

This single-page presentation contains information, either explicitly or implicitly, for discussion of a wide range of chemical reactions.

Periodic trends in bond energies are all readily seen in the figure.

The Most Important Main Group Trends

Groups one and 14 show the steady decrease in single self-bond energies with increasing atomic size and overlap capabilities. Groups 15 to 17 show the same trend in periods three through six; however, the unusually weak single self-bonds of the period two elements in these groups, attributed to interatomic lone-pair repulsions at short distance, are clear.

Groups 14–16 show a steady decrease in multiple self-bond strength down the group. It is clear that the period two element forms by far the strongest double and triple self-bonds. Subsequent elements in these families show substantially weaker multiple self-bonds.

In period two, single self-bonds become smoothly stronger from Li to C, as effective nuclear charge increases. A sharp decrease at N that persists through F again shows the interatomic lone-pair repulsion effect. Periods three through five continue this trend: A smooth increase in single self-bond energy through group 14, a “jog” between groups 14 and 15, and again a smooth increase from group 15 to group 17.

In groups 15 and 16, the period two elements form double self-bonds that are more than twice as strong as their single self-bonds, due to a large reduction in interatomic lone-pair repulsions when hybridization changes from sp^3

¹Several of these examples were suggested by my colleague, Ladislav H. Berka.

to sp^2 . In all other cases, double bond energies are less than twice the corresponding single bond energies.

Using this presentation of bond energy data, one can explain with little more than a glance why, for example,¹

1. HF etches glass;
2. most explosives involve single bonds between nitrogen and itself or some other element;
3. elemental nitrogen is so unreactive;
4. fluorine is so reactive toward almost all other elements;
5. halogens are unreactive toward oxygen, which reacts readily with most other elements;
6. ozone is unstable with respect to dioxygen;
7. the stable allotrope of oxygen is $O=O$, while that of S involves S–S bonds;
8. acidity of the binary hydrides of groups 16 and 17 increases down the group;
9. boron hydrides spontaneously ignite in air;
10. expected “hydroxides” (i.e., oxoacids) of period two elements are unstable (e.g., $C(OH)_4$ decomposes to H_2CO_3 ; $N(OH)_3$ decomposes to HNO_2 (4));
11. phosphorous acid has the structure $P(H)(=O)(OH)_2$, rather than $P(OH)_3$;
12. borates and silicates are so abundant in nature;
13. the stable oxide of carbon is monomeric $O=C=O$; whereas, that of silicon is the polymeric SiO_2 , containing single Si–O bonds;
14. silicon forms a limited series of catenated silanes; and
15. the chemistry of carbon is extensive.

Although most of the data in the figure were taken from reference 3, the source of data is not critical. The figure can be updated readily as new data become available.

Literature Cited

1. Bowser, J. R. *Inorganic Chemistry*; Brooks/Cole Publishing: Pacific Grove, CA, 1993.
2. Shriver, D. F.; Atkins, P. W.; Langford, C. H. *Inorganic Chemistry*; W. H. Freeman and Company: New York, 1990.
3. Huheey, J. E.; Keiter, E. A.; Keiter, R. L. *Inorganic Chemistry—Principles of Structure and Reactivity*, 4th ed., Harper Collins College Publishers: New York, 1993.
4. Kildahl, N. K. *J. Chem. Educ.* **1991**, *68*, 1001.