

Introduction: Biophysics and Its Applications:

Biophysics is a discipline at the interface of physics and biology, where tools for observing physical phenomena are applied to molecules of biological interest. The progressive discovery of the unity of physical processes in all living cells has been a major driver for the development of biophysics. Physicists aim to explain most observations by proposing synthetic theories. The greatest successes occur when multiple observations in different contexts and organisms are linked to the same physical explanation.

Several fields of biology, in the broadest sense, have benefited from advances made in biophysics, including medicine, cell biology, and molecular biology.

1. Aqueous and Organic Solutions:

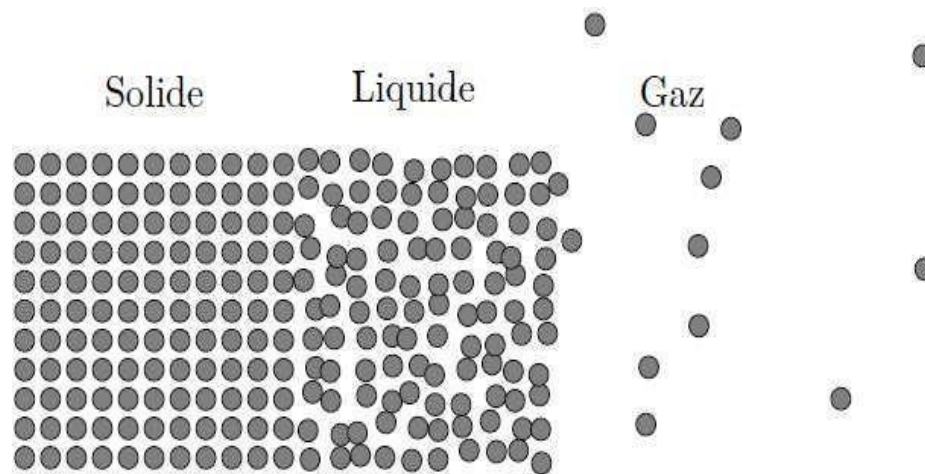
A solution is defined as any homogeneous mixture in a condensed phase (liquid or solid), with only one phase, comprising:

- The solvent and the solute (or dissolved substance). If one of the components is liquid, it is considered the solvent.
- For example: In an aqueous solution of copper (solid) and ammonia (gas), the solvent is water. If two or more components are liquids, the most abundant liquid is the solvent. There are two types of solutions: neutral solutions, where all solutes are neutral molecules or particles, and electrolytic solutions, which are solutions of salts, bases (OH⁻), and acids (H⁺, H₃O⁺). Electrolytic solutions conduct electricity

2. States of Matter:

Matter is composed of atoms, molecules, or ions, which are bound by various forces (covalent, ionic, metallic, Van der Waals, or hydrogen bonds). Depending on the strength of the bonding forces, there are three states of matter: solid, liquid, and gaseous. In practice, the term "fluid" refers to both liquids and gases.

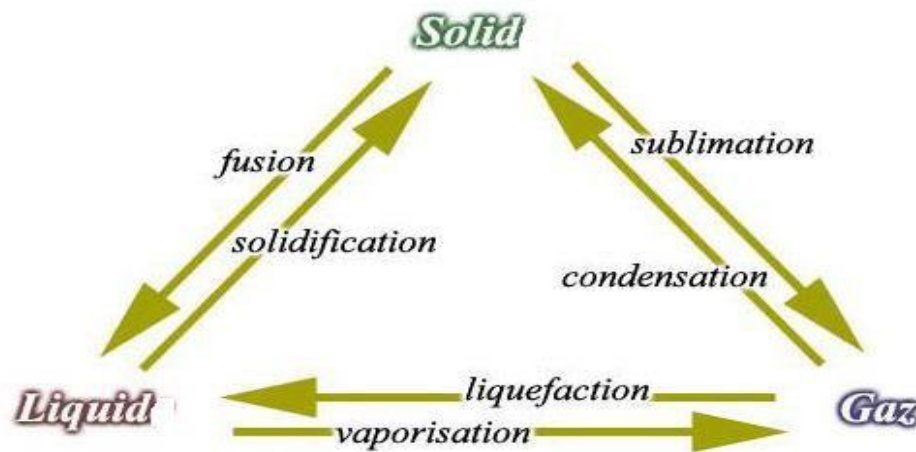
- **Solid State** (ordered: long-range order or amorphous: no order



The transition from one state of matter to another occurs through phase transitions during a thermodynamic state change caused by a modification of its pressure, temperature, and/or volume. Phase transformations or state changes are as follows:

- 1) **Melting**: This is the transformation of matter from a solid state to a liquid state.
- 2) **Liquefaction (condensation)**: This is the transformation from a gaseous state to a liquid state.
- 3) **Condensation or (deposition)**: This is the transformation from a gaseous state to a solid state.
- 4) **Gasification (vaporization)**: This is the transformation of a liquid into vapor.
- 5) **Solidification**: This is the opposite of melting. It is also called freezing when the transformation occurs at low temperatures.
- 6) **Sublimation**: This is the transition from the solid state to the gaseous state.

Transformations 1, 4, and 6 occur with the absorption of energy, while transformations 2, 5, and 3 are accompanied by the release of heat.



The Different States of Matter and Their Phase Changes

To separate two molecules (to break a molecular bond between two molecules), an energy E_L , called bond energy E_L , must be provided. E_L varies from 0.5 to 20 kJ/mol depending on the nature of the substance but remains significantly lower than the covalent bond energy between two atoms of the same molecule. Covalent bond energy ranges from 400 to 800 kJ/mol.

However, molecules are not fixed but are subject to a state of disordered agitation called thermal agitation, the intensity of which is measured by the average kinetic energy of a molecule. This kinetic energy, denoted by E_c is proportional to the absolute temperature and is expressed by:

$$E_c = \frac{3}{2} K_B T (\text{K.J/molecule})$$

Where:

K_B : is Boltzmann's constant, equal to 1.38×10^{-23} J.K

T : is the absolute temperature

According to the predominance of E_c in relation to E_L , three fundamental physical states of matter are distinguished: the gaseous state, the liquid state, and the solid state.

- ❖ If EC is much lower than EL ($EC \ll EL$) the molecules cannot separate from one another, and their kinetic energy only translates into rotations and vibrations around a fixed average position. This is the solid state.
- ❖ If EC is of the same order of magnitude as EL ($EC = EL$), the molecules can slide past each other without detaching and remain grouped together. This is the liquid state.
- ❖ As the temperature increases, the molecules rotate and vibrate faster, their EC increases, and when it becomes sufficiently greater than EL ($EC > EL$) the intermolecular bonds eventually break. The molecules then become independent of one another in a disordered state (Brownian motion). This is the gaseous state.

2.1.Gases:

A gas occupies all the space available to it. When a gas is contained in a closed chamber, the molecules, which are numerous, randomly collide with the walls of the chamber, which is the origin of the pressure (force exerted on the walls). This pressure is greater when EC is large (the collisions are more violent) and when the number of molecules is higher. For example, when you blow air into an elastic balloon, you increase the number of air molecules, which increases the number of collisions, and consequently, the pressure increases, causing the balloon to expand.

An ideal gas defines a state where the interactions between gas molecules are negligible, and the gas molecules themselves do not occupy a significant portion of the volume of space occupied by the gas. Therefore, in an ideal gas, the molecules only have kinetic energy, which remains conserved. Under normal conditions (temperature = 20°C, pressure = 10^5 Pa = 1 bar), most real gases (e.g., air) are considered ideal gases.

In the case of an ideal gas, statistical thermodynamics shows that the pressure P is directly proportional to the absolute temperature T, the number Nm of molecules, and inversely proportional to the volume V of the space occupied by the gas:

$$P = K_B \cdot T \frac{Nm}{V}$$

The number N_m being extremely large, it is customary to divide it by Avogadro's number N_A . The number of molecules is then counted in moles (mol).

$$N_m/N_A = n \text{ mol}$$

The expression for pressure becomes :

$$P = K_B \cdot T \frac{n \cdot N_A}{V} = n \cdot N_A \cdot K_B \frac{T}{V}$$

The product of the two constants N_A and K_B is replaced by the letter R , called the "ideal gas constant." With $N_A = 6.023 \times 10^{23} \text{ mol}^{-1}$ and $K_B = 1.38 \times 10^{-23} \text{ J.K}^{-1}$ we get $R = 8.32 \text{ J.K}^{-1} \cdot \text{mol}^{-1}$

The equation for pressure then becomes:

$$P = n \cdot R \frac{T}{V}$$

Finally, we obtain the equation : $PV = n \cdot R \cdot T$ called the "ideal gas law."

Example:

Initially, an ideal gas with volume $V_1 = 5 \text{ m}^3$ is at pressure $P_1 = 500 \text{ Pa}$. It is compressed, while maintaining a constant temperature, to a volume $V_2 = 2 \text{ m}^3$

What is the final pressure P_2 ?

Answer:

We apply the ideal gas law:

At the initial state:

$$P_1 V_1 = nRT_1$$

At the final state:

$$P_2 V_2 = nRT_2$$

Since during compression the temperature remains constant, we have:

$$T_2 = T_1$$

Therefore:

$$P_2V_2=P_1V_1$$

From this:

$$P_2=P_1V_1V_2$$

$$P_2 = \frac{V_1}{V_2} P_1 = \frac{5}{2} \times 500 = 1250 \text{ Pa}$$

Thus, we find that $P_2 > P_1$

When the pressure of a real gas exceeds normal conditions ($10^5 \text{ Pa} = 1 \text{ bar}$), the interactions between molecules and their volumes are not negligible and depend heavily on the nature of the gas.

Van der Waals established the equation of state for real gases by introducing correction terms into the ideal gas law:

$$\left(P + \frac{an^2}{V^2}\right) (V - nb) = nRT$$

Where:

- P = measured pressure,
- V = volume of the container,
- n = number of moles of gas,
- R = ideal gas constant,
- T = absolute temperature.

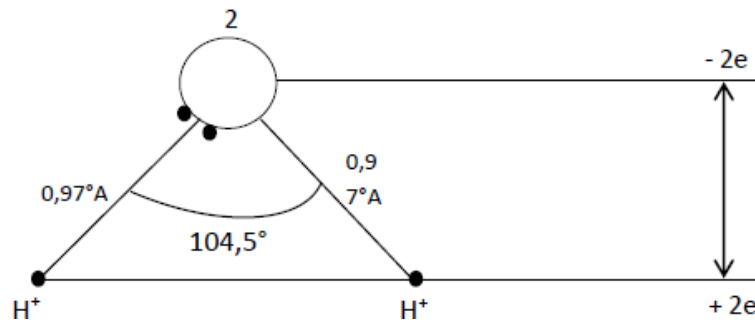
The term an^2/V^2 represents the pressure correction due to the interaction between gas molecules.

The volume correction, nb , subtracts the volume of the gas molecules from the total volume of space occupied by the gas.

a and b are measured constants specific to each gas.

2.2 Liquids Structure of Water:

The water molecule is composed of one oxygen atom (O), which has six peripheral electrons (predisposing it to two covalent bonds), and two hydrogen atoms (H), each of which has one electron in its shell. The electrical interactions between the oxygen and hydrogen atoms cause the water molecule, H_2O , to have the following triangular structure:

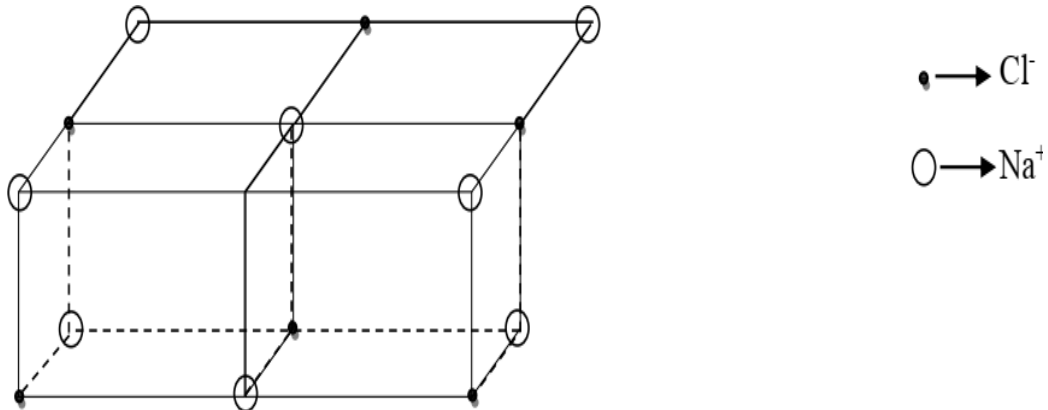


This structure shows that the H–O–H angle is 104.5° , and the interatomic distance $d_{\text{O–H}} = 0.97 \text{ \AA}$.

Since oxygen is much more electronegative (it attracts negative charges) than hydrogen, the electron pair of each O–H bond shifts toward the oxygen atom. Thus, the oxygen atom carries a negative electric charge ($-2e$), and the two hydrogen atoms, having given up their electrons, form a positive electric charge ($+2e$). This makes the H_2O molecule an electric dipole. This dipole is permanent, which is why water is very stable and has the property of being an excellent solvent for other elements, whether solid, liquid, or gas. For example, it dissolves polar molecules like HCl or ionic crystalline solids like salts (NaCl, KCl, etc.). In fact, water can dissolve all ionic crystalline solids.

2.3 Solids : Types of Structure:

In solids, atoms and molecules are arranged in three-dimensional space either in a fundamentally random (disordered) manner or in a regular and periodic arrangement. In the disordered state, the solid is called an "amorphous solid," such as glass, while in the ordered state, the solid is called a "crystalline solid" or "crystal," such as diamond or sodium chloride.



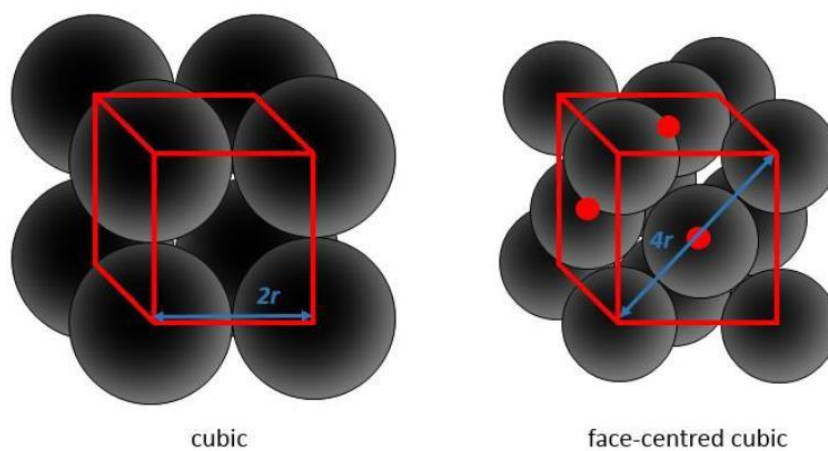
Thus, the crystals of sodium chloride (NaCl) have a cubic structure. Graphite and diamond are two different solids made up of carbon (C) atoms. In graphite, the carbon atoms form parallel layers of hexagonal lattices. In diamond, the lattice is face-centered cubic. Although the atoms are the same, the properties of diamond and graphite are very different.

The lattice of a crystalline structure is defined by the lengths of the sides (a , b , c), the angles between the sides (α , β , γ), and the positions of the atoms within the lattice. The side lengths vary from one material to another.

The different types of structures are as follows:

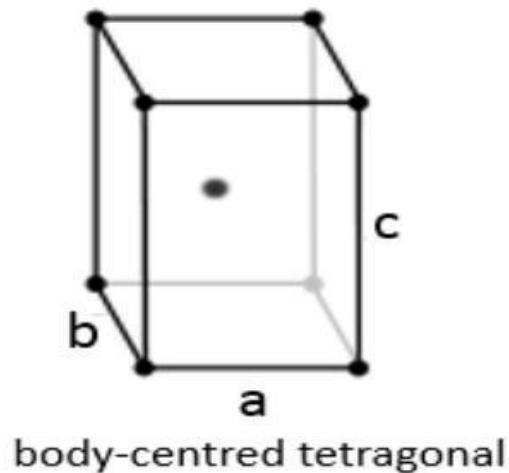
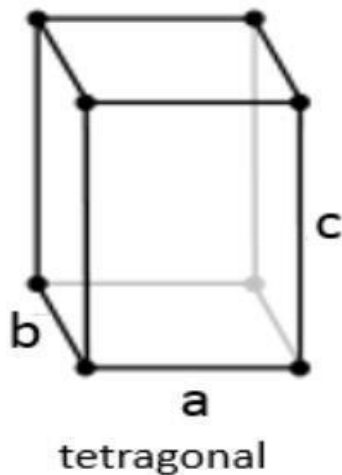
✓ **Cubic:**

In the cubic structure, the lengths of the sides are equal, $a = b = c$ and $\alpha = \beta = \gamma = 90^\circ$. If an atom is in the center of the cube, we call this structure body-centered. If an atom is in the center of each face, we call it face-centered.



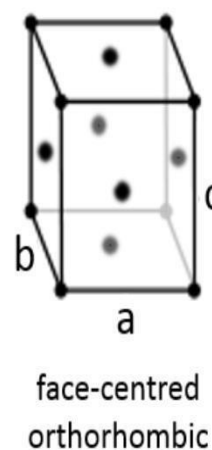
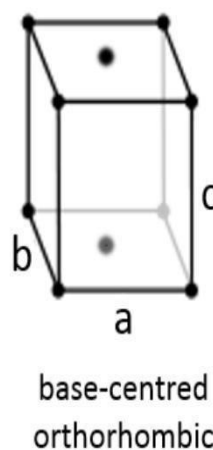
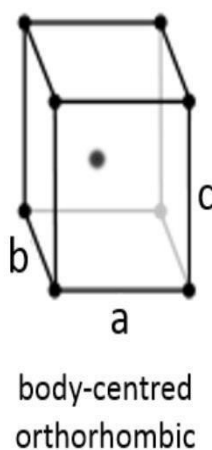
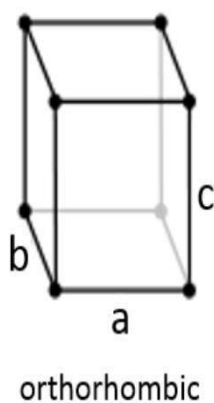
✓ **Quadrilateral:**

When one side has a different length from the other two ($a=b \neq c$), the structure is called a quadrilateral. There is also a body-centered version of this structure.



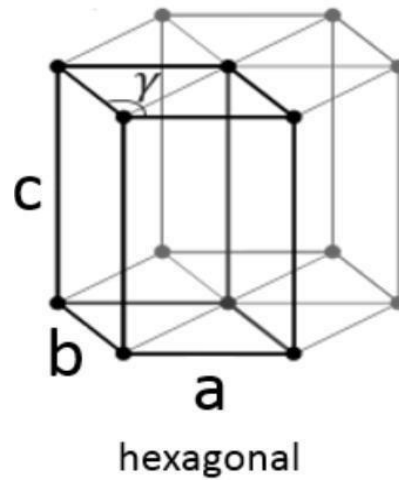
✓ **Orthorhombic:**

The angles are still 90° , but none of the three characteristic lengths are equal ($a \neq b \neq c$). In the case of orthorhombic structures, there is an additional structure: the base-centered structure. In this structure, two opposite sides have an atom at their center, while the other two sides do not.



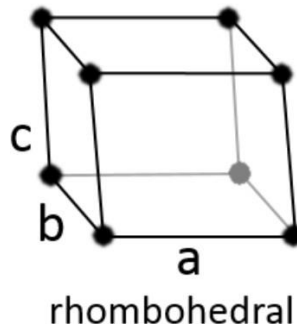
✓ **Hexagonal :**

This structure is called hexagonal, but none of its faces are hexagonal. However, if you assemble three of them, you get a hexagonal prism. The base and top faces are diamond-shaped ($a=b$) with an angle $\gamma=120^\circ$. The other faces of the structure are rectangular ($a=b \neq c$) or square ($a=b=c$).



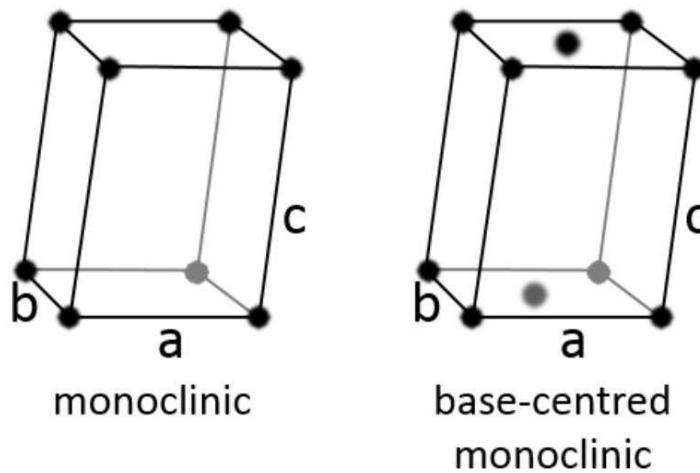
✓ **Rhombohedral :**

As with cubic structures, $a=b=c$ and the angles are equal, but in a rhombohedral structure, the angles are not right angles: $\alpha = \beta = \gamma \neq 90^\circ$



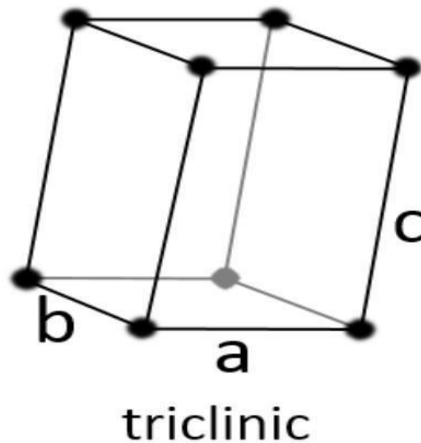
✓ **Monoclinic:**

In this case, two angles are equal to 90° , but the third angle is not a right angle. The sides have different lengths $a \neq b \neq c$.



✓ **Triclinic:**

The angles $\alpha \neq \beta \neq \gamma \neq 90^\circ$ are different, and there are no right angles. The lengths are also different $a \neq b \neq c$



1. Intermediate States:

1.1 Liquid Crystals: These are complex fluids that combine properties of both the liquid phase and the crystallized solid phase.

1.2 Granular States: Granular materials can exhibit behaviors similar to those of a solid, liquid, or gas, depending on the amount of energy they receive.

1.3 Polymer: A polymer is a macromolecule formed by the covalent bonding of a large number of repeating units derived from one or more monomers (also called motifs), and is prepared from molecules known as monomers.

