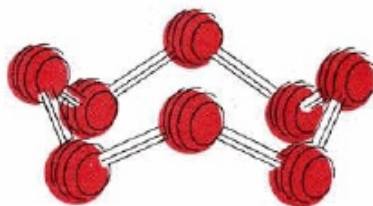


## 12.1: Prelude to Descriptive Chemistry

So far we have been devoted to explaining fundamental concepts and principles such as the atomic theory, electronic structure and chemical bonding, intermolecular forces and their effects on solids, liquids, or gases, and classes of reactions such as redox or acid-base. It is well to remember, though, that all these concepts and principles have been developed and used by chemists in order to better understand, recall, and systematize macroscopic laboratory observations. In other words, although the concepts we have described so far have their own inherent beauty as great ideas, they are primarily important because they reduce the quantity of memorization which is necessary to master descriptive chemistry, allowing people to recall facts that otherwise might be forgotten.



Such a reduction in memory work is only possible if you know how to apply principles to specific elements, their compounds, and the reactions they undergo. This is not as easy as it may seem, but neither is it impossibly difficult. We will describe some of the chemistry of the representative elements, showing as we do so how their properties may be rationalized on the basis of concepts and principles. Explanation of these properties is organized according to the concept of periodicity, with each subsequent section corresponding to one of the eight groups of representative elements. As you read, you should try to see why a certain reaction occurs, as well as what actually happens, and instead of memorizing each specific equation, you should try to organize the chemistry of these elements according to the generalizations you have already learned.

At this point it is useful to look and consolidate some of the general trends observed for the representative elements. First, metals on the far left of the periodic table are good reducing agents, while nonmetals on the far right (excluding noble gases) are strong oxidizing agents. Thus these elements are quite reactive, especially when one from the left combines with one from the right. Hydrogen compounds (hydrides) of the alkali and alkaline-earth metals contain strongly basic  $\text{H}^-$  ions and produce basic solutions. Toward the middle of the periodic table acid-base properties of hydrogen compounds are harder to predict. Some, like  $\text{CH}_4$  in Group IVA, are neither acids nor bases, but others, like  $\text{NH}_3$ , have lone pairs of electrons and can accept protons. Protons can be easily donated and are acidic only when they are bound to halogens or oxygen.

The acidic behavior of oxides also increases from left to right across the periodic table and decreases from top to bottom. The situation is complicated by the fact that the higher the oxidation state of an atom, the more covalent its oxide and the more acidic it will be. Thus  $\text{SO}_3$  dissolves in water to give a strong acid, while  $\text{SO}_2$  gives a weak one. Taking account of both of these trends, one can fairly well predict which oxides are likely to be basic, which amphoteric, and which acidic.

General rules can also be used to predict which oxidation states will be most common. On the left of the periodic table the group number gives the most common oxidation state. From group IIIA on, the group number minus 2 (for the  $ns^2$  electrons) is also common, especially for elements near the bottom of the table. The group number is a good choice when an element combines with a highly electronegative element, but the group number minus 2 is more common when one element is bonded to another element of intermediate electron-withdrawing power. For example from the Chalcogens,  $\text{SF}_4$  and  $\text{SF}_6$  are both stable, but  $\text{SF}_4$  is the most stable sulfur chloride.

From group VA on to the right of the table, the group number minus 8 is an important oxidation state, especially for the first member of a group. Oxidation numbers other than those already mentioned usually differ by increments of 2. For example, chlorine exhibits  $-1$ ,  $+1$ ,  $+3$ ,  $+5$  and  $+7$  oxidation states in stable compounds, and sulfur is found in  $-2$ ,  $+2$ ,  $+4$ , and  $+6$  states.

In conclusion, do not overlook the forest by concentrating too much on individual trees. Look for and try to understand and use the generalizations and correlations that have been developed in this chapter. If you do this, you will retain the facts presented here much more efficiently.

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