

Electrolytes and the Dissolving Process

What happens when a substance dissolves, and why do some solutions conduct electricity?

●○○ Beginner

Solutions

Last reviewed 6 October 2026

Before you start: it helps to understand [Ionic and Covalent Bonds](#) and [Intermolecular Forces](#).

What is it?

A **solution** is a uniform mixture. The substance present in the larger amount is the **solvent** (usually water); the dissolved substance is the **solute**.

When a substance dissolves, its particles separate and become surrounded by solvent molecules. Water is a **polar** molecule, with a slightly negative oxygen and slightly positive hydrogens, so:

- **Ionic compounds** split into separate ions: $\text{NaCl (s)} \longrightarrow \text{Na}^+ \text{ (aq)} + \text{Cl}^- \text{ (aq)}$. Each ion is surrounded by water molecules (it is **hydrated**): the oxygen side faces positive ions and the hydrogen side faces negative ions.
- **Polar molecular compounds** such as sugar and ethanol dissolve as **whole molecules**, held by hydrogen bonds and dipole attractions.

A solution that conducts electricity contains **ions** that can move. Solutes are classified by how many ions they give:

Type	In water	Conducts?	Examples
Strong electrolyte	Completely split into ions	Strongly	Soluble ionic compounds (NaCl , CaCl_2), strong acids (HCl , HNO_3), strong bases (NaOH)
Weak electrolyte	Only a small fraction forms ions	Weakly	Weak acids (ethanoic acid), weak bases (ammonia)
Non-electrolyte	Stays as molecules	No	Sugar, ethanol, urea

KEY IDEA

Electrical conductivity in solution comes from **moving ions**. The more ions a solution contains, the better it conducts. That is why salt water conducts, sugar water doesn't, and vinegar conducts only weakly.

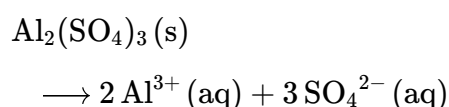
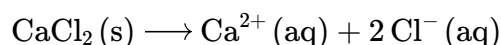
Why does it matter?

- **Your body runs on electrolytes.** Nerve signals and heartbeats depend on Na^+ , K^+ and Ca^{2+} ions; sports drinks replace the ions lost in sweat.
- **Counting particles.** One mole of CaCl_2 gives three moles of ions. This matters for reactions, for colligative properties and for net ionic equations.
- **Hot and cold packs.** Some substances absorb heat as they dissolve, others release it.

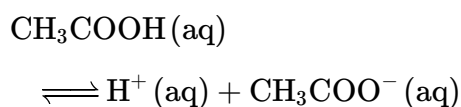
How does it work?

1. Writing dissociation equations

For a strong electrolyte, show every ion with its coefficient:



A weak electrolyte is shown with an equilibrium arrow, because most of it stays as molecules:



2. Ion concentrations

Multiply the concentration of the compound by the number of each ion in its formula. In 0.150 mol/L CaCl_2 : $[\text{Ca}^{2+}] = 0.150 \text{ mol/L}$ and $[\text{Cl}^-] = 2 \times 0.150 = 0.300 \text{ mol/L}$.

3. Energy changes on dissolving

Dissolving an ionic solid takes two steps:

1. **Breaking up the lattice** (separating the ions): needs energy, **endothermic**.
2. **Hydrating the ions** (attracting water molecules): releases energy, **exothermic**.

If hydration releases **more** energy than the lattice needs, dissolving is exothermic and the solution warms up (CaCl_2 , NaOH). If it releases **less**, dissolving is endothermic and the solution cools (NH_4NO_3 , used in instant cold packs).

∞ THINK OF IT LIKE THIS

Dissolving salt is like a crowd breaking up at a party. Each guest (ion) leaves their own group and is soon surrounded by a circle of new friends (water molecules). If the new friendships are stronger than the old ones, energy is released; if they are weaker, energy has to be put in.

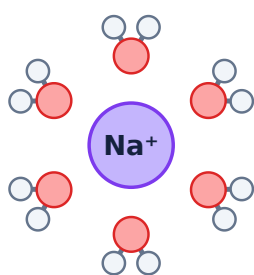
🔍 MORE PRECISELY

Even strong electrolytes are not perfectly “free” ions in concentrated solutions: oppositely charged ions attract each other and form ion pairs, so solutions behave as if they contain slightly fewer particles. The “strong” and “weak” labels describe the **extent** of ionization, not the concentration: a dilute strong acid can be less acidic than a concentrated weak one. Whether a substance dissolves also depends on entropy (the spreading out of particles), which favours dissolving even when it is slightly endothermic.

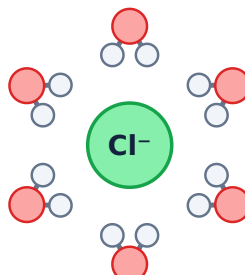
Visualise it

Hydrated ions

Red: oxygen (δ^-) · grey: hydrogen (δ^+)



O (δ^-) faces Na^+



H (δ^+) faces Cl^-

Each ion is surrounded by water molecules.

Hydrated ions: water molecules turn their oppositely charged ends towards each ion.

Worked example

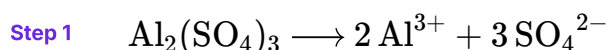
✏️ WORKED EXAMPLE: Classifying solutes

Question: Classify each as a strong, weak or non-electrolyte: (a) KNO_3 (b) glucose (c) HCl (d) NH_3 (e) ethanol

- Step 1** (a) Soluble ionic compound: **strong**.
- Step 2** (b) Molecular, does not ionize: **non-electrolyte**.
- Step 3** (c) Strong acid, fully ionized: **strong**.
- Step 4** (d) Weak base, only partly ionized: **weak**.
- Step 5** (e) Molecular, does not ionize: **non-electrolyte**.

 WORKED EXAMPLE: Ion concentrations

Question: What are the ion concentrations in 0.0500 mol/L $\text{Al}_2(\text{SO}_4)_3$?



Step 2 $[\text{Al}^{3+}] = 2 \times 0.0500 \text{ mol/L} = 0.100 \text{ mol/L}$

Step 3 $[\text{SO}_4^{2-}] = 3 \times 0.0500 \text{ mol/L} = 0.150 \text{ mol/L}$

Step 4 Total ions: 0.250 mol/L, five times the concentration of the compound.

 WORKED EXAMPLE: Mixing two solutions

Question: 25.0 mL of 0.200 mol/L NaCl is mixed with 25.0 mL of 0.100 mol/L CaCl_2 . What is $[\text{Cl}^-]$ in the mixture?

Step 1 Moles of Cl^- from NaCl: $0.200 \text{ mol/L} \times 0.0250 \text{ L} = 5.00 \times 10^{-3} \text{ mol}$

Step 2 From CaCl_2 (2 Cl^- each): $2 \times 0.100 \text{ mol/L} \times 0.0250 \text{ L} = 5.00 \times 10^{-3} \text{ mol}$

Step 3 Total volume 0.0500 L:

$$\begin{aligned} [\text{Cl}^-] &= \frac{1.000 \times 10^{-2} \text{ mol}}{0.0500 \text{ L}} \\ &= 0.200 \text{ mol/L} \end{aligned}$$

 WORKED EXAMPLE: A cold pack

Question: 10.0 g of NH_4NO_3 dissolves in 100. g of water. $\Delta H(\text{solution}) = +25.7 \text{ kJ/mol}$. By how much does the temperature fall?

Step 1 Moles: $\frac{10.0 \text{ g}}{80.05 \text{ g/mol}} = 0.1249 \text{ mol}$

Step 2 Heat absorbed: $0.1249 \text{ mol} \times 25.7 \text{ kJ/mol} = 3.210 \text{ kJ} = 3210 \text{ J}$

Step 3 This heat comes from the 110. g of solution:

$$\begin{aligned} \Delta T &= \frac{q}{mc} \\ &= \frac{3210 \text{ J}}{110. \text{ g} \times 4.184 \frac{\text{J}}{\text{g}^\circ\text{C}}} \\ &= 6.98^\circ\text{C} \end{aligned}$$

Step 4 The solution cools by about 7°C .

Common mistake

COMMON MISTAKE: Thinking all dissolved substances conduct

Sugar dissolves completely but forms **no ions**, so sugar solution does not conduct. Dissolving and ionizing are different things.

COMMON MISTAKE: Forgetting the subscripts

CaCl_2 gives **two** chloride ions per formula unit, so $[\text{Cl}^-]$ is twice the concentration of the compound.

COMMON MISTAKE: Splitting polyatomic ions

Na_2SO_4 gives $2\text{Na}^+ + \text{SO}_4^{2-}$. The sulfate ion stays together; it does not break into S and O.

NOTATION NOTE

- Square brackets mean concentration in mol/L: $[\text{Cl}^-]$.
- (aq) means “dissolved in water” (aqueous).

Remember this

REMEMBER THIS

- Solute dissolves in solvent; water is polar and hydrates ions.
- Strong electrolytes: soluble ionic compounds, strong acids and bases (fully ionized). Weak: weak acids and bases. Non-electrolytes: molecular solutes such as sugar.
- Ion concentration = compound concentration $\times$ number of that ion in the formula.
- Dissolving: lattice breaking (endothermic) + hydration (exothermic). The balance decides hot or cold.

References

1. Brown, T. L.; LeMay, H. E., Jr.; Bursten, B. E.; Murphy, C. J.; Woodward, P. M.; Stoltzfus, M. W. Chemistry: The Central Science, 15th ed.; Pearson, 2022.

Practise this topic with flashcards and a quiz at chemistryclarity.com/chemistry/electrolytes-and-dissolving/