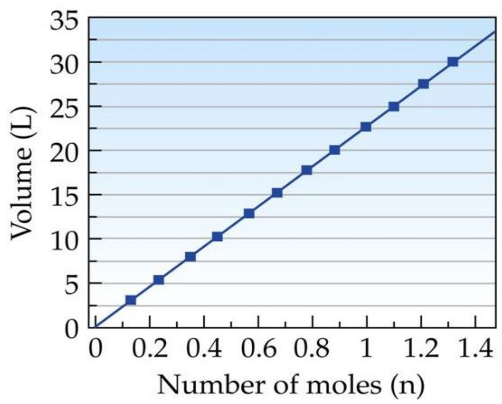
Point to be noted: Although the quantity of the gas in the above situation is the same, the mass of the gas is different.

Mass of 1 mole of Argon (Ar), Oxygen (O_2) and Nitrogen (N_2) gas is 40g, 32g and 28g respectively.

If the pressure and temperature conditions remain the same, the number of moles of a gas will vary directly as the volume of the gas.

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

(Here V = volume of the gas and n = number of moles)



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EXAMPLE

Find the volume occupied by 1 mol of an ideal gas at a temperature of 0°C and a pressure of 1 atm?

SOLUTION

Using the ideal gas equation:

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

$$= \frac{1\text{mol x } 0.0821 \text{ L atm / (mol·K) x } 273}{1atm}$$

$$= \quad 22.4 \text{ litres}$$

So 1 mole of an ideal gas at NTP always occupies 22.4L

EXAMPLE

A room is 5m×5m×4m. If the air pressure in the room is 1 atm and the temperature is 300 K,

a. Find the number of moles of air in the room.
b. If the temperature of the room increases by 6 K how many moles leave the room, if the pressure of the room remains 1 atm?

SOLUTION

a. Considering the air in the room to be following the ideal gas equation,

Pressure of the air, P= 1 atm
Volume of the air, V= 5 x 5x 4 = 100m^3 = 100 x 1000 Litres
Temperature of the room, T = 300K
Gas constant, R = 0.0821 L atm / (mol·K)

Number of moles, $n = \frac{PV}{RT} = \frac{1 \times 100 \times 1000}{0.0821 \times 300}$ = 4060 (approx.)

So the room contains 4060 moles of air molecules.

b. Since the pressure of the room does not change with 6K increase in temperature and the volume of the room is also constant.

We have, $nT = \frac{PV}{R}$ = constant

$$n_1T_1 = n_2T_2$$

$$\frac{4060 \text{ x}300}{306} = n_2 = 3980 \text{ moles}$$

So the number of moles of gas which leave the room = 4060 -3980 = 80 moles

9. DEVIATION FROM IDEAL GAS BEHAVIOUR

All the gas laws stated above are followed only when the gas is an ideal gas. But in reality there is nothing called an ideal gas.

In none of the gases the intermolecular forces between the molecules of the gas is zero and the molecules of any gas do occupy some volume.

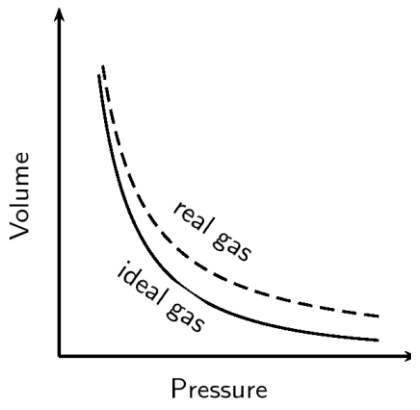
But all is not lost.

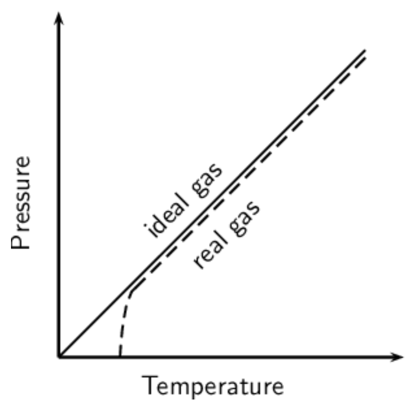
A gas at very **low pressure and high temperature** very closely approximates the ideal gas behaviour. A gas at low pressure has its molecules so widely separated that the intermolecular forces between them is negligible. Also the volume occupied by the molecule is very small compared to the total volume of the gas.

And at high temperature, the molecules of the gas have a large kinetic energy. This also helps in reducing the intermolecular forces between the gas molecules.

At **high pressure and low temperature** on the other hand a gas shows a lot of deviation from the ideal gas behaviour. **The close proximity of the molecules causes the intermolecular forces to increase significantly. Also the volume occupied by the molecules becomes significant in comparison to the total volume.**

So for the study of ideal gas behaviour gases at very low pressures and high temperatures are preferred.





Deviation in Boyles law **Deviation in the pressure law (Gay Lussac'slaw)**

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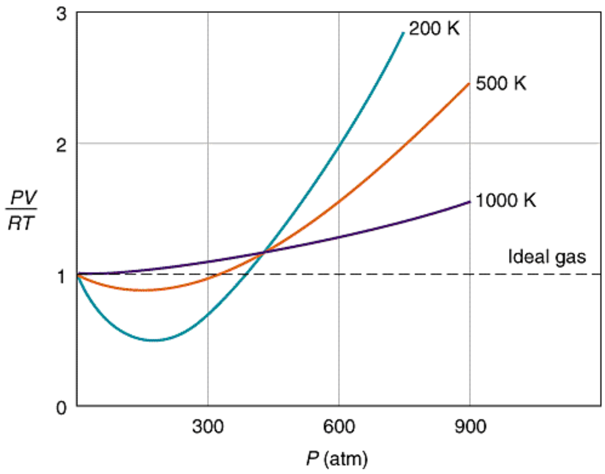
Interpretation of the graphs:

- The graph of volume versus pressure shows a greater deviation from ideal behaviour at low volume and high pressure. This is because the molecules of the gas come closer when they are compressed. So the intermolecular forces between the molecules become larger causing a deviation from Boyle's law.



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Graph showing the deviation from ideal gas behaviour.

- The graph of pressure versus temperature shows a deviation at very low temperatures. This is because the gas starts to liquefy at very low temperatures.

Understanding the graph

- **An ideal gas follows the ideal gas equation PV= nRT, where the symbols have the usual meaning. For one mole of a gas the equation becomes PV = RT. So the ratio $\frac{PV}{RT}$ = 1 for an ideal gas. Deviation from this ratio = 1 show a deviation from the ideal gas behaviour.**
- **At very low pressures all gases approach the ideal gas behaviour. So at around 1 atmospheric pressure real gases are very close to ideal gases in their behaviour.**
- **The gas maintained at 1000K shows least deviation from the ideal gas behaviour while the gas maintained at 200K shows maximum deviation from ideal gas behaviour**

10. SUMMARY:

1. **Ideal gas** :

i) Molecules of the gas occupy negligible volume as compared to the total volume of the gas

ii) Intermolecular forces between the molecules of the gas is absent

iii) The collisions between the molecules and the collision of the molecules with the walls of the container are elastic collisions

iv) The gas follows the ideal gas equation PV = nRT

2. Gas laws

i) **Boyle's law:** At constant temperature, the pressure of a fixed amount of ideal gas is inversely proportional to it volume

ii) **Charles's law:** If the pressure of an ideal gas is held constant, the volume of the gas is directly proportional to the temperature (in Kelvin) of the gas.

iii) **Gay lussac's law**: If the volume of a certain mass of an ideal gas is fixed, its pressure varies directly as the temperature of the gas

iv) **Dalton's law of partial pressure:** In a mixture of non-reacting gases, the total pressure exerted is equal to the sum of the partial pressures exerted by each of the individual gases forming the mixtures**.**

v) **Avogadro's law:** Equal volume of all gases, at the same temperature and pressure, have equal number of molecules.

3. **Deviation from ideal gas behaviour**: No real gas follows the gas laws strictly. But real gases approximate the ideal gas behaviour at very low pressures and high temperatures. The molecules are far apart and the intermolecular forces are negligible at low pressures and high temperature. So the deviation from the ideal gas behaviour is minimum in this case.