Dougherty Valley HS Chemistry
Oxidation Numbers

WORKSHEET #00

Directions: Use the rules for Assigning Oxidation numbers to determine the oxidation number assigned to each element in each of the given formulas.

	Formula					Formula			
1.	Cl_2	Cl			16.	Na_2O_2	Na	O	
2.	Cl^-	Cl			17.	HNO_3	H	N	O
3.	Na	Na			18.	$CaCl_2$	Ca	Cl	
4.	Na^+	Na			19.	PO_4^{3-}	P	O	
5.	KCl	K	Cl		20.	MnO_2	Mn	O	
6.	H_2S	H	S		21.	K_3PO_4	K	P	O
7.	CaO	Ca	O		22.	Fe_2O_3	Fe	O	
8.	H_2O	H	O		23.	KNO_2	K	N	O
9.	NO_3^-	N	O		24.	N_2	N		
10.	NO_2	N	O		25.	Al^{3+}	Al		
11.	$Cr_2O_7^{2-}$	Cr	O		26.	H_2O_2	H	O	
12.	O_2	O			27.	H_2SO_4	H	S	O
13.	NH_3	N	H		28.	NH_4Cl	N	H	Cl
14.	CaH_2	Ca	H		29.	FeO	Fe	O	
15.	SO_4^{2-}	S	O		30.	SiO_2	Si	O	

Rules for Assigning Oxidation Numbers

1. The oxidation number of any uncombined element is 0.
2. The oxidation number of a monatomic ion equals the charge on the ion.
3. The more-electronegative element in a binary compound is assigned the number equal to the charge it would have if it were an ion.
4. The oxidation number o fluorine in a compound is always -1
5. Oxygen has an oxidation number of -2 unless it is combined with F, when it is +2, or it is in a peroxide, such as H_2O_2, when it is -1
6. The oxidation state of hydrogen in most of its compounds is +1 unless it is combined with a metal, in which case it is -1
7. In compounds, the elements of groups 1 and 2 as well as aluminum have oxidation numbers +1, +2 and +3 respectively.
8. The sum of the oxidation numbers of all atoms in a neutral compound is 0.
9. The sum of the oxidation numbers of all atoms in a polyatomic ion equals charge of the ion.

Dougherty Valley HS Chemistry
More Redox

WORKSHEET #0

$Li + F_2 \rightarrow 2F^- + Li^+$

Balance Oxidation half reaction:

Balance Reduction half reaction:

$Pb^{2+} + Mn^{2+} \rightarrow MnO_2 + Pb$

Balance Oxidation half reaction:

Balance Reduction half reaction:

$Cl_2 + 2Br^- \rightarrow 2Cl^- + Br_2$

Balance Oxidation half reaction:

Balance Reduction half reaction:

$Mg + NO_3^- \rightarrow Mg^{2+} + NO$

Balance Oxidation half reaction:

Balance Reduction half reaction:

$MnO_4^- + Pb \rightarrow Pb^{2+} + Mn^{2+}$

Balance Oxidation half reaction:

Balance Reduction half reaction:

$Fe_2O_3(s) + 2Al(s) \rightarrow 2Fe(l) + Al_2O_3(s)$

Balance Oxidation half reaction:
Balance Reduction half reaction:

$2Ag + Ce^{4+} \rightarrow Ag_2O_2 + Ce^{3+}$

Balance Oxidation half reaction:
Balance Reduction half reaction:

$PbO_2 + Ag \rightarrow Ag^{+} + Pb^{2+}$

Balance Oxidation half reaction:
Balance Reduction half reaction:

$Hg_2^{2+} + Cu \rightarrow Cu^{2+} + 2Hg$

Balance Oxidation half reaction:
Balance Reduction half reaction:

Dougherty Valley HS Chemistry
Oxidation State Worksheet

WORKSHEET #0a

In each of the following chemical compounds, determine the oxidation states of each element:

1) Sodium nitrate ________________________________

2) Ammonia ________________________________

3) Zinc oxide ________________________________

4) Water ________________________________

5) Calcium hydride ________________________________

6) Carbon dioxide ________________________________

7) Nitrogen ________________________________

8) Sodium sulfate ________________________________

9) Aluminum hydroxide ________________________________

10) Magnesium phosphate ________________________________

In each of the following reactions, determine what was oxidized and what was reduced.

11) $Ca + H_2O \rightarrow CaO + H_2$

Element oxidized: ________________________________

Element reduced: ________________________________

12) $2\ H_2 + O_2 \rightarrow 2\ H_2O$

Element oxidized: ________________________________

Element reduced: ________________________________

Dougherty Valley HS Chemistry
Redox Reaction Practice

WORKSHEET #1

1. Consider the following redox reaction: **NO*(g)* + $KMnO_4$*(aq)* → MnO_2*(s)* + KNO_3*(aq)***

a. Give the ***oxidation number*** for each element in the compounds below:

In **$KMnO_4$**: K = _____, Mn=_____, and O = _____.

In **KNO_3**: K = _____, N=_____, and O = _____.

Indicate the ***neutral element*** or the ***ion (****or* ***element with its oxidation number) in a compound*** that is oxidized or reduced and the ***reactants*** that served as the oxidizing and the reducing agents below:

b. The reactant oxidized is ______________, and the oxidizing agent is _______________.

The reactant reduced is ______________, and the reducing agent is _______________.

c. The total number of electrons transferred in this reaction is _________.

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2. Consider the following redox reaction: **Fe_2O_3(s) + 3 CO (g) → 2 Fe (s) + 3 CO_2(g)**
a. Give the ***oxidation number*** for each element in the compounds below:

In **CO_2**: C = _____, and O = _____.

b. The reactant oxidized is ______________, and the oxidizing agent is _______________.

The reactant reduced is ______________, and the reducing agent is _______________.

c. The total number of electrons transferred in this reaction is _________.

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3. Consider the following : **H_2SO_3(aq) + 2 Mn(s) + 4HCl(aq) → S(s) + 2 $MnCl_2$(aq) + 3 H_2O(l)**
a. Give the ***oxidation number*** for each element in the compounds below:

In **H_2SO_3**: H=_____, S = _____, and O = _____.

In **$MnCl_2$**: Mn = _____ and Cl = _____.

b. The reactant oxidized is ______________, and the oxidizing agent is _______________.

The reactant reduced is ______________, and the reducing agent is _______________.

c. The total number of electrons transferred in this reaction is _________.

4. Consider the following reaction: $\mathbf{Pb}(s) + \mathbf{PbO_2}(s) + \mathbf{H_2SO_4}(aq) \rightarrow \mathbf{2\ PbSO_4}(s) + \mathbf{2\ H_2O}(l)$
Indicate the ***neutral element*** or the ***ion (or element with its oxidation number) in a compound*** that is oxidized or reduced and the ***reactants*** that served as the oxidizing and the reducing agents below:

a. The reactant oxidized is _______________, and the oxidizing agent is ________________.

The reactant reduced is _______________, and the reducing agent is ________________.

b. The total number of electrons transferred in this reaction is __________.

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5. Consider the following reaction: $\mathbf{2\ H_2O_2}(aq) \rightarrow \mathbf{2\ H_2O}(l) + \mathbf{O_2}(aq)$
Indicate the ***neutral element*** or the ***ion (or element with its oxidation number) in a compound*** that is oxidized or reduced and the ***reactants*** that served as the oxidizing and the reducing agents below:

a. The reactant oxidized is _______________, and the oxidizing agent is ________________.

The reactant reduced is _______________, and the reducing agent is ________________.

b. The total number of electrons transferred in this reaction is __________.

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6. Consider the following reaction: $\mathbf{3\ HNO_2}(aq) \rightarrow \mathbf{H_2O}(l) + \mathbf{HNO_3}(aq) + \mathbf{2NO}(g)$
Indicate the ***neutral element*** or the ***ion (or element with its oxidation number) in a compound*** that is oxidized or reduced and the ***reactants*** that served as the oxidizing and the reducing agents below:

a. The reactant oxidized is _______________, and the oxidizing agent is ________________.

The reactant reduced is _______________, and the reducing agent is ________________.

b. The total number of electrons transferred in this reaction is __________.

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7. Consider the following reaction: $\mathbf{Zn + 2MnO_2 + 2NH_4Cl \rightarrow ZnCl_2 + Mn_2O_3 + 2NH_3 + H_2O}$
Indicate the ***neutral element*** or the ***ion (or element with its oxidation number) in a compound*** that is oxidized or reduced and the ***reactants*** that served as the oxidizing and the reducing agents below:

a. The reactant oxidized is _______________, and the oxidizing agent is ________________.

The reactant reduced is _______________, and the reducing agent is ________________.

b. The total number of electrons transferred in this reaction is __________.

Dougherty Valley HS Chemistry
Oxidation and Reduction Practice

WORKSHEET #2

In each of the following equations, indicate the element that has been oxidized and the one that has been reduced. You should also label the oxidation state of each before and after the process:

1) $2\ Na + FeCl_2 \rightarrow 2\ NaCl + Fe$

2) $2\ C_2H_2 + 5\ O_2 \rightarrow 4\ CO_2 + 2\ H_2O$

3) $2\ PbS + 3\ O_2 \rightarrow 2\ SO_2 + 2\ PbO$

4) $2\ H_2 + O_2 \rightarrow 2\ H_2O$

5) $Cu + HNO_3 \rightarrow CuNO_3 + H_2$

6) $AgNO_3 + Cu \rightarrow CuNO_3 + Ag$

Dougherty Valley HS Chemistry
Balancing Redox reactions

WORKSHEET #3

Complete the following reactions on a separate sheet of paper, showing all steps involved, neatly and clearly.

For Half-Reactions in Acidic Solution

Step One: Balance the atom being reduced/oxidized. In our example, there is already one Mn on each side of the arrow, so this step is already done.

$$MnO_4^- \text{ ---> } Mn^{2+}$$

Step Two: Balance the oxygen's. Do this by adding water molecules (as many as are needed) to the side needing oxygen. In our case, the left side has 4 oxygen's, while the right side has none, so:

$$MnO_4^- \text{ ---> } Mn^{2+} + 4H_2O$$

Notice that, when the water is added, hydrogen's also come along. There is nothing that can be done about this; we'll take care of it in the next step. A common question is: "Why can't I just add 4 oxygen atoms to the right side?" Quick answer: don't do it, it's wrong. The "why" will be left to another day.

Step Three: Balance the hydrogen's. Do this by adding hydrogen ions (as many as are needed) to the side needing hydrogen. In our example, we need 8 (notice the water molecule's formula, then consider 4 x 2 = 8).

$$8H^+ + MnO_4^- \text{ ---> } Mn^{2+} + 4H_2O$$

Step Four: Balance the total charge. This will be done using electrons. It is ALWAYS the last step.

First, a comment. You do not need to look at the oxidation number for each atom. You only need to look at the charge on the ion or molecule, then sum those up.

Left side of the reaction, total charge is +7. There are 8 H^+, giving 8 x +1 = +8 and a minus one from the permanganate. (A very typical wrong answer for the left side is zero. The person sees only the +1 and the -1, they forget the 8. When you do this step in the parallel example, don't forget to multiply 2 times 3. I'll leave you to figure out where in the problem that is.)

Right side of the reaction, total charge is +2. The water molecule is neutral (zero charge) and the single Mn is +2.

$$5e^- + 8H^+ + MnO_4^- \text{ ---> } Mn^{2+} + 4H_2O$$

Five electrons reduces' the +7 to a +2 and the two sides are EQUAL in total charge. The half-reaction is now correctly balanced.

For Half-Reactions in Basic Solution

Step One to Four: Balance the half-reaction AS IF it were in acid solution. I hope you got that. The half-reaction is actually in basic solution, but we are going to start out as if it were in acid solution. Here are the 4 acid steps:

1) Balance the atom being reduced/oxidized.
2) Balance the oxygens.
3) Balance the hydrogens.
4) Balance the charge.

When you do that to the above half-reaction, you get:

$$2e^- + 2H^+ + PbO_2 \text{ ---> } PbO + H_2O$$

Step Five: Convert all H^+ to H_2O. Do this by adding OH^- ions to both sides. The side with the H^+ will determine how many hydroxide to add. In our case, the left side has 2 hydrogen ions, while the right side has none, so:

$$2e^- + 2H_2O + PbO_2 \text{ ---> } PbO + H_2O + 2OH^-$$

Notice that, when the two hydroxide ions on the left were added, they immediately reacted with the hydrogen ion present. The reaction is:

$$H^+ + OH^- \text{ ---> } H_2O$$

Step Six: Remove any duplicate molecules or ions. In our example, there are two water molecules on the left and one on the right. This means one water molecule may be removed from each side, giving:

$$2e^- + H_2O + PbO_2 \text{ ---> } PbO + 2OH^-$$

The half-reaction is now correctly balanced.

By the way, notice the $2OH^-$. Be careful to read that as two hydroxide ions ($2\ OH^-$) and NOT twenty hydride ions ($20\ H^-$). People have been known to do that.

Balance each half-reaction, the reaction being in acidic solution.

[1] $Re \rightarrow ReO_2$

$2H_2O + Re \rightarrow ReO_2 + 4H^+ + 4e^-$

[2] $Cl_2 \rightarrow HClO$

$2H_2O + Cl_2 \rightarrow 2HClO + 2H^+ + 2e^-$

[3] $NO_3^- \rightarrow HNO_2$

$2e^- + 3H^+ + NO_3^- \rightarrow HNO_2 + H_2O$

[4] $H_2GeO_3 \rightarrow Ge$

$4e^- + 4H^+ + H_2GeO_3 \rightarrow Ge + 3H_2O$

[5] $H_2SeO_3 \rightarrow SeO_4^{2-}$

[6] $Au \rightarrow Au(OH)_3$ (this one is a bit odd!)

[7] $H_3AsO_4 \rightarrow AsH_3$

[8] $H_2MoO_4 \rightarrow Mo$

[9] $NO \rightarrow NO_3^-$

Balance each half-reaction, the reaction being in basic solution.

[10] $NiO_2 \rightarrow Ni(OH)_2$

$2e^- + 2H_2O + NiO_2 \rightarrow Ni(OH)_2 + 2OH^-$

[11] $BrO_4^- \rightarrow Br^-$

$8e^- + 4H_2O + BrO_4^- \rightarrow Br^- + 8OH^-$

[12] $SbO_3^- \rightarrow SbO_2^-$

$2e^- + H_2O + SbO_3^- \rightarrow SbO_2^- + 2OH^-$

[13] $Cu_2O \rightarrow Cu$

$2e^- + H_2O + Cu_2O \rightarrow 2Cu + 2OH^-$

[14] $S_2O_3^{2-} \rightarrow SO_3^{2-}$

[15] $Tl^+ \rightarrow Tl_2O_3$

[16] $Al \rightarrow AlO_2^-$

[17] $Sn \rightarrow HSnO_2^-$

[18] $CrO_4^{2-} \rightarrow Cr(OH)_3$

Balance the following Redox reactions in Acidic Solution

[19] $ClO_3^- + SO_2 \rightarrow SO_4^{2-} + Cl^-$

$ClO_3^- + 3H_2O + 3SO_2 \rightarrow 3SO_4^{2-} + Cl^- + 6H^+$

[20] $H_2S + NO_3^- \rightarrow S_8 + NO$

$24H_2S + 16H^+ + 16NO_3^- \rightarrow 3S_8 + 16NO + 32H_2O$

[21] $MnO_4^- + H_2S \rightarrow Mn^{2+} + S_8$

$40H_2S + 48H^+ + 16MnO_4^- \rightarrow 5S_8 + 16Mn^{2+} + 64H_2O$

[22] $Cu + SO_4^{2-} \rightarrow Cu^{2+} + SO_2$

$Cu + 4H^+ + SO_4^{2-} \rightarrow Cu^{2+} + SO_2 + 2H_2O$

[23] $MnO_4^- + CH_3OH \rightarrow CH_3COOH + Mn^{2+}$

[24] $Cr_2O_7^{2-} + Fe^{2+} \rightarrow Cr^{3+} + Fe^{3+}$

[25] $HNO_2 \rightarrow NO + NO_2$

[26] $H_2C_2O_4 + MnO_4^- \rightarrow CO_2 + Mn^{2+}$

[27] $O_2 + As \rightarrow HAsO_2 + H_2O$

[28] $NO_2 \rightarrow NO_3^- + NO$

[29] $ClO_4^- + Cl^- \rightarrow ClO^- + Cl_2$

[30] $H_5IO_6 + Cr \rightarrow IO_3^- + Cr^{3+}$

[31] $Fe + HCl \rightarrow HFeCl_4 + H_2$

[32] $NO_3^- + H_2O_2 \rightarrow NO + O_2$

[33] $BrO_3^- + Fe^{2+} \rightarrow Br^- + Fe^{3+}$

[34] $Cr_2O_7^{2-} + C_2H_4O \rightarrow CH_3COOH + Cr^{3+}$

[35] $MnO_4^{-} + C_2H_4O \rightarrow CH_3COOH + MnO_2$

[36] $Zn + NO_3^{-} \rightarrow NH_4^{+} + Zn^{2+}$

[37] $HBr + SO_4^{2-} \rightarrow SO_2 + Br_2$

[38] $NO_3^{-} + I_2 \rightarrow IO_3^{-} + NO_2$

[39] $CuS + NO_3^{-} \rightarrow NO + Cu^{2+} + HSO_4^{-}$

Balance the following Redox Reactions in Basic Solution

[40] $NH_3 + ClO^{-} \rightarrow N_2H_4 + Cl^{-}$

$2HN_3 + ClO^{-} \rightarrow N_2H_4 + Cl^{-} + H_2O$

[41] $Au + O_2 + CN^{-} \rightarrow Au(CN)_2^{-} + H_2O_2$

$4CN^{-} + 2Au + 2H_2O + O_2 \rightarrow 2Au(CN)_2^{-} + H_2O_2 + 2OH^{-}$

[42] $Br^{-} + MnO_4^{-} \rightarrow MnO_2 + BrO_3^{-}$

[43] $AlH_4^{-} + H_2CO \rightarrow Al^{3+} + CH_3OH$

$AlH_4^{-} + 4H_2O + 4H_2CO \rightarrow Al^{3+} + 4CH_3OH + 4OH^{-}$

[44] $Se + Cr(OH)_3 \rightarrow Cr + SeO_3^{2-}$

$6OH^{-} + 3Se + 4Cr(OH)_3 \rightarrow 4Cr + 3SeO_3^{2-} + 9H_2O$

[45] $H_2O_2 + Cl_2O_7 \rightarrow ClO_2^{-} + O_2$

[46] $Fe + NiO_2 \rightarrow Fe(OH)_2 + Ni(OH)_2$

[47] $MnO_4^{-} + H_2O_2 \rightarrow MnO_2 + O_2$

[48] $Zn + BrO_4^{-} \rightarrow [Zn(OH)_4]^{2-} + Br^{-}$

[49] $MnO_4^{-} + S^{2-} \rightarrow MnO_2 + S_8$

[50] $Pb(OH)_4^{2-} + ClO^{-} \rightarrow PbO_2 + Cl^{-}$

[51] $Tl_2O_3 + NH_2OH \rightarrow TlOH + N_2$

[52] $Fe(OH)_2 + CrO_4^{2-} \rightarrow Fe_2O_3 + Cr(OH)_4^{-}$

[53] $Br_2 \rightarrow Br^{-} + BrO_3^{-}$

[54] $ClO_2^{-} + H_2O \rightarrow ClO_2^{-} + OH^{-}$

Dougherty Valley HS Chemistry
Redox Reactions More

WORKSHEET #4

Redox Reactions in <u>Acidic</u> Solution:

1. I^- (aq) + ClO^- (aq) $\longrightarrow$ I_3^- (aq) + Cl^- (aq)

2. As_2O_3 (s) + NO_3^- (aq) $\longrightarrow$ H_3AsO_4 (aq) + NO (g)

3. Br^- (aq) + MnO_4^- (aq) $\longrightarrow$ Br_2 (l) + Mn^{2+} (aq)

4. CH_3OH (aq) + $Cr_2O_7^{2-}$ (aq) $\longrightarrow$ CH_2O(l) + Cr^{3+}(aq)

5. Mn^{2+}(aq) + BiO_3^- (aq) $\longrightarrow$ Bi^{3+}(aq) + MnO_4^- (aq)

6. S_8(s) + NO_3^- (aq) $\longrightarrow$ SO_3^{2-} (aq) + NO(g)

7. H_3AsO_4(aq) + Zn(s) $\longrightarrow$ AsH_3(g) + Zn^{2+}(aq)

8. P_4(s) + $Cr_2O_7^{2-}$ (aq) $\longrightarrow$ H_3PO_4(aq) + Cr^{3+}(aq)

Redox Reactions in <u>Basic</u> Solution

1. Al(s) + MnO_4^- (aq) $\longrightarrow$ MnO_2(s) + $Al(OH)_4^-$ (aq)

2. NO_2^- (aq) + Al(s) $\longrightarrow$ NH_3(aq) + AlO_2^- (aq)

3. Cr(s) + CrO_4^{2-} (aq) $\longrightarrow$ $Cr(OH)_3$(s)
   Note: $Cr(OH)_3$ is found in BOTH half reactions!

4. MnO_4^- (aq) + S^{2-} (aq) $\longrightarrow$ MnO_2(s) + SO_3^{2-} (aq)

5. Cl_2(aq) + Br_2(l) $\longrightarrow$ OBr^- (aq) + Cl^- (aq)

6. H_2O_2(aq) + I^- (aq) $\longrightarrow$ IO_3^- (aq)
   Note: IO_3^- is found in both half reactons!

7. NO_3^- (aq) + NH_3(aq) $\longrightarrow$ NO_2^- (aq)

8. S_8(aq) + MnO_4^- (aq) $\longrightarrow$ SO_4^{2-} (aq) + MnO_2(s)