Chapter 8

Basic Concepts of Chemical Bonding

Chapter 8 suggested problems
10th Ed.: 7, 11, 13, 17, 29, 33, 35, 45, 49, 53, 57, 59, 61, 84
11th Ed.: 7, 11, 13, 17, 29, 33, 35, 45, 49, 53, 57, 59, 61, 88

Class Notes

I. Lewis electron dot structures (Lewis structures) for atoms

   A. Octet rule - generally governs observed behavior

   B. Lewis structures generally consist of the elemental symbol surrounded by one dot for each valence electron of the substance

      1. Valence electrons are the outer shell s and p electrons

      2. Electrons in filled d shells behave as inner core electrons

      3. In partially filled d shells the d electrons are valence electrons (transition metals)

   C. Examples - 2nd and 3rd period elements

II. Ionic bonds and covalent bonds

   A. Ionic bonds and electron transfer

      1. Lewis structures of ions

      2. Bonding and non-bonding electron pairs

   B. Covalent bonds and electron sharing

      1. Show Cl_2 as an example

      2. Bonding and non-bonding electron pairs

3. Coordinate covalent bonds (Lewis acid-base reactions) - $H^+ + NH_3 \Rightarrow NH_4^+$

III. Lewis structures for molecules

A. Rules

1. Sum the valence electrons from all atoms
2. Draw the skeleton structure
3. Subtract 2 e^-/bond from the total electrons and divide the remainder by 2
4. Assign nbp to terminal atoms
5. Assign remaining nbp to central atom
6. Form multiple bonds if necessary
7. Resonance structures if appropriate

B. Examples

1. CH_4
2. PCl_3
3. HCN
4. BrO_3^-
5. NH_4^+
6. SO_3
7. CO_3^{2-}

IV. Exceptions to the octet rule

A. Molecules with an odd number of electrons (ClO_2, NO, NO_2)

B. Molecules with an atom with less than an octet (H, Be, B in BH_3)

C. Molecules with an atom with more than an octet

1. Most common: P (10 electrons) and S (12 electrons)
2. Other elements in the 3rd Period and higher take advantage of empty d orbitals to accommodate the extra electrons

a. ICl_4^-

b. PF_5

V. Resonance

A. In the real world delocalized bonding is a common phenomena

B. Delocalized bonding: a bonding pair of electrons is shared by more than one pair of atoms

C. A single Lewis structure cannot adequately describe delocalized bonding, so resonance structures are drawn to represent all of the possible Lewis structures

1. The stability of various resonance structures is affected by any formal charges on atoms in the structures

2. Each contributing structure must have the same number of electrons and the same net charge

3. Each contributing structure must have the same number of non-bonding electrons

D. The actual molecule is the average of all possible structures; the molecule does not "flip" from one structure to another

E. Delocalized bonding (and hence resonance) results in stabilization of a compound; a molecule with delocalized electrons possesses greater stability than than of any of the contributing individual structures

1. Benzene is 152 kJ/mol more stable than 1,3,5-cyclohexatriene

2. The difference is called "resonance energy"

VI. Formal Charge

A. Formal charge - "the *hypothetical charge* (just as oxidation numbers are hypothetical charges) on an atom in a Lewis structure, obtained by assuming that all bonding electrons are shared equally by the atom and that the electrons in all non-bonding pairs (lone pairs) belong exclusively to the atom"

B. Formal charge = (total valence electrons) - (½ x total bonding electrons) - (total non-bonding electrons)

1. Rule 1: if there is more than one possible Lewis structure, the one with the lowest magnitudes of formal charges is the most stable

2. Rule 2: If two or more structures have formal charges of the same magnitude, the structure with negative formal charges on the more electronegative atom is the most stable

C. Examples

1. ClNO

2. $POCl_3$

3. N_2O

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