CHAPTER I

DIFFERENT STATES OF MATTER **2026/2027**

I.1. Introduction

Matter is everything that makes up the universe and occupies space. It exists in various forms, ranging from solids to liquids and gases. In our daily lives, we encounter matter in diverse forms: the water we drink, the air we breathe, and even our own body. Matter is composed of particles, such as atoms and/or molecules, which interact with one another according to physical and chemical laws. Understanding matter is essential in many fields, such as chemistry, physics, and biology, as it helps us explain natural phenomena and develop new technologies. Matter is a fundamental concept that allows us to better understand the world around us.

I.2. Different States of Matter

Matter can exist in several states, each having distinct (different) characteristics. There are three classical states of matter, which are:

- The solid state;
- The liquid state;
- The gaseous state.

A. The Solid State

The particles of a substance in the solid state are very close to one another because they are strongly bound together and are also highly ordered, as shown in Figure 1. This gives the solid a definite (fixed) shape and volume. This means that it does not necessarily occupy all the space (volume) available in its container.

Examples of solids

- Wood;
- Metal;
- Ice cubes;
- Pieces of chocolate, etc.

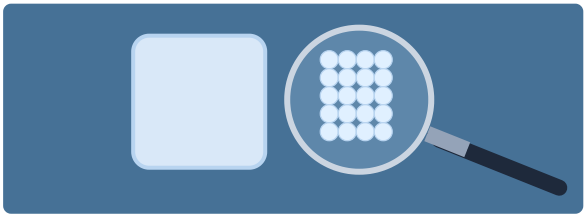


Figure 1. Particles of a solid (ice cube).

> **Remark**
>
> Solids are incompressible (such as wood, iron, etc.) or very difficult to compress (such as beverage cans made of aluminum, metal sheets, etc.).

B. The Liquid State

A liquid is a form of fluid; the particles are therefore farther apart, less strongly bound, and less ordered than in a solid, and they can move relative to one another, as shown in Figure 2. This allows the liquid to flow and take the shape of its container while maintaining a constant volume under the effect of gravity g. Thus, a liquid has no definite shape of its own, but it possesses a definite volume, which changes only with temperature. Two other important characteristics of a liquid are that its free surface at rest is flat and horizontal (even when the container is tilted) and it is incompressible.

Examples of liquids

- Water;
- Oil;
- Honey.

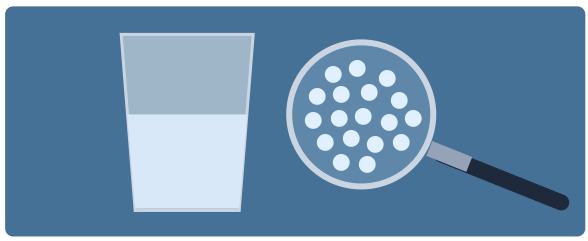


Figure 2. Particles of a liquid.

> **Remark**
>
> At constant pressure, when the temperature increases, the molecules of the fluid move farther apart from one another, the volume increases, and the density decreases. The opposite occurs when the temperature decreases.

C. The Gaseous State (or Gas)

A gas is a fluid just like a liquid; in this state, the particles (atoms and/or molecules) are very far apart (they spread out widely) from one another because they are very weakly bound and move freely (quasi-independently), as shown in Figure 3. Gases have an indefinite shape and an indefinite volume (they have no fixed shape or volume of their own); they take the shape of their container or vessel and occupy its entire volume (they tend to fill all available space). A very important characteristic of gases is that they are compressible, unlike liquids which are incompressible.

Examples of gases

- Air;
- Water vapor.

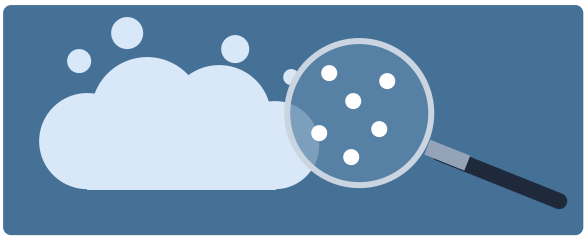


Figure 3. Particles of a gas.

At low pressure (and high temperature), real gases have properties similar to those of ideal gases. The ideal gas model is a theoretical model whose particles (atoms and/or molecules) satisfy the following conditions (according to the kinetic theory of gases):

1. The particles (atoms and/or molecules) are hard point-like spheres whose diameter is negligible compared to interatomic and/or intermolecular distances.
2. The particles are in constant random motion and undergo collisions (impacts) with one another and against the walls of the container holding the gas. These collisions are elastic collisions.
3. Between two collisions, the particles follow straight-line (rectilinear) trajectories.
4. Outside of collisions, there are no interactions between the particles.
5. The particles are distributed uniformly throughout the entire available volume.
6. All directions are equally probable. The distribution of velocities is isotropic: it does not depend on any particular direction.

Macroscopically (what can be observed with the naked eye), an ideal gas is defined as any gas that simultaneously obeys the following laws:

1. Boyle–Mariotte Law: The product of the two parameters P and V of an ideal gas is constant at constant temperature. This law is expressed by the equation:

$$\mathbf{P \times V = Cst \quad (at\ constant\ T)} \qquad \mathbf{(1)}$$

2. Charles's Law: The ratio of the two parameters V and T of an ideal gas is constant at constant pressure.

$$\mathbf{\frac{V}{T} = Cst \quad (at\ constant\ P)} \qquad \mathbf{(2)}$$

3. Avogadro–Ampère Law: This law states that at a given pressure and temperature, all ideal gases have the same molar volume V_m. That is, for any two gases 1 and 2 at the same pressure and temperature:

$$\mathbf{\frac{V_1}{n_1} = \frac{V_2}{n_2} = V_m = Cst \quad (at\ constant\ pressure\ and\ temperature)} \qquad \mathbf{(3)}$$

where n_1 and n_2 are the numbers of moles of gases 1 and 2 respectively, and V_1 and V_2 are their volumes.

- **Standard Molar Volume of Ideal Gases – Values calculated from the ideal gas law:**

Conditions	Pressure	Temperature	Molar Volume
Normal conditions or NTP (CNTP)	1 atm	0 °C (273.15 K)	22.4 liters/mole
Standard conditions or STP (CSTP)	1 bar	25 °C (298.15 K)	24.8 liters/mole

NB: 1 atmosphere = 101,325 Pa = 1.01325 bar; 1 bar = 100,000 Pa

4. Gay-Lussac's Law: The ratio of the two parameters *P* and *T* of an ideal gas is constant at constant volume.

$$\frac{P}{T} = \text{Cst} \quad \text{(at constant V)} \tag{4}$$

5. Dalton's Law – Law of Partial Pressures: This is a law of thermodynamics stating that the total pressure within a mixture of ideal gases is equal to the sum of the partial pressures of its constituents:

$$P_{Total} = \Sigma_i P_i \tag{5}$$

Combining the preceding laws leads to the well-known general ideal gas law:

$$P \cdot V = n \cdot R \cdot T \tag{6}$$

Where:

P is the pressure measured in (Pa);

V is the volume measured in (m^3);

n is the number of moles or the amount of gas in (mole);

R is the ideal gas constant measured in ($Pa \cdot m^3/mol \cdot K$ or $J/mol \cdot K$). Its empirical value is:

$R = 8.314\ J \cdot mol^{-1} \cdot K^{-1}$ (SI)

$R = 0.082\ L \cdot atm \cdot mol^{-1} \cdot K^{-1}$ (CGS)

T is the absolute temperature measured in (K).

The ideal gas law (Eq. 6) can also be expressed in another form:

$$P \cdot V = N \cdot k_B \cdot T \tag{7}$$

In this equation:

N is the number of particles constituting the gas ($N = n \times N_{Av}$), where N_{Av} is Avogadro's number given by the value:

$$N_{Av} = 6.023 \times 10^{23}\ /\text{mol}$$

k_B is the Boltzmann constant: $k_B = \frac{R}{N_{Av}} = \frac{8.314}{6.023 \times 10^{23}} = 1.381 \times 10^{-23}\ \text{J}\cdot\text{K}^{-1}$

* Behavior of a Real Gas – J. D. van der Waals Law

A real gas differs from an ideal gas in that its molecules have a (non-negligible) volume and interact with one another through forces called van der Waals forces. Therefore, the pressure exerted by a real gas on the wall of the container will be lower than that of an ideal gas and is given by the equation:

$$\mathbf{P_{Wall} = P_{Kinetic} - a\ \frac{n^2}{V^2}}$$

Where:

$P_{Kinetic}$ is the pressure of the gas on the wall when neglecting interactions between particles (ideal gas model – kinetic theory of ideal gases);

a is a corrective parameter that accounts for interactions between molecules;

n is the amount of substance in moles;

V is the volume of the real gas.

In this case, the volume of the real gas that must be taken into account is:

$$\mathbf{V = V_{Kinetic} + n \cdot b}$$

Where:

$V_{Kinetic}$ is the volume of the gas when neglecting the volume of its particles (ideal gas model – kinetic theory of gases);

n is the amount of substance in moles;

b is a corrective parameter called the covolume, which accounts for the volume of the particles.

Correcting Equation (6) for a real gas leads to the van der Waals equation of state for real gases:

$$\mathbf{P_{Kinetic} \cdot V_{Kinetic} = \left(P_{Wall} + a\ \frac{n^2}{V^2} \right) (V - n \cdot b) = n \cdot R \cdot T}$$

* Kinetic Energy of a Gas

If a gas molecule of mass *m* moves at a velocity *v* inside a container, its kinetic energy is:

$$E_c = \frac{1}{2} m v^2 \tag{8}$$

For a gas composed of N identical particles, its total kinetic energy is:

$$E_c = N \frac{1}{2} m v^2 = N \frac{3}{2} k_B T \tag{9}$$

D. Other States of Matter

Plasma: It is an ionized gas resulting from heating matter to very high temperatures (around 2,000 °C or higher). As a result, the atoms of matter lose electrons and become positive ions. Plasma is present in stars, including our Sun, and in lightning.

I.3. Changes of State

Matter can transform from one state to another depending on the conditions of the surrounding environment, notably the temperature and pressure to which it is subjected. The transformation of matter from one state to another is always accompanied by the release or absorption of energy as well as a change in volume. These different states contain the exact same molecules.

The difference in behavior is mainly caused by variations in intermolecular bonds, which act to hold molecules together. Conversely, thermal agitation and Brownian motion act to dissociate them. The diagram in Figure 4 identifies the different changes of state.

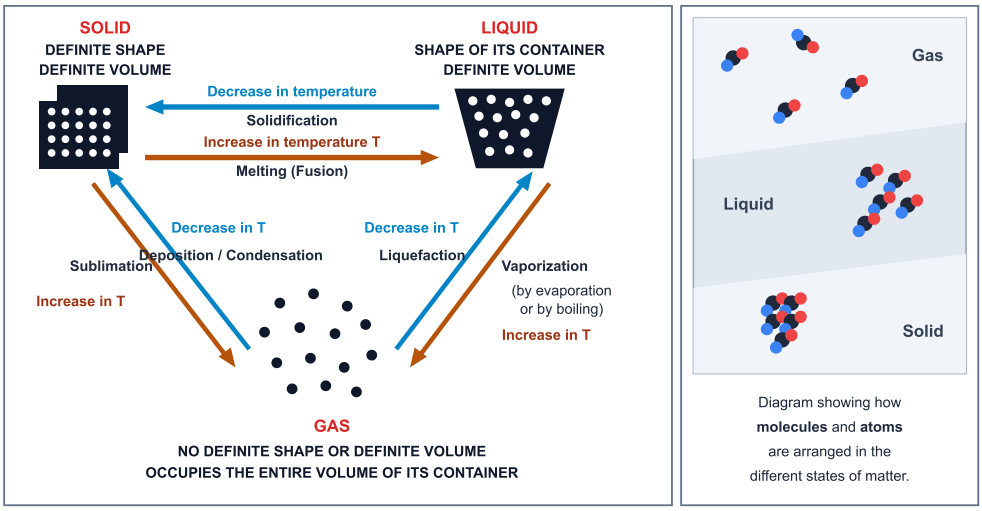


Figure 4. Transformations of matter.

Remarks and Clarifications

1. A fluid is a substance that has no definite shape of its own and takes the shape of its container.
2. The classification into solid, liquid, and gas applies only to pure substances.
3. Mass does not change during a change of state, although volume may vary.
4. A substance is called **"pure"** when it is composed exclusively of a single type of atom and/or molecule.
5. From a structural point of view (organization and arrangement of atoms), materials can be classified into three categories corresponding to their states:
   - **Ordered state:** such as crystals (or crystalline solids); the atoms are arranged in a highly regular or ordered manner according to a 3-dimensional lattice.
     Examples: ice, table salt (Na^+Cl^-), iron, alumina (Al_2O_3), etc.
   - **Disordered state:** such as amorphous solids (a state that is neither truly solid nor liquid).
     Examples: glasses and plastics.
   - **Intermediate or mesomorphic state:** a state between solid and liquid, such as soap (liquid crystals).