Dougherty Valley HS AP Chemistry
Electrochemistry – Oxidation and Reduction

WORKSHEET #1 KEY

Name: **Date:** **Period:** **Seat #:**

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Assigning oxidation numbers:

Determine the oxidation number of the underlined element.

1.	$\underline{Be}Cl_2$	2.	$\underline{N}O$	3.	$Na_2\underline{S}O_3$	4.	$H_2\underline{O}_2$
5.	$Ag\underline{Br}$	6.	$\underline{Au}Cl_3$	7.	$H\underline{N}O_3$	8.	$H_2\underline{Sn}O_3$
9.	$\underline{S}O_3$	10.	$\underline{U}F_6$	11.	$Ba\underline{Cr}O_4$	12.	$Ca\underline{Se}O_4$
13.	$H\underline{I}$	14.	$H_2\underline{Se}$	15.	$K_2\underline{Pt}Cl_6$	16.	$\underline{Ni}SO_4$
17.	$\underline{N}H_3$	18.	$H\underline{Cl}O$	19.	$\underline{N}H_4Cl$	20.	$(NH_4)_2\underline{Te}$

1.	+2	2.	+2	3.	+4	4.	-1
5.	-1	6.	+3	7.	+5	8.	+4
9.	+6	10.	+6	11.	+6	12.	+6
13.	-1	14.	+2	15.	+4	16.	+2
17.	-3	18.	+1	19.	-3	20.	-2

The Half-Reaction Method:

1. Write the equation as two half-reactions. Include the particles (atoms, ions, molecules) that are involved in change of oxidation state.
2. Balance each half-reaction with respect to atoms and charges; first atoms other than H and O, then O with H_2O and H with H^+, and ionic charges with electrons (e^-).
3. Equalize the number of electrons lost in the oxidation half-reaction with the number of electrons gained in the reduction half-reaction.
4. Add the two half-reactions to form a balanced net ionic equation.

[a]

$$HCl + K_2Cr_2O_7 \rightarrow KCl + CrCl_3 + Cl_2$$

14, 1, 2, 2, 3, $7H_2O$

Reduction half-reaction	Oxidation half-reaction
$14H^+ + K_2Cr_2O_7 + 6e^- \rightarrow 2K^+ + 2Cr^{3+} + 7H_2O$	$6HCl \rightarrow 3Cl_2 + 6H^+ + 6e^-$

Overall:

$$14HCl + K_2Cr_2O_7 \rightarrow 2KCl + 2CrCl_3 + 3Cl_2 + 7H_2O$$

[b]

$$FeCl_2 + KMnO_4 + HCl \rightarrow FeCl_3 + KCl + MnCl_2 + H_2O$$

5, 1, 8, 5, 1, 1, $4H_2O$

Reduction half-reaction	Oxidation half-reaction
$5Fe^{2+} \rightarrow 5Fe^{3+} + 5e^-$	$KMnO_4 + 8H^+ + 5e^- \rightarrow Mn^{2+} + K^+ + 4H_2O$

Overall: note → Chlorines are omitted in half reactions b/c their oxidation state does not change

$$5FeCl_2 + KMnO_4 + 8HCl \rightarrow F5eCl_3 + KCl + MnCl_2 + 4H_2O$$

[c]

$CuS + NO_3^- \rightarrow Cu^{2+} + S + NO$

3, 2, $8H^+$, 3, 3, 2, $4H_2O$

Reduction half-reaction	Oxidation half-reaction
$3CuS \rightarrow 3Cu^{2+} + 3S + 6e^-$	$8H^+ + 2NO_3^- + 6e^- \rightarrow 2NO + 4H_2O$

Overall:

$3CuS + 2NO_3^- + 8H^+ \rightarrow 3Cu^{2+} + 3S + 2NO + 4H_2O$

Balance the following redox reactions in acidic solutions:

[d] $HNO_3 + S \rightarrow NO_2 + H_2SO_4 + H_2O$

6$HNO_3 + S \rightarrow$ **6**$NO_2 + H_2SO_4 +$ **2**H_2O

[e] $KMnO_4 + HCl + H_2S \rightarrow KCl + MnCl_2 + S$

2$KMnO_4 +$ **6**$HCl +$ **5**$H_2S \rightarrow$ **2**$KCl +$ **2**$MnCl_2 +$ **5**$S +$ **8**H_2O

[f] $FeCl_3 + H_2S \rightarrow FeCl_2 + HCl + S$

2$FeCl_3 + H_2S \rightarrow$ **2**$FeCl_2 +$ **2**$HCl + S$

[g] $Cu + HNO_3 \rightarrow Cu(NO_3)_2 + NO_2$

$Cu +$ **4**$HNO_3 \rightarrow Cu(NO_3)_2 +$ **2**$NO_2 +$ **2**H_2O