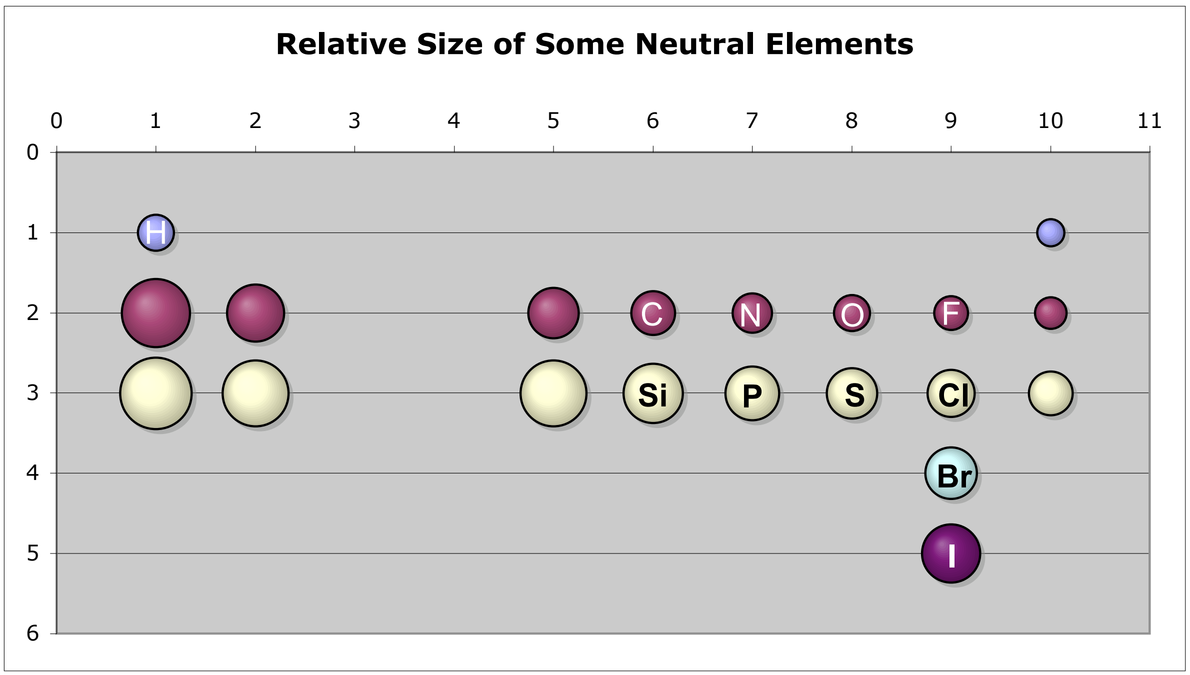
van der Waals Radii

Atomic numbe r	Element	Van der Waals radii (Å)
1	H	1.20
6	C	1.70
7	N	1.55
8	O	1.52
9	F	1.47
14	Si	2.10
15	P	1.80
16	S	1.80
17	Cl	1.75
35	Br	1.85
53	I	1.98



http://periodictable.com/Properties/A/VanDerWaalsRadius.v.html

The relative sizes of the neutral atoms are shown as a function of their position within the periodic table. Within a given row, the atoms get smaller as the atomic number increases. Within a given column, the atoms get larger as the atomic number increases.

Stability Trends for Negative Charge

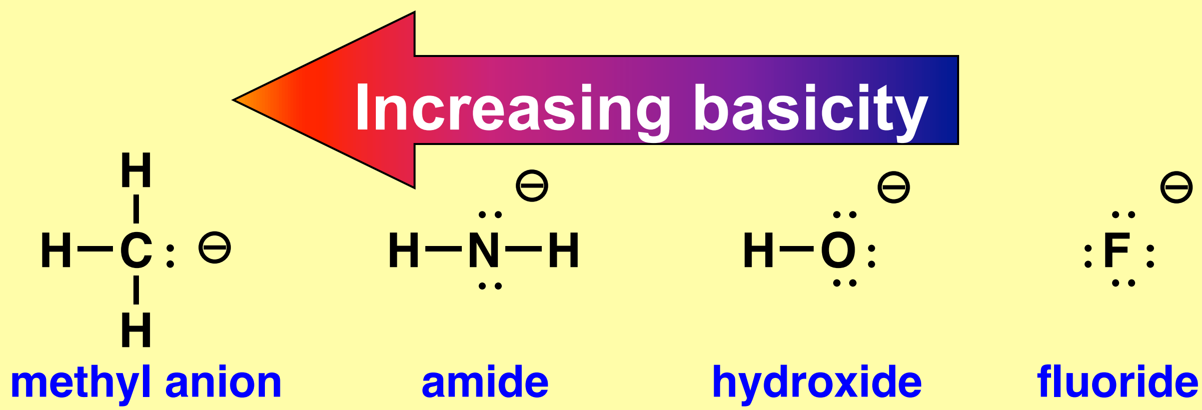
There is considerable energy associated with charged species; consequently, the relative stability of charged atoms often provides important information about chemical reactivity. Size and electronegativity are important factors that influences the chemical potential of ions. **These trends correlate to an atom's position in the periodic table.**

Ability to Stabilize Negative Charge

⇒ For atoms in the *same row*, stability is de termined by *electrognegativity* (the larger the electronegativity, the greater the stability).

⇒ For atoms in the *same column*, stability is determined by *size* (the larger the size, the greater the stability).

What's the rationale behind these rules? Within a row, the stability trends with electronegativity as expected. However, within a column, stability trend runs counter to electronegativity because size dominates (within a row atomic sizes are similar but between rows the atomic size varies significantly). Here's why size matters. The valence electron cloud carries the net negative charge. The more this charge cloud is spread out (i.e., the lower the charge density), the greater the stability. Thus, the anion with the largest size (iodide) is most effective at accomplishing this task.



Of the anions listed above, we are all familiar with hydroxide ion, so we will use this as our reference. Although me may think of hydroxide as a fairly strong base, it turns out to be far less basic than the nitrogen or carbon atoms that bear a formal charge of -1. Why? The rules state that within a row of the periodic table, the stability of negative charge tracks with electronegativity. Thus, the negative charge on the fluorine atom is the most stable. In contrast, charge is least stable on carbon. The methyl anion therefore most strongly seeks a proton in order to annihilate its unfavorable negative charge. In other words, methyl anion is the strongest base and fluoride is the weakest.

Stability Trends for Positive Charge

Ability to Stabilize Positive Charge

⇒ For atoms in the *same row*, stability is determined by *electrognegativity* (the <u>smaller</u> the electronegativity, the greater the stability).

⇒ For atoms in the *same column*, stability is determined by *size* (the <u>smaller</u> the size, the greater the stability).

The electronegativity trends and size factors also rationalize the trends for cations. The example shown here uses the rules to explain the relative acidity of cations: i.e., why hydronium is a stronger acid than the ammonium. The rules suggest that nitrogen stabilizes a positive charge better than oxygen. Hence, the positively charged oxygen of the hydronium ion has a higher chemical potential and is more strongly driven to give up a proton than the positively charged nitrogen of ammonium. Therefore, hydronium is the stronger acid.