Chem 151 Pre-lab Practice 3A

Experiment 3: REDOX OR NOT?

(see next page for answers)

This pre-lab provides you with six examples of how to balance a redox reaction equation. Listed in the box is an outline of the procedure. Try to balance the reactions given in the six examples and check the answers (with explanations) on the following pages:

The procedure:

1. ***Identify which species are oxidized and reduced using the change in their oxidation numbers.***
2. ***Write the oxidation and reduction half-reaction equations. For each half-reaction,***
   a. ***Balance the elements except for O and H.***
   b. ***Balance the electrons gained or lost based on the changes in oxidation number.***
   c. ***Balance the net charges by adding H^+ (when reaction takes place in an acidic medium) or OH^- (when reaction takes place in a basic medium) to either side of the equation.***
   d. ***Balance O and H by adding H_2O to either side of the equation.***
3. ***Make the number of electrons gained form the reduction half reaction equal to the electrons lost in the oxidation half-reaction. This can be achieved by multiplying one or both of the half-reaction equations by a factor.***
4. ***Add the balanced oxidation and reduction half-reactions together to obtain the balanced redox equation.***

Example 1. $NH_4^+ + MnO_4^-$ ® $NO_3^- + Mn^{2+}$ **in acidic solution**

Example 2. $CuS + NO_3^-$ ® $Cu^{2+} + HSO_4^- + NO$ **in acidic solution**

Example 3. $V^{2+} + H_2O_2$ ® $H_4VO_4^+ + H_2O$ **in acidic solution**

Example 4. $PbO_2 + Cl^-$ ® $ClO^- + Pb(OH)_3^-$ **in basic solution**

Example 5. $S^{2-} + Cl_2$ ® $SO_4^{2-} + Cl^-$ **in basic solution**

Example 6. $ClO^- + Cr(OH)_4^-$ ® $Cl^- + CrO_4^{2-}$ **in basic solution**

Example 1. **NH_4^+ + MnO_4^- ® NO_3^- + Mn^{2+} in acidic solution**

Step 1 Assign oxidation numbers to the elements involved in redox reaction

$$NH_4^+ + MnO_4^- \rightarrow NO_3^- + Mn^{2+}$$

-3 **+7** **+5** **+2**

N in NH_4^+ loses 8 electrons to become NO_3^- (oxidation)
Mn in MnO_4^- gains 5 electrons to become Mn^{2+} (reduction)

Step 2 The unbalanced half-reaction equations are:

Oxidation: $NH_4^+ \rightarrow NO_3^- + 8\,e^-$
Reduction: $MnO_4^- + 5\,e^- \rightarrow Mn^{2+}$
} The elements except for O are balanced.

To balance the oxidation half-reaction equation $NH_4^+ \rightarrow NO_3^- + 8\,e^-$
the net charge on the left is +1, but on the right is -9.
Add 10 H^+ to the right to balance the net charge: $NH_4^+ \rightarrow NO_3^- + 8\,e^- + 10\,H^+$
Add 3 H_2O to the left to balance the H and O atoms: $3\,H_2O + NH_4^+ \rightarrow NO_3^- + 8\,e^- + 10\,H^+$

To balance the reduction half-reaction equation $MnO_4^- + 5\,e^- \rightarrow Mn^{2+}$
the net charge on the left is -6, but on the right is +2.
Add 8 H^+ to the left to balance the net charge: $8\,H^+ + MnO_4^- + 5\,e^- \rightarrow Mn^{2+}$
Add 4 H_2O to the right to balance the H and O atoms: $8\,H^+ + MnO_4^- + 5\,e^- \rightarrow Mn^{2+} + 4\,H_2O$

Step 3 Balance the 5 e^- gained in the reduction with the 8 e^- lost in the oxidation:
Oxidation: $(3\,H_2O + NH_4^+ \rightarrow NO_3^- + 8\,e^- + 10\,H^+) \times 5$
Reduction: $(8\,H^+ + MnO_4^- + 5\,e^- \rightarrow Mn^{2+} + 4\,H_2O) \times 8$

Step 4. Add the two balanced half-reactions together:

Oxidation: $15\,H_2O + 5\,NH_4^+ \rightarrow 5\,NO_3^- + 40\,e^- + 50\,H^+$
Reduction: $64\,H^+ + 8\,MnO_4^- + 40\,e^- \rightarrow 8\,Mn^{2+} + 32\,H_2O$ (+

Overall **$5\,NH_4^+ + 8\,MnO_4^- + 14\,H^+$ ® $5\,NO_3^- + 8\,Mn^{2+} + 17\,H_2O$**

Example 2. **CuS + NO_3^- ® Cu^{2+} + HSO_4^- + NO in acidic solution**

Step 1 Assign oxidation numbers to the elements involved in redox reaction

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

-2 **+5** **+6** **+2**

S in CuS loses 8 electrons to become HSO_4^- (oxidation)
N in NO_3^- gains 3 electrons to become NO (reduction)

Step 2 Balance the half reaction equations:

Oxidation:
$CuS \rightarrow Cu^{2+} + HSO_4^- + 8\,e^-$
$\Downarrow$
$CuS \rightarrow Cu^{2+} + HSO_4^- + 8\,e^- + 7\,H^+$
$\Downarrow$
$4\,H_2O + CuS \rightarrow Cu^{2+} + HSO_4^- + 8\,e^- + 7\,H^+$
$\Downarrow$

Reduction:
$NO_3^- + 3\,e^- \rightarrow NO$
$\Downarrow$
$4\,H^+ + NO_3^- + 3\,e^- \rightarrow NO$
$\Downarrow$
$4\,H^+ + NO_3^- + 3\,e^- \rightarrow NO + 2H_2O$
$\Downarrow$

Step 3 $(4\,H_2O + CuS \rightarrow Cu^{2+} + HSO_4^- + 8\,e^- + 7\,H^+) \times 3$ $(4\,H^+ + NO_3^- + 3\,e^- \rightarrow NO + 2H_2O) \times 8$

Step 4 The overall balanced equation is: **$3\,CuS + 11\,H^+ + 8\,NO_3^-$ ® $3\,Cu^{2+} + 3\,HSO_4^- + 8\,NO + 4\,H_2O$**

Example 3. V^{2+} + H_2O_2 ® $H_4VO_4^+$ **in acidic solution**

Step 1 Assign oxidation numbers to the elements involved in redox reaction

V^{2+} + H_2O_2 $\rightarrow$ $H_4VO_4^+$ + H_2O

+2 **-1** **+5** **-2**

V^{2+} loses 3 electrons to become $H_4VO_4^+$ (oxidation)

O in H_2O_2 gains one electron to become H_2O (reduction)*

Step 2 Balance the half reaction equations:

Oxidation: $V^{2+} \rightarrow H_4VO_4^+ + 3\,e^-$

$\Downarrow$

$V^{2+} \rightarrow H_4VO_4^+ + 3\,e^- + 4\,H^+$

$\Downarrow$

$4\,H_2O + V^{2+} \rightarrow H_4VO_4^+ + 3\,e^- + 4\,H^+$

$\Downarrow$

Reduction: $H_2O_2 + 2\,e^- \rightarrow 2\,H_2O$

$\Downarrow$

$2\,H^+ + H_2O_2 + 2\,e^- \rightarrow 2\,H_2O$

$\Downarrow$

$2\,H^+ + H_2O_2 + 2\,e^- \rightarrow 2\,H_2O$

$\Downarrow$

Step 3 $(4\,H_2O + V^{2+} \rightarrow H_4VO_4^+ + 3\,e^- + 4\,H^+) \times 2$ $(2\,H^+ + H_2O_2 + 2\,e^- \rightarrow 2\,H_2O) \times 3$

Step 4 The overall balanced equation is: **$2\,V^{2+} + 3\,H_2O_2 + 2\,H_2O$ ® $2\,H_4VO_4^+ + 2\,H^+$**

*** Since there are two oxygen atoms in a H_2O_2 molecule, each H_2O_2 gains two electrons to become 2 H_2O.**

Example 4. PbO_2 + Cl^- ® ClO^- + $Pb(OH)_3^-$ **in basic solution**

Step 1 Assign oxidation numbers to the elements involved in redox reaction

PbO_2 + Cl^- $\rightarrow$ ClO^- + $Pb(OH)_3^-$

+4 **-1** **+1** **+2**

Cl^- loses 2 electrons to become ClO^- (oxidation)

Pb in PbO_2 gains 2 electrons to become $Pb(OH)_3^-$ (reduction)

Step 2 Balance the half reaction equations:

Oxidation: $Cl^- \rightarrow ClO^- + 2\,e^-$

$\Downarrow$

$2\,OH^- + Cl^- \rightarrow ClO^- + 2\,e^-$

$\Downarrow$

$2\,OH^- + Cl^- \rightarrow ClO^- + 2\,e^- + H_2O$

Reduction: $PbO_2 + 2\,e^- \rightarrow Pb(OH)_3^-$

$\Downarrow$

$PbO_2 + 2\,e^- \rightarrow Pb(OH)_3^- + OH^-$

$\Downarrow$

$2\,H_2O + PbO_2 + 2\,e^- \rightarrow Pb(OH)_3^- + OH^-$

Step 3 This step can be skipped because the number of electrons gained in the oxidation reaction is same as the number of electrons lost in the reduction reaction.

Step 4 The overall balanced equation is: **$Cl^- + PbO_2 + H_2O + OH^-$ ® $ClO^- + Pb(OH)_3^-$**

Example 5. **S^{2-} + Cl_2 ® SO_4^{2-} + Cl^- in basic solution**

Step 1 Assign oxidation numbers to the elements involved in redox reaction

$S^{2-} + Cl_2 \rightarrow SO_4^{2-} + Cl^-$

-2 **0** **+6** **-1**

S^{2-} loses 8 electrons to become SO_4^{2-} (oxidation)
Each of the Cl in Cl_2 gains one electron to become Cl^- (reduction)

Step 2 Balance the half reaction equations:

Oxidation: $S^{2-} \rightarrow SO_4^{2-} + 8\,e^-$ Reduction: $Cl_2 + 2\,e^- \rightarrow 2\,Cl^-$ (This equation is balanced)

$\Downarrow$ $\Downarrow$

$8\,OH^- + S^{2-} \rightarrow SO_4^{2-} + 8\,e^-$

$\Downarrow$ $\Downarrow$

$8\,OH^- + S^{2-} \rightarrow SO_4^{2-} + 8\,e^- + 4\,H_2O$

$\Downarrow$ $\Downarrow$

Step 3 $(8\,OH^- + S^{2-} \rightarrow SO_4^{2-} + 8\,e^- + 4\,H_2O) \times 1$ $(Cl_2 + 2\,e^- \rightarrow 2\,Cl^-) \times 4$

Step 4 The overall balanced equation is: **$S^{2-} + 4\,Cl_2 + 8\,OH^-$ ® $SO_4^{2-} + 8\,Cl^- + 4\,H_2O$**

Example 6. **$ClO^- + Cr(OH)_4^-$ ® Cl^- + CrO_4^{2-} in basic solution**

Step 1 Assign oxidation numbers to the elements involved in redox reaction

$ClO^- + Cr(OH)_4^- \rightarrow Cl^- + CrO_4^{2-}$

+1 **+3** **-1** **+6**

Cr in $Cr(OH)_4^-$ loses 3 electrons to become CrO_4^{2-} (oxidation)
Cl in ClO^- gains 2 electrons to become Cl^- (reduction)

Step 2 Balance the half reaction equations:

Oxidation: $Cr(OH)_4^- \rightarrow CrO_4^{2-} + 3\,e^-$ Reduction: $ClO^- + 2\,e^- \rightarrow Cl^-$

$\Downarrow$ $\Downarrow$

$4\,OH^- + Cr(OH)_4^- \rightarrow CrO_4^{2-} + 3\,e^-$ $ClO^- + 2\,e^- \rightarrow Cl^- + 2\,OH^-$

$\Downarrow$ $\Downarrow$

$4\,OH^- + Cr(OH)_4^- \rightarrow CrO_4^{2-} + 3\,e^- + 4\,H_2O$ $H_2O + ClO^- + 2\,e^- \rightarrow Cl^- + 2\,OH^-$

$\Downarrow$ $\Downarrow$

Step 3 $(4\,OH^- + Cr(OH)_4^- \rightarrow CrO_4^{2-} + 3\,e^- + 4\,H_2O) \times 2$ $(H_2O + ClO^- + 2\,e^- \rightarrow Cl^- + 2\,OH^-) \times 3$

Step 4 The overall balanced equation is: 3 **$ClO^- + 2\,Cr(OH)_4^- + 2\,OH^-$ ® $3\,Cl^- + 2\,CrO_4^{2-} + 5\,H_2O$**