COLLEGE OF THE CANYONS

Activity 201 – 3
Balancing Oxidation-Reduction Reactions

Directions: This GLA worksheet goes over the half-reaction method of balancing oxidation-reduction (redox) reactions. Part A introduces the oxidation states of individual atoms and identifying oxidation and reduction reactions. Part B discusses balancing redox reactions that occur in acidic solutions. Part C discusses balancing redox reactions that occur in basic solutions. The worksheet is accompanied by an in-depth key. See www.canyons.edu/departments/CHEM/gla for additional materials.

Part A – Oxidation States and Redox Reactions

Oxidation-Reduction (redox) reactions are a type of reaction that involves a transfer of electrons. Some types of redox reactions are very obviously redox reactions, such as reactions of a metal and a nonmetal to form an ionic compound. The transfer of electrons in some types of redox reactions, such as combustion reactions, is less obvious.

To determine whether or not a specific reaction is a redox reaction, we assign oxidation states (or oxidation numbers) to elements that compose both reactants and products. **Oxidation states** are charges that we assign to atoms as if the elements had charges.

The following are specific rules for assigning oxidation states:

1. The oxidation state of an atom in its elemental form is always zero. Most metals are typically solids and some elements are diatomic (such as O_2, N_2, F_2, H_2, etc.) in their elemental form.
2. The oxidation state of a monoatomic ion is equal to the charge of the ion.
3. The oxidation state of ions within ionic compounds are equal to their charge.
4. Oxidation states for elements within molecular compounds or polyatomic ions should be assigned in the following order:
   a) Fluorine is always -1
   b) Hydrogen is +1
   c) Oxygen is -2
   d) Group 7A is -1
   e) Group 6A is -2
   f) Group 5A is -3
   g) Anything not on this list should be assigned according to rule #5 below
5. The sum of the oxidation states for all atoms in a molecule or polyatomic ion equals the charge on the molecule or ion.

<u>Practice</u>

Assign oxidation numbers for each of the elements in each of the following atoms, compounds, or ions:

a) $Ag(s)$
Since silver is in its elemental form, its oxidation number is 0 (Rule #1 above).

b) $Ag^+(aq)$
Silver in this form is a monoatomic ion, so its oxidation number is +1 (Rule #2 above).

c) $N_2(g)$

Since nitrogen is in its elemental form, its oxidation number is 0 (Rule #1 above).

d) $C_2H_4(g)$

In this molecule, we assign the oxidation state of each element, C and H.

-2 +1
C_2H_4

Atom	Oxidation State	Reasoning
H	**+1** C_2H_4	**In this molecular compound, we assign the oxidation state of hydrogen first as +1 (Rule #4).**
C	**-2** C_2H_4	**We then assign the oxidation state of carbon depending on the charge of the whole molecule (Rule #5). Since the sum of all the oxidation states need to be equal to 0 (the charge of the entire molecule) then each carbon atom must have an oxidation state of -2. We can determine this with an equation:** **2(oxidation of C) + 4(oxidation of H) = 0** **2(oxidation of C) + 4(+1) = 0** **oxidation of C = -2**

e) $MnO_4^-(aq)$

In this ion, we assign the oxidation state of each element, Mn and O.

+7 -2
MnO_4^-

Atom	Oxidation State	Reasoning
O	**-2** MnO_4^-	**In this polyatomic ion, the oxidation state of oxygen is assigned first as -2 (Rule #4).**
Mn	**+7** MnO_4^-	**The oxidation state of manganese is assigned depending on the charge of the whole ion (Rules #4 and #5 above). The oxidation states of all the atoms in this ion must be equal to -1.** **1(oxidation of Mn) + 4(oxidation of O) = -1** **1(oxidation of Mn) + 4(-2) = -1** **oxidation of Mn = +7**

Example #1
Assign oxidation numbers for each of the elements in each of the following atoms, compounds, or ions:

a) C(s, graphite)

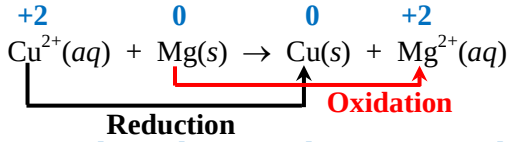
b) H_2O_2

c) SO_3^{2-}

d) Na_3PO_4

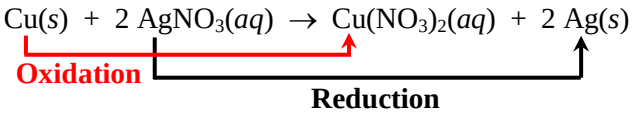
Once oxidation states are assigned, we can determine whether a reactant lost electrons (called **oxidation**) or gained electrons (called **reduction**) during a chemical reaction. When the oxidation states of reactants and products of a redox reaction are assigned, the oxidation state of one element will become more positive, meaning that it lost electrons, or was **oxidized**. This species is known as the **reducing agent**, because it allows another species to be reduced. (If no species lost electrons, then no species could gain electrons.) Conversely, the oxidation state of another element will become more negative, indicating that the species gained electrons, or was **reduced**. This species is called the **oxidizing agent**, because it allows another species to be oxidized.

Practice
Determine the oxidation states of each of the following elements in the chemical reaction below and determine which species is oxidized and which is reduced.



The Cu^{2+}(aq) ion is reduced. It gained two electrons, changing its oxidation state from +2 to 0.
The solid Mg(s) is oxidized. It lost two electrons, changing its oxidation state from 0 to +2.

Determine the oxidation states of each of the elements in the chemical reaction below and determine which species is oxidized and which is reduced.



The Cu (s) metal is oxidized. It lost two electrons, changing its oxidation state from 0 to +2. The Ag^+ (aq) ions are reduced. Each ion gained one electron, changing the oxidation state from +1 to 0.