CHEM 191

Module 1
Chemical Reactions in Aqueous Solution

Lecture 2

Aqueous solutions

Brown (15th) Chapter 11.2

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Module 1 Lecture 2

Learning objectives

- **To understand how hydrogen bonding leads to the peculiar properties of water and its solvent properties**
- **To understand how interactions between the water molecules and solutes enable the solute to dissolve**
- **To understand the difference between a strong and a weak electrolyte**

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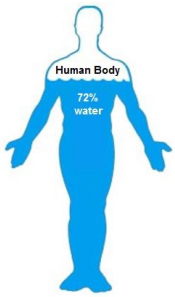
Aqueous solutions



- Aqueous solutions are important in biological systems.
- The solubility of substances in aqueous biological solutions is crucial in
  - Transport of gases
  - Transport of inorganic materials
  - Application and transport of drugs
- Why are some substances soluble in water and some not?
- Consider the nature of the water molecule

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The water molecule

- A water molecule contains two O-H bonds. Because the oxygen atom is much more electronegative than the hydrogen atom, the pair of electrons in the bond are not equally shared, but instead spend more of the time surrounding the oxygen atom.
- This gives the oxygen atom a partial negative charge (δ-), and consequently the hydrogen atom has a partial positive charge (δ+).

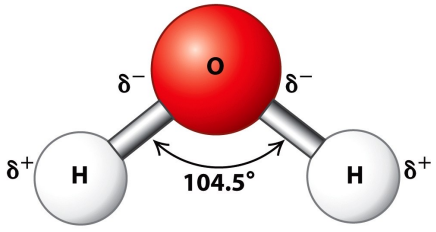

Figure 2-5
Molecular Cell Biology, Sixth Edition

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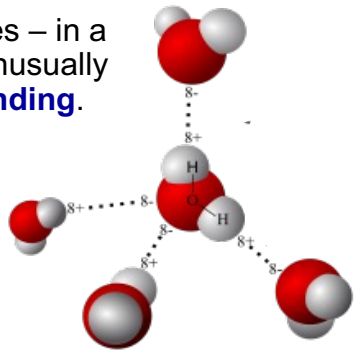
The water molecule

- Water molecules are attracted to each other by electrostatic interactions between these partial charges – a dipole-dipole interaction.

- Dipole-Dipole interactions are found in many molecules – in a small number, including water, these attractions are unusually strong and are given a special name – **Hydrogen Bonding**.



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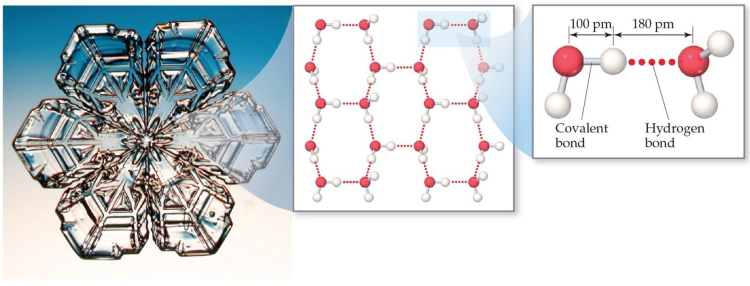
Water

Water is a polar molecule which engages in extensive **hydrogen bonding** interactions in both the solid and solution states.



▲ FIGURE 11.10 **Hydrogen bonding in ice.** The empty channels in the structure of ice make water less dense as a solid than as a liquid.

Polar molecules and ionic solids tend to be soluble in water.

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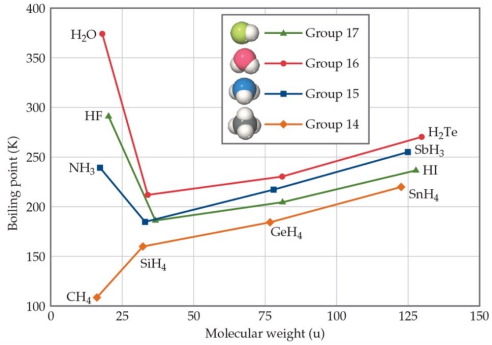
Unusual properties of water

- Very high melting and boiling points compared to analogous molecules

	T_m / °C	T_b / °C
HF	-83	19.5
H_2O	**0**	**100**
NH_3	-77.7	-33.4
CH_4	-183	-161
H_2S	-85.5	-60.7



- Due to extensive H-bonding

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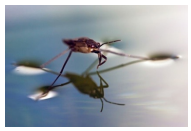
Unusual properties of water

- Water gets *more* dense when it melts. Its maximum density occurs at 4 °C. Therefore, ice floats and lakes freeze from the top down.

- The freezing point of water *decreases* as pressure increases
- There are 14 different forms of ice. One form (ice X) melts at 2100 °C (at 800,000 times atmospheric pressure)
- Water has a relatively large surface tension and heat capacity



- High concentration **55.5** mol L^{-1} at 25°C **!!**

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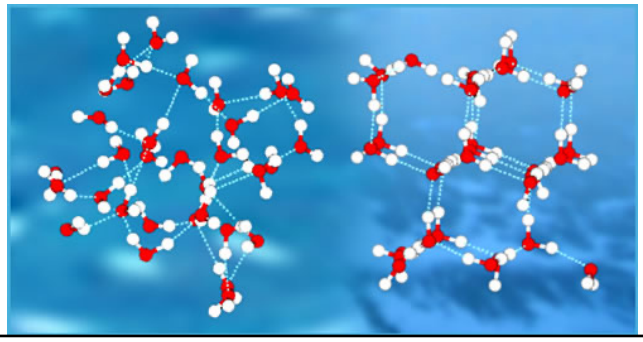
Unusual properties of water

- In liquid water, the molecules move around close to each other, continuously making and breaking hydrogen bonds with their neighbours. In solid water (ice) the molecules are rigidly held in a structure which maximised the possible number of hydrogen bonds – this structure is more open and so ice is less dense than liquid water.



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Water as a solvent

- Water is an excellent solvent for electrolytes, substances that give ions when dissolved.

  e.g. **$NaCl(s) \rightarrow Na^+(aq) + Cl^-(aq)$**

- The dissolution occurs spontaneously, despite the fact that the electrostatic forces holding NaCl(s) together are strong [NaCl(s) melts at 801 ^oC]

- Ions, by themselves, are very high-energy species. However, solvation helps stabilise the ions.

- Solvation involves interactions of the charged ions with the dipoles of the water molecules. All ions in water are surrounded by a number of water molecules.

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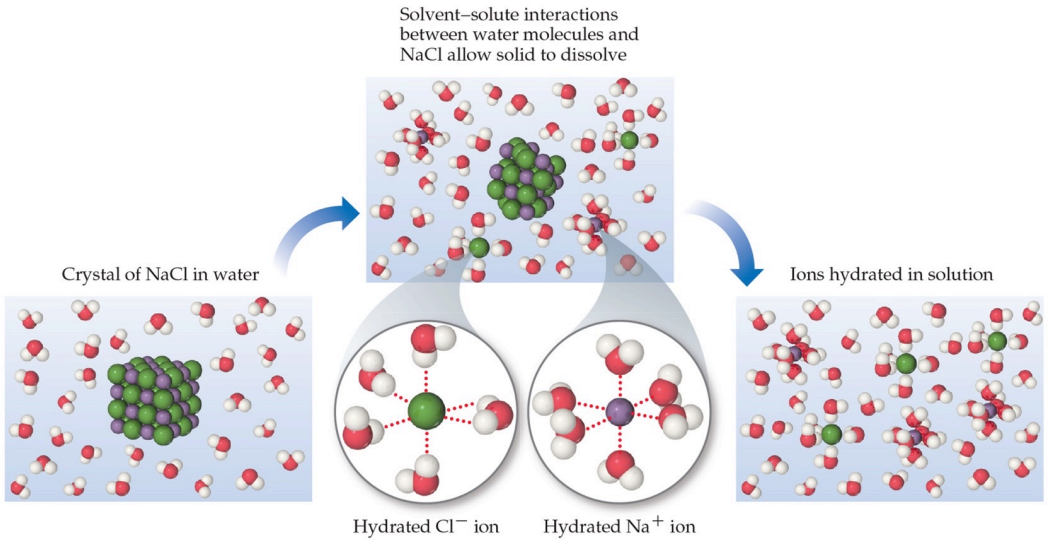
Solvation

When water is the solvent, this is called ***hydration***



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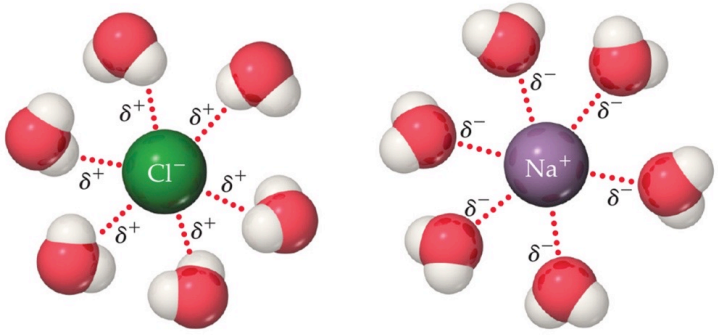
Solvation

When water is the solvent, this is called ***hydration***



Positive ends of polar molecules are oriented towards negatively charged anion

Negative ends of polar molecules are oriented towards positively charged cation

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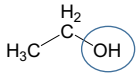
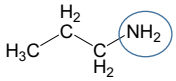
Aqueous solubility

- Water is a good solvent for **polar molecules** – ones which also have polar bonds due to large differences in electronegativity between atoms in a bond.
- These include solids and liquids such as

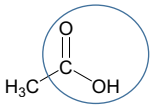
Alcohols



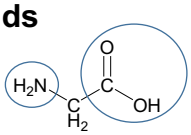
Amines



Carboxylic acids



Amino Acids



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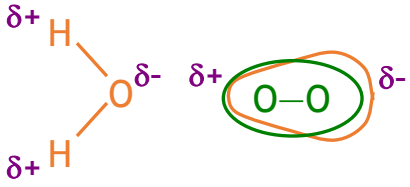
Aqueous solubility

- Non-polar molecules such as $O_2(g)$ and $N_2(g)$ are generally not very soluble in water. However, their solubility is not zero
- This is due to the fact that the permanent dipole of water can induce a dipole in the non-polar molecule.



- The magnitude of the induced dipole depends on the size (number of electrons) and shape of the gas molecule.

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Aqueous solubility

- Non-polar gases are not very soluble in water

gas	Solubility / g kg^{-1} (298 K)
He	0.00150
O_2	0.0393
N_2	0.0175
Cl_2	6.41
CO_2	1.45
NH_3	489

- Ammonia (NH_3) can hydrogen-bond and is very soluble in water. *Note though, that it also undergoes a chemical reaction with water – see later*

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Definitions

- A **solution** is formed when 1 or more chemical species dissolve in a liquid
- A **solute** is a substance that dissolves in a liquid. It can be solid, liquid or gas
-
- A **solvent** is a liquid in which the solute dissolves
- A **dilute solution** has *n*(solute) << *n* (solvent)
- An **electrolyte** dissolves to give ions in solution – e.g. NaCl, $AgNO_3$, $CuSO_4 \cdot 5H_2O$
- A **non-electrolyte** dissolves without the formation of ions – e.g. glucose ($C_6H_{12}O_6$), urea ($(NH_2)_2CO$)

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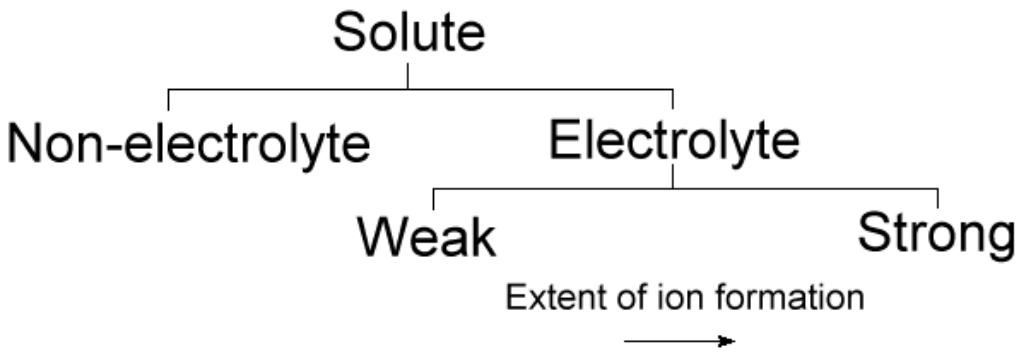
Classification of solutes

- Can divide solutes into different types



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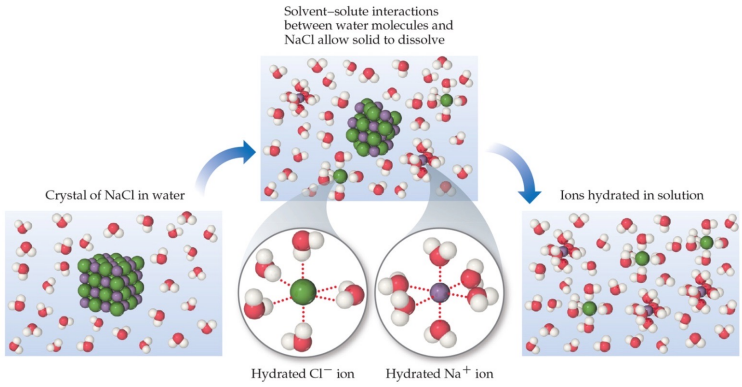
Electrolytes and non-electrolytes

- Sodium chloride (salt): Electrolyte. An ionic solid - Na^+ and Cl^- occupy lattice sites. Dissolution spontaneous in water

$$NaCl(s) \rightarrow Na^+(aq) + Cl^-(aq)$$

- In solution, Na^+ and Cl^- ions are uniformly distributed throughout H_2O



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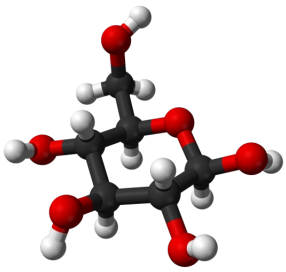
Electrolytes and non-electrolytes

- Glucose: Non-electrolyte. A molecular solid - molecules held together by Hydrogen-bonds. Dissolution spontaneous in water

$C_6H_{12}O_6(s) \rightarrow C_6H_{12}O_6(aq)$



- In solution, molecules uniformly distributed throughout H_2O. The molecules are unchanged in their bonding – no ions are formed.

- **Electrolytes include all ionic solids, acids, bases**

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Strong Electrolytes

- These undergo **complete** dissociation in water.

- When the compound dissolves, all the dissolved species are ions – in other words, the dissociation reaction 'goes to completion'.

- Even ionic solids which are not very soluble in water (e.g. AgCl) are classified as strong electrolytes, as all of the solid that dissolves is completely dissociated into Ag^+ and Cl^- ions.

$AgCl(s) \rightarrow Ag^+(aq) + Cl^-(aq)$

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Weak Electrolytes

- These undergo incomplete dissociation in water.
- When the compound dissolves, most of the dissolved species are molecules, some of which then dissociate to give ions
- Here, the dissociation reaction does not 'go to completion'.
- Weak acids, such as acetic acid, are the classic examples of weak electrolytes.

$$CH_3COOH(l) + H_2O(l) \rightarrow CH_3COOH(aq) \rightleftharpoons CH_3COO^-(aq) + H_3O^+(aq)$$

- How can we quantify this?

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✱ Homework ✱

NO HOMEWORK TODAY

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