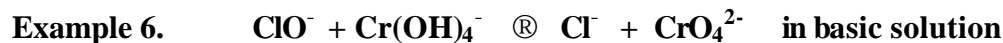
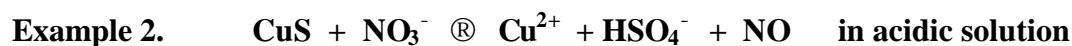


Chem 151 Pre-lab Practice 3A
Experiment 3: REDOX OR NOT?
(see next page for answers)

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:

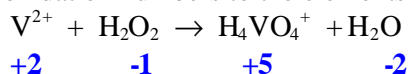
The procedure:

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 from 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 3. $V^{2+} + H_2O_2 \text{ ® } H_4VO_4^+$ in acidic solution

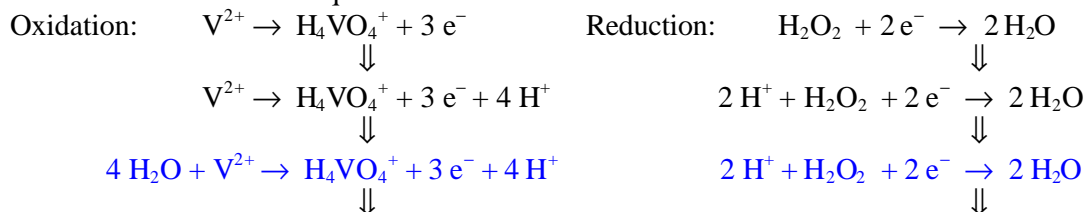
Step 1 Assign oxidation numbers to the elements involved in redox reaction



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:



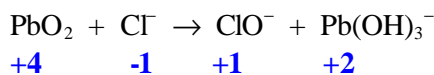
Step 3 $(4 H_2O + V^{2+} \rightarrow H_4VO_4^+ + 3 e^- + 4 H^+) \times 2$ $(2 H^+ + H_2O_2 + 2 e^- \rightarrow 2 H_2O) \times 3$

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

* 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. $PbO_2 + Cl^- \text{ ® } ClO^- + Pb(OH)_3^-$ in basic solution

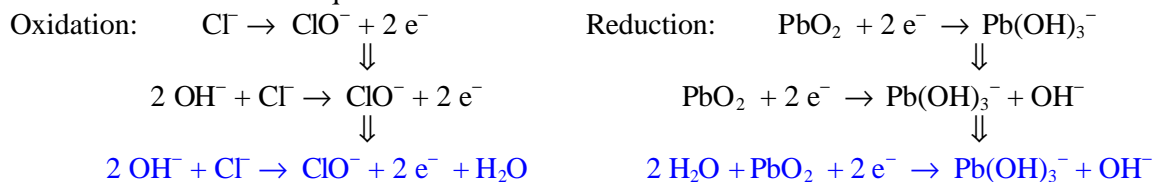
Step 1 Assign oxidation numbers to the elements involved in redox reaction



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:

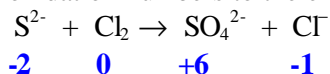


Step 3 This step can be skipped because the number of electrons gained in the oxidation reaction is same as the number of electrons lost in the reduction reaction.

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

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

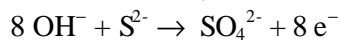
Step 1 Assign oxidation numbers to the elements involved in redox reaction



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:

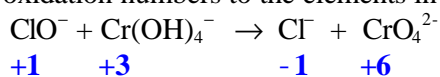


Step 3 $(8 OH^- + S^{2-} \rightarrow SO_4^{2-} + 8 e^- + 4 H_2O) \times 1$ $(Cl_2 + 2 e^- \rightarrow 2 Cl^-) \times 4$

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

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

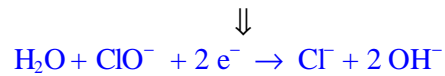
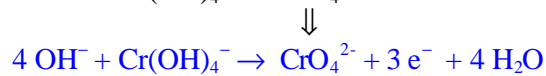
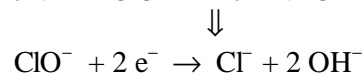
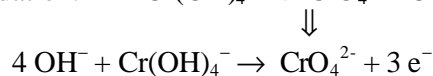
Step 1 Assign oxidation numbers to the elements involved in redox reaction



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:



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: $3 ClO^- + 2 Cr(OH)_4^- + 2 OH^- \rightarrow 3 Cl^- + 2 CrO_4^{2-} + 5 H_2O$