5.1 Recognize a redox reaction by assigning oxidation numbers. [Reading 5.1 Problems 3, 5, 6, 26, 28, 30, 32, & 34]
Oxidation-Reduction (Redox) Reactions

- These are reactions in which there is a transfer of electrons
- LEO - loss of electrons is oxidation
- GER - gain of electrons is reduction
- Oxidation and reduction must occur together

Redox Reactions

Oxidizing agent - the electron accepting substance
The oxidizing agent is REDUCED
Reducing agent - electron donating substance
The reducing agent is OXIDIZED

Oxidation-Reduction Reactions

Clorox bleach contains a strong oxidizing agent, NaClO, which oxidizes this red dye to a colorless product.
Oxidation Numbers

• A convenient method of bookkeeping for electrons
• Oxidation numbers do not imply ionic charge on an atom
• A change in the oxidation number implies a transfer of electrons
• Redox reactions can be recognized by the change in the oxidation numbers

Assigning Oxidation Numbers

1. An atom in its elemental state has an oxidation number of 0
2. A monatomic ion has an oxidation number equal to its charge
3. The sum of the oxidation numbers of the atoms must equal the net charge for a molecule or polyatomic ion.

Assigning Oxidation Numbers

4. Fluorine has an ox. # of -1 in compounds
5. Hydrogen has an ox. # number of +1 in compounds
6. Oxygen has an ox. # of -2 in compounds
7. Halogens usually have an oxidation number of -1
   – Except when Cl, Br, or I is bonded to oxygen.
Assign Oxidation Numbers

- Fe
- O₂
- CO₂
- MnO₄⁻
- Fe₂O₃
- Na₂Cr₂O₇

When conflicts occur

When conflicts in the rules occur, assign the first rule and ignore the later, conflicting, rule.

Oxidation Numbers

Molybdenum disulfide, MnS₂, is a black powder used as a lubricant.

Hydrogen peroxide, H₂O₂, shown as a 3% aqueous H₂O₂ is a strong oxidizing agent.

What are the oxidation numbers of each element in these two compounds?
5.2 Determine which molecule is oxidized and which is reduced in a chemical reaction. [Reading 5.1 Problems 24, & 25]

Oxidation Numbers in a Reaction

- Identify which species is oxidized and which is reduced in the following reaction.
- \[ 2 \text{Fe}_2\text{O}_3(s) + 3\text{C}(s) \rightarrow 4\text{Fe}(s) + 3\text{CO}_2(g) \]

\[
\begin{array}{cccc}
\text{Fe}_2\text{O}_3 & \text{C} & \text{Fe} & \text{CO}_2 \\
-2 & 0 & 0 & -2 \\
x \cdot 3 & x \cdot 2 & & \\
-6 & & -4 & \\
x = +3 & x = +4 \\
\end{array}
\]
Oxidation Numbers in a Reaction

- Identify which species is oxidized and which is reduced in the following reaction.
- \( 2 \text{Fe}_2\text{O}_3(s) + 3 \text{C}(s) \rightarrow 4 \text{Fe}(s) + 3 \text{CO}_2(g) \)
  +3  -2  0  0  +4 -2
- Iron gains electrons, thus, is reduced
- C loses electrons, thus, is oxidized
- Oxidation and reduction must occur together
5.3 Balance Redox Equations by the Ion-Electron Method
[Reading 5.2 Problems 8, 36, 38, 40, 42, & 44]

Balancing Redox Reactions
1. Write the two unbalanced half reactions separately
2. Balance each half reaction for all atoms other than H and O
3. Balance both half-reactions for oxygen by adding H₂O to the side with less O

Balancing Redox Reactions
4. Balance for H by adding H⁺ to the side with less H
5. Balance the charge in both half-reactions by adding electrons to the side of the equation which is more positive
6. Multiply each half reaction by a scalar to make the number of electrons gained in one half reaction equal to the number lost in the other half reaction.
Balancing Redox Reactions

7. Add the two reactions together and cancel electrons and molecules which occur on both sides of the equation.
8. In a basic solution, add OH\(^-\) in the final step to neutralize the H\(^+\) ions.
9. Combine H\(^+\) with OH\(^-\) to form water
10. Cancel any water possible

Example problem

\[ \text{H}_2\text{C}_2\text{O}_4 + \text{MnO}_4^- \rightarrow \text{Mn}^{2+} + \text{CO}_2 \]

Example problem

\[ \text{H}_2\text{C}_2\text{O}_4 + \text{MnO}_4^- \rightarrow \text{Mn}^{2+} + \text{CO}_2 \]

1. Write two unbalanced half reactions
Example problem

H₂C₂O₄ + MnO₄⁻ → Mn²⁺ + CO₂

1) Unbalanced Half reactions:
   H₂C₂O₄ → ___ CO₂
   MnO₄⁻ → Mn²⁺

Example problem

2) Balance the Half reactions for elements other than H and O:
   H₂C₂O₄ → ___ CO₂
   MnO₄⁻ → Mn²⁺

Example problem

2) Balance the Half reactions:
   H₂C₂O₄ → 2 CO₂
   MnO₄⁻ → Mn²⁺
Example problem

3 & 4) Balance the Half reactions (O & H):

\[ \text{H}_2\text{C}_2\text{O}_4 \rightarrow 2\text{CO}_2 + \text{needs } 2 \text{H} \]

\[ \text{MnO}_4^- \rightarrow \text{Mn}^{2+} \]

Example problem

3 & 4) Balance the Half reactions (O & H):

\[ \text{H}_2\text{C}_2\text{O}_4 \rightarrow 2\text{CO}_2 + 2\text{H}^+ \]

\[ \text{MnO}_4^- \rightarrow \text{Mn}^{2+} \]

Example problem

3 & 4) Balance the Half reactions (O & H):

\[ \text{H}_2\text{C}_2\text{O}_4 \rightarrow 2\text{CO}_2 + 2\text{H}^+ \]

\[ \text{MnO}_4^- \rightarrow \text{Mn}^{2+} + (\text{needs } 4 \text{O atoms}) \]
Example problem

3 & 4) Balance the Half reactions (O & H):

\[ \text{H}_2\text{C}_2\text{O}_4 \rightarrow 2\text{ CO}_2 + 2 \text{ H}^+ \]
\[ \text{MnO}_4^- \rightarrow \text{Mn}^{2+} + 4 \text{ H}_2\text{O} \]

Example problem

3 & 4) Balance the Half reactions (H & O):

\[ \text{H}_2\text{C}_2\text{O}_4 \rightarrow 2\text{ CO}_2 + 2 \text{ H}^+ \]
needs 8H + \text{MnO}_4^- \rightarrow \text{Mn}^{2+} + 4 \text{ H}_2\text{O}
Example problem

5) Balance the Half reactions (Charge):

\[ \text{H}_2\text{C}_2\text{O}_4 \rightarrow 2\text{CO}_2 + 2\text{H}^+ \]
\[ 8\text{H}^+ + \text{MnO}_4^- \rightarrow \text{Mn}^{2+} + 4\text{H}_2\text{O} \]

Example problem

5) Balance the Half reactions (Charge):

\[ \text{H}_2\text{C}_2\text{O}_4 \rightarrow 2\text{CO}_2 + 2\text{H}^+ + 2\text{e}^- \]
\[ 5\text{e} + 8\text{H}^+ + \text{MnO}_4^- \rightarrow \text{Mn}^{2+} + 4\text{H}_2\text{O} \]

Example problem

6) Balance the Half reactions (Charge):

\[ 5[\text{H}_2\text{C}_2\text{O}_4 \rightarrow 2\text{CO}_2 + 2\text{H}^+ + 2\text{e}^-] \]
\[ 2[5\text{e} + 8\text{H}^+ + \text{MnO}_4^- \rightarrow \text{Mn}^{2+} + 4\text{H}_2\text{O}] \]
Example problem

6) Balance the Half reactions (Charge):

\[ 5\text{H}_2\text{C}_2\text{O}_4 \rightarrow 10 \text{CO}_2 + 10\text{H}^+ + 10 \text{e}^- \]
\[ 10 \text{e}^- + 16 \text{H}^+ + 2\text{MnO}_4^- \rightarrow 2 \text{Mn}^{2+} + 8\text{H}_2\text{O} \]

Example problem

6 & 7) Cancel and add the two reactions

\[ 5\text{H}_2\text{C}_2\text{O}_4 \rightarrow 10 \text{CO}_2 + 10\text{H}^+ +10 \text{e}^- \]
\[ 10 \text{e}^- + 16 \text{H}^+ + 2\text{MnO}_4^- \rightarrow 2 \text{Mn}^{2+} + 8\text{H}_2\text{O} \]

\[ 5\text{H}_2\text{C}_2\text{O}_4 + 6 \text{H}^+ + 2\text{MnO}_4^- \rightarrow 10 \text{CO}_2 + 2 \text{Mn}^{2+} + 8\text{H}_2\text{O} \]

Oxidation-Reduction Reactions

Complete and balance net ionic equation for reaction in acidic solution

\[ \text{MnO}_4^-(aq) + \text{Fe}^{2+}(aq) \rightarrow \]
\[ \text{Mn}^{2+}(aq) + \text{Fe}^{3+}(aq) \]

Colors of aqueous solutions:
- \( \text{MnO}_4^- \) - dark purple
- \( \text{Fe}^{2+} \) - pale blue-green
- \( \text{Mn}^{2+} \) - colorless
- \( \text{Fe}^{3+} \) - pale red-orange

Addition of \( \text{KMnO}_4(aq) \) to acidic solution of \( \text{FeSO}_4(aq) \).
Redox Practice

Balance the following reaction equation in a basic solution:

\[
\text{CrO}_2^- + \text{S}_2\text{O}_8^{2-} \rightarrow \text{CrO}_4^{2-} + \text{SO}_4^{2-}
\]