OXIDATION NUMBERS[1]

SECTION A - Assigning Oxidation Numbers

This worksheet presents the material in a slightly different fashion from the lessons, but the outcome is the same. ***This is not meant to replace your lessons nor does it cover everything in your lessons***. The worksheet is broken down into four sections: (A) assigning oxidation numbers, (B) determining oxidation numbers for ionic compounds, (C) determining oxidation numbers for transition metals, and (D) determining oxidation numbers in reactions.

Oxidation numbers, sometimes called oxidation states, are signed numbers assigned to atoms in molecules and ions. They allow us to keep track of the electrons associated with each atom. Oxidation numbers are frequently used to write chemical formulas, to help us predict properties of compounds, and to help balance equations in which electrons are transferred. Knowledge of the oxidative state of an atom gives us an idea about its positive or negative character. In themselves, oxidation numbers have no physical meaning; they are used to simplify tasks that are more difficult to accomplish without them.

Rules For Assigning Oxidation Numbers

Rule 1: All pure elements are assigned the oxidation number of zero.
Rule 2: All monatomic (single element) ions are assigned oxidation numbers equal to their charges.
Rule 3: Certain elements usually possess a fixed oxidation number in compounds.

- The oxidation number of O in most compounds is -2.
- The oxidation number of H in most compounds is +1.
- The oxidation number of halogens in many, but not all, binary compounds is -1
- The oxidation numbers of alkali metals (Group I) and alkaline earth metals (Group II) are +1 and +2, respectively.

Rule 4: The sum of all oxidation numbers in a compound equals zero, and the sum of oxidation numbers in a polyatomic ion equals the ion's charge.

Let's now discuss each rule.

[1] Wolfe, Drew H. Introduction to College Chemistry. McGraw-Hill Publishing Company. pp. 222-223.

Rule 1 states that all uncombined elements are assigned the oxidation number of zero, regardless of how they exist in nature - by themselves or diatomically. It is common practice to write oxidation numbers above the symbols of the atoms.

$$\overset{0}{Na} \quad \overset{0}{Fe} \quad \overset{0}{H}_2 \quad \overset{0}{O}_2$$

Rule 2 states that all monatomic ions are assigned oxidation numbers equal to their charge. When you write an oxidation number, the sign comes before the magnitude of the charge (-2). Charges on ions are written with the magnitude before the sign (2-). Some examples of oxidation numbers of monatomic ions are as follows:

$$\overset{-1}{Cl}{}^{1-} \qquad \overset{+3}{Al}{}^{3+} \qquad \overset{-2}{O}{}^{2-}$$

Rule 3 states that some atoms have fixed oxidation numbers. These atoms usually have oxidation numbers that correspond to the number of electrons that they lose or gain when forming a binary ionic compound. For example, halogens (GroupVII) gain one electron, alkali metals lose one electron, and alkaline earth metals lose two electrons to obtain noble gas configurations when they form ionic compounds. There are exceptions to Rule 3 but I won't go into those here, because you don't need to know those right now.

Through the application of **Rule 4**, we can identify the oxidation numbers of all elements in a compound. For example, we can calculate the oxidation number of N in NO. Since we know that the oxidation number of O is -2, and the sum of the oxidation numbers in NO must equal zero, then the oxidation number of N in NO must be +2. (*To keep track of oxidation numbers, write the known oxidation states for individual atoms below the formula and the total oxidation number for all elements of that type above the symbol.*)

$$\underset{+2}{\overset{?}{N}}\,\underset{-2}{\overset{-2}{O}}\;{}^{=0}$$

What is the oxidation number of S in SO_2? Again following **Rule 4**, we assign -2 to each O atom, making a total of -4; hence, the oxidation number of S is +4 in order for the sum to equal zero.

$$\underset{+4}{\overset{+4}{S}}\;\underset{-2}{\overset{-4}{O_2}}\;{}^{=0}$$

For a polyatomic ion, the reasoning is the same except that the sum of the oxidation numbers equals the charge on the ion. To illustrate this, let us calculate the oxidation number of a P atom in PO_4^{3-}. Four O atoms have a total oxidation number of -8, so the oxidation state of P must be +5 for the sum to equal the charge of -3.

$$\underset{+5}{\overset{+5}{P}}\ \underset{-2}{\overset{-8}{O_4^{3-}}}$$

With the exception of the metals in groups IA, IIA, and IIIA, metals generally can exist in more than one oxidation state. Chromium, Cr, for example, is found in the +6, +3, and +2 oxidation states. Gold is found in both +3 and +1 oxidation states. Nonmetals, except Fluorine, F, also exhibit a range of oxidation states. For example, the oxidation states of Sulfur, S, include +6, +4, +2, and -2.

SECTION B - Determining Oxidation Numbers for Ionic Compounds

The oxidation number for ionic compounds is the same as the ionic charge.

Example #1: $AlCl_3$ Al^{3+} Cl^{1-}

SECTION C - Determining Oxidation Numbers for Transition Metals

- Assign the non-metal, the one that carries a negative charge, its usual anion charge. (O = -2, Br = -1, etc.)
- Set the net charge of the compound to equal zero.
- Calculate the oxidation number needed to balance the total charge.

Example #2: What is the oxidation number of Cu in Cu_3N_2?
The oxidation number of N is -3.
There are 2 N atoms: (2)(-3) = -6
The net charge is 0.
There are 3 Cu atoms: (3)(?) = +6
Therefore, the oxidation state of Cu is +2.

SECTION D - Determining Oxidation Numbers in Reactions

An oxidation-reduction reaction, also called redox reaction, is the transfer of electrons from one species or substance to another. When we figure out what species or substance underwent oxidation and reduction, we are talking about the reactants.

OXIDATION	**REDUCTION**
Gain of oxygen *or*	Loss of oxygen *or*
Loss of hydrogen *or*	Gain of hydrogen *or*
Loss of electrons	Gain of electrons

There are two ways to figure out which has undergone reduction and oxidation. The first is a continuation of using what you have learned in Section A.

Example #3: $Zn_{(s)} + Cu^{2+}{}_{(s)} \rightarrow Zn^{2+}{}_{(aq)} + Cu_{(s)}$

Using **Rule #1** from **Section A**, all pure elements are assigned the oxidation number of zero. This would mean that Zn and Cu have a zero oxidation number in the above reaction. Zinc metal (reactant) loses electrons to become more positive - going from zero which is neutral to 2+ (product side) meaning it has lost electrons. The copper ion with its charge of 2+ (reactant) which is also the oxidation #, has gone from 2+ to zero or neutral. This means the copper has become more negative or gained electrons. From the definitions above then, because the zinc lost electrons, it underwent oxidation and because the copper gained electrons, then it underwent reduction. This can be related to a number line (draw). When going from a +2 to zero, you are headed in a negative direction.

Example #4: $CH_{4(g)} + 2O_{2(g)} \rightarrow CO_{2(g)} + 2H_2O$

Using **Rule #3** from **Section A**, the oxidation # of hydrogen is +1, oxygen is -2 and rule #4, where the sum of all oxidation #s in a compound equals zero. We can calculate the oxidation numbers for the above reactants - Because hydrogen always has a +1 charge, then the carbon will be -4; oxygen is in its elemental form so it will be 0. On the product side, because O is -2, the C would become +4. Water does not participate in oxidation reduction reactions.

Because the carbon went from -4 to +4, we know it lost electrons because it became more positive, then the CH_4

Example #5: What is the oxidation number of C in CO_2 ?
The oxidation number of O is -2.
There are two O atoms: (2)(-2) = -4
The net charge is 0.
There is one C atom: (1)(?) = +4
Therefore, the oxidation state of C is +4.

Example #6: What is the oxidation number of N in NO_3^- ?
The oxidation number of O is -2.
There are three O atoms: (3)(-2) = -6
The net charge is -1.
There is one N atom: (1)(?) = +5
The oxidation state of N is +5.

Example #7: What is the oxidation number of Cl in HClO?
The oxidation number of H = +1.
There is one H atom: (1)(+1) = +1
The oxidation number of O is -2.
There is one O atom: (1)(-2) = -2
Therefore, the oxidation state of Cl is +1.

Example #8: What is the oxidation number of Cl in $HClO_2$?
The oxidation number of H = +1.
There is 1 H atom: (1)(+1) = +1
The oxidation number of O is -2.
There are two O atoms: (2)(-2) = -4
Therefore, the oxidation state of Cl is +3.

Example #9: What is the oxidation number of Cl in $HClO_4$?
The oxidation number of H = +1.
There is one H atom: (1)(+1) = +1
The oxidation number of O is -2.
There are four O atoms here: (4)(-2) = -8
Therefore, the oxidation state of Cl is +7.

NOTE: The maximum positive oxidation number for chlorine is +7, the same as its group number (VII). This is generally true because an atom cannot share more electrons than it actually has in its valence shell.

Also, in covalent compounds, non-metals other than oxygen have positive oxidation numbers, indicating an unequal sharing of electrons. However, in ionic compounds, the non-metals other than oxygen have negative oxidation numbers, indicating they have gained electrons.

The following are practice problems for you to try.

1)	CuF_2	2)	HNO_3	3)	SO_4^{2-}
4)	$C_{12}H_{22}O_{11}$	5)	PBr_3	6)	$C_2O_4^{2-}$
7)	CO_3^{2-}	8)	H_2O	9)	$FeCl_2$

Answers**:**

(1) CuF_2 Halogens such as F have an oxidation number equal to -1; two F atoms have a total oxidation number of -2; therefore, it follows that the oxidation number of Cu is +2**.**

(2) HNO_3 The oxidation number of H is +1, and three O atoms have a total oxidation number of -6; thus, the oxidation number of N is +5 in order for the total to be zero.

(3) SO_4^{2-} The oxidation number of each O atom is -2, so four O atoms have a total oxidation number of -8. To make the sum of the oxidation numbers -2, the charge on SO_4^{2-}, the oxidation number of S is +6.

(4) $C_{12}H_{22}O_{11}$ The total oxidation number of 22 H atoms is +22 (22 x +1), and the total for 11 O atoms is -22 (11 x -2). Therefore, the 12 C atoms have a total oxidation number of 0, which means that each C has an oxidation number equal to zero.

(5) PBr_3 P +3, Br -1

(6) $C_2O_4^{2-}$ C +3, O -2

(7) CO_3^{2-} C +4, O -2

(8) H_2O H +1, O -2

(9) $FeCl_2$ Fe +2, Cl -1