Lesson #2 Electronegativity & Ionic Bonds

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Trends

SPECIAL NOTE: The following ***must*** be remembered when examining trends down or across the periodic table.

When going **DOWN** a family in the periodic table, properties are affected by the **INCREASING SIZE** of the atoms and the **INCREASING DISTANCE** between the nuclei and the valence electrons.

When going **ACROSS** a period in the periodic table, properties are affected by the **DIFFERING VALENCE, NUCLEAR CHARGE** and **CHARGE ON THE SPECIES.**

B. ELECTRONEGATIVITY (χ)

Electronegativity (χ) is defined as the ability of an atom in a molecule to "steal" electrons from another atom (attract e^- to itself).

Electronegativity is a function of two properties of isolated atoms:

i. The atom's ionization energy (how strongly an atom holds onto its own electrons)

ii. The atom's electron affinity (how strongly the atom attracts other electrons)

For example, an element which has a large (negative) electron affinity and a high ionization *(always endothermic, or positive for neutral atoms)*...

...Will attract e^- from other atoms and resist having electrons taken away. eg. Halogens

We say these atoms are *"highly electronegative"* → atoms that form $\ominus$ ions.

- atoms with ↑ χ are very attractive to the valence electrons of neighbouring atoms
- atoms with ↑ χ strongly attract their own valence e^- (almost never lose e^-)
- values of χ range from a low of 0.7 (Fr) to a high of 3.9 (F):

- The difference in electronegativity between two atoms, $\Delta\chi$, tells you what type of bond will form

Electronegativity increases across a period because:

H 2.2							He –
Li 1.0	Be 1.6	B 2.0	C 2.6	N 3.0	O 3.4	F 4.0	Ne –
Na 0.9	Mg 1.3	Al 1.6	Si 1.9	P 2.2	S 2.6	Cl 3.2	Ar –
K 0.8						Br 3.0	Kr –

1. The atomic radius decreasing (↓)
2. The charge on the nucleus increasing without significant extra shielding. New electrons do not contribute much to shielding because they are added to the same principal energy level across the period.

Electronegativity decrease down a group because:

H 2.2							He –
Li 1.0	Be 1.6	B 2.0	C 2.6	N 3.0	O 3.4	F 4.0	Ne –
Na 0.9	Mg 1.3	Al 1.6	Si 1.9	P 2.2	S 2.6	Cl 3.2	Ar –
K 0.8						Br 3.0	Kr –

1. The atomic radius increasing
2. Although the charge on the nucleus increases, shielding also **increases** significantly. This is because electrons added down the group fill new principal energy levels.

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in Data booklet.

N/A

ΔEN = change in Electroneg.

Electronegativities

1 H 2.1																	2 He
3 Li 1.0	4 Be 1.5											5 B 2.0	6 C 2.5	7 N 3.0	8 O 3.5	9 F 4.0	10 Ne
11 Na 0.9	12 Mg 1.2											13 Al 1.5	14 Si 1.8	15 P 2.1	16 S 2.5	17 Cl 3.0	18 Ar
19 K 0.8	20 Ca 1.0	21 Sc 1.3	22 Ti 1.5	23 V 1.6	24 Cr 1.6	25 Mn 1.5	26 Fe 1.8	27 Co 1.8	28 Ni 1.8	29 Cu 1.9	30 Zn 1.6	31 Ga 1.6	32 Ge 1.8	33 As 2.0	34 Se 2.4	35 Br 2.8	36 Kr
37 Rb 0.8	38 Sr 1.0	39 Y 1.2	40 Zr 1.4	41 Nb 1.6	42 Mo 1.8	43 Tc 1.9	44 Ru 2.2	45 Rh 2.2	46 Pd 2.2	47 Ag 1.9	48 Cd 1.7	49 In 1.7	50 Sn 1.8	51 Sb 1.9	52 Te 2.1	53 I 2.5	54 Xe
55 Cs 0.7	56 Ba 0.9																

Table 6.3.3 *Relationship of ΔEN and Bond Designation*

ΔEN	Bond Designation
0	non-polar covalent
< 0.4	mostly covalent
0.4 – 1.7	polar covalent
> 1.7	ionic

covalent

metals + non-metals

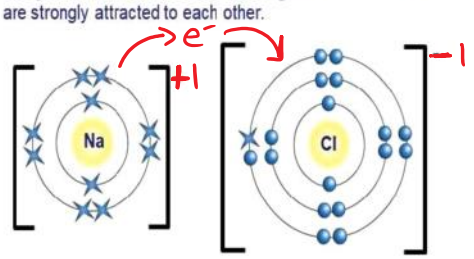
C. TYPES OF CHEMICAL BONDS

Force

An **IONIC BOND** is formed by the electrostatic attraction between a cation (+) and an anion (−).

- results when $\Delta\chi \geq 1.7$ (difference in Elec.)
- *usually* involves a metal (low χ) + non-metal (high χ)
- relatively strong bonds (strength depends on ionic radii and charges of the ions).

The **positive** sodium ions and the **negative** chloride ions are strongly attracted to each other.



It is this electrostatic attraction that forms ionic bonds in sodium chloride and other ionic compounds.

e.g. NaCl is an ionic bond because...

χ_{Na}: 0.9

χ_{Cl}: 3.0

subtract BIG - small

$\Delta\chi$: 3.0 − 0.9 = 2.1

$\Delta\chi > 1.7$ ∴ ionic bond.

Electronegativity & Ionic Bonds:

- Formed between two atoms with large differences in their ionization energies and electronegativities
- An ***electronegativity difference of greater than*** >1.7 can be classified as an **ionic bond**
- In this case we can essentially say electrons are transferred from one atom to another.

(+) e⁻ (−)

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Practice Problems — Ionic Compounds $\Delta\chi > 1.7$

Write formulas for the ionic compounds formed when the following elements combine. Using the ΔEN values, arrange the compounds in order of the increasing ionic character of the bonds in each compound.

(a) Na and N	(b) Sr and Br	(c) Li and Cl	(d) Cs and F	(e) Rb and O
0.9 3.0	1.0 2.8	1.0 3.0	0.7 4.0	0.8 3.5
$\Delta\chi = 2.1$ ionic ✓	$\Delta\chi = 1.8$ ionic ✓	$\Delta\chi = 2.0$ ionic ✓	$\Delta\chi = 3.3$ ionic ✓	$\Delta\chi = 2.7$ ionic ✓

How do ionic bonds form?

a) Pre-existing Ions Get Together *e.g.* synthesis, single replacement, double replacement (precipitation)

b) Atoms React To Form Ions Which Then Combine (Electron transfer)

- The **metal atom** loses one (or more) electrons to establish a closed outer shell ⊕ cation
- The **non-metal atom** gains one (or more) electrons to establish a closed outer shell ⊖ anion
- **both** ions are then isoelectronic ← same # of e^- to their nearest noble gas (*The Octet Rule*) 8 e^-
- the ions are held together by the attraction of their opposite charges.

⊕⊖ electrostatic forces.

Lewis Structures show the electrons in the valence shell of an atom, ion, or molecule.

example 1: Li + F → $[Li]^+[F]^-$ ionic bond.

isoelectronic He–Ne

atomic radii: 152pm 64pm

68pm 136pm

example 2: K + Br → $[K]^+[Br]^-$

isoelectronic: Ar Kr

227 pm 114pm

133pm 196pm

For Ionic Bonds: Pretend that each ion's charge is at its centre. Compare distances between the centres.

Which ionic bond is stronger? LiF or KBr

Why? smaller distance between ions => stronger / larger force of attraction

Predicting Ion Charges

Group (Family)	1	2	13	14	15	16	17	18
Charge on the Ion	+1	+2	+3	///	−3	−2	−1	0

When atoms form ions, they lose or gain electrons to get a closed shell.

Groups 14 and 15 are on only partially predictable. (covalent bonds)

- C, Si, and Ge usually do not form ions while Sn and Pb (metals) form ions => generally +2 or +4
- Bi forms +3 or +5 cations.

non-metals (metalloid)

Properties of Ionic Compounds

What is an ionic lattice?

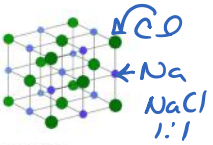
In an **ionic compound** when any macroscopic sample ions react together, countless atoms will transfer electrons to form countless **oppositely charged ions**. (+) (−) attraction

These oppositely charged species being produced in close proximity are drawn together into an giant, 3D, ordered, solid, three-dimensional array of cations and anions called a crystalline lattice

The small whole number *cation-anion ratio* in this structure represents the chemical formula for the ionic compound.

The structure of the ionic lattice affects the ***properties*** of the ionic compound.

All ionic compounds form lattices and crystals when solid (s).



NaCl 1:1

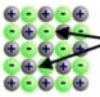
Heating ionic compounds

Why are ionic compounds solid at room temperature and have high melting points and boiling points?

Compound	Ion charges	Melting point (°C)	Boiling point (°C)
sodium chloride	1^+ and 1^-	801	1,413
magnesium oxide	2^+ and 2^-	2,852	3,600

NaCl MgO

9 magnitude of charge

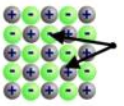


- Ionic bonds are strong and a lot of heat (E) is needed to break them.
- ***Larger ionic charges*** produce stronger ionic bonds and so much more heat is required to break the ionic bonds in magnesium oxide than in sodium chloride.

+/-2 stronger > +/- 1 weaker

How can ionic compounds conduct electricity?

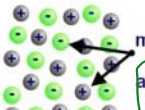
- As **solids**, ionic compounds CANNOT conduct electricity because their ions are bonded together in the lattice.



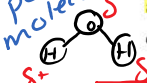
- When *liquid (molten)*, the ions can break free of the lattice and are able to move. The ions are charged particles and so can carry an *electric current.*

(aq) solution

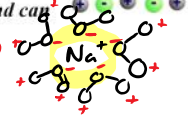
- Ionic compounds are usually soluble (dissolve) **in water** because water molecules have a *slight electrical charge* and so can attract the ions away from the lattice. ***When dissolved, the ions are free to move and can carry an electric current.***

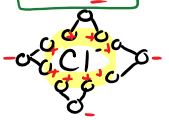


polar molecule

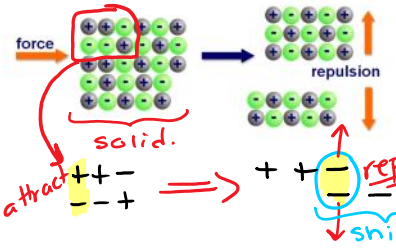


hydration shell =>





Why are ionic compounds brittle? - they shatter when they are hit.



- When the lattice is hit, a layer of ions is shifted so that ***ions with the same charges*** are lined up together.
- These *like charges* repel each other and so split the ionic lattice causing it to break / shatter

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Summary Comments ions individually carry a charge

1. Ionic compounds are electrically neutral. For this reason, the total amount of ⊕ must equal total ⊖ charge

2. All ionic compounds are solid + brittle at room temperature. How does the model above help to explain this fact?
↳ crystalline lattice

3. The higher the melting point of an ionic compound, the stronger the bond.

Homework **Assignment #3 Hebden pg. 172 #57ace, pg. 173 #58-61 pg 175 #63-64adf pg. 176 #65-67**

all assignments are to be completed on a separate page with the assignment number & heading

EXERCISE:

57. Which of the following atom pairs would you expect to form ionic bonds when they join?
(a) Ba and S (b) P and Cl (c) Ca and O (d) Rb and I (e) O and H (f) S and O

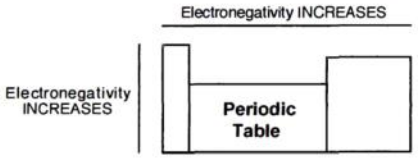
58. This exercise compares Li and F.
(a) Which atom is larger: Li or F? (Hint: see exercise 41.)
(b) Which atom has the stronger attraction to the outer electrons on a neighbouring atom, based only on the atomic radius?
(c) Which atom has the greater nuclear charge?
(d) Which atom can attract electrons from an adjacent atom most strongly, based on both size and nuclear charge?
(e) Summarize the above by filling in the blank below.

IN GENERAL, when going from ***left*** to ***right*** across the periodic table the electronegativity of the atoms will _____.

59. This exercise compares F and I.
(a) Which atom is larger: F or I?
(b) Which atom has a stronger attraction to the outer electrons of another atom?
(c) Summarize the above by filling in the blank below.

IN GENERAL, when going down a family of the periodic table the electronegativity of the atoms will _____.

60. Place arrowheads in the correct direction on the horizontal and vertical arrows below.



61. (a) Ignoring the noble gases, which atom is the most electronegative?
(b) Ignoring the noble gases, which atom is the least electronegative?
(c) Which is more electronegative: K or Be?
(d) Which is more electronegative: Pb or S?

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HW.

63. Mg^{2+} and Na^{+} have roughly the same ionic radius. O^{2-} and F^{-} have roughly the same ionic radius. Which substance should have a higher melting temperature: NaF or MgO? Why?

64. Which member of each following pair would you expect to have the higher melting point?
(a) CaO or RbI (c) LiF or NaCl (e) RbI or KCl
(b) BeO or BN (d) CsCl or BaS (f) BeO or MgS

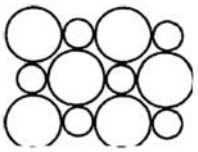
65. ***NEGATIVE IONS:*** Assume extra electrons are added to a neutral atom of O to make O^{2-}. The resulting ion has the same positive nuclear charge and an increased number of negative electrons surrounding the nucleus.
(a) What happens to the amount of electrostatic repulsion existing between the electrons?
(b) What happens to the volume occupied by the electrons due to the change in the amount of electron–electron repulsion?
(c) Fill in the appropriate word.

NEGATIVE IONS are _____ than the corresponding neutral atom.

66. ***POSITIVE IONS:*** Assume electrons are removed from a neutral atom of Mg to make Mg^{2+}. The resulting ion has the same positive nuclear charge and a decreased number of negative electrons surrounding the nucleus.
(a) What happens to the amount of electrostatic repulsion existing between the electrons?
(b) What happens to the volume occupied by the electrons due to the change in repulsion?
(c) Fill in the appropriate word.

POSITIVE IONS are _____ than the corresponding neutral atom.

67. Examine the diagram below, which shows a section of a crystal of NaCl.



Which circles represent Na^{+}: the larger or smaller ones?

HW (2)