CHM 101
Chemical Bondings

Reference:

General Chemistry
Principles and Modern Applications TENTH EDITION,
Pearson Canada
Toronto

Contents

Some fundamental ideas associated with Lewis s theory follow:

1. Electrons, especially those of the outermost (valence) electronic shell, playa fundamental role in chemical bonding.

2. In some cases, electrons are *transferred* from one atom to another. Positive and negative ions are formed and attract each other through electrostatic forces called **ionic bonds**.

3. In other cases, one or more pairs of electrons are *shared* between atoms. A bond formed by the sharing of electrons between atoms is called a **covalent bond**.

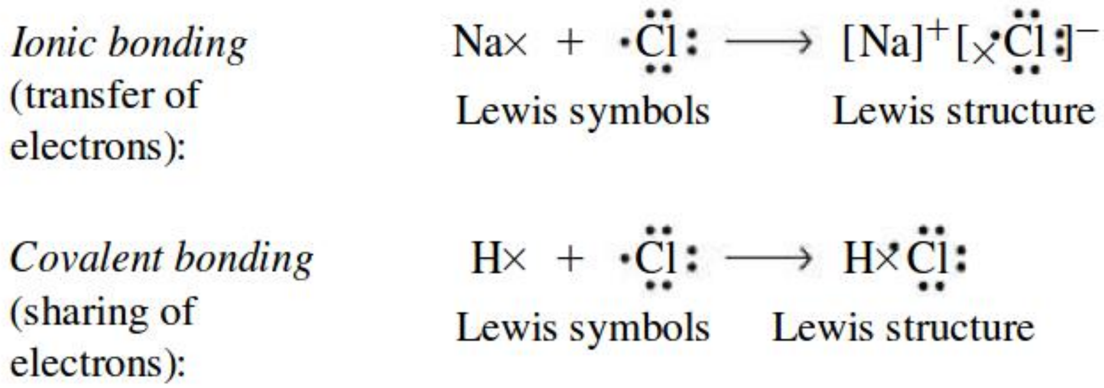
4. Electrons are transferred or shared in such a way that each atom acquires an especially stable electron configuration. Usually this is a noble gas configuration, one with eight outer-shell electrons, or an **octet**.

Lewis symbols and Lewis structure

- A **Lewis symbol** consists of a chemical symbol to represent the nucleus and *core* (inner-shell) *electrons* of an atom, together with dots placed around the symbol to represent the *valence* (outer-shell) *electrons.*
- the Lewis symbol for silicon, which has the electron configuration [Ne]$3s^2 3p^6$, is

- We will write Lewis symbols in the way Lewis did.
- We will place single dots on the sides of the symbol, up to a maximum of four.
- Then we will pair up dots until we reach an octet. Lewis symbols are commonly written for main-group elements but much less often for transition elements.

A **Lewis structure** is a combination of Lewis symbols that represents either the transfer or the sharing of electrons in a chemical bond



IONic Bonding

- electrons are transferred between valence shells of atoms
- ionic compounds are made of ions

NOT MOLECULES

- ionic compounds are called Salts or Crystals

IONic Bonding

- Electronegativity difference **> 2.0**
  - Look up e-neg of the atoms in the bond and subtract

    $NaCl$

    $CaCl_2$
- **Compounds** with polyatomic ions

  $NaNO_3$

Properties of Ionic Compounds

SALTS
Crystals

- hard solid @ 22^{o}C
- high mp temperatures
- **nonconductors** of electricity in **solid** phase
- **good** conductors in liquid phase or dissolved in water (aq)

Covalent Bonding

molecules

- **Pairs** of e- are **shared** between **non-metal** atoms
- electronegativity *difference* **< 2.0**
- forms polyatomic ions

Covalent Bonding

- The broken circles represent the outermost electron shells of the bonded atoms.
- The number of dots lying on or within each circle represents the effective number of electrons in each valence shell. The H atom has two dots, as in the electron configuration of He. The Cl atom has eight dots, corresponding to the outershell configuration of Ar.
- Note that we counted the two electrons between H and Cl *twice*. These two electrons are shared by the H and Cl atoms. This shared pair of electrons constitutes the covalent bond.

- As was the case for Cl in HCl, the O atom in the Lewis structure of H_2O and Cl_2O in is surrounded by eight electrons (when the bond-pair electrons are double counted). In attaining these eight electrons, the O atom conforms to the **octet rule** a requirement of eight valence-shell electrons for the atoms in a Lewis structure. Note, however, that the H atom is an exception to this rule. The H atom can accommodate only two valence-shell electrons.

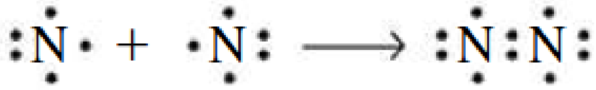
- The sharing of a single pair of electrons between bonded atoms produces a **single covalent bond**. To underscore the importance of electron pairs in the Lewis theory the term **bond pair** applies to a pair of electrons in a covalent bond, while **lone pair** applies to electron pairs that are not involved in bonding. Also, in writing Lewis structures it is customary to replace bond pairs with lines (—). These features are shown in the following Lewis structures.

- The bond formed between the N atom of and the ion in structure is a *coordinate covalent bond*. It is important to note, however, that once the bond has formed, it is impossible to say which of the four bonds is the coordinate covalent bond. Thus, a coordinate covalent bond is indistinguishable from a regular covalent bond.

Multiple Covalent Bonds

- More than one pair of electrons must be shared if an atom is to attain an octet (noble gas electron configuration).
- CO_2 and N_2 are two molecules in which atoms share more than one pair of electrons.
- C atom can share a valence electron with each O atom, thus forming two carbon-to-oxygen single bonds.

- Lewis structure for the molecule.
- Our first attempt might again involve a single covalent bond and the incorrect structure shown below:

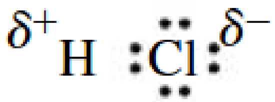


(*Incorrect*)

- Each N atom appears to have only six outer-shell electrons, not the expected eight. The situation can be corrected by bringing the four unpaired electrons into the region between the N atoms and using them for additional bond pairs.
- In all, we now show the sharing of *three* pairs of electrons between the N atoms.

Polar Covalent Bonds

- A covalent bond in which electrons are not shared equally between two atoms is called a **polar covalent bond.**
- The unequal sharing of the electrons leads to a partial negative charge on the more nonmetallic element, signified by $\delta-$ and a corresponding partial positive charge on the more metallic element, designated by $\delta+$



Electronegativity

- **Electronegativity (EN)** describes an atom s ability to compete for electrons with other atoms to which it is bonded. As such, electronegativity is related to ionization energy (*I*) and electron affinity (*EA*).

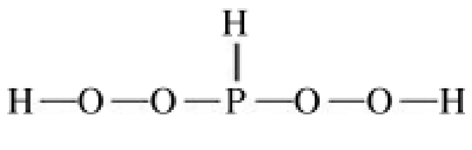
$$\mathrm{A} + \mathrm{B} \longrightarrow \mathrm{A}^{+}\mathrm{B}^{-} \qquad \Delta E_1 = (I_{\mathrm{A}} + EA_{\mathrm{B}})$$

$$\mathrm{A} + \mathrm{B} \longrightarrow \mathrm{A}^{-}\mathrm{B}^{+} \qquad \Delta E_2 = (I_{\mathrm{B}} + EA_{\mathrm{A}})$$

Writing Lewis Structures

- *All* the valence electrons of the atoms in a Lewis structure must appear in the structure.
- *Usually*, all the electrons in a Lewis structure are paired.
- *Usually*, each atom acquires an outer-shell octet of electrons. Hydrogen, however, is limited to two outer-shell electrons.
- *Sometimes*, multiple covalent bonds (double or triple bonds) are needed.
- Multiple covalent bonds are formed most readily by C, N, O, P, and S atoms.

- Skeletal structure all the atoms in the structure arranged in the order in which they are bonded to one another. In a skeletal structure with more than two atoms, we generally need to distinguish between central and terminal atoms.
- A central atom is bonded to two or more atoms, and a terminal atom is bonded to just one other atom.
- Hydrogen atoms are always terminal atoms. This is because an H atom can accommodate only two electrons in its valence shell, so it can form only one bond to another atom.
- Central atoms are generally those with the lowest electronegativity.
- Carbon atoms are always central atoms.
- Molecules and polyatomic ions generally have compact, symmetrical structures



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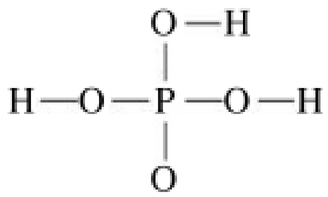

(*Correct*)

Resonance

- When we apply the usual rules for Lewis structures for ozone, we come up with these *two* possibilities.

- The situation in which two or more plausible Lewis structures contribute to the correct structure is called **resonance**.