# Biochem 03
## *November 11, 2009*

- Lewis Structures
- Redox Review
- Case Studies

# Five Steps for Lewis

- 1. Where do the nucleii go?

- 2. How many electrons are to be distributed?
  - only count electrons in outermost or valence shell
  - subtract one electron for each plus charge, add one electron for each minus charge.

- 3. How are the electrons to be distributed?
  - Octet rule:  try to make completed shells (usually octets, except hydrogen, which has two, and atoms in 3rd or higher row which can have up to 18, though usually prefer 8); sometimes, you cannot possibly get octets, like when you have odd number of electrons, but that's ok.

# Five Steps for Lewis

- 3. Cont'd Distribution recipe which is "foolproof"
  - a) Draw a line representing a single bond containing two electrons between joined atoms.
  - b) Distribute the remaining electrons evenly in pairs on the outer atoms so these have up to eight electrons (except for hydrogen). Any still not used after this should be placed on the central atom.
  - c) If the central atom is now surrounded by fewer than eight electrons, move sufficient nonbonding pairs from outer atoms other than halogens to between joined atoms, thus making them nonbonding, to bring the number on the central atom up to a maximum of eight (second period) or eighteen (third and higher periods).

# Five Steps for Lewis

- **4. What are the formal charges on the individual atoms?**
  - Count the number of electrons "owned" by each atom, pretending bonding electrons are evenly shared. To evaluate the formal charge at that atom, compare the result with the number of valence electrons of the neutral atom. Show only nonzero charges.

  - **F.C. = valence electrons - [unshared $e^-$ +$\frac{1}{2}$(shared $e^-$)]**

  - Remember that the formal charges on the atoms must add up to the total charge on the molecule. If they don't, then you have miscounted electrons. This is a good check of your structure.

# Five Steps for Lewis

- **5. Choosing among competing Structures**: There may be several structures which all are "correct" but which is the best Lewis Structure?
  - The best structure is one which minimizes charge distribution while maintaining closed shell configurations.
  - If you have a choice, put the negative charge on the more electronegative atom.
  - Sometimes, it is impossible to choose, and in those cases, the structures may in fact ALL contribute to the "real" structure. Example is resonance

# Five Steps for Lewis

- Draw the Lewis structures of the following molecules or ions, starting by counting the total number of valence electrons and include the formal charge on each atom.
- CO
- $CO_2$
- $[CO_3]^{-2}$

# Conservation of Charge

- *What is it?* Fundamental particles cannot be created or destroyed in a chemical reaction, therefore, the net charge must be conserved
- The charges of atoms and complex ions must be known in order to write correct molecular formula and balance charges in a chemical equation
  - Molecules and ionic compounds form through sharing or trading electrons. Though we don't write these electrons down explicitly, they are understood in the equation
  - Charges on both sides must be balanced

# Why must charges be balanced in a compound or reaction?

- If a charge in nature is not balanced, **enormous** quantities of energy are required. The lowest energy state of a system always has positive ions balanced by negative ions
- *Example*: human DNA, polyanions of the phosphate backbone of the DNA are balanced by positively charged metals such as $Zn^{2+}$, $Mg^{2+}$, $Ca^{2+}$, $Na^{+}$, $K^{+}$.
  - In fact, with $4 \times 10^9$ base pairs, or $8 \times 10^9$ (–) charges from phosphate backbone, many of these charges are also charge balanced by proteins which bear (+) charges called histones.

# Oxidation States

- Oxidation states only represent true charges on atoms in ionic compounds.  In covalent compounds, they represent an *effective* charge.
- Assigning oxidation states allows us to keep track of electrons in chemical reactions
  - This is essential in order to ensure that the conservation of charge is maintained
- Oxidation states depends upon the electronegativity of the elements in a covalent compound
- The electronegativity of an element measures its tendency to  attract electrons (pp. 581-583, Zumdahl)
  - Ranges from 0.7 (Fr and Cs with little ability to

# Common Charges of Groups

| COLUMN | TENDS to...... | OXIDATION | NAME |
| --- | --- | --- | --- |
| Group 1A | loose an electron | +1 | Alkali Metals |
| Group 2A | loose two electrons | +2 | Alkaline Earth |
| Group 3A,13 | share 3 electrons | (+3) | B, Al,....… |
| Group 4A,14 | share 4 electrons | (+4) | C, Si,....… |
| Group 5A,15 | share 3 electrons | (-3) | N, P......… |
| Group 6A,16 | gain two electrons | -2 | O, S......… |
| Group 7A,17 | gain one electron | -1 | Halides |
| Group 8A,18 | doesn’t react | 0 | Noble Gases |

# Rules for Assigning Oxidation States

1. The oxidation state of an atom in an element is 0.

2. The oxidation numbers of the atoms in a neutral molecule must add up to zero, and those in an ion must add up to the charge on the ion.

3. Alkali metals have an oxidation number of +1, and alkaline earth metals +2, in their compounds.

4. Fluorine always has an oxidation number of –1 in its compounds. The other halogens have –1 in their compounds *except* in compounds with oxygen and other halogens, where they can have positive oxidation #s.

5. Hydrogen has an oxidation number of +1 in its compounds *except* metal hydrides (i.e., LiH), where rule #3 takes precedence, and H has oxidation number –1

6. Oxygen is assigned an oxidation number of –2 in its compounds except when bound to fluorine (rule #4 takes precedence) and in compounds containing O–O bonds (such as $Na_2O_2$ and $H_2O_2$) where rules #3 and #5 take precedence.

# Assigning Oxidation States: Examples

What are the oxidation states of the atoms in the following compounds/ions?

| | | |
|---|---|---|
| $CO_3^{2-}$ | $Fe_2(SO_4)_3$ | $SO_2$ |
| $NH_3$ | $ClO^-$ | $CaH_2$ |
| $KMnO_4$ | $I_2$ | $N_2O$ |

# Assigning Oxidation States: Examples

What are the oxidation states of the atoms in the following compounds/ions?

$\overset{+4}{C}\overset{-2}{O}{}_3^{2-}$

$\overset{+3}{Fe}_2(\overset{+6}{S}\overset{-2}{O}_4)_3$

$\overset{+4}{S}\overset{-2}{O}_2$

$\overset{-3}{N}\overset{+1}{H}_3$

$\overset{+1}{Cl}\overset{-2}{O}{}^{-}$

$\overset{+2}{Ca}\overset{-1}{H}_2$

$\overset{+1}{K}\overset{+7}{Mn}\overset{-2}{O}_4$

$\overset{0}{I}_2$

$\overset{+1}{N}_2\overset{-2}{O}$

# Oxidation, Reduction, and Charge Balance in Chemical Equations

- If a material is oxidized, it loses electron (acts as a reducing agent).
- If a material is reduced, it gains electrons (acts as an oxidizing agent).
- **Remember LEO the GERman, Lose Electrons Oxidized, Gain Electrons Reduced**

# Igniting Magnesium in Air

$$\underline{\ \ 2\ \ }Mg(s) + \underline{\ \ 1\ \ }O_2(g) \rightarrow \underline{\ \ 2\ \ }MgO(s)$$

$$Mg^0 \rightarrow Mg^{2+} + 2e^-$$

$$O_2 + 4e^- \rightarrow 2O^{2-}$$

$$\underline{\ \ 2\ \ }Mg(s) + \underline{\ \ 1\ \ }O_2(g) \rightarrow \underline{\ \ 2\ \ }MgO(s)$$

**Mg is oxidized to $Mg^{2+}$**

**O is reduced from O to $O^{2-}$**