MP Jansen
12/AP/IB Chemistry

Guidelines for Determining Oxidation Numbers

Oxidation numbers are used to keep track of electron movement in oxidation-reduction (redox) reactions.

An element that is **oxidized** loses electrons (<u>L</u>oss of <u>E</u>lectrons due to <u>O</u>xidation, LEO); when an element is **reduced** it gains electrons (<u>G</u>ain of <u>E</u>lectrons due to <u>R</u>eduction, GER).

Think of oxidation numbers as an accounting system for electrons.

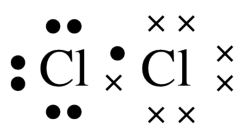
Oxidation Number Rules—and an explanation of the rules

1. The oxidation number (ON) of any element in its standard state (SATP) is 0.

   eg. The ON of metallic iron, Fe, is zero, 0

   the oxidation number of each chlorine atom in Cl_2, the standard state of chlorine at SATP, is 0.

   Think of it like this: the Cl–Cl bond contains two shared electrons. Since each Cl atom has the same electronegativity value (3.0), the bonding electrons are shared equally. Therefore, each Cl "has" 7 valence electrons ($3s^2 3p^5$), as it does in the periodic table—no extra, no deficiency; hence the ON of each Cl atom is 0.

2. The oxidation number of a monatomic ion is simply the charge on the ion.
   eg. The ON of sodium in Na^+ is +I. The sodium cation has one fewer electron that neutral sodium; hence the ON of +I. (Recall: each electron carries a negative charge.)

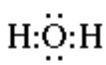
   The ON of the sulfide ion, S^{2-}, is –II. The sulfide ion has two "extra" electrons.
   NB. IUPAC states that oxidation numbers be represented by the appropriate charge (+ or –), followed by the **Roman numeral**.

3. As you might have guessed, the sum of the oxidation numbers of all atoms in a neutral compound must equal zero; for a polyatomic ion, the ONs add to give the charge on the ion.

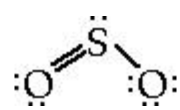
4. To find the ON of an element in a molecular compound or in a polyatomic ion it is best to use a few "rules". The rules are based on the electronegativity concept and on Lewis structures, introduced in #1, above. Here's how to determine the ON of an element in a compound or polyatomic ion: Look at the Lewis structure of the compound or ion. The bonding electrons "belong" to the more electronegative element. Compare this number of electrons to the number of electrons in the neutral atom to arrive at the ON for that atom. Extra electrons mean a negative ON. This explains why the ON of Cl in Cl_2 is 0; see #1 above.

a) The ON of oxygen in almost all compounds is –II. Oxygen is extremely electronegative (EN = 3.5), second only to fluorine (EN = 4.0). This means that when O is bonded to anything other than F or to another O, it will "take" the bonding electrons.

eg. $H_2\underline{O}$

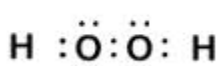


$S\underline{O}_2$



Now look at cases where O is bonded to another O (H_2O_2) or to F (OF_2).

eg. H_2O_2



OF_2

F O F

b) The ON of H in all compounds, except metal hydrides (eg NaH), is +I. Since H has a low electronegativity value (EN = 2.1), it will "lose" its electron when bonded to anything more electronegative. Since metals are less electronegative that hydrogen, H will "take" the bonding electrons when it bonds to alkali metals or to alkali earth metals.

eg. $C\underline{H}_4$ NaH with H underlined: $Na\underline{H}$

Some more practice. Remember that ONs for the atoms in a neutral compound must total zero; for a polyatomic ion, the ONs must total the charge on the ion.
eg. $\underline{C}_2H_6$

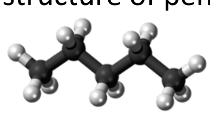
$\underline{S}O_3$

$\underline{P}O_4^{3-}$

$\underline{Cr}_2O_7^{2-}$ (find the ON of *each* Cr atom)

$\underline{P}_4O_{10}$

N.B. Oxidation numbers are not always whole numbers; For example, in pentane, C_5H_{12}, the two terminal C atoms are in different bonding environments than the three C atoms in the middle of the chain. Use the ON Rules to find the *average* ON of C in pentane, then draw the Lewis structure of pentane and determine individual C's ONs in that manner.



—*fin*—