Oxidation Numbers

In many reactions, atoms break and form bonds with other atoms, but otherwise the atoms don't change much. In some ionic reactions, however, atoms alter their charge over the course of the reaction. An atom might get more electrons, and so its charge goes down. We say that the atom has been **reduced**. In other case, the atom might lose electrons, and we say that the atom has been **oxidized**. When we try to balance reactions where oxidation and reduction has occurred, we must be careful to balance the change in charge as well. One way to keep track of the charge of an element through a reaction is to assign it an **oxidation number**, which represents what the charge on the atom would be if it were an ion all by itself. This worksheet will summarize how to assign oxidation numbers to atoms.

There are three important principles to keep in mind before you assign numbers:

[1] **Oxidation numbers are assigned to *atoms*, not elements**. This means that if there is more than one atom of a particular element in a compound or ion, the oxidation number must be multiplied by the number of atoms. So, if hydrogen has an oxidation number of +1 in a water molecule, then there are two atoms with that number, and the *total* for hydrogen is +2.

[2] **The total of the oxidation numbers in a neutral compound is zero**. Since oxidation numbers represent charges, their total should be zero in a neutral molecule.

[3] **The total of the oxidation numbers in an ion is equal to the charge of the ion**. This is a similar idea to the previous principle. This fact, in particular, is useful when you have ions consisting of more than one atom.

Keeping these principles in mind, we can assign oxidation numbers using these facts (in order of importance):

A. The oxidation number for an **elemental compound** is **0**. In Ag, the oxidation number of silver is 0. In S_8, the oxidation number for sulfur is 0.

B. In compounds with other elements, the oxidation number for **fluorine** is **−1**.

C. In most compounds, the oxidation number for **other halogens** is always **−1**. In oxyanions (such as $C\ell O_2^-$) it can vary.

D. **Metals** that form only one ion will always have that charge for an oxidation number. Sodium only forms Na^+, so it's always +1. Magnesium is always Mg^{2+}, so it's always +2.

E. The oxidation number for **hydrogen** is **+1** when bonded to a nonmetal and **−1** when bonded to a metal.

F. The oxidation number for **oxygen** is **−2** unless it is in oxygen gas, ozone (O_3), or a peroxide (a compound containing the O_2^{2-} ion, where the oxidation number for oxygen is −1).

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G. All other oxidation numbers can be found using logic. Remember that in an ionic compound, the ions will still have their individual charges, and you can use that information to determine oxidation numbers.

Example 1: Determine the oxidation numbers of chlorine in (a) $C\ell_2$ (b) $HC\ell$ (c) $NaC\ell O_4$.

Solution: (a) This is an elemental compound. The oxidation number of chlorine is 0.

(b) We look through the facts from top to bottom. Nothing applies until we come to fact C about halogens. We see that halogens in binary compounds have an oxidation number of −1. That applies here. Chlorine has −1.

(c) We look through again. The facts tell us directly that the oxygen atoms should each have an oxidation number of −2, and sodium is +1. We don't know what $C\ell$ is yet because it's in an oxyanion. The three principles tell us that ions should get their charges as oxidation numbers. Since the perchlorate ion ($C\ell O_4^-$) has a charge of −1, the total of the numbers on oxygen atoms and chlorine atoms must be −1. Each of the oxygen atoms already has an oxidation number of −2, so:

$$\begin{aligned} x + 4(-2) &= -1 \\ x &= +7 \end{aligned}$$

Chlorine has an oxidation number of +7 here.

EXERCISES

A. Assign an oxidation number to nitrogen in each of these compounds:
1) ammonia: NH_3
2) nitrogen gas: N_2
3) nitrous acid: HNO_2
4) nitric oxide: NO
5) dinitrogen tetroxide: N_2O_4
6) palladium nitrate: $Pd(NO_3)_2$ [*Hint: ionic compound.*]

B. Determine the oxidation numbers of the indicated elements:
1) ferrous oxide: $\underline{Fe}O$
2) xenon hexafluoride: $\underline{Xe}F_6$
3) uranyl chloride: $\underline{U}O_2C\ell_2$
4) aluminum hydroxide: $\underline{A\ell}(OH)_3$
5) buckminsterfullerene: $\underline{C}_{60}$
6) sodium peroxide: $Na_2\underline{O}_2$
7) boric acid: $H_3\underline{B}O_3$
8) silane: $\underline{Si}H_4$

C. Determine the oxidation numbers of all the elements in these ions:
1) hydroxide: OH^-
2) bicarbonate: HCO_3^-
3) iodate: IO_3^-
4) ammonium: NH_4^+
5) molybdate: MoO_4^{2-}
6) oxalate: $C_2O_4^{2-}$
7) dichromate: $Cr_2O_7^{2-}$
8) formate: HCO_2^-

SOLUTIONS

A. (1) −3 (2) 0 (3) +3 (4) +2 (5) +4 (6) +5

B. (1) +2 (2) +6 (3) +6 (4) +3 (5) 0 (6) −1 (7) +3 (8) +4

C. (1) O: −2, H: +1 (2) H: +1, C: +4, O: −2 (3) I: +5, O: −2 (4) N: −3, H: +1
(5) Mo: +6, O: −2 (6) C: +3, O: −2 (7) Cr: +6, O: −2 (8) H: +1, C: +2, O: −2