Electronegativity

The ability to bind bonding electron pairs during the formation of a chemical bond is different for different atoms of elements. A measure of this ability is the dimensionless **electronegativity** of the elements **χ** . Simply put, the higher the electronegativity of an element, the more it attracts bonding electrons to itself.

The opposite of electronegativity is electropositivity , which expresses the ability to donate an electron to another atom

Electronegativity of elements in the periodic table increases from left to right in periods and decreases from top to bottom in groups. The most electronegative are the halogens (led by fluorine), on the contrary, the least electronegative are the alkali metals.

The most widely used definition of the electronegativity of elements comes from Linus Pauling . It is based on the observation that the covalent bond in the two-element compound A–B is stronger than we would expect based on the properties of the bonds in compounds A–A and B–B. Pauling attributed this "additional stabilization" of the bond to the electrical force that acts between the partially ionized atoms involved in the bond.

As we mentioned, it can be experimentally verified that the dissociation energy of the covalent bond in the diatomic compound A–B is greater than the average of the dissociation energies of the bonds in the compounds A–A and B–B:

$$E_d(\mathrm{AB}) > \frac{E_d(\mathrm{AA}) + E_d(\mathrm{BB})}{2}$$

The magnitude of this difference corresponds to Pauling's definition of electronegativity difference:

$$\Delta\chi = \sqrt{E_d(\mathrm{AB}) - \frac{E_d(\mathrm{AA}) + E_d(\mathrm{BB})}{2}}$$

See also: Electronegativity#Methods of calculation (https://en.wikipedia.org/wiki/Electronegativity#Methods_of_calculation) Based on empirical measurements, only the electronegativity difference of the atoms in the compound can be calculated. In order to be able to tabulate electronegativity for each element, a reference point is arbitrarily set in the so-called Pauling scale. Originally it was the electronegativity of fluorine (established as 4.0 by Pauling), today it is the electronegativity of hydrogen (2.20 in the latest revision).

Strictly speaking, the electronegativity of an atom depends on the bonding context - in different compounds the electronegativity of the same element varies slightly.

Dependance of electronegativity on the position of the element in the periodic table

→ Atomic radius decreases → Ionization energy increases → Electronegativity increases →

Group / Period	1	2	3	4	5	6	7	8	9	10	11	12	13	14	15	16	17	18
1	H 2.20																	He 3.89
2	Li 0.98	Be 1.57											B 2.04	C 2.55	N 3.04	O 3.44	F 3.98	Ne 3.67
3	Na 0.93	Mg 1.31											Al 1.61	Si 1.90	P 2.19	S 2.58	Cl 3.16	Ar 3.3
4	K 0.82	Ca 1.00	Sc 1.36	Ti 1.54	V 1.63	Cr 1.66	Mn 1.55	Fe 1.83	Co 1.88	Ni 1.91	Cu 1.90	Zn 1.65	Ga 1.81	Ge 2.01	As 2.18	Se 2.55	Br 2.96	Kr 3.00
5	Rb 0.82	Sr 0.99	Y 1.22	Zr 1.33	Nb 1.6	Mo 2.16	Tc 1.9	Ru 2.2	Rh 2.28	Pd 2.20	Ag 1.93	Cd 1.69	In 1.78	Sn 1.96	Sb 2.05	Te 2.1	I 2.66	Xe 2.6
6	Cs 0.79	Ba 0.89	*	Hf 1.3	Ta 1.5	W 2.36	Re 1.9	Os 2.2	Ir 2.20	Pt 2.28	Au 2.54	Hg 2.00	Tl 1.62	Pb 2.33	Bi 2.02	Po 2.0	At 2.2	Rn 2.2
7	Fr 0.7	Ra 0.9	**	Rf	Db	Sg	Bh	Hs	Mt	Ds	Rg	Cn	Nh	Fl	Mc	Lv	Ts	Og
Lanthanoids	*	La 1.1	Ce 1.12	Pr 1.13	Nd 1.14	Pm 1.13	Sm 1.17	Eu 1.2	Gd 1.2	Tb 1.1	Dy 1.22	Ho 1.23	Er 1.24	Tm 1.25	Yb 1.1	Lu 1.27		
Actinoids	**	Ac 1.1	Th 1.3	Pa 1.5	U 1.38	Np 1.36	Pu 1.28	Am 1.13	Cm 1.28	Bk 1.3	Cf 1.3	Es 1.3	Fm 1.3	Md 1.3	No 1.3	Lr 1.291		

Bond polarity

Most often, the electronegativity of elements is used to estimate the properties of the bonds in the compounds they form. The greater the difference in electronegativities of atoms connected by a bond, the more polar this bond is.

Type of bond according to electronegativity difference

Binding type	Elektronegativity difference
Nonpolar covalent	< 0,4
Polar covalent	0,4 – 1,7
Ionic	> 1,7

Bonds between the same elements, e.g. C–C bonds between carbons in organic compounds, are non-polar – both atoms have the same electronegativity, so the difference is zero.

C–H bonds are also non-polar (electronegativity difference just below 0.4). In contrast, C–O, C–N or N–H bonds are polar (electronegativity differences 1.0; 0.6; 1.4).

Sodium chloride NaCl is an ionic compound (electronegativity difference 1.9).

Links

References

- ŠRÁMEK, Vratislav. *General and inorganic chemistry.* 2nd edition. Olomouc. 2005.
- GÄRTNER, Harald, et al. *Compendium of Chemistry.* 1st edition. Universe, 2007

Recommended literature

- Wikipedia article on electronegativity