

Conductivity is one of the most important tools available to chemists for studying the properties of aqueous solutions. Current electricity is the flow of charge and thus for a solution to conduct it must contain mobile charges. In an aqueous solution these charge carriers are ions. By measuring the ability of a solution to conduct electricity a chemist can infer the number of charges (ions) present in a solution. In order to carry enough current to support the lighting of a 150 W incandescent bulb there needs to be a very significant number of ions present (~ 0.1 M). To support the lighting of a 7.5 W bulb or to dimly light a 25 W bulb there needs only be ~ 0.001 M ion concentration. Chemists conveniently classify solutes based on their ability of conduct when placed in an aqueous solution.

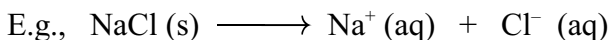
Electrolyte: A solute that conducts electricity when in aqueous solution.

Non-electrolyte: A solute that does not conduct electricity when dissolved in aqueous solution.

Question: If a solution is found to conduct, what can chemists infer about the nature of the bonding within the original undissolved solute?

Clearly the solute is an electrolyte but, that doesn't mean that the solute must be an ionic compound. That's because there are three different methods by which ions can end up in the final solution.

Method I. Dissociation of Ionic Solids: During the dissolution process the ions that make up the ionic compound are separated by the water molecule. This separation of the ions allows the individual ions to move independent of other ions and thus are available to act as charge carriers.



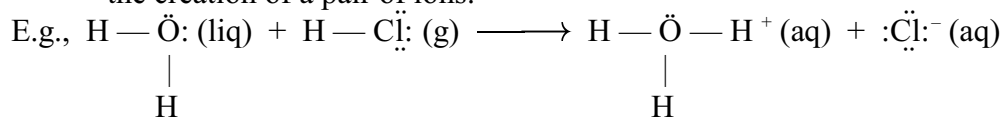
Note: In a solid crystal of NaCl there are ions (Na^+ and Cl^-) but, they cannot move and therefore NaCl (s) is not a conductor of electricity. It is only after the dissolution process that the ions become mobile and thus the solution is conductive.

Solute-Solvent Reaction: When a molecular solute and a molecular solvent are mixed a reaction **might** occur between the solute and solvent which creates ions. In order to create ions from molecular substances a charged particle of some kind must be broken off from one molecule and added to the other molecule to create a pair of oppositely charged ions. There are two types of solute-solvent reactions.

Method II

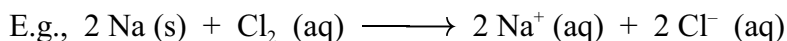
Acid/Base Reaction:

A covalently bonded hydrogen on either the solute or solvent is attracted to an unshared pair of electrons on the other component. If this attraction is great enough, the hydrogen nucleus will release its grip on the pair of electrons in the original covalent bond and grab hold of the lone pair of electrons on the other component. This results in the transfer of an H^+ from one component to the other and the creation of a pair of ions.



Method III
Redox Reaction:

One or more valence electrons on either the solute or solvent is transferred to the other component. This results in the creation of a pair of ions. This type of reaction is called an oxidation/reduction reaction or redox reaction for short



Thus, the answer to the question previously posed, “IF a solution is found to conduct, what can chemists infer about the nature of the bonding within the original undissolved solute?” is that chemists cannot infer anything definite about the nature of bonding within the original solute. The solute could be salt held together by ionic bonding which dissociated its constituent ions to create a conductive solution. Or, the solute could have been a molecular substance which underwent either an acid/base reaction or an oxidation/reduction reaction to create ions which then caused the resulting solution to be conductive. So the bonding within the solute could be either ionic or covalent.

Question: IF a solution is found to be nonconductive, what can chemists infer about the nature of the bonding within the original undissolved solute?

In this instance, the inferences drawn by chemists are rather straight forward. Since the solution is nonconductive, it follows that the solute is a non-electrolyte. Furthermore, since the solution contains no ions in any measurable amounts, the solute cannot be an ionic compound because ionic compounds would dissociate during the dissolution process to liberate ions and produce conductive solutions. Thus, the solute **must** be a molecular compound and therefore the bonding is covalent.

Question: IF a solute is known to be an ionic solid, what can chemists infer about the electrolytic nature of the solute?

All ionic solids dissociate ions during the dissolution process. The resulting solution would therefore, contain ions and thus must be conductive. The solute **must** be an electrolyte.

Question: IF a solute is known to be a molecular substance, what can chemists infer about the electrolytic nature of the solute?

As a molecular substance, the solute itself does not contain ions. Some molecular compounds simply dissolve in the solvent without reacting to produce a solution devoid of ions. In this case, the solute would be a non-electrolyte. On the other hand, some molecular solute react with the solvent to produce ions via either an acid/base reaction or an oxidation/reduction reaction. Under these conditions the resulting solution would conduct and the solute would be classified as an electrolyte. Thus, no conclusive inferences can be drawn about the electrolytic nature of the solute simply from the fact the solute is molecular.

Dissociation of Ionic Solids: (Method I)

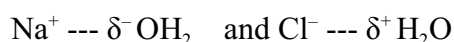
The forces of attraction between the cations and anions that hold ionic solids together obey Coulomb's Law.

I.E., $F \propto \frac{q_1 \times q_2}{r^2}$ where q_1 is the charge on one ion, q_2 is the charge on the other ion, and r is the distance between the centers of the ions.

These forces in NaCl are strong enough to give NaCl a melting point of 812°C . The forces of attraction between water molecules are much weaker than the forces in NaCl and are only sufficiently strong to give water a melting point of 0°C .

Question: How can water cause the ions in NaCl to come apart as the salt dissolves in water?

Recall that water has a bent structure and thus is a polar molecule. The δ^- end of the dipole is on the more electronegative oxygen while the δ^+ end of the dipole is shared by the two hydrogens. Thus, the ions in the ionic solid can interact with the appropriate ends of the water molecules to produce ion-dipole attractions.



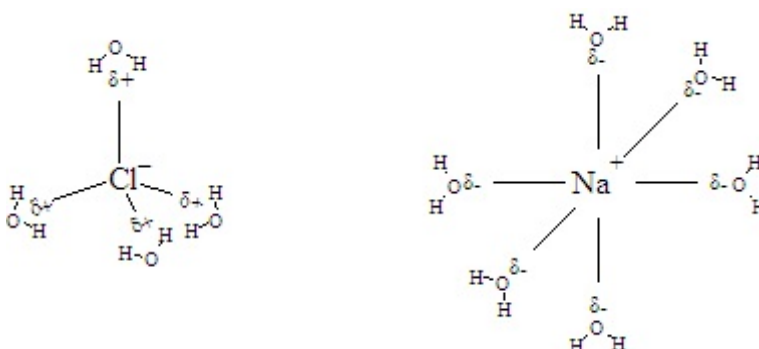
However, the δ^- on the O's in H_2O are only a fraction of the $-$ charges on the Cl^- ions and the δ^+ on the H's in H_2O are only a fraction of the $+$ charges on Na^+ ions. Thus, the forces of attraction between the Na^+ ions and Cl^- ions must be larger than the ion-dipole attractions between the ions and water (the radii of the ions and water are similar).

I.E., Force of $\text{Na}^+ \text{ --- } \text{Cl}^- > \text{Force of } \text{Na}^+ \text{ --- } \delta^- \text{OH}_2$ and
Force of $\text{Na}^+ \text{ --- } \text{Cl}^- > \text{Force of } \text{Cl}^- \text{ --- } \delta^+ \text{H}_2\text{O}$

So, again we ask, "How can water cause the ions in NaCl to come apart as the salt dissolves in water?"

The answer of course is to realize that when a pair of Na^+ and Cl^- ions dissolve they interact with MORE than one water molecule each! The individual ions are surrounded by water molecules. In the case of Na^+ , the ion is spherical and thus the attractions to water are the same in every direction. Due to the sizes of water molecules and the Na^+ ion each Na^+ can accommodate 6 water molecules. The Cl^- ion has 4 pairs of electrons arrayed in the shape of a tetrahedron and thus can interact with 4 water molecules.

I.E.,



Even though the individual ion-dipole interactions are weaker than the individual ion-ion interactions in salts, the cumulative effects of having many ion-dipole interactions with each ion allows water to dissolve ion compounds. Thus, the ions are separated as they dissolve into a solution.

INVENTORY:

In chemistry, an inventory is NOT an all inclusive list of everything that is present in the solution. It is simply a list of the MAJOR solute species (both molecules and ions) that are present in a solution in relatively high concentrations. Why is it necessary to take an inventory of a solution?

Chemists usually refer to solutions based on the components that were combined to make the solution and NOT to what is actually present in the solution. Recall, that there are several different possible results when a solute and solvent are mixed. The solute could: merely dissolve; it could undergo an acid/base reaction with the solvent; or it could undergo an oxidation/reduction reaction with the solvent. Some of these may have even occurred without the chemists knowing it! Thus, chemists prefer to list the original solution components and allow the user of the solution to deduce the actual solute species present. The process of deducing the current solution components is referred to as “inventorying.”

For example, 1M NaCl. In this case, the 1M NaCl “label” merely means that NaCl was mixed with water and that there was originally 1 mole of NaCl for every liter of final solution. It doesn’t mean that was only 1 mole of NaCl or that there was only 1 liter of solution. It doesn’t even mean that there is **any** NaCl left in the final solution!

Indeed, since the ions of the ionic solid are completely separated upon dissolution in water there are no “NaCl” units left but rather Na^+ (aq) and Cl^- (aq) ions.

I.e., $\text{NaCl (s)} \longrightarrow \text{Na}^+ \text{ (aq)} + \text{Cl}^- \text{ (aq)}$ Large Extent (100%).

What, then, is the “inventory” for this solution? Clearly, the solute species existing in the final solution are Na^+ (aq) and Cl^- (aq) ions. The inventory of this solution is written as $[\text{Na}^+] = 1\text{M}$ and $[\text{Cl}^-] = 1\text{M}$ where [xxx] is read as “concentration of”.

Thus, the equation for **any** reaction involving this solution **should not** include NaCl but rather Na^+ and/or Cl^- .

Question: How should 0.25M CaCl_2 be inventoried? Upon dissolution, the ions will be separated:

$\text{CaCl}_2 \text{ (s)} \longrightarrow \text{Ca}^{+2} \text{ (aq)} + 2 \text{Cl}^- \text{ (aq)}$ Large Extent (100%).

Thus, the inventory should be: $[\text{Ca}^{+2}] = 0.25\text{M}$ and $[\text{Cl}^-] = 0.50\text{M}$ (i.e., $2 \times 0.25\text{M} = 0.50\text{M}$)

Finally, what should logically be the inventory of solutions of **any** ionic solid?
(The inventory should be as constituent ions!)