# Gases

All matter exists in three states as gas, liquid and solid under certain conditions of pressure and temperatures. Water for examples, can be solid ice, liquid water, steam, or water vapour. The physical properties of a substance often depend on its state. Gases are simpler than liquids and solids in many ways. Molecular motion in gases is totally random and the forces of attraction between gas molecules are so small that each molecule moves freely and essentially independently of other molecules. When subjected to change in temperature and pressure, it is easier to predict the behaviour of gases.

A gas consists of molecules separated wide apart in empty space. The molecules are free to move about throughout the container.

A liquid has molecules touching each other. However the intermolecular space, permit the movement of molecules throughout the liquid.

A solid has molecules, atoms or ions arranged in a certain order in fixed positions in the crystal lattice. The particles in solids are not free to move about but vibrate in their fixed positions.

**General characteristics of gases**

1. Expansibility: gases have limitless expansibility. They expand to fill the entire vessel they are placed in. I,e Gases assume the volume and shape of their containers.
2. Compressibility: gases are easily compressed by application of pressure to a movable piston fitted in the container, i.e Gases are the most compressible of the states of matter.

3. Diffusibility: gases can diffuse rapidly through each other to form a homogeneous mixture. i.e Gases will mix evenly and completely when confined to the same container.
4. Pressure: Gases exert pressure on the walls of the container in all directions.
5. Effect of heat: when gas, confined in a vessel is heated, its pressure increases. Upon heating in a vessel fitted with a piston, volume of the gas increases.
6. Gases have much lower densities than liquids and solids.

**Parameters of A gas**

A gas can be described in terms of four parameters (measurable properties)

a. The volume, V of the gas
b. Its pressure, P
c. Its temperature, T
d. The number of moles, n, of gas in the container.

The volume, V

The volume of the container is the volume of the gas sample. It is usually given in liter (L) or milliliters (mL)

| | | |
|---|---|---|
| | $1\ litre = 1000\ mL$ | |
| | $1\ mL = 10^{-3} L$ | |

One millilitter is practically equal to one cubic centimeter (cc). actually

| | | |
|---|---|---|
| | $1\ litre = 1000.028\ cc$ | |

The SI unit for volume is cubic meter ($m^3$) and the smaller unit is decimeter ($dm^3$).

**The pressure, P**

**Temperature, T**

The temperature of a gas may be measured in **centigrade degrees** ($^0$C) or **Celsius degrees**. The SI units of temperature is **kelvin (K)** or **absolute degree**. The centigrade degrees can be converted to kelvins by using the equation.

$$K = \quad {}^0\mathrm{C} + 273$$

The kelvin temperature or absolute temperature) is always used in calculations of other parameters of gases.

**The moles of a Gas sample, n**

The number of moles, n, of a sample of a gas in a container can be found by dividing the mass, m, of the sample by the molar mass, M (molecular mass).

$$moles\ of\ gas\ (n) = \frac{mass\ of\ gas\ sample\ (m)}{molecular\ mass\ of\ gas\ (M)}$$

**Substances That Exist as Gases**

We live at the bottom of an ocean of air whose composition by volume is roughly 78 percent $N_2$ 21 percent $O_2$, and 1 percent other gases, including $CO_2$.

The elements that are gases under normal atmospheric conditions is shown in fig. 3. Note that hydrogen, nitrogen, oxygen, fluorine, and chlorine exist as gaseous diatomic molecules: $H_2$, $N_2$, $O_2$, $F_2$, and $Cl_2$. An allotrope of oxygen, ozone ($O_3$), is also a gas at room temperature. All the elements in Group 8A, the noble gases, are monatomic gases: He, Ne, Ar, Kr, Xe, and Rn.

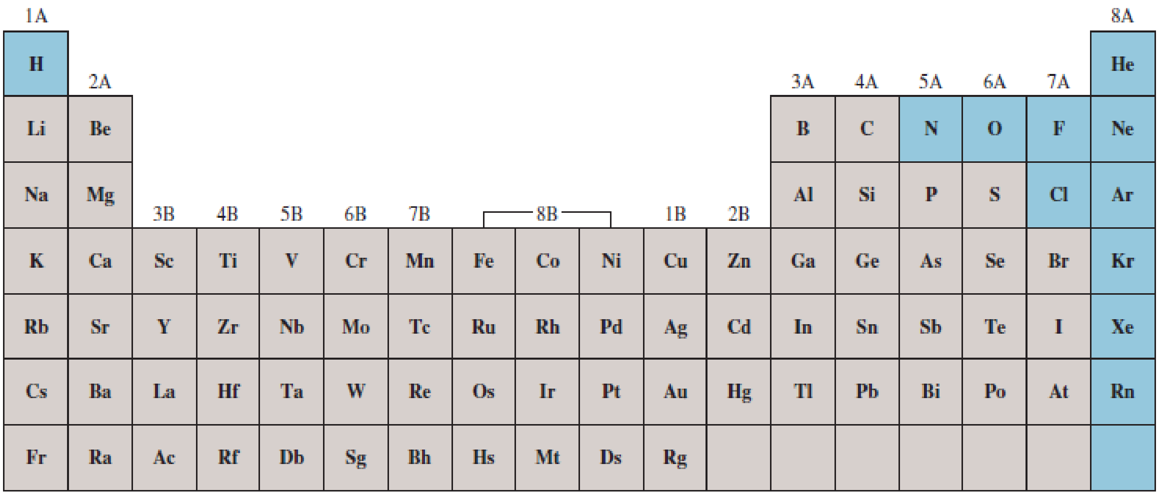



*Figure 1:Elements that exist as gases at 25ºC and 1 atm. The noble gases (the Group 8A elements) are monatomic species; the other elements exist as diatomic molecules. Ozone ($O_3$) is also a gas.*

Ionic compounds do not exist as gases at 25°C and 1 atm, because cations and anions in an ionic solid are held together by very strong electrostatic forces; that is, forces between positive and negative charges. To overcome these attractions, we must apply a large amount of energy, which in practice means strongly heating the solid. Under normal conditions, all we can do is melt the solid; for example, NaCl melts at the rather high temperature of 801°C. In order to boil it, we would have to raise the temperature to well above 1000°C.

The behaviour of molecular compounds is more varied. Some—for example, CO, $CO_2$, HCl, $NH_3$, and $CH_4$ (methane)—are gases, but the majority of molecular compounds are liquids or solids at room temperature. However, on heating they are converted to gases much more easily than ionic compounds. In other words, molecular compounds usually boil at much lower temperatures than ionic compounds do.

Note that, the stronger the attractive forces among the molecules, (called *intermolecular forces*), the less likely a compound can exist as a gas at ordinary temperatures.

*Table 1:Some substances found as gases at 1 atm and 25°C*

| Elements | Compounds |
|---|---|
| $H_2$ (molecular hydrogen) | HF (hydrogen fluoride) |
| $N_2$ (molecular nitrogen) | HCl (hydrogen chloride) |
| $O_2$ (molecular oxygen) | HBr (hydrogen bromide) |
| $O_3$ (ozone) | HI (hydrogen iodide) |
| $F_2$ (molecular fluorine) | CO (carbon monoxide) |
| $Cl_2$ (molecular chlorine) | $CO_2$ (carbon dioxide) |
| He (helium) | $NH_3$ (ammonia) |
| Ne (neon) | NO (nitric oxide) |
| Ar (argon) | $NO_2$ (nitrogen dioxide) |
| Kr (krypton) | $N_2O$ (nitrous oxide) |
| Xe (xenon) | $SO_2$ (sulfur dioxide) |
| Rn (radon) | $H_2S$ (hydrogen sulfide) |
| | HCN (hydrogen cyanide)* |

*The boiling point of HCN is 26°C, but it is close enough to qualify as a gas at ordinary atmospheric conditions.

Of the gases listed in Table 1, only $O_2$ is essential for our survival. Hydrogen sulphide ($H_2S$) and hydrogen cyanide (HCN) are deadly poisons. Several others, such as CO, $NO_2$, $O_3$, and $SO_2$, are somewhat less toxic. The gases He, Ne, and Ar are chemically inert; that is, they do not react with any other substance. Most gases are colourless. Exceptions are $F_2$, $Cl_2$, and $NO_2$. The dark brown colour of $NO_2$ is sometimes visible in polluted air. All gases have the following physical characteristics:

**Pressure of a Gas**

Gases exert pressure on any surface with which they come in contact, because gas molecules are constantly in motion. The pressure of a gas is defined as the force exerted by the impacts of its molecules per unit surface area in contact. The pressure of a gas sample can be measured with the help of a mercury manometer (fig 1). Similarly, the atmospheric pressure can be determined with a mercury barometer (fig. 2)

The pressure of air that can support 760 mmHg column at sea level, is called one atmosphere (1 atm). The unit of pressure, millimetre of mercury, is also called **torr**.

**SI Units of Pressure**

Pressure is one of the most readily measurable properties of a gas. In order to understand how we measure the pressure of a gas; it is helpful to know how the units of measurement are derived. We begin with velocity and acceleration. Velocity is defi ned as the change in distance with elapsed time; that is,

$$velocity = \frac{distance\ moved}{elapsed\ time}$$

The SI unit for velocity is m/s, although we also use cm/s.

*Acceleration* is the change in velocity with time, or

$$acceleration = \frac{change\ in\ velocity}{elapsed\ time}$$

Acceleration is measured in $m/s^2$ (or $cm/s^2$).

The second law of motion, formulated by Sir Isaac Newton in the late seventeenth century, defines another term, from which the units of pressure are derived, namely, *force*. According to this law,

$$force = mass \times acceleration$$

In this context, the *SI unit of force* is the ***newton (N),*** where

$$1N = 1kgms^{-1}$$

Finally, we define **pressure** as force applied per unit area:

$$pressure = \frac{force}{area}$$

The SI unit of pressure is the **pascal (Pa),** defined as one newton per square meter: $1Pa = Nm^2$

**Atmospheric pressure** is the pressure exerted by Earth's atmosphere. The actual value of atmospheric pressure depends on location, temperature, and weather conditions.

At the molecular level, air pressure results from collisions between the air molecules and any surface with which they come in contact. The magnitude of pressure depends on how often and how strongly the molecules impacts the surface.

The ***barometer*** is probably the most familiar *instrument for measuring atmospheric pressure.* A simple barometer consists of a long glass tube, closed at one end and filled with mercury.

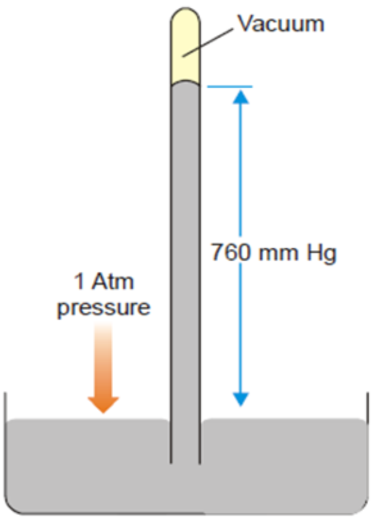



*Figure 2: mercury barometer*

If the tube is carefully inverted in a dish of mercury so that no air enters the tube, some mercury will flow out of the tube into the dish, creating a vacuum at the top (Figure 2). The weight of the mercury remaining in the tube is supported by atmospheric pressure acting on the surface of the mercury in the dish. ***Standard atmospheric pressure (1 atm*** is equal to *the pressure that supports a* *column of mercury exactly 760 mm (or 76 cm) high at 0°C at sea level.* In other words, the standard atmosphere equals a pressure of 760 mmHg, where mmHg represents the pressure exerted by a column of mercury 1 mm high. The mmHg unit is also called the *torr*, after the Italian scientist Evangelista Torricelli, who invented the barometer. Thus,

$$1\ atm = 760\ mmHg = 760\ torr$$

$$1\ torr = 1\ mmHg$$

The SI unit of pressure is the Pascal (Pa). the relation between atmosphere, torr and pascal is

$$1\ atm = 760\ torr = 1.01325 \times 10^{5} Pa$$

$$1000\ \mathrm{Pa} = 1\ \ \mathrm{k}Pa(kilopascal)$$

$$1\ atm = 1.01325 \times 10^{2} kPa$$

The unit of pressure 'Pascal' is not in common use.

The pressure outside a jet plane flying at high altitude falls considerably below standard atmospheric pressure. Therefore, the air inside the cabin must be pressurized to protect the passengers. What is the pressure in atmospheres in the cabin if the barometer reading is 688 mmHg?

Solution

$$1\ atm = 760\ mmHg$$

$$pressure\ in\ the\ cabin = 688\ \cancel{mmHg} \times \frac{1\ atm}{760\ \cancel{mmHg}}$$

$$pressure\ in\ the\ cabin = 0.905\ atm$$

Practice: convert 749 mmHg to atmospheres.

The atmospheric pressure in San Francisco on a certain day was 732 mmHg. What was the pressure in kPa?

$$1\ atm = 760\ mmHg$$

$$pressure\ in\ the\ cabin = 688\ \cancel{mmHg} \times \frac{1\ atm}{760\ \cancel{mmHg}}$$

$$pressure\ in\ the\ cabin = 0.905\ atm$$

Example 2

The atmospheric pressure in San Francisco on a certain day was 732 mmHg. What was the pressure in kPa?

$$1\ atm = 1.01325 \times 10^5 Pa = 760\ mmHg$$

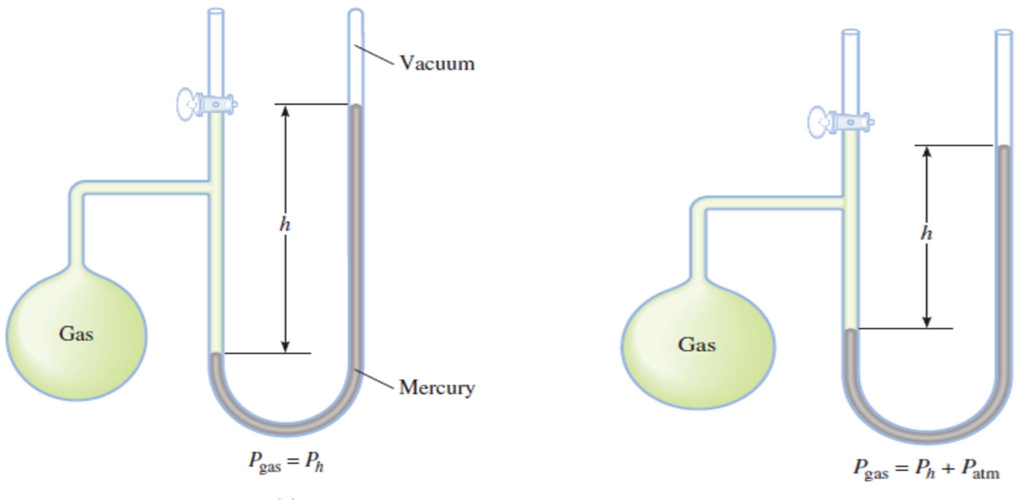
$$pressure\ in\ san\ Feancisco = 732\ \cancel{mmHg} \times \frac{1.01325 \times 10^5 Pa}{760\ \cancel{mmHg}}$$

$$pressure\ in\ san\ Feancisco = 9.76 \times 10^4\ Pa$$

$$pressure\ in\ san\ Feancisco = 97.6kPa$$

Practice Exercise Convert 295 mmHg to kilopascals.

A **manometer** is a device used to measure the pressure of gases other than the atmosphere.




(a) (b)

*Figure 3 Two types of manometers used to measure gas pressures. (a) Gas pressure is less than atmospheric pressure. (b) Gas pressure is greater than atmospheric pressure.*

The principle of operation of a manometer is like that of a barometer. There are two types of manometers, shown in Figure 3. The ***closed-tube manometer*** is normally used to measure pressures below atmospheric pressure [ Figure 3 (a)], whereas the ***open-tube manometer*** is better suited for measuring pressures equal to or greater than atmospheric pressure [ Figure 3 (b)].

Nearly all barometers and many manometers use mercury as the working fluid, even though it is a toxic substance with a harmful vapor. The reason is that mercury has a remarkably high density (13.6 g/mL) compared with most other liquids.

**The Gas Laws**

The volume of a given sample of gas depends on the temperature and pressure applied to it. Any change in temperature or pressure will affect the volume of the gas. The relationships, which describe the general behaviour of gases, are called the **gas laws**.

**The Pressure-Volume Relationship: Boyle's Law**

In 1660, Robert Boyle, investigated the pressure-volume relationship of a gas sample. He notes that as the pressure (P) is increased at constant temperature, the volume (V) occupied by a given amount of gas decreases. Conversely, if the applied pressure is decreased, the volume the gas occupies increases. This relationship is now known as Boyle's law, ***which states that the pressure of a fixed amount of gas at a constant temperature is inversely proportional to the volume of the gas***.

The Boyle's law may be expressed mathematically as

$$V \propto \frac{1}{P} \qquad (T, n\ are\ constant)$$

$$V = k \times \frac{1}{P}$$

Where k is a constant called the proportionality constant

$$PV = K$$

If $P_1V_1$ are the initial pressure and volume of a given sample of gas and $P_2$, $V_2$ the changed pressure and volume, we can write.

$$P_1V_1 = P_2V_2$$

The relationship is useful for the determination of the volume of a gas at any pressure, if its volume at any other pressure is known.

Graphically Boyle's law can be expressed in the following ways:

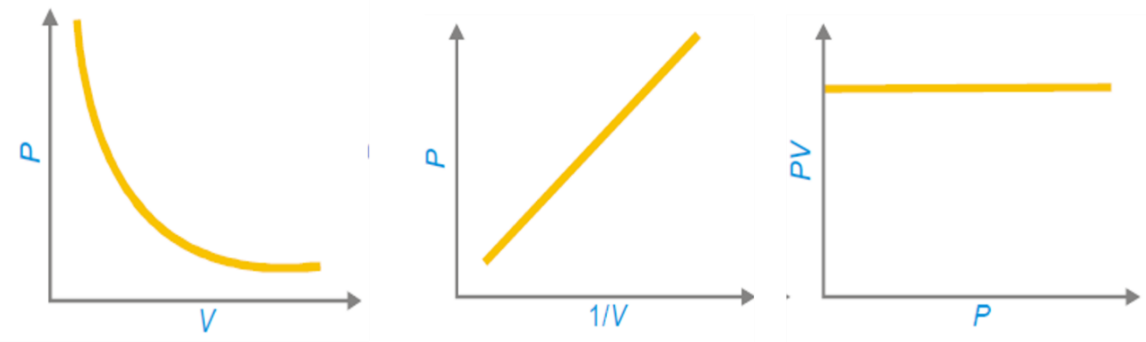



*Figure 4: Graphical representation of Boyle's law*

**The Temperature-Volume Relationship: Charles's and Gay-Lussac's Law**

In 1787 Jacques Charles and Joseph Gay-Lussac, investigated the effect of change of temperature on the volume of a fixed amount of gas at constant pressure. Their studies showed that, at constant pressure, the volume of a gas sample expands when heated and contracts when cooled.

He established a generalisation which is called the Charles' Law. which states that the volume of a fixed amount of gas maintained at constant pressure is directly proportional to the absolute temperature of the gas.

Charles's Law may be expressed mathematically as

$$V \propto T \qquad \text{(P, n are constant)}$$

$$V = kT$$

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

Where k is a constant.

If $V_1$, $T_1$ are the initial volume and temperature of a given mass of a gas at constant pressure and $V_2$, $T_2$ be the new values, we can write

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

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

Graphically, Charles's law can be represented as

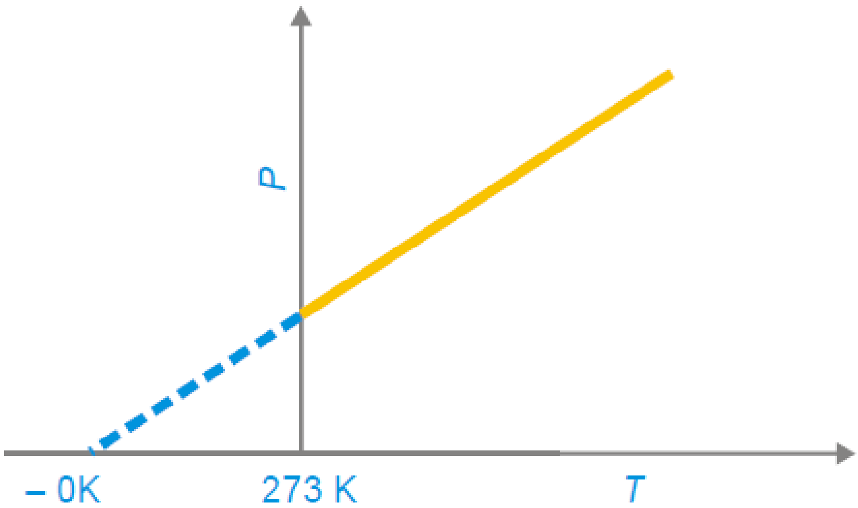



*Figure 5: Graphical representation of Charle's law.*

The Volume-Amount Relationship: Avogadro's Law

The Italian scientist Amedeo Avogadro complemented the studies of Boyle, Charles, and Gay-Lussac. In 1811 he published a hypothesis stating that at the same temperature and pressure, equal volumes of different gases contain the same number of molecules (or atoms if the gas is monatomic). It follows that the volume of any given gas must be proportional to the number of moles of molecules present; that is,

$$V \propto n \qquad\qquad \text{P,T constant}$$

$$V = kn$$

where $n$ represents the number of moles and $k_4$ is the proportionality constant.

**Avogadro's law,** states that at constant pressure and temperature, the volume of a gas is directly proportional to the number of moles of the gas present.

**The Ideal Gas Equation**

The simultaneous effect of change of pressure and temperature of a gas can be studied by combining Boyle's law and Charles' law. The derived new equation is called combined gas law or ideal gas equation.

| | | |
|---|---|---|
| $Boyle's\ law$: | $V \propto \frac{1}{P}$ | $(T, n\ are\ constant)$ |
| $Charles'\ law$ | $V \propto T$ | $(P, n\ are\ constant)$ |
| $Avogadro's\ law$ | $V \propto \mathrm{n}$ | $(P, T\ are\ constant)$ |

We can combine all three expressions to form a single master equation for the behaviour of gases:

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

$$V = R\frac{\mathrm{nT}}{P}$$

$$PV = nRT\ (ideal\ gas\ equation)$$

where **R,** the proportionality constant, is called the **gas constant.** The **ideal gas equation** describes the relationship among the four variables $P, V, T$, and $n$. An **ideal gas** is a hypothetical gas whose pressure-volume-temperature behaviour can be completely accounted for by the ideal gas equation.

Examples

Sulphur hexafluoride ($SF_6$) is a colourless, odourless, very unreactive gas. Calculate the pressure (in atm) exerted by 1.82 moles of the gas in a steel vessel of volume 5.43 L at 69.5°C.

$$PV = nRT$$

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

$$P = \frac{(1.82\ \cancel{moles}) \times (0.0821 \cancel{L} atm \cancel{K^{-1}} \cancel{moles^{-1}}) \times (69.5 + 273)\cancel{K}}{5.43\ \cancel{L}}$$

$$P = 9.42\ atm$$

Calculate the volume (in litres) occupied by 2.12 moles of nitric oxide (NO) at 6.54 atm and 76°C.