

Unit I **Water and Mineral
Translocation in Plants**

Chapter |



Plant–Water Relations

1.1 Water – A Universal Solvent

Water is the most abundant resource for nearly all organisms. Covering over 70 per cent of the earth's surface and making up as much as 95 per cent of the matter of living organisms, it is virtually unique among all liquids. The water content of plants is in a continuous state of flux, based on the degree of metabolic activity, the water status of the surrounding air and soil, and a variety of other factors.

Water plays a crucial role in the physiological processes of plants; roles for which it is well-suited because of its special physical and chemical properties, some of which are:

- Water is an excellent solvent, making it a suitable medium for the absorption and translocation of mineral elements and other solutes necessary for normal plant growth and development.
- Most of the 'biochemical reactions', that are so characteristic of life, occur in water and water itself participates directly or indirectly in all the metabolic reactions.
- The 'thermal properties' of water, i.e., high specific heat, high latent heat of vaporization and high heat of fusion, ensure that water remains in the liquid form over the range of temperatures where most biological reactions take place.
- The combined properties of '**cohesion**' (capacity to bind strongly with itself), '**adhesion**' (capacity to bind strongly with other molecules containing oxygen such as glass and cell walls) and '**tensile strength**' (a measure of the maximum tension a material can

Keywords

- Adhesion
- Aquaporins
- Chemical potential
- Cohesion
- Dialysis
- Electrochemical gradient
- Gated channel
- Hydration shell
- Hydraulic conductivity
- Hydrogen bond
- Ion channel
- Proton motive force
- Relative water content
- Tensile strength
- Transport proteins
- Water potential

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withstand before breaking) are especially significant for maintaining the continuity of water columns in plants.

- The 'transparency' of water facilitates penetration of sunlight into the aqueous medium of cells thus enabling photosynthesis and other physiological processes to function in the cells.

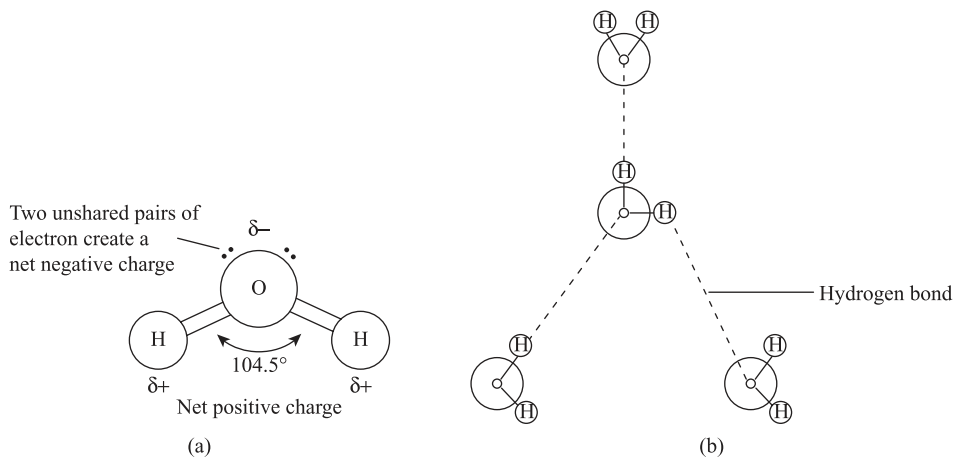


Fig. 1.1 (a) Structural configuration of a water molecule. Because of asymmetrical distribution of charge, the water molecule is polar. The bonding angle between the two hydrogen atoms is 104.5° rather than 180° . The oxygen atom shares one electron with each hydrogen atom. (b) Hydrogen bonds among the water molecules resulting from the electrostatic attraction between the partial positive charge on one molecule and the partial negative charge on the next.

These properties derive primarily from the 'polar' structure of the water molecule (Figure 1.1a). The water (H_2O) molecule (molecular weight, 18) is composed of an oxygen atom covalently bonded to two hydrogen atoms that are attached asymmetrically to one side. The two O–H bonds form an angle of 104.5° rather than 180° . The oxygen atom being more electronegative than hydrogen, has a tendency to attract the electrons of the covalent bond thus resulting in a partial negative charge at the oxygen molecule and a partial positive charge at each hydrogen. Due to equal partial charges, the water molecule has no net charge.

The polarity and the tetrahedral shape of water molecules permit them to form **hydrogen bonds** that give water its unique physical properties (Figure 1.1b). Hydrogen bonding is a key feature that makes certain compounds soluble in water. Besides interactions among water molecules, hydrogen bonding is also accountable for the attraction between water and other molecules. However, when the groups responsible for hydrogen bonding are linked to a group other than water, they lose their solubility. This commonly occurs when hydrogen bonding exists within the same molecule, i.e., intra-molecular hydrogen bonding. For example: cellulose—a polymer of glucose found in plant cell walls. Hydrogen bonding is the basis for formation of **hydration shells** that develop around biological macromolecules such as proteins, lipids, nucleic acids and carbohydrates.

When water freezes below 0°C , it produces crystal 'lattice work', which we know as snow and ice. Ice is much less dense (a density of about 0.9179 g cm^{-3}) than liquid water (0.999 g cm^{-3}). This enables the ice to float on the lakes and ponds surfaces rather than

sinking to the bottom where it might remain there throughout the year. This is extremely important for the survival of all types of aquatic life. If it were otherwise, it would have disastrous consequences, such as killing aquatic life. This unusual feature enables icebergs to float. The maximum density of water is 1.0 g cm^{-3} at 4°C . This happens because molecules in the liquid state are tighter packed than in the crystalline state of ice.

The water expands when it freezes, resulting in the bursting of water pipes, and in plants, the inter- or intra-cellular risk of tissue damage increases manifold during extreme cold conditions.

The ‘**surface tension**’ – the tendency of the surface to contract and resemble an elastic membrane—is high for water (Figure 1.2). And so, water droplets tend to form spheres as the water molecules at the surface are subjected to unequal forces and are pulled in towards the centre. Because of surface tension, the molecules tend to occupy the least possible space. In a deeper part of a mass of water, the molecules of water are attracted with equal force in different directions while at the air–water interface the water molecules, however, are attracted downward and sideways by other water molecules. Here, all of the hydrogen bonds in water face downward, causing the water molecules to cling together, resembling an elastic membrane. Some insects and spiders can glide on water surface due to the surface tension property of water, which is stronger than the force that one foot brings to bear.

The various physiological activities of the cell take place in an aqueous medium. They actually operate in dilute aqueous solutions, suspensions and colloidal phases. A study of their properties is, therefore, necessary for a better understanding of the different physiological processes.

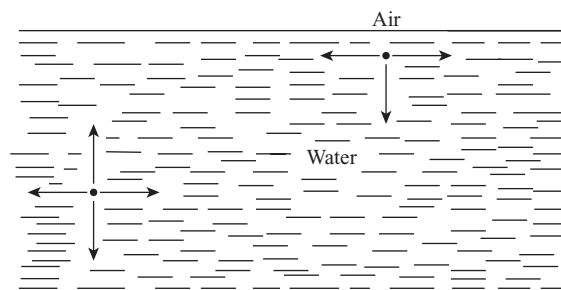


Fig. 1.2 Diagrammatic representation explaining surface tension and forces in a liquid. The particle is pulled downward because of unbalanced forces. A particle deeper in the liquid is attracted by forces from all directions.

1.1.1 Solution

A ‘solution’ is a homogenous mixture composed of two or more components. A component that is present in greater proportion is termed the ‘solvent’ while the other component that is present in smaller amounts is called the ‘solute’.

A solution may exist in gaseous, liquid or solid state. It is convenient to distinguish solutions in terms of the physical states of the mixing components. The various types of solutions with examples are given in Table 1.1.

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Most significant types of solution are: (i) solids-in-liquids; (ii) liquids-in-liquids; and (iii) gas-in-liquids.

Table I.1 Various types of solutions.

Solute	Solvent	Examples
Gas	Gas	Air
Gas	Liquid	CO ₂ in aerated water
Gas	Solid	H ₂ in platinum or palladium
Liquid	Gas	Moisture in air during mist
Liquid	Liquid	Alcohol in water
Liquid	Solid	Jellies, mercury in silver
Solid	Gas	Camphor in air
Solid	Liquid	Sugar in water
Solid	Solid	Solid solutions, alloys, etc.

Terminology connected with the concentration of solutions

The concentrations of solutions are expressed in different units.

Molarity (M): The 'molarity' of a solution gives the number of moles of the solute present per litre of the solution. Thus if one mole of a solute is present in one litre of solution, the concentration of the solution is said to be one molar. The term molar is generally used for salts, sugars and other compounds.

$$\text{Molarity} = \frac{\text{Weight of a solute in g/L of solution}}{\text{Molecular weight of a solute}}$$

Molality (m): The 'molality' of a solution gives the number of moles of the solute present in one kilogram (1000 g) of the solvent. Thus, if one mole of a solute is dissolved in 1000 grams of the solvent, the concentration of the solution is said to be one molal. Molality is used in situations, where it is desirable to express the ratio of solute to solvent molecules, for example, in discussions of osmotic pressure.

$$\text{Molality} = \frac{\text{Weight of a solute in g/kg of solvent}}{\text{Molecular weight of a solute}}$$

Normality (N): In volumetric work, we express the concentrations in terms of normality of the solutions.

The normality of a solution gives the number of gram equivalents of the solute per litre of the solution. Thus, if one gram equivalent of a solute is dissolved per litre of the solution, the concentration of the solution is said to be one normal. A normal solution is generally used for acids and bases.

$$\text{Normality} = \frac{\text{Amount of a solute in g/L of solution}}{\text{Equivalent weight of solute}}$$

Per cent Solution (%): The concentration of substances that do not have a uniformly defined composition, such as proteins, nucleic acids, and such, is expressed in weight per unit volume rather than moles per unit volume. Thus, a 5 per cent solution would imply 5g of solute dissolved in solvent to give a final volume of 100ml of the solution. When either solute or solvent is taken by weight or volume, then per cent solution results in v/v (volume/volume), v/w (volume/weight), w/w (weight/weight) or w/v (weight/volume).

ppm Solution (mg/L or µg/ml solution): ppm solution needs to be prepared when a solute is present in a very small quantity. It represents one gram of a solute per million grams of the solution or one gram of a solute per million ml of the solution.

$$\text{ppm} = \frac{\text{mass of the component}}{\text{total mass of the solution}} \times 10^6$$

or

$$\text{ppm} = \frac{\text{g or ml of solute}}{\text{g or ml of solution}} \times 10^6$$

When 1 mg of a solute is dissolved in water and the volume is made up to 1000 ml, it results in 1 ppm (parts per million) solution.

In the true solution, the molecules of the solute do not settle down and cannot be visibly distinguished from the solvent even under the highest power of microscope. A true solution is transparent and has a higher boiling point, and lower freezing point as well as water potential, as compared to pure water. The number of solvent and solute molecules is constant in a sample of given solution. If to a given amount of solvent maintained at a constant temperature, a solute is added gradually in increasing amounts, a stage will be reached when some of the solute will remain undissolved, no matter how vigorously we stir. The solution is then said to be **saturated**. The excess solute is in dynamic equilibrium with the solute that has gone in solution.

1.1.2 Factors affecting solubility

- 1. Nature of Solute and Solvent:** In general, if a solute and solvent have similar chemical characters, the solubility is very high and if they have entirely different chemical character, the solubility is extremely low. For example, sugars are largely soluble in water but not in benzene.
- 2. Temperature:** The solubility of a solute in a given solvent varies greatly with temperature. It may increase or decrease with rise in temperature. Most of the salts (e.g., KNO_3 , NH_4Br) show marked increase in solubility with rise in temperature, whereas others (such as, NaCl) show only a small increase. There are only a few substances (e.g., anhydrous sodium sulphate, Na_2SO_4 , caesium sulphate) which show decrease in solubility with rise in temperature.

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'Standard solution' contains a known weight of the substance dissolved in a known volume of the solvent.

1.1.3 Suspension

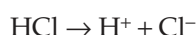
A 'suspension' is a two-phase system where solute particles are not dispersed as molecules or ions. The larger solute particles ($>0.1 \mu\text{m}$ in diameter) remain suspended throughout the liquid. The system is unstable and the particles settle down with the passage of time.

1.1.4 Emulsion

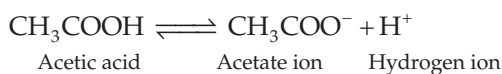
Unstable emulsions are prepared by vigorously shaking two immiscible liquids together. Small droplets (dispersed phase) of one of the liquids will be dispersed throughout the other (dispersion medium). An emulsion can be made stable by the addition of an emulsifying agent. These substances generally function either by decreasing the surface tension of the liquids or by forming a protective layer around the droplets, thus making it impossible for them to combine with each other. Milk is an example of emulsion containing butter fat dispersed in water, with casein acting as an emulsifying agent.

1.1.5 Acids, Bases and Salts

Acids are ionizable compounds that give up protons. Strong acids like HCl, H_2SO_4 and HNO_3 tend to ionize completely when in solution.



Weak acids are the organic acids and tend to ionize only partially when in solution. Thus, acetic acid ionizes to form acetate ions and protons, but some acetic acid still remains unionized.



Bases are ionizable compounds that tend to take up protons. Strong bases such as sodium hydroxide (NaOH) dissociate into its ions, when in solution.



In contrast, ammonia is a relatively weak base. Occasionally, substances that yield OH^- ions are considered bases. For example: acetate (CH_3COO^-) is the conjugate base to acetic acid (CH_3COOH).

Salts are compounds formed from the conjugate base of an acid and the conjugate acid of a base, e.g. on mixing a solution of NaOH with acetic acid, the salt of sodium acetate is formed.

1.1.6 Electrolytes and Non-electrolytes

Electrolytes are those substances that can conduct an electric current and undergo ionization when dissolved in water. Acids, bases and salts are electrolytes. Some common strong electrolytes with complete ionization in plant tissues are the inorganic ions K^+ , Na^+ , Ca^{2+} , Mg^{2+} , Mn^{2+} , Zn^{2+} , Cl^- , SO_4^{2-} and PO_4^{3-} . Common weak electrolytes with incomplete ionization are the organic cations and anions, amino acids, organic acids and the complex charged polymers, such as proteins and nucleic acids. The behaviour of electrolytes was first discussed by Arrhenius in 1880.

Non-electrolytes are substances that do not undergo ionization on dissolving in water. Examples are: gases (oxygen and hydrogen) and most of the organic compounds such as sugars and alcohols and others. However, they do carry groups having positive or negative potential, such as, amino NH_2 , imino NH , amide $CONH_2$, carbonyl CO , carboxyl $COOH$, hydroxyl OH and so on. These groups combine with water molecules which protect them from coming together by forming water shells.

1.1.7 Polar and Non-polar Compounds

Jacobs (1924) emphasised the importance of recognizing two classes of compounds, polar and non-polar, from the stand point of permeability. Polar compounds are compounds that have slight charges (+ve as well as -ve) within the compound. For example, water (H_2O), hydrogen fluoride (HF), hydrogen chloride (HCl), and ammonia (NH_3). Slight charges are due to electro-negativity differences between 0.3 and 1.4. In polar molecules, two atoms do not share electrons equally in a covalent bond. In addition to water, acids, bases and salts, there are many polar organic compounds such as sugars and alcohols.

Non-polar compounds do not have slightly negative and positive charges within the compound (in other words, the electrical charges are evenly distributed across the molecule), such as hydrocarbons, organic molecules: benzene, methane, ethylene, carbon tetrachloride, and the noble gases: helium, neon, argon, krypton, xenon and so on. Electronegativity differences of non-polar compounds lie between 0 and 0.2. In non-polar molecules, electrons are shared equally by the atoms of the molecules and there is no net electrical charge. Some molecules such as those of the fatty acids are mostly non-polar (the long hydrocarbon chain) but have polarity at one end ($-COOH$ group).

1. Polar substances tend to be soluble in water and other polar solvents while non-polar substances are soluble in fats and organic solvents.
2. Polar compounds, with the exception of water, penetrate cells with difficulty, whereas non-polar compounds penetrate cells rapidly.
3. Polar compounds ionize readily and are highly reactive (except sugars and alcohols), whereas non-polar compounds do not ionize and are electrically neutral.

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1.1.8 pH and Buffers

pH

The term 'pH' has been derived from the French word 'Pouvoir Hydrogene' (power of hydrogen) and is used to quantitatively measure the acidity or alkalinity of aqueous or other liquid solutions. The concept of 'pH' was introduced by a Danish chemist, Søren Peder Lauritz Sørensen (1909), to express concentration of the hydrogen ion (H^+) in an aqueous solution. The 'pH' can be defined as the 'negative logarithm' of the hydrogen ion concentration.

$$pH = -\log [H^+]$$

or

$$\text{mathematically speaking, } pH = \log \frac{1}{[H^+]}$$

Water dissociates into equal numbers of H^+ and OH^- ions and is therefore neither an acid nor a base. At ordinary temperature (22 °C), the actual concentration of H^+ (and of OH^-) ions in pure water is very small, i.e., 10^{-7} mol/L each of H^+ (protons) and OH^- (hydroxyl) or [0.0000001 M]. The product of the two ($H^+ \times OH^-$) is 10^{-14} and this remains constant for all aqueous solutions. By knowing the concentration of H^+ , the concentration of OH^- can be calculated or vice-versa.



$$[H^+][OH^-]$$

$$[10^{-7}][10^{-7}] = 10^{-14} \text{ (to multiply } 10^{-7} \text{ by } 10^{-7}, \text{ exponents are added to give } 10^{-14} \text{.)}$$

The pH of pure water, (with H^+ concentration of 1.0×10^{-7} M), will be 7 as calculated below:

$$\begin{aligned} pH &= \log \frac{1}{[H^+]} = \log \frac{1}{1 \times 10^{-7}} \\ &= \log [1 \times 10^7] \\ &= \log 1.0 + \log 10^7 \\ &= 0 + 7 \\ pH &= 7 \end{aligned}$$

Hence, pure water with equal numbers of protons and hydroxyl ions will have a pH of 7. However, the pH of water is lowered down at higher temperature, e.g., at 50 °C, the pH of water is 6.55 ± 0.01 .

The pH scale covers a range of values from 0 to 14, in which the midpoint is 7.0 (Figure 1.3). Solutions having pH less than 7.0 are acidic (having more H^+ ions), while those having pH greater than 7.0 are basic or alkaline (having more OH^- ions than H^+ ions). Thus a solution of 6.0 pH will have 10 times more H^+ ions than that of a 7.0 pH solution. Similarly a solution of 8.0 pH will have 10 times less H^+ than that of a 7.0 pH solution.

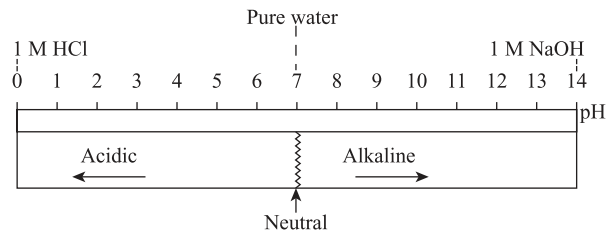


Fig. 1.3 The pH scale.

There is always an inverse relationship between the H^+ and OH^- ions concentration in a solution (Table 1.2). One needs to know only the value $[H^+]$ to determine $[OH^-]$ as well. Even a very strong acid of say pH 1 contains some OH^- ions ($10^{-13}N$) and a strong base of pH 14 has a few H^+ ions ($10^{-14}N$).

Table 1.2 Relationship between H^+ and OH^- ions concentration in a solution and the corresponding pH values.

H^+ mol/L	OH^- mol/L	$\text{Log } [H^+]$	$-\text{Log } [H^+]$ = pH
1×10^0	1×10^{-14}	0	0
1×10^{-1}	1×10^{-13}	-1	1
1×10^{-4}	1×10^{-10}	-4	4
1×10^{-7}	1×10^{-7}	-7	7
1×10^{-10}	1×10^{-4}	-10	10
1×10^{-13}	1×10^{-1}	-13	13
1×10^{-14}	1×10^0	-14	14

Thus, 1N HCl or any other strong acid that dissociates (100 per cent) shall have a pH of 0 because

$$\begin{aligned}
 \text{pH} &= -\log [H^+] \text{ or } [H] \text{ would be equal to 1.} \\
 &= -\log [1] \\
 &= 0
 \end{aligned}$$