

Lewis Structures and Molecular Geometry

Models of Covalent Bonding

Introduction

Molecules have shape! The structure and shape of a molecule influences its physical properties and affects its chemical behavior as well. Lewis structures and VSEPR theory offer useful models for visualizing the structures of covalent compounds.

Concepts

- Valence electrons
- Covalent bonding
- Lewis structures
- VSEPR theory

Background

Covalent bonds are defined as the net attractive forces between nonmetal atoms that share one, two or three pairs of electrons. In general, only the *valence electrons*, those in the highest energy levels that are farthest away from the nucleus, are available for bonding. The number of valence electrons influences the number of bonds that an atom will form. The Periodic Table offers a convenient shortcut for determining the number of valence electrons in an atom. Remember that the position of an element in the modern Periodic Table reflects its electron configuration. When the representative elements are arranged in columns from IA to VIIA (see Figure 1), the number of valence electrons for an element is equal to its group number. Thus, potassium in Group IA has one valence electron, carbon in Group IVA has four valence electrons, and chlorine in Group VIIA has seven valence electrons.

IA	IIA							IIIA	IVA	VA	VIA	VIIA
									C			Cl
K												

Figure 1. Numbering of representative elements in Groups IA–VIIA.

In 1916, G. N. Lewis, an American chemist, proposed arranging dots around the symbols of the elements to represent the valence electrons. *Lewis electron dot symbols* (see Figure 2) remain the most popular way to picture the valence electrons that are available for bonding.



Figure 2. Lewis electron-dot symbols for representative elements.

Lewis structures built on the Lewis electron-dot symbols of the elements to show the bonding arrangement of atoms in a molecule and the distribution of all valence electrons. The Lewis structure of a molecule thus shows all of the atoms and how they are connected. A single covalent bond between two atoms, corresponding to a pair of electrons, is represented using a dash (—). Sometimes atoms share more than one pair of electrons between them in order to form stable molecules. Two dashes (corresponding to two pairs of electrons) and three dashes (corresponding to three pairs of electrons) are used to represent double and triple bonds, respectively.

G. N. Lewis offered a simple theory, based on the known stability of the noble gases (He, Ne, etc.), to predict how many bonds an atom will form and how many atoms of a particular type will come together to form a stable molecule. According to Lewis, nonmetals may share electrons in order to achieve a valence shell electron “count,” similar to that of the noble gases:

“Most atoms may conform to the rule of eight, or the octet rule... by sharing one or more pairs of electrons. The electrons which are held in common by two atoms may be considered to belong to the outer shell of both atoms.”

The noble gases have filled s and p orbitals with eight electrons. The octet rule assumes that atoms form molecules to achieve this stable, noble gas electron configuration. In counting valence electrons in order to predict the structure of a covalent compound, we will distinguish between two kinds of electron pairs. *Sharing pairs* of electrons are shared between atoms and thus “belong” to both atoms in the bond. *Nonsharing or unshared pairs* of electrons are not shared between atoms and are therefore “counted” toward only one of the atoms.

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Geometry of Molecules George E. Ryschkewitsch, 1963 *Molecules and Models* Arne Haaland, 2008-03-06 This book
provides a systematic description of the molecular structures and bonding in simple compounds of the main group elements
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structures of hydrogen halogen and methyl derivatives of the elements in each group some of these molecules are ionic some
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have not yet been studied The book provides a lucid discussion of a number of models such as the Lewis electron pair bond
and the VSEPR models the spherical and polarizable ion models and molecular orbital calculations and it outlines the
successes and failures of each *A Pictorial Approach to Molecular Bonding* John G. Verkade, 2012-12-06 With the
development of accurate molecular calculations in recent years useful predictions of molecular electronic properties are
currently being made It is therefore becoming increasingly important for the non theoretically oriented chemist to appreciate

the underlying principles governing molecular orbital formation and to distinguish them from the quantitative details associated with particular molecules. It seems highly desirable then that the non-theoretician be able to deduce results of general validity without esoteric mathematics. In this context pictorial reasoning is particularly useful. Such an approach is virtually indispensable if bonding concepts are to be taught to chemistry students early in their careers. Undergraduate chemistry majors typically find it difficult to formulate molecular orbital schemes especially delocalized ones for molecules more complicated than diatomics. The major reason for this regrettable situation is the general impracticability of teaching group theory before students take organic and inorganic courses wherein the applications of these concepts are most beneficial. Consequently many students graduate with the misconception that the ground rules governing bonding in molecules such as NH_3 are somehow different from those which apply to aromatic systems such as C_6H_6 . Conversely seniors and many graduate students are usually only vaguely if at all aware that sigma bonding like extended pi bonding can profitably be described in a delocalized manner when discussing the UV photoelectron spectrum of CH_2 for example.

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