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Lithium chloride is ionic or covalent

Ionic bonds are atomic bonds created by the attraction of two differently charged ions. The bond is typically between a metal and a non-metal. The structure of the bond is rigid, strong and often crystalline and solid. Ionic bonds also melt at high temperatures. Dissolved in water, ionic bonds are aqueous, that is, they can conduct. As solids they are insulators. Ionic bonds can also be called electrovalent bonds. An ionic bond is formed when ions interact to create an ionic compound with the positive and negative charges in balance. Ionic bond examples include: LiF - Lithium FluorideLiCl - Lithium ChlorideLiBr - Lithium BromideLiI - Lithium IodideNaF - Sodium FluorideNaCl - Sodium ChlorideNaBr - Sodium BromideNaI - Sodium IodideKF - Potassium FluorideKCl - Potassium ChlorideKBr - Potassium BromideKI - Potassium IodideCsF - Cesium FluorideCsCl - Cesium ChlorideCsBr - Cesium BromideCsI - Cesium IodideBeO - Beryllium OxideBeS - Beryllium SulfideBeSe - Beryllium SelenideMgO - Magnesium OxideMgS - Magnesium SulfideMgSe - Magnesium SelenideCaO - Calcium OxideCaS - Calcium SulfideCaSe - Calcium SelenideBaO - Barium OxideBaS - Barium SulfideBaSe - Barium SelenideCuF - Copper(I) FluorideCuCl - Copper(I) ChlorideCuBr - Copper(I) BromideCuI - Copper(I) IodideCuO - Copper(II) OxideCuS - Copper(II) SulfideCuSe - Copper(II) SelenideFeO - Iron(II) OxideFeS - Iron(II) SulfideFeSe - Iron(II) SelenideCoO - Cobalt(II) OxideCoS - Cobalt(II) SulfideCoSe - Cobalt(II) SelenideNiO - Nickel(II) OxideNIS - Nickel(II) SulfideNiSe - Nickel(II) SelenidePbO - Lead(II) OxidePbS - Lead(II) SulfidePbSe - Lead(II) SelenideSnO - Tin(II) OxideSnS - Tin(II) SulfideSnSe - Tin(II) SelenideLi2O - Lithium OxideLi2S - Lithium SulfideLi2Se - Lithium SelenideNa2S - Sodium OxideNa2S - Sodium SulfideNa2Se - Sodium SelenideK2O - Potassium OxideK2S - Potassium SulfideK2Se - Potassium SelenideCs2O - Cesium OxideSc2S - Cesium SulfideCs2Se - Cesium SelenideBeF2 - Beryllium FluorideBeCl2 - Beryllium ChlorideBeBr2 - Beryllium BromideBeI2 - Beryllium IodideMgF2 - Magnesium FluorideMgCl2 - Magnesium ChlorideMgBr2 - Magnesium BromideMgI2 - Magnesium IodideCaF2 - Calcium FluorideCaCl2 - Calcium ChlorideCaBr2 - Calcium BromideCaI2 - Calcium IodideBaF2 - Barium FluorideBaCl2 - Barium ChlorideBaBr2 - Barium BromideBaI2 - Barium IodideFeF2 - Iron(II) FluorideFeCl2 - Iron(II) ChlorideFeBr2 - Iron(II) BromideFeI2 - Iron(II) IodideCoF2 - Cobalt(II) FluorideCoCl2 - Cobalt(II) ChlorideCoBr2 - Cobalt(II) BromideCoI2 - Cobalt(II) IodideNiF2 - Nickel(II) FluorideNiCl2 - Nickel(II) ChlorideNiBr2 - Nickel(II) BromideNiI2 - Nickel(II) IodideCuF2 - Copper(II) FluorideCuCl2 - Copper(II) ChlorideCuBr2 - Copper(II) BromideCuI2 - Copper(II) IodideSnF2 - Tin(II) FluorideSnCl2 - Tin(II) ChlorideSnBr2 - Tin(II) BromideSnI2 - Tin(II) IodidePbF2 - Lead(II) FluoridePbCl2 - Lead(II) ChloridePbBr2 - Lead(II) BromidePbI2 - Lead(II) IodideFeF3 - Iron(III) FluorideFeCl3 - Iron(III) ChlorideFeBr3 - Iron(III) BromideFeI3 - Iron(III) IodideCoF3 - Cobalt(III) FluorideCoCl3 - Cobalt(III) ChlorideCoBr3 - Cobalt(III) BromideCoI3 - Cobalt(III) IodideNiF3 - Nickel(III) FluorideNiCl3 - Nickel(III) ChlorideNiBr3 - Nickel(III) BromideNiI3 - Nickel(III) IodideSnF4 - Tin(IV) FluorideSnCl4 - Tin(IV) ChlorideSnBr4 - Tin(IV) BromideSnI4 - Tin(IV) IodidePbF4 -Lead(IV) FluoridePbCl4 - Lead(IV) ChloridePbBr4 - Lead(IV) BromidePBI4 - Lead(IV) IodideLi3N - Lithium NitrideLi3P - Lithium PhosphideNa3N - Sodium NitrideNa3P - Sodium PhosphideK3N - Potassium NitrideK3P - Potassoum PhosphideCs3N - Cesium NitrideCs3P - Cesium PhosphideHC2H3O2 - Hydrogen AcetateLiC2H302 - Lithium AcetateLiHCO3 - Lithium Hydrogen CarbonateLiOH - Lithium HydroxideLiNO3 - Lithium NitrateLiMnO4 - Lithium PermanganateLiClO3 - Lithium ChlorateNaC2H3O2 - Sodium AcetateNaHCO3 - Sodium Hydrogen CarbonateNaOH - Sodium HydroxideNaNO3 - Sodium NitrateNaMnO4 - Sodium PermanganateNaClO3 - Sodium ChlorateKC2H3O2 - Potassium AcetateKHCO3 - Potassium Hydrogen CarbonateKOH - Potassium HydroxideKNO3 - Potassium NitrateKMnO4 - Potassium PermanganateKClO3 - Potassium ChlorateCsC2H3O2 - Cesium AcetateCsHCO3 - Cesium Hydrogen CarbonateCsOH - Cesium HydroxideCsNO3 - Cesium NitrateCsMnO4 - Cesium PermanganateCsClO3 - Cesium ChlorateAgF - Silver FluorideAgCl - Silver ChlorideAgBr - Silver BromideAgI - Silver IodideAgC2H3O2 - Silver AcetateAgHCO3 - Silver Hydrogen CarbonateAgOH - Silver HydroxideAgNO3 - Silver NitrateAgMnO4 - Silver PermanganateAgClO3 - Silver ChlorateAg2O - Silver OxideAg2S - Silver SulfideAg2Se - Silver SelenideNH4C2H3O2 - Ammonium AcetateNH4HCO3 - Ammonium Hydrogen CarbonateNH4OH - Ammonium HydroxideNH4MnO4 - Ammonium PermanganateNH4ClO3 - Ammonium Chlorate(NH4)2O - Ammonium Oxide(NH4)2S - Ammonium Sulfide(NH4)2Se - Ammonium SelenideBe(C2H3O2)2 - Beryllium AcetateBe(HCO3)2 - Beryllium Hydrogen CarbonateBe(OH)2 - Beryllium HydroxideBe(MnO4)2 - Beryllium PermanganateBeCO3 - Beryllium CarbonateBeCrO4 - Beryllium ChromateBeCr2O7 - Beryllium DichromateBeSO4 - Beryllium SulfateBe3(PO4)2 - Beryllium PhosphateBe(ClO3)2 - Beryllium ChlorateMg(C2H3O2)2 = Magnesium AcetateMg(HCO3)2 - Magnesium Hydrogen CarbonateMg(OH)2 - Magnesium HydroxideMg(MnO4)2 - Magnesium PermanganateMg(ClO3)2 - Magnesium ChlorateMgCO3 - Magnesium CarbonateMgCrO4 - Magnesium ChromateMgCr2O7 - Magnesium DichromateMgSO4 - Magnesium SulfateMg3(PO4)2 - Magnesium PhosphateCa(C2H3O2)2 - Calcium AcetateCa(HCO3)2 - Calcium Hydrogen CarbonateCa(OH)2 - Calcium HydroxideCa(MnO4)2 - Calcium PermanganateCa(ClO3)2 - Calcium ChlorateCaCO3 - Calcium CarbonateCaCrO4 - Calcium ChromateCaCr2O7 - Calcium DichromateCaSO4 - Calcium SulfateCa3(PO4)2 - Calcium PhosphateBa(C2H3O2)2 - Barium AcetateBa(HCO3)2 - Barium Hydrogen CarbonateBa(OH)2 - Barium HydroxideBa(MnO4)2 - Barium PermanganateBa(ClO3)2 - Barium ChlorateBaCO3 - Barium CarbonateBaCrO4 - Barium ChromateBaCr2O7 - Barium DichromateBaSO4 - Barium SulfateBa3(PO4)2 - Barium PhosphateZnF2 - Zinc FluorideZnCl2 - Zinc ChlorideZnBr2 - Zinc BromideZnI2 - Zinc IodideZnO - Zinc OxideZnS - Zinc SulfideZnSe - Zinc SelenideZn(C2H3O2)2 - Zinc AcetateZn(HCO3)2 - Zinc Hydrogen CarbonateZn(OH)2 - Zinc HydroxideZn(MnO4)2 - Zinc PermanganateZn(ClO3)2 - Zinc ChlorateZnCO3 - Zinc CarbonateZnCrO4 - Zinc ChromateZnCr2O7 - Zinc DichromateZnSO4 - Zinc SulfateZn3(PO4)2 - Zinc PhosphateFe(C2H3O2)2 - Iron(II) AcetateFe(HCO3)2 - Iron(II) Hydrogen CarbonateFe(OH)2 - Iron(II) HydroxideFe(MnO4)2 - Iron(II) PermanganateFe(ClO3)2 - Iron(II) ChlorateFeCO3 - Iron(II) CarbonateFeCrO4 - Iron(II) ChromateFeCr2O7 - Iron(II) DichromateFeSO4 - Iron(II) SulfateFe3(PO4)2 - Iron(II) PhosphateFe(NO3)2 - Iron(II) NitrateZn(NO3)2 - Zinc NitrateBa(NO3)2 - Barium NitrateCa(NO3)2 - Calcium NitrateMg(NO3)2 - Magnesium NitrateBe(NO3)2 - Beryllium NitrateAgNO3 - Silver Nitrate 7th grade 8th grade 9th grade 10th grade middle school high school college Ionic or covalent? An ionic bond is an electrostatic force between positive and negative ions. This electrostatic force in non-directional, it acts in all directions. The strength of the force depends on the magnitude of the charges on the ions and the sum of the ionic radii. But what determines whether a compound is ionic or covalent? If the bonding atoms have a large difference in electronegativity then this causes transfer of electrons and the formation of ionic compounds. As the difference in electronegativity decreases, the bond develops covalent character until eventually it becomes essentially covalent. Notice that this process is not black and white. The bond type changes gradually from pure ionic to pure covalent, passing through all degrees. Pure ionic >>> ionic with covalent character >>> polarised covalent >>> pure covalent Pure ionic compounds are formed by group 1 metals when combining with non-metals. These are highly electropositive, having electronegativity values of between 0.7 and 1.0. However, when the electronegativity value of the metal is higher, the bonding has a degree of covalent character. This means that the negative ion electron density is distorted and drawn towards the metal ion. There is some electron density between the two particles, typical of covalent bonding. This can be seen on electron density maps produced in X-ray crystallography. Instead of the electron density being symmetrical around the ions, it is distorted towards the positive ions. The classic example is aluminium chloride. Aluminium is a metal from group 13 and consequently forms Al3+ ions. However, it is not very electropositive and the high charge density of the small Al3+ ion allows it to polarise the negative charge cloud on negative ions formed from atoms of lesser electronegativity. Aluminium oxide is an ionic compound, but aluminium chloride is only ionic in the solid state at low temperatures. At higher temperatures it becomes covalent. This is because the high charge density Al3+ ion can polarise the Cl- charge cloud, making an ionic bond with a high degree of covalent character, so much so that AlCl3 is usually considered to be covalent. The difference in electronegativity between aluminium (1.5) and chlorine (3.0) is 1.5 units. This could be taken as a rough guide for the limit between ionic and covalent bonding. When metals bond to non-metals if the difference in electronegativity is greater than 1.5, then the compound would be expected to be ionic, less that 1.5 and covalency is expected. It should be stressed that this is only an approximation and it is easy to find exceptions. top Consequences of covalent character Covalent substances are molecular and usually have low melting and boiling points. They tend to dissolve in non-polar solvents, forming solutions that do not conduct electricity. Many covalent chorides are hydrolysed by water, forming hydrogen chloride. All of these facts can be used as pointers towards covalent character in ionic compounds. Aluminium chloride is a covalent substance whose characteristics reflect this nature. compound sodium chloride aluminium chloride silicon chloride bonding ionic covalent covalent melting point /ºC 801 190 (sublimes) -69 reaction with water dissolves and dissociates into ions hydrolysed HCl hydrolysed HCl Magnesium chloride is an ionic substance. However, the small size and double charge of the magnesium ion causes a degree of covalent character. For example, evaporation of a solution of magnesium chloride produces 'basic' magnesium chloride - Mg(OH)Cl - showing that the magnesium chloride has been partially hydrolysed by water. Even lithium chloride has a small degree of covalent character, being soluble in non-polar solvents. Example: Magnesium chloride and silicon(IV) chloride have very different properties. Give the formula and physical state at room temperature of each compound. State the conditions under which each compound conducts electricity (if at all). Each chloride is added to water in separate experiments. Suggest an approximate pH value for the solution formed and give an equation for the reaction. 1. The valency comes from the group number, hence MgCl2 and SiCl4. 2. Magnesium chloride is ionic. It conducts electricity when molten or in aqueous solution. Silicon chloride does not conduct electricity. 3. Magnesium chloride is ionic, it dissolves in water and the ions dissociate. The pH is about 5 as magnesium chloride is weakly acidic by hydrolysis (it is the salt of a weak base and a strong acid) MgCl2 + xH2O Mg2+(aq) + 2Cl-(aq) Silicon(IV) chloride is hydrolysed by water: The solution formed has a pH of 1 as the HCl formed is a strong acid. SiCl4 + 4H2O Si(OH)4 + 4HCl top Page 2 Page 3 Syllabus ref: 4.2 Syllabus ref: 14.1 "Co" means sharing, "valent" refers to the electrons in the outer, or valence, shell. Hence, the term covalent bonding means to share electron pairs between two outer shells of atoms in order to bond the atoms together, making a more complex particle. There are two ways that this can occur, either each atom provides one electron for the pair, or both of the electrons are provided (donated) by one of the atoms. In the following chapter we examine the process and the consequences of both variations of covalent bonding. Nature of science: Looking for trends and discrepancies. Compounds containing non-metals have different properties to compounds that contain non-metals and metals. Use theories to explain natural phenomena. Lewis introduced a class of compounds which share electrons. Pauling used the idea of electronegativity to explain unequal sharing of electrons. Principle of Occam's razor-bonding theories have been modified over time. Newer theories need to remain as simple as possible while maximizing explanatory power, for example the idea of formal charge. Understandings - SL A covalent bond is formed by the electrostatic attraction between a shared pair of electrons and the positively charged nuclei. Single, double and triple covalent bonds involve one, two and three shared pairs of electrons respectively. Bond length decreases and bond strength increases as the number of shared electrons increases. Understandings - HL Essential idea: Larger structures and more in-depth explanations of bonding systems often require more sophisticated concepts and theories of bonding. Covalent bonds result from the overlap of atomic orbitals. A sigma bond (s) is formed by the direct head-on/end-to-end overlap of atomic orbitals, resulting in electron density concentrated between the nuclei of the bonding atoms. A pibond (p) is formed by the sideways overlap of atomic orbitals, resulting in electron density above and below the plane of the nuclei of the bonding atoms. Formal charge (FC) can be used to decide which Lewis (electron dot) structure is preferred from several. The FC is the charge an atom would have if all atoms in the molecule had the same electronegativity. FC = (Number of valence electrons)-½(Number of bonding electrons)-(Number of non-bonding electrons). The Lewis (electron dot) structure with the atoms having FC values closest to zero is preferred. Exceptions to the octet rule include some species having incomplete octets and expanded octets. Delocalization involves electrons that are shared by/between all atoms in a molecule or ion as opposed to being localized between a pair of atoms. Resonance involves using two or more Lewis (electron dot) structures to represent a particular molecule or ion. A resonance structure is one of two or more alternative Lewis (electron dot) structures for a molecule or ion that cannot be described fully with one Lewis (electron dot) structure alone. Applications and skills - SL Applications and skills - HL Prediction whether sigma (s) or pi (p) bonds are formed from the linear combination of atomic orbitals. Deduction of the Lewis (electron dot) structures of molecules and ions showing all valence electrons for up to six electron pairs on each atom. Application of FC to ascertain which Lewis (electron dot) structure is preferred from different Lewis (electron dot) structures. Deduction using VSEPR theory of the electron domain geometry and molecular geometry with five and six electron domains and associated bond angles. Explanation of the wavelength of light required to dissociate oxygen and ozone. Description of the mechanism of the catalysis of ozone depletion when catalysed by CFCs and NOx. Page 4 The covalent bond Atoms are held together in covalent bonding by means of shared pairs of electrons. This constitutes a single covalent bond. The various theories as to how and why these bonds are formed are discussed below. It is important to recognise that formation of covalent bonds (or any other type of bond) is an exothermic process, one that releases energy. Similarly, to break a bond always requires energy. Any cleavage (breakage) of covalent bonds involves a chemical reaction and new substances are necessarily formed. top Bonding theory Atoms are unstable unless they have fully occupied outer shells of electrons. We know from observation and inference that atoms bond together to make larger structures. In order to explain how this happens, different theories are proposed that explain these observations. All of these theories revolve around the accepted idea that opposite charges attract (electrostatic attraction). The scientific method A theory helps to explain all of the observable facts. If it fits all of the evidence, then it is adopted as a useful explanation of how the universe works. However, if further observations throw up anomalies, or observations that cannot be explained by the current theory, then it must be either adapted, or even discarded completely. top Linear combination of atomic orbitals This theory of bonding proposes that molecules are formed by overlapping regions of space, allowing atoms to mutually share pairs of electrons. This pair of electrons is then held in a position between the two nuclei of the atoms holding them together. We can consider this theory by looking at the hydrogen molecule. top The hydrogen molecule The hydrogen molecule is the simplest structure formed between atoms. In this case two hydrogen atoms share one pair of electrons between them, forming a diatomic molecule. The negative charge clouds (atomic orbitals) overlap, placing a region of negative charge between the two hydrogen nuclei. The electrons in each atom is also attracted to the nucleus of the neighbouring atom. The atoms are pulled together. This view of chemical bonding is known as covalent bonding as the valence electrons of both the hydrogen atoms are shared. ('co' = shared and 'valent' = valence electrons) The regions of space in which electrons are found are called orbitals and we say that this bond arises from direct overlap of atomic orbitals. This is sometimes called the Linear Combination of Atomic Orbital Theory (LCAO). This model works very well for a simple picture of bonding between atoms. The nuclei of the atoms are held to one another by the positive - negative - positive attractions caused by electron pairs between the nuclei. top Full outer energy shells The only atoms that occur in nature not combined into some kind of structure are the inert gases. These all have full outer (valence) shells. It seems to be that this particular electronic configuration is stable for some reason. Atom configuration helium 2 neon 2,8 argon 2,8,8 krypton 2,8,18,8 xenon 2,8,18,18,8 Covalent bonding occurs between atoms so that they can attain a full outer shell by SHARING electrons. In the case of hydrogen molecules, above, the atom would have a full outer shell if it were to attain two electrons. By an arrangement in which two atoms share one pair of electrons, each atom achieves the requirement. The system is now stable, it does not change further. Covalent bonding usually occurs between non-metal atoms; they attain a full outer shell of electrons by sharing electrons. However, as we shall see in the following section there are exceptions. top Summary of covalent bonding Covalent bonding happens when non-metal atoms combine Electron pairs are shared to give a full outer shell of valence electrons to each atom The final units are molecules of atoms joined together by shared pairs of electrons The covalent bond (shared pair) may be represented by a line to show typical stick structures such as: ethyne propanal propyne A single line is drawn to represent one pair of electrons. A double bond containing two pairs of electrons between atoms is shown as a double line, and three pairs, a triple bond, as a three lines. It is not possible for two atoms to share more than three pairs of electrons. top