Chapter 12 Oxidation-Reduction Reactions

Solutions to Practice Problems

1.
Problem
Balance each of the following ionic equations for acidic conditions. Identify the oxidizing agent and the reducing agent in each case:
a) $MnO_4^-(aq) + Ag(s) \rightarrow Mn^{2+}(aq) + Ag^+(aq)$
b) $Hg(l) + NO_3^-(aq) + Cl^-(aq) \rightarrow HgCl_4^{2-}(s) + NO_2(g)$
c) $AsH_3(g) + Zn^{2+}(aq) \rightarrow H_3AsO_4(aq) + Zn(s)$
d) $I_2(s) + OCl^-(aq) \rightarrow IO_3^-(aq) + Cl^-(aq)$

What Is Required?
You must write a balanced net ionic equation for the given reactions that occur in acidic conditions and identify the oxidizing agent and the reducing agent in each case.

What Is Given?
You know the formulas of the reactants and products, and that the reactions take place in acidic solution.

Plan Your Strategy
Use the following the steps for balancing a net ionic equation by the half-reaction method in acidic conditions.

Step 1. Write unbalanced half-reactions that show the formulas of the given reactant(s) and product(s). Write only those compounds that contain the atom that is oxidized or the atom that is reduced.
Step 2. Balance any atoms other than oxygen and hydrogen first.
Step 3. Balance any oxygen atoms by adding water molecules.
Step 4. Balance any hydrogen atoms by adding hydrogen ions.
Step 5. Balance the charges by adding electrons.
Step 6. Determine the lowest common multiple (LCM) of the numbers of electrons in the oxidation and reduction half-reactions.
Step 7. Multiply one or both half-reactions by the number that will bring the number of electrons to the LCM and therefore an equal number of electrons will be lost and gained.
Step 8. Add the balanced half-reactions.
Step 9. Cancel the electrons and any other identical molecules or ions present on both sides of the equation.
Step 10. If spectator ions were removed when forming half-reactions, add them back to the equation.

Identify the oxidizing agent and the reducing agent from the half-reactions.

Act on Your Strategy
a) $MnO_4^-(aq) + Ag(s) \rightarrow Mn^{2+}(aq) + Ag^+(aq)$

Step 1. The two unbalanced half-reactions are
Oxidation: $Ag(s) \rightarrow Ag^{+}(aq)$
Reduction: $MnO_4^{-}(aq) \rightarrow Mn^{2+}(aq)$

Step 2. Atoms other than oxygen and hydrogen are balanced.

Step 3. Add $4H_2O(l)$ to the right side of the reduction half-reaction to balance the oxygen.
Reduction: $MnO_4^{-}(aq) \rightarrow Mn^{2+}(aq) + 4H_2O(l)$

Step 4. Add $8H^{+}(aq)$ to the left side of the reduction half-reaction to balance the hydrogen.
Reduction: $MnO_4^{-}(aq) + 8H^{+}(aq) \rightarrow Mn^{2+}(aq) + 4H_2O(l)$

Step 5. Balance the charges by adding electrons.
Oxidation: $Ag(s) \rightarrow Ag^{+}(aq) + 1e^{-}$
Reduction: $MnO_4^{-}(aq) + 8H^{+}(aq) + 5e^{-} \rightarrow Mn^{2+}(aq) + 4H_2O(l)$

Step 6. The LCM of electrons is 5.

Step 7. Multiply the oxidation half-reaction by 5 so that the loss of electrons in the oxidation half-reaction is equal to the gain of electrons in the reduction half-reaction.
Oxidation: $5Ag(s) \rightarrow 5Ag^{+}(aq) + 5e^{-}$

Step 8. Add the two half-reactions

$$5Ag(s) \rightarrow 5Ag^{+}(aq) + 5e^{-}$$
$$MnO_4^{-}(aq) + 8H^{+}(aq) + 5e^{-} \rightarrow Mn^{2+}(aq) + 4H_2O(l)$$
$$\overline{5Ag(s) + MnO_4^{-}(aq) + 8H^{+}(aq) + 5e^{-} \rightarrow 5Ag^{+}(aq) + 5e^{-} + Mn^{2+}(aq) + 4H_2O(l)}$$

Step 9. Cancel the electrons on each side of the equation.
$5Ag(s) + MnO_4^{-}(aq) + 8H^{+}(aq) \rightarrow 5Ag^{+}(aq) + Mn^{2+}(aq) + 4H_2O(l)$
Since the MnO_4^{-} gains electrons, it is reduced and it is the oxidizing agent.
Since the Ag loses electrons, it is oxidized and it is the reducing agent.

b) $Hg(l) + NO_3^{-}(aq) + Cl^{-}(aq) \rightarrow HgCl_4^{2-}(s) + NO_2(g)$

Assign an oxidation number to each element to identify the oxidation and reduction half-reactions.

$$\overset{0}{Hg}(l) + \overset{+5}{N}\overset{-2}{O_3^{-}}(aq) + \overset{-1}{Cl^{-}}(aq) \rightarrow \overset{+2}{Hg}\overset{-1}{Cl_4^{2-}}(s) + \overset{+4}{N}\overset{-2}{O_2}(g)$$

The oxidation number of mercury increases (oxidation).
The oxidation number of nitrogen decreases (reduction).

Step 1. The two unbalanced half-reactions are
Oxidation: $Hg(l) + Cl^{-}(aq) \rightarrow HgCl_4^{2-}(s)$

Reduction: $NO_3^-(aq) \rightarrow NO_2(g)$

Step 2. Balance the chlorine in the oxidation half-reaction.
Oxidation: $Hg(l) + 4Cl^-(aq) \rightarrow HgCl_4^{2-}(s)$

Step 3. Add $H_2O(l)$ to the right side of the reduction half-reaction to balance the oxygen.
Reduction: $NO_3^-(aq) \rightarrow NO_2(g) + H_2O(l)$

Step 4. Add $2H^+(aq)$ to the left side of the reduction half-reaction to balance the hydrogen.
Reduction: $2H^+(aq) + NO_3^-(aq) \rightarrow NO_2(g) + H_2O(l)$

Step 5. Balance the charges by adding electrons.
Oxidation: $Hg(l) + 4Cl^-(aq) \rightarrow HgCl_4^{2-}(s) + 2e^-$
Reduction: $2H^+(aq) + NO_3^-(aq) + 1e^- \rightarrow NO_2(g) + H_2O(l)$

Step 6. The LCM of electrons is 2.

Step 7. Multiply the reduction half-reaction by 2 so that the loss of electrons in the oxidation half-reaction is equal to the gain of electrons in the reduction half-reaction.
Reduction: $4H^+(aq) + 2NO_3^-(aq) + 2e^- \rightarrow 2NO_2(g) + 2H_2O(l)$

Step 8. Add the two half-reactions

$$\begin{array}{l} \qquad Hg(l) + 4Cl^-(aq) \rightarrow HgCl_4^{2-}(s) + 2e^- \\ \underline{4H^+(aq) + 2NO_3^-(aq) + 2e^- \rightarrow 2NO_2(g) + 2H_2O(l)} \\ Hg(l) + 4Cl^-(aq) + 4H^+(aq) + 2NO_3^-(aq) + 2e^- \rightarrow HgCl_4^{2-}(s) + 2e^- + 2NO_2(g) + 2H_2O(l) \end{array}$$

Step 9. Cancel the electrons on each side of the equation.
$Hg(l) + 4Cl^-(aq) + 4H^+(aq) + 2NO_3^-(aq) \rightarrow HgCl_4^{2-}(s) + 2NO_2(g) + 2H_2O(l)$
Since the NO_3^- is reduced it is the oxidizing agent.
Since the Hg is oxidized it is the reducing agent.

c) $AsH_3(g) + Zn^{2+}(aq) \rightarrow H_3AsO_4(aq) + Zn(s)$

Step 1. The two unbalanced half-reactions are
Oxidation: $AsH_3(g) \rightarrow H_3AsO_4(aq)$
Reduction: $Zn^{2+}(aq) \rightarrow Zn(s)$

Step 2. Atoms other than oxygen and hydrogen are balanced.

Step 3. Add $4H_2O(l)$ to the left side of the oxidation half-reaction to balance the oxygen.
Oxidation: $AsH_3(g) + 4H_2O(l) \rightarrow H_3AsO_4(aq)$

Step 4. Add $8H^+(aq)$ to the right side of the oxidation half-reaction to balance the hydrogen.
Oxidation: $AsH_3(g) + 4H_2O(l) \rightarrow H_3AsO_4(aq) + 8H^+(aq)$

Step 5. Balance the charges by adding electrons.
Oxidation: $AsH_3(g) + 4H_2O(l) \rightarrow H_3AsO_4(aq) + 8H^+(aq) + 8e^-$
Reduction: $Zn^{2+}(aq) + 2e^- \rightarrow Zn(s)$

Step 6. The LCM of electrons is 8.

Step 7. Multiply the reduction half-reaction by 4 so that the loss of electrons in the oxidation half-reaction is equal to the gain of electrons in the reduction half-reaction.
Reduction: $4Zn^{2+}(aq) + 8e^- \rightarrow 4Zn(s)$

Step 8. Add the two half-reactions
$AsH_3(g) + 4H_2O(l) \rightarrow H_3AsO_4(aq) + 8H^+(aq) + 8e^-$
$4Zn^{2+}(aq) + 8e^- \rightarrow 4Zn(s)$

$AsH_3(g) + 4H_2O(l) + 4Zn^{2+}(aq) + 8e^- \rightarrow H_3AsO_4(aq) + 8H^+(aq) + 8e^- + 4Zn(s)$

Step 9. Cancel the electrons on each side of the equation.
$AsH_3(g) + 4H_2O(l) + 4Zn^{2+}(aq) \rightarrow H_3AsO_4(aq) + 8H^+(aq) + 4Zn(s)$
Since the $Zn^{2+}(aq)$ gains electrons, it is reduced and it is the oxidizing agent.
Since the $AsH_3(g)$ loses electrons, it is oxidized and it is the reducing agent.

d) $I_2(s) + OCl^-(aq) \rightarrow IO_3^-(aq) + Cl^-(aq)$

Assign an oxidation number to each element to identify the oxidation and reduction half-reactions.

$$\overset{0}{I_2}(s) + \overset{-2}{O}\overset{+1}{Cl^-}(aq) \rightarrow \overset{+5}{I}\overset{-2}{O_3^-}(aq) + \overset{-1}{Cl^-}(aq)$$

The oxidation number of iodine increases (oxidation).
The oxidation number of chlorine decreases (reduction).

Step 1. The two unbalanced half-reactions are
Oxidation: $I_2(s) \rightarrow IO_3^-(aq)$
Reduction: $OCl^-(aq) \rightarrow Cl^-(aq)$

Step 2. Balance the iodine in the oxidation half-reaction.
Oxidation: $I_2(s) \rightarrow 2IO_3^-(aq)$

Step 3. Add $6H_2O(l)$ to the left side of the oxidation half-reaction to balance the oxygen. Add $H_2O(l)$ to the right side of the reduction half-reaction to balance the oxygen.
Oxidation: $I_2(s) + 6H_2O(l) \rightarrow 2IO_3^-(aq)$
Reduction: $OCl^-(aq) \rightarrow Cl^-(aq) + H_2O(l)$

Step 4. Add $12H^+(aq)$ to the right side of the reduction half-reaction to balance the hydrogen. Add $2H^+(aq)$ to the left side of the reduction half-reaction to balance the hydrogen.
Oxidation: $I_2(s) + 6H_2O(l) \rightarrow 2IO_3^-(aq) + 12H^+(aq)$

Reduction: $OCl^-(aq) + 2H^+(aq) \rightarrow Cl^-(aq) + H_2O(l)$

Step 5. Balance the charges by adding electrons.
Oxidation: $I_2(s) + 6H_2O(l) \rightarrow 2IO_3^-(aq) + 12H^+(aq) + 10e^-$
Reduction: $OCl^-(aq) + 2H^+(aq) + 2e^- \rightarrow Cl^-(aq) + H_2O(l)$

Step 6. The LCM of electrons is 10.

Step 7. Multiply the reduction half-reaction by 5 so that the loss of electrons in the oxidation half-reaction is equal to the gain of electrons in the reduction half-reaction.
Oxidation: $I_2(s) + 6H_2O(l) \rightarrow 2IO_3^-(aq) + 12H^+(aq) + 10e^-$
Reduction: $5OCl^-(aq) + 10H^+(aq) + 10e^- \rightarrow 5Cl^-(aq) + 5H_2O(l)$

Step 8. Add the two half-reactions.
$I_2(s) + 6H_2O(l) \rightarrow 2IO_3^-(aq) + 12H^+(aq) + 10e^-$
$5OCl^-(aq) + 10H^+(aq) + 10e^- \rightarrow 5Cl^-(aq) + 5H_2O(l)$

$I_2(s) + 6H_2O(l) + 5OCl^-(aq) + 10H^+(aq) + 10e^- \rightarrow 2IO_3^-(aq) + 12H^+(aq) + 10e^- + 5Cl^-(aq) + 5H_2O(l)$

Step 9. Cancel the electrons on each side of the equation. Remove identical ions or molecules present on both sides of the equation.

$I_2(s) + H_2O(l) + 5OCl^-(aq) \rightarrow 2IO_3^-(aq) + 2H^+(aq) + 5Cl^-(aq)$

Since $I_2(s)$ is oxidized, it is a reducing agent. Since $OCl^-(aq)$ is reduced, it is an oxidizing agent.

2.
Problem
Balance each of the following ionic equations for basic conditions. Identify the oxidizing agent and the reducing agent in each case:
a) $MnO_4^-(aq) + I^-(aq) + \rightarrow MnO_4^{2-}(aq) + IO_3^-(aq)$
b) $H_2O_2(aq) + ClO_2(aq) \rightarrow ClO^-(aq) + O_2(g)$
c) $ClO^-(aq) + CrO_2^-(aq) \rightarrow CrO_4^{2-}(aq) + Cl_2(g)$
d) $Al(s) + NO^-(aq) \rightarrow NH_3(g) + AlO_2^-(aq)$

What Is Required?
You must write a balanced net ionic equation for the given reactions that occur in basic conditions and identify the oxidizing agent and the reducing agent in each case.

What Is Given?
You know the formulas of the reactants and products, and that the reactions take place in basic solution.

Plan Your Strategy

Use the following the steps for balancing a net ionic equation by the half-reaction method in basic conditions.

Step 1. Write unbalanced half-reactions that show the formulas of the given reactant(s) and product(s). Write only those compounds that contain the atom that is oxidized or the atom that is reduced.
Step 2. Balance any atoms other than oxygen and hydrogen first.
Step 3. Balance any oxygen atoms by adding water molecules.
Step 4. Balance any hydrogen atoms by adding hydrogen ions.
Step 5. Adjust for basic conditions by adding to both sides the same number of hydroxide ions as the number of hydrogen ions already present.
Step 6. Simplify the equation by combining hydrogen ions and hydroxide ions on the same side of the equation into water molecules.
Step 7. Cancel any water molecules present on both sides of the equation.
Step 8 Balance the charges by adding electrons.
Step 9. Determine the least common multiple of the numbers of electrons in the oxidation and reduction half-reactions.
Step 10. Multiply one or both half-reactions by the number that will bring the number of electrons to the least common multiple and therefore an equal number of electrons are lost and gained.
Step 11. Add the balanced half-reactions.
Step 12. Cancel the electrons on both sides of the equation and simplify by combining $H_2O(l)$ and $OH^-(aq)$ on each side of the equation.
Step 13. If spectator ions were removed when forming half-reactions, add them back to the equation.

Act on Your Strategy
a) $MnO_4^-(aq) + I^-(aq) \rightarrow MnO_4^{2-}(aq) + IO_3^-(aq)$
Assign an oxidation number to each element to determine the oxidation and reduction half-reactions.

$$\overset{+7\ -2}{MnO_4^-}(aq) + \overset{-1}{I^-}(aq) \rightarrow \overset{+6\ -2}{MnO_4^{2-}}(aq) + \overset{+5\ -2}{IO_3^-}(aq)$$

Step 1. The two unbalanced half-reactions are
Oxidation: $I^-(aq) \rightarrow IO_3^-(aq)$
Reduction: $MnO_4^-(aq) \rightarrow MnO_4^{2-}(aq)$

Step 2. Atoms other than oxygen and hydrogen are balanced.

Step 3. Add $3H_2O(l)$ to the left side of the oxidation half-reaction to balance the oxygen.
Oxidation: $I^-(aq) + 3H_2O(l) \rightarrow IO_3^-(aq)$
Reduction: $MnO_4^-(aq) \rightarrow MnO_4^{2-}(aq)$

Step 4. Add $6H^+(aq)$ to the right side of the oxidation half-reaction to balance the hydrogen.
Oxidation: $I^-(aq) + 3H_2O(l) \rightarrow IO_3^-(aq) + 6H^+(aq)$
Reduction: $MnO_4^-(aq) \rightarrow MnO_4^{2-}(aq)$

Step 5. Add $6OH^-(aq)$ to each side of the oxidation half-reaction to adjust for basic conditions.
Oxidation: $I^-(aq) + 3H_2O(l) + 6OH^-(aq) \rightarrow IO_3^-(aq) + 6H^+(aq) + 6OH^-(aq)$

Step 6. Combine $H^+(aq)$ and $OH^-(aq)$ on the same side of each equation into water molecules.
Oxidation: $I^-(aq) + 3H_2O(l) + 6OH^-(aq) \rightarrow IO_3^-(aq) + 6H_2O(l)$

Step 7. Cancel any water molecules common to both sides of the equation.
Oxidation: $I^-(aq) + 6OH^-(aq) \rightarrow IO_3^-(aq) + 3H_2O(l)$

Step 8. Balance the charges by adding electrons.
Oxidation: $I^-(aq) + 6OH^-(aq) \rightarrow IO_3^-(aq) + 3H_2O(l) + 6e^-$
Reduction: $MnO_4^-(aq) + 1e^- \rightarrow MnO_4^{2-}(aq)$

Step 9. The LCM of electrons is 6.

Step 10. Multiply the reduction half-reaction by 6 so that the loss of electrons in the oxidation half-reaction is equal to gain of electrons in the reduction half-reaction.
Oxidation: $I^-(aq) + 6OH^-(aq) \rightarrow IO_3^-(aq) + 3H_2O(l) + 6e^-$
Reduction: $6MnO_4^-(aq) + 6e^- \rightarrow 6MnO_4^{2-}(aq)$

Step 11. Add the two half-reactions.
$I^-(aq) + 6OH^-(aq) \rightarrow IO_3^-(aq) + 3H_2O(l) + 6e^-$
$6MnO_4^-(aq) + 6e^- \rightarrow 6MnO_4^{2-}(aq)$

$I^-(aq) + 6OH^-(aq) + 6MnO_4^-(aq) + 6e^- \rightarrow IO_3^-(aq) + 3H_2O(l) + 6e^- + 6MnO_4^{2-}(aq)$

Step 12. Cancel the electrons on each side of the equation and simplify by combining $H_2O(l)$ and $OH^-(aq)$ on each side of the equation.
$I^-(aq) + 6OH^-(aq) + 6MnO_4^-(aq) \rightarrow IO_3^-(aq) + 3H_2O(l) + 6MnO_4^{2-}(aq)$

Since the $MnO_4^-(aq)$ is reduced, it is the oxidizing agent. Since the $I^-(aq)$ is oxidized it is the reducing agent.

b) $H_2O_2(aq) + ClO_2(aq) \rightarrow ClO_2^-(aq) + O_2(g)$

Assign oxidation numbers to the elements to determine the oxidation and reduction half-reactions.

$$\overset{+1\,-1}{H_2O_2}(aq) + \overset{+4\,-2}{ClO_2}(aq) \rightarrow \overset{+3\,-2}{ClO_2^-}(aq) + \overset{0}{O_2}(g)$$

Step 1. The two unbalanced half-reactions are
Oxidation: $H_2O_2(aq) \rightarrow O_2(g)$
Reduction: $ClO_2(aq) \rightarrow ClO_2^-(aq)$

Step 2. Atoms other than oxygen and hydrogen are balanced.

Step 3. The oxygen atoms are balanced in each equation.

Step 4. Add $2H^+(aq)$ to the right side of the oxidation half-reaction to balance the hydrogen.
Oxidation: $H_2O_2(aq) \rightarrow O_2(g) + 2H^+(aq)$

Step 5. Add $2OH^-(aq)$ to each side of the oxidation half-reaction to adjust for basic conditions.
Oxidation: $H_2O_2(aq) + 2OH^-(aq) \rightarrow O_2(g) + 2H^+(aq) + 2OH^-(aq)$

Step 6. Combine $H^+(aq)$ and $OH^-(aq)$ on the same side of each equation into water molecules.
Oxidation: $H_2O_2(aq) + 2OH^-(aq) \rightarrow O_2(g) + 2H_2O(l)$

Step 7. Balance the charges by adding electrons.
Oxidation: $H_2O_2(aq) + 2OH^-(aq) \rightarrow O_2(g) + 2H_2O(l) + 2e^-$
Reduction: $ClO_2(aq) + 1e^- \rightarrow ClO_2^-(aq)$

Step 8. The LCM of electrons is 2.

Step 9. Multiply the reduction half-reaction by 2 so that the loss of electrons in the oxidation half-reaction is equal to gain of electrons in the reduction half-reaction.
Reduction: $2ClO_2(aq) + 2e^- \rightarrow 2ClO_2^-(aq)$

Step 10. Add the two half-reactions.
$H_2O_2(aq) + 2OH^-(aq) \rightarrow O_2(g) + 2H_2O(l) + 2e^-$
$2ClO_2(aq) + 2e^- \rightarrow 2ClO_2^-(aq)$

$H_2O_2(aq) + 2OH^-(aq) + 2ClO_2(aq) + 2e^- \rightarrow O_2(g) + 2H_2O(l) + 2e^- + 2ClO_2^-(aq)$

Step 11. Cancel the electrons on each side of the equation.
$H_2O_2(aq) + 2OH^-(aq) + 2ClO_2(aq) \rightarrow O_2(g) + 2H_2O(l) + 2ClO_2^-(aq)$

Since the ClO_2 (aq) is reduced, it is the oxidizing agent. Since the $H_2O_2(aq)$ is oxidized it is the reducing agent.

c) $ClO^-(aq) + CrO_2^-(aq) \rightarrow CrO_4^{2-}(aq) + Cl_2(g)$
Assign an oxidation number to each element to determine the oxidation and reduction half-reactions.

$$\overset{+1\,-2}{ClO^-}(aq) + \overset{+3\,-2}{CrO_2^-}(aq) \rightarrow \overset{+6\,-2}{CrO_4^{2-}}(aq) + \overset{0}{Cl_2}(g)$$

Step 1. The two unbalanced half-reactions are
Oxidation: $CrO_2^-(aq) \rightarrow CrO_4^{2-}(aq)$
Reduction: $ClO^-(aq) \rightarrow Cl_2(g)$

Step 2. Balance the chlorine in the reduction half-reaction**.**
Reduction: $2ClO^{-}(aq) \rightarrow Cl_2(g)$

Step 3. Add $2H_2O(l)$ to the left side of the oxidation half-reaction and $2H_2O(l)$ to the right side of the reduction half-reaction to balance the oxygen.
Oxidation: $CrO_2^{-}(aq) + 2H_2O(l) \rightarrow CrO_4^{2-}(aq)$
Reduction: $2ClO^{-}(aq) \rightarrow Cl_2(g) + 2H_2O(l)$

Step 4. Add $4H^{+}(aq)$ to the right side of the oxidation half-reaction and $4H^{+}(aq)$ to the left side of the reduction half-reaction to balance the hydrogen.
Oxidation: $CrO_2^{-}(aq) + 2H_2O(l) \rightarrow CrO_4^{2-}(aq) + 4H^{+}(aq)$
Reduction: $2ClO^{-}(aq) + 4H^{+}(aq) \rightarrow Cl_2(g) + 2H_2O(l)$

Step 5. Add $4OH^{-}(aq)$ to each side of the oxidation and reduction half-reactions to adjust for basic conditions.
Oxidation: $CrO_2^{-}(aq) + 2H_2O(l) + 4OH^{-}(aq) \rightarrow CrO_4^{2-}(aq) + 4H^{+}(aq) + 4OH^{-}(aq)$
Reduction: $2ClO^{-}(aq) + 4H^{+}(aq) + 4OH^{-}(aq) \rightarrow Cl_2(g) + 2H_2O(l) + 4OH^{-}(aq)$

Step 6. Combine $H^{+}(aq)$ and $OH^{-}(aq)$ on the same side of each equation into water molecules.
Oxidation: $CrO_2^{-}(aq) + 2H_2O(l) + 4OH^{-}(aq) \rightarrow CrO_4^{2-}(aq) + 4H_2O(l)$
Reduction: $2ClO^{-}(aq) + 4H_2O(l) \rightarrow Cl_2(g) + 2H_2O(l) + 4OH^{-}(aq)$

Step 7. Cancel any water molecules common to both sides of the equation.
Oxidation: $CrO_2^{-}(aq) + 4OH^{-}(aq) \rightarrow CrO_4^{2-}(aq) + 2H_2O(l)$
Reduction: $2ClO^{-}(aq) + 2H_2O(l) \rightarrow Cl_2(g) + 4OH^{-}(aq)$

Step 8. Balance the charges by adding electrons.
Oxidation: $CrO_2^{-}(aq) + 4OH^{-}(aq) \rightarrow CrO_4^{2-}(aq) + 2H_2O(l) + 3e^{-}$
Reduction: $2ClO^{-}(aq) + 2H_2O(l) + 2e^{-} \rightarrow Cl_2(g) + 4OH^{-}(aq)$

Step 9. The LCM of electrons is 6.

Step 10. Multiply the oxidation half-reaction by 2 and the reduction half-reaction by 3 so that the loss of electrons in the oxidation half-reaction is equal to gain of electrons in the reduction half-reaction.
Oxidation: $2CrO_2^{-}(aq) + 8OH^{-}(aq) \rightarrow 2CrO_4^{2-}(aq) + 4H_2O(l) + 6e^{-}$
Reduction: $6ClO^{-}(aq) + 6H_2O(l) + 6e^{-} \rightarrow 3Cl_2(g) + 12OH^{-}(aq)$

Step 11. Add the two half-reactions

$$2CrO_2^{-}(aq) + 8OH^{-}(aq) \rightarrow 2CrO_4^{2-}(aq) + 4H_2O(l) + 6e^{-}$$
$$\underline{6ClO^{-}(aq) + 6H_2O(l) + 6e^{-} \rightarrow 3Cl_2(g) + 12OH^{-}(aq)}$$
$$2CrO_2^{-}(aq) + 8OH^{-}(aq) + 6ClO^{-}(aq) + 6H_2O(l) + 6e^{-} \rightarrow 3Cl_2(g) + 12OH^{-}(aq) + 2CrO_4^{2-}(aq) + 4H_2O(l) + 6e^{-}$$

Step 12. Cancel the electrons on each side of the equation and simplify by combining $H_2O(l)$ and $OH^-(aq)$ on each side of the equation.
$2CrO_2^-(aq) + 6ClO^-(aq) + 2H_2O(l) \rightarrow 3Cl_2(g) + 4OH^-(aq) + 2CrO_4^{2-}(aq)$

Since the $ClO^-(aq)$ is reduced, it is the oxidizing agent.
Since the $CrO_2^-(aq)$ is oxidized it is the reducing agent.

d) $Al(s) + NO^-(aq) \rightarrow NH_3(g) + AlO_2^-(aq)$
Assign an oxidation number to each element to determine the oxidation and reduction half-reactions.

0 +1 −2 −3 +1 +3 −2
$Al(s) + NO^-(aq) \rightarrow NH_3(g) + AlO_2^-(aq)$

Step 1. The two unbalanced half-reactions are
Oxidation: $Al(s) \rightarrow AlO_2^-(aq)$
Reduction: $NO^-(aq) \rightarrow NH_3(g)$

Step 2. Atoms other than oxygen and hydrogen are balanced.

Step 3. Add $2H_2O(l)$ to the left side of the oxidation half-reaction and $H_2O(l)$ to the right side of the reduction half-reaction to balance the oxygen.
Oxidation: $Al(s) + 2H_2O(l) \rightarrow AlO_2^-(aq)$
Reduction: $NO^-(aq) \rightarrow NH_3(g) + H_2O(l)$

Step 4. Add $4H^+(aq)$ to the right side of the oxidation half-reaction and $5H^+(aq)$ to the left side of the reduction half-reaction to balance the hydrogen.
Oxidation: $Al(s) + 2H_2O(l) \rightarrow AlO_2^-(aq) + 4H^+(aq)$
Reduction: $NO^-(aq) + 5H^+(aq) \rightarrow NH_3(g) + H_2O(l)$

Step 5. Add $4OH^-(aq)$ to each side of the oxidation half-reaction and $5OH^-(aq)$ to each side of the reduction half-reaction to adjust for basic conditions.
Oxidation: $Al(s) + 2H_2O(l) + 4OH^-(aq) \rightarrow AlO_2^-(aq) + 4H^+(aq) + 4OH^-(aq)$
Reduction: $NO^-(aq) + 5H^+(aq) + 5OH^-(aq) \rightarrow NH_3(g) + H_2O(l) + 5OH^-(aq)$

Step 6. Combine $H^+(aq)$ and $OH^-(aq)$ on the same side of each equation into water molecules.
Oxidation: $Al(s) + 4OH^-(aq) + 2H_2O(l) \rightarrow AlO_2^-(aq) + 4H_2O(l)$
Reduction: $NO^-(aq) + 5H_2O(l) \rightarrow NH_3(g) + 5OH^-(aq) + H_2O(l)$

Step 7. Cancel any water molecules common to both sides of the equation.
Oxidation: $Al(s) + 4OH^-(aq) \rightarrow AlO_2^-(aq) + 2H_2O(l)$
Reduction: $NO^-(aq) + 4H_2O(l) \rightarrow NH_3(g) + 5OH^-(aq)$

Step 8. Balance the charges by adding electrons.
Oxidation: $Al(s) + 4OH^-(aq) \rightarrow AlO_2^-(aq) + 2H_2O(l) + 3e^-$

Reduction: NO^{-}(aq) + $4H_2O$(l) + $4e^{-} \rightarrow NH_3$(g) +$5OH^{-}$(aq)

Step 9. The LCM of electrons is 12.

Step 10. Multiply the oxidation half-reaction by 4 and the reduction half-reaction by 3 so that the loss of electrons in the oxidation half-reaction is equal to gain of electrons in the reduction half-reaction.
Oxidation: 4Al(s) + $16OH^{-}$(aq)$\rightarrow$ $4AlO_2^{-}$(aq) + $8H_2O$(l) + $12e^{-}$
Reduction: $3NO^{-}$(aq) + $12H_2O$(l) + $12e^{-} \rightarrow 3NH_3$(g) +$15OH^{-}$(aq)

Step 11. Add the two half-reactions

4Al(s) + $16OH^{-}$(aq)$\rightarrow$ $4AlO_2^{-}$(aq) + $8H_2O$(l) + $12e^{-}$
$3NO^{-}$(aq) + $12H_2O$(l) + $12e^{-} \rightarrow 3NH_3$(g) +$15OH^{-}$(aq)

4Al(s) + $16OH^{-}$(aq) + $3NO^{-}$(aq) + $12H_2O$(l) + $12e^{-} \rightarrow 3NH_3$(g) +$15OH^{-}$(aq) + $4AlO_2^{-}$(aq) + $8H_2O$(l) + $12e^{-}$

Step 12. Cancel the electrons on each side of the equation and simplify by combining H_2O(l) and OH^{-}(aq) on each side of the equation.
4Al(s) + OH^{-}(aq) + $3NO^{-}$(aq) + $4H_2O$(l) $\rightarrow$ $3NH_3$(g) + $4AlO_2^{-}$(aq)

Since the NO^{-}(aq) is reduced, it is the oxidizing agent.
Since the Al(s) is oxidized it is the reducing agent.

Check Your Solutions
Since the number of atoms is the same on both sides of the equation and the net charge on each side of the equation is the same, the equations are balanced.

3.
Problem
Balance the following disproportionation reaction**:**
$PbSO_4$(aq) $\rightarrow$ Pb(s) + PbO_2(aq) + SO_4^{2-}(aq) (acidic solution; $PbSO_4$(aq) and PbO_2(aq) are soluble in an acidic solution)

What is Required?
You must balance a disproportionation reaction that occurs in acidic conditions.

What is Given?
You know the formulas for reactants and products, and that the reaction occurs in acidic conditions.

Plan Your Strategy
Write the equation as a net ionic equation.
Cancel spectator ions common to both sides of the reaction.
Use the rules for balancing a redox reaction in acid conditions.

Act on Your Strategy
$Pb^{2+}(aq) + SO_4^{2-}(aq) \rightarrow Pb(s) + PbO_2(aq) + SO_4^{2-}(aq)$
$Pb^{2+}(aq) \rightarrow Pb(s) + PbO_2(aq)$

Assign an oxidation number to each element to identify oxidation and reduction half-reactions.
$\overset{+2}{Pb^{2+}}(aq) \rightarrow \overset{0}{Pb}(s) + \overset{+4}{Pb}\overset{-2}{O_2}(aq)$

Step 1. The two unbalanced half-reactions are
Oxidation: $Pb^{2+}(aq) \rightarrow PbO_2(aq)$
Reduction: $Pb^{2+}(aq) \rightarrow Pb(s)$

Step 2. Atoms other than oxygen and hydrogen are balanced

Step 3. Add $2H_2O(l)$ to the left side of the oxidation half-reaction to balance the oxygen.
Oxidation: $Pb^{2+}(aq) + 2H_2O(l) \rightarrow PbO_2(aq)$
Reduction: $Pb^{2+}(aq) \rightarrow Pb(s)$

Step 4. Add $4H^+(aq)$ to the right side of the oxidation half-reaction to balance the hydrogen.
Oxidation: $Pb^{2+}(aq) + 2H_2O(l) \rightarrow PbO_2(aq) + 4H^+(aq)$
Reduction: $Pb^{2+}(aq) \rightarrow Pb(s)$

Step 5. Balance the charges by adding electrons.
Oxidation: $Pb^{2+}(aq) + 2H_2O(l) \rightarrow PbO_2(aq) + 4H^+(aq) + 2e^-$
Reduction: $Pb^{2+}(aq) + 2e^- \rightarrow Pb(s)$

Step 6. The LCM of electrons is 2.
Since the gain and loss of electrons is the same, go to **Step 8**.

Step 8. Add the two half-reactions
$Pb^{2+}(aq) + 2H_2O(l) \rightarrow PbO_2(aq) + 4H^+(aq) + 2e^-$
$Pb^{2+}(aq) + 2e^- \rightarrow Pb(s)$

$2Pb^{2+}(aq) + 2H_2O(l) + 2e^- \rightarrow Pb(s) + PbO_2(aq) + 4H^+(aq) + 2e^-$

Step 9. Cancel the electrons on each side of the equation.
$2Pb^{2+}(aq) + 2H_2O(l) \rightarrow Pb(s) + PbO_2(aq) + 4H^+(aq)$

Rewrite the equation to include the spectator ions.
$2PbSO_4(aq) + 2H_2O(l) \rightarrow Pb(s) + PbO_2(aq) + 2H_2SO_4(aq)$

Check Your Solution
Since there is the same number atoms of each element are on both sides of the equation and the net charge on each side of the equation is zero, the equation is balanced.

4.

Problem
Balance the following disproportionation reaction:
$NO_2(g) + H_2O(l) \rightarrow HNO_3(aq) + NO(g)$ (acidic solution)

What is Required?
You must balance a disproportionation reaction that occurs in acidic conditions.

What is Given?
You know the formulas for reactants and products, and that the reaction occurs in acidic conditions.

Plan Your Strategy
Write the equation in the form of a net ionic reaction.
Use the rules for balancing a redox reaction in acid conditions.

Act on Your Strategy
The ionic equation for this reaction is
$NO_2(g) \rightarrow NO_3^-(aq) + NO(g)$

Assign an oxidation number to each element to identify oxidation and reduction half-reactions.

$$\overset{+4\,-2}{NO_2}(g) \rightarrow \overset{+5\,-2}{NO_3^-}(aq) + \overset{+2\,-2}{NO}(g)$$

Step 1. The two unbalanced half-reactions are
Oxidation: $NO_2(g) \rightarrow NO_3^-(aq)$
Reduction: $NO_2(g) \rightarrow NO(g)$

Step 2. Atoms other than oxygen and hydrogen are balanced

Step 3. Add $H_2O(l)$ to the left side of the oxidation half-reaction and $H_2O(l)$ to the right side of the reduction half-reaction to balance the oxygen.
Oxidation: $NO_2(g) + H_2O(l) \rightarrow NO_3^-(aq)$
Reduction: $NO_2(g) \rightarrow NO(g) + H_2O(l)$

Step 4. Add $2H^+(aq)$ to the right side of the oxidation half-reaction and $2H^+(aq)$ to the left side of the reduction half-reaction to balance the hydrogen.
Oxidation: $NO_2(g) + H_2O(l) \rightarrow NO_3^-(aq) + 2H^+(aq)$
Reduction: $NO_2(g) + 2H^+(aq) \rightarrow NO(g) + H_2O(l)$

Step 5. Balance the charges by adding electrons.
Oxidation: $NO_2(g) + H_2O(l) \rightarrow NO_3^-(aq) + 2H^+(aq) + e^-$
Reduction: $NO_2(g) + 2H^+(aq) + 2e^- \rightarrow NO(g) + H_2O(l)$

Step 6. The LCM of electrons is 2.

Step 7. Multiply the oxidation half-reaction by 2 so that the loss of electrons in the oxidation is equal to the gain of electrons in the reduction.
Oxidation: $2NO_2(g) + 2H_2O(l) \rightarrow 2NO_3^-(aq) + 4H^+(aq) + 2e^-$
Reduction: $NO_2(g) + 2H^+(aq) + 2e^- \rightarrow NO(g) + H_2O(l)$

Step 8. Add the two half-reactions

$$\begin{array}{l} \quad 2NO_2(g) + 2H_2O(l) \rightarrow 2NO_3^-(aq) + 4H^+(aq) + 2e^- \\ \underline{NO_2(g) + 2H^+(aq) + 2e^- \rightarrow NO(g) + H_2O(l)} \\ 3NO_2(g) + 2H_2O(l) + 2H^+(aq) + 2e^- \rightarrow NO(g) + H_2O(l) + 2NO_3^-(aq) + 4H^+(aq) + 2e^- \end{array}$$

Step 9. Cancel the electrons on each side of the equation and combine $H_2O(l)$ from each side of the equation.
$3NO_2(g) + H_2O(l) \rightarrow NO(g) + 2NO_3^-(aq) + 2H^+(aq)$

Rewrite the equation in its original non-ionic form.
$3NO_2(g) + H_2O(l) \rightarrow NO(g) + 2HNO_3(aq)$

Check Your Solution
Since there is same number of atoms of each element on both sides of the equation and the net charge on both sides of the equation is zero, the equation is balanced.

5.
Problem
Balance the following disproportionation reaction**:**
$Cl_2(g) \rightarrow ClO^-(aq) + Cl^-(aq)$ (basic conditions)

What is Required?
You must balance a disproportionation reaction that occurs in basic conditions.

What is Given?
You know the formulas for reactants and products, and that the reaction occurs in basic conditions.

Plan Your Strategy
Use the rules for balancing a redox reaction in basic conditions.

Assign an oxidation number to each element to determine the oxidation and reduction half-reactions.

$$\overset{0}{Cl_2}(g) \rightarrow \overset{+1}{Cl}\overset{-2}{O^-}(aq) + \overset{-1}{Cl^-}(aq)$$

Step 1. The two unbalanced half-reactions are
Oxidation: $Cl_2(g) \rightarrow ClO^-(aq)$
Reduction: $Cl_2(g) \rightarrow Cl^-(aq)$

Step 2. Balance the chlorine atoms in each half-reaction.
Oxidation: $Cl_2(g) \rightarrow 2ClO^-(aq)$
Reduction: $Cl_2(g) \rightarrow 2Cl^-(aq)$

Step 3. Add $2H_2O(l)$ to the left side of the oxidation half-reaction to balance the oxygen.
Oxidation: $Cl_2(g) + 2H_2O(l) \rightarrow 2ClO^-(aq)$
Reduction: $Cl_2(g) \rightarrow 2Cl^-(aq)$

Step 4. Add $4H^+(aq)$ to the right side of the oxidation half-reaction to balance the hydrogen.
Oxidation: $Cl_2(g) + 2H_2O(l) \rightarrow 2ClO^-(aq) + 4H^+(aq)$
Reduction: $Cl_2(g) \rightarrow 2Cl^-(aq)$

Step 5. Add $4OH^-(aq)$ to each side of the oxidation to adjust for basic conditions.
Oxidation: $Cl_2(g) + 2H_2O(l) + 4OH^-(aq) \rightarrow 2ClO^-(aq) + 4H^+(aq) + 4OH^-(aq)$
Reduction: $Cl_2(g) \rightarrow 2Cl^-(aq)$

Step 6. Combine $H^+(aq)$ and $OH^-(aq)$ on the same side of each equation into water molecules.
Oxidation: $Cl_2(g) + 2H_2O(l) + 4OH^-(aq) \rightarrow 2ClO^-(aq) + 4H_2O(l)$
Reduction: $Cl_2(g) \rightarrow 2Cl^-(aq)$

Step 7. Cancel any water molecules common to both sides of the equation.
Oxidation: $Cl_2(g) + 4OH^-(aq) \rightarrow 2ClO^-(aq) + 2H_2O(l)$
Reduction: $Cl_2(g) \rightarrow 2Cl^-(aq)$

Step 8. Balance the charges by adding electrons.
Oxidation: $Cl_2(g) + 4OH^-(aq) \rightarrow 2ClO^-(aq) + 2H_2O(l) + 2e^-$
Reduction: $Cl_2(g) + 2e^- \rightarrow 2Cl^-(aq)$

Step 9. The LCM of electrons is 2.
Since the gain and loss of electrons is the same, go to **Step 11**.

Step 11. Add the two half-reactions
$Cl_2(g) + 4OH^-(aq) \rightarrow 2ClO^-(aq) + 2H_2O(l) + 2e^-$
$\underline{Cl_2(g) + 2e^- \rightarrow 2Cl^-(aq)}$
$2Cl_2(g) + 4OH^-(aq) + 2e^- \rightarrow 2Cl^-(aq) + 2ClO^-(aq) + 2H_2O(l) + 2e^-$

Step 12. Cancel the electrons on each side of the equation and reduce the coefficients to the LCM.
$Cl_2(g) + 2OH^-(aq) \rightarrow Cl^-(aq) + ClO^-(aq) + H_2O(l)$

Check Your Solution
Since there is the same number of atoms of each element on both sides of the equation and the net charge on both sides of the equation is −2, the equation is balanced.

6.

Problem

Balance the following disproportionation reaction:

$I_3^-(aq) \rightarrow I^-(aq) + IO_3^-(aq)$ (acidic conditions)

Assign an oxidation number to each element to identify oxidation and reduction half-reactions.

$-1/3 \quad -1 \quad +5\ -2$

$I_3^-(aq) \rightarrow I^-(aq) + IO_3^-(aq)$

Step 1. The two unbalanced half-reactions are

Oxidation: $I_3^-(aq) \rightarrow IO_3^-(aq)$

Reduction: $I_3^-(aq) \rightarrow I^-(aq)$

Step 2. Balance the iodine in the oxidation and reduction half-reactions.

Oxidation: $I_3^-(aq) \rightarrow 3IO_3^-(aq)$

Reduction: $I_3^-(aq) \rightarrow 3I^-(aq)$

Step 3. Add $9H_2O(l)$ to the left side of the oxidation half-reaction to balance the oxygen.

Oxidation: $I_3^-(aq) + 9H_2O(l) \rightarrow 3IO_3^-(aq)$

Step 4. Add $18H^+(aq)$ to the right side of the oxidation half-reaction to balance the hydrogen.

Oxidation: $I_3^-(aq) + 9H_2O(l) \rightarrow 3IO_3^-(aq) + 18H^+(aq)$

Step 5. Balance the charges by adding electrons.

Oxidation: $I_3^-(aq) + 9H_2O(l) \rightarrow 3IO_3^-(aq) + 18H^+(aq) + 16e^-$

Reduction: $I_3^-(aq) + 2e^- \rightarrow 3I^-(aq)$

Step 6. The LCM of electrons is 16.

Step 7. Multiply the reduction half-reaction by 8 so that the loss of electrons in the oxidation half-reaction is equal to the gain of electrons in the reduction half-reaction.

Reduction: $8I_3^-(aq) + 16e^- \rightarrow 24I^-(aq)$

Step 8. Add the two half-reactions

$I_3^-(aq) + 9H_2O(l) \rightarrow 3IO_3^-(aq) + 18H^+(aq) + 16e^-$

$8I_3^-(aq) + 16e^- \rightarrow 24I^-(aq)$

$9I_3^-(aq) + 9H_2O(l) + 16e^- \rightarrow 3IO_3^-(aq) + 18H^+(aq) + 24I^-(aq) + 16e^-$

Step 9. Cancel the electrons on each side of the equation and divide by 3.

$3I_3^-(aq) + 3H_2O(l) \rightarrow IO_3^-(aq) + 6H^+(aq) + 8I^-(aq)$

Check Your Solution

Since the number of atoms of each element is the same on both sides of the equation and the net charge on both sides of the equation is the same, the equations are balanced.

7.
Problem
Determine the oxidation number of the atoms of the specified element in each of the following:
a) N in $NF_3(g)$
b) S in $S_8(s)$
c) Cr in $CrO_4^{2-}(aq)$
d) P in $P_2O_5(s)$
e) C in $C_{12}H_{22}O_{11}(s)$
f) C in $CHCl_3(g)$

What is Required?
You must determine the oxidation number of the specified element.

What is Given?
The formula of the compound is known.

Plan Your Strategy
Follow the rules for assigning oxidation number (O.N.).

Act on Your Strategy
a) N in $NF_3(g)$
The sum of the O.N. = 0
(1 atom N × O.N. of N) + (3 atoms F × O.N.of F) = 0
(1 atom N × O.N. of N) + (3 atoms F × –1) = 0
O.N. of N = +3

b) S in $S_8(s)$ S = 0 since this is the elemental form of sulfur

c) Cr in $CrO_4^{2-}(aq)$
The sum of the O.N. = charge on the ion = 2–
(1 atom Cr × O.N. of Cr) + (4 atoms of O × O.N. of O) = 2–
(1 atom Cr × O.N. of Cr) + (4 atoms of O × –2) = 2–
O.N. of Cr = +6

d) P in $P_2O_5(s)$
The sum of the O.N. = 0
(2 atom P × O.N. of P) + (5 atoms of O × O.N. of O) = 0
(2 atom P × O.N. of P) + (5 atoms of O × –2) = 0
O.N. of P = +5

e) C in $C_{12}H_{22}O_{11}(s)$
The sum of the O.N. = 0
(12 atom C × O.N. of C) + (22 atoms of H × O.N. of H) + (11 atoms of O × O.N. of O) = 0
(12 atom C × O.N. of C) + (22 atoms of H × +1) + (11 atoms of O × –2) = 0
O.N. of C = 0

f) C in $CHCl_3(g)$
The sum of the O.N. = 0
(1 atom C × O.N. of C) + (1 atom of H × O.N. of H) + (3 atoms of Cl × O.N. of Cl) = 0
(1 atom C × O.N. of C) + (1 atom of H × +1) + (3 atoms of Cl × −1) = 0
O.N. of C = +2

Check Your Solution
In each case the calculated value of the O.N. of the specific element gives the expected sum of the O.N. for the molecule or ion.

8.
Problem
Determine the oxidation number of each of the atoms in each of the following compounds:
a) $H_2SO_3(aq)$
b) $OH^{-}(aq)$
c) $HPO_4^{2-}(aq)$

What is Required?
You must determine the oxidation number of each element in the given compounds.

What is Given?
The formula of the compound is known.

Plan Your Strategy
Follow the rules for assigning oxidation number (O.N.).

Act on Your Strategy
a) $H_2SO_3(aq)$
The sum of the O.N. = 0
From the rules for assigning O.N., H = +1 and O = –2
(2 atoms H × O.N. of H) + (1 atom of S × O.N. of S) + (3 atoms of O × O.N. of O) = 0
(2 atoms H × +1) + (1 atom of S × O.N. S) + (3 atoms of O × –2) = 0
O.N. of S = +4

b) $OH^{-}(aq)$
The sum of the O.N. = net charge on $OH^{-}(aq)$ = 1–
From the rules for assigning O.N., H = +1 and O = –2

c) $HPO_4^{2-}(aq)$
The sum of the O.N. = net charge on the $HPO_4^{2-}(aq)$ = 2–
From the rules for assigning O.N. H = +1 and O = –2
(1 atom H × O.N. of H) + (1 atom of P × O.N. of P) + (4 atoms of O × O.N. of O) = 2 –
(1 atom H × +1) + (1 atom of P × O.N. of P) + (4 atoms of O × –2) = 2–
O.N. for P = +5

Check Your Solution
In each case the calculated values of the oxidation number of the specific elements gives the expected sum of the O.N. for the molecule or ion.

9.
Problem
As stated in rule 4, oxygen does not always have its usual oxidation number of −2. Determine the oxidation number of oxygen in each of the following:
a) the compound oxygen difluoride, $OF_2(g)$
b) the peroxide ion, O_2^{2-} (aq)

What is Required?
You must determine the oxidation number of oxygen in each compound.

What is Given?
The formula of the compound is known.

Plan Your Strategy
Follow the rules for assigning oxidation number (O.N.) except the O.N. for oxygen ≠ −2.

Act on Your Strategy
a) $OF_2(g)$
The sum of the O.N. = 0
(1 atom O × O.N. of O) + (2 atoms F × O.N. F) = 0
(1 atom O × O.N. of O) + (2 atoms F × −1) = 0
O.N. of O = +2

b) O_2^{2-} (aq)
The sum of the O.N. = charge on O_2^{2-} (aq) = 2–
2 atoms O × O.N. of O) = 2–
O.N of O = −1

Check Your Solution
In each case, the calculate O.N. for oxygen gives the sum of the O.N. for the molecule or ion.

10.
Problem
Determine the oxidation number of each element in each of the following ionic compounds by considering the ions separately. (**Hint:** One formula unit of the compound in part (c) contains two identical monatomic ions and one polyatomic ion.)
a) $Al(HCO_3)_3(s)$
b) $(NH_4)_3PO_4(aq)$
c) $K_2H_3IO_6(aq)$

What is Required?
You must determine the oxidation number of each element in the given compounds.

What is Given?
The formula of the compound is known.

Plan Your Strategy
Follow the rules for assigning oxidation number (O.N.).

Act on Your Strategy
a) $Al(HCO_3)_3(s)$
The sum of the O.N. = 0
From the rules for assigning O.N., H = +1 and O = –2, Al = +3
(1 atomAl × O.N. of Al) + (3 atoms of H × O.N. of H) + (3 atoms of C × O.N. of C) + (9 atoms O × O.N. of O) = 0
1 atomAl × O.N. of +3) + (3 atoms of H × O.N. of +1) + (3 atoms of C × O.N. of C) + (9 atoms O × O.N. of –2) = 0
O.N. of C = +4

b) $(NH_4)_3PO_4(aq)$
The sum of the O.N. = 0
From the rules for assigning O.N., H = +1 and O = –2
The sum of the O.N. for NH_4^+ = charge on NH_4^+ = 1+
(1 atom N × O.N. N) + 4 atom H × O.N. of H) = 1+
O.N. of N = –3
The sum of the O.N. for PO_4^{3-} = charge on PO_4^{3-} = 3–
(1 atom P × O.N. of P) + (4 atoms O × O.N. of O) = 3–
O.N. of P = +5

c) $K_2H_3IO_6(aq)$
$K_2H_3IO_6(aq) = 2K^+(aq) + H_3IO_6^{2-}(aq)$
From the rules for assigning oxidation numbers, O.N., H = +1, O = –2, K = +1
The sum of the O.N. for $H_3IO_6^{2-}(aq)$ = net charge on $H_3IO_6^{2-}(aq)$ = 2–
(3 atoms of H × O.N. of H) + (1 atom of I × O.N. of I) + (6 atoms O × O.N. of O) = 2–
(3 atoms of H × +1) + (1 atom of I × O.N. of I) + (6 atoms O × O.N. of –2) = 2–
O.N. of I = +7

Check Your Solution
In each case, the sum of the oxidation numbers gives a net charge of zero for each compound.

11.
Problem
Which of the following are redox reactions? Identify any disproportionation reactions:
a) $H_2O_2(aq) + 2Fe(OH)_2(s) \rightarrow 2Fe(OH)_3(s)$
b) $PCl_3(l) + 3H_2O(l) \rightarrow H_3PO_3(aq) + 3HCl(aq)$
c) $2C_2H_6(g) + 7O_2(g) \rightarrow 4CO_2(g) + 6H_2O(l)$
d) $3NO_2(g) + H_2O(l) \rightarrow 2HNO_3(aq) + NO(g)$

What is Required?
You must determine if the reactions are oxidation-reduction and identify any disproportionation reactions.

What is Given?
Balanced equations showing the formulas for the reactants and products are given.

Plan Your Strategy
Follow the rules for assigning oxidation numbers and identify the oxidation number for each element. In a redox reaction, oxidation can be identified by an increase in oxidation number and reduction can be identified by a decrease in oxidation number. In a disproportionation reaction, the same element undergoes both oxidation and reduction.

Act on Your Strategy

a) +1 −1 +2 −2 +1 +3 −2 +1

$H_2O_2(aq) + 2Fe(OH)_2(s) \rightarrow Fe(OH)_3(s).$

O.N. of Fe increases from +2 in $Fe(OH)_2(s)$ to +3 in $Fe(OH)_3(s)$ (oxidation)
O.N. of O decreases from −1 in $H_2O_2(aq)$ to −2 in $Fe(OH)_3(s)$ (reduction)
This is a redox reaction.

b) +3 −1 +1 −2 +1 +3 −2 +1 −1

$PCl_3(l) + 3H_2O(l) \rightarrow H_3PO_3(aq) + 3HCl(aq)$

There is no change in the O.N. of any element. This is not a redox reaction.

c) −3 +1 0 +4 −2 +1 −2

$2C_2H_6(g) + 7O_2(g) \rightarrow 4CO_2(g) + 6H_2O(l)$

The O.N. of C increases from −3 in $C_2H_6(g)$ to +4 in $CO_2(g)$ (oxidation)
The O.N. of oxygen in decreases from 0 in $O_2(g)$ to −2 in $CO_2(g)$ (reduction)
This is a redox reaction.

d) +4 −2 +1 −2 +1 +5 −2 +2 −2

$3NO_2(g) + H_2O(l) \rightarrow 2HNO_3(aq) + NO(g)$

The O.N of N increases from +4 in $NO_2(g)$ to +5 in $HNO_3(aq)$ (oxidation)
The O.N of N decreases from +4 in $NO_2(g)$ to +2 in NO(g) (reduction)
This is a redox reaction. Since N in $NO_2(g)$ undergoes an increase and a decrease in O.N., this is a disproportionation reaction.

Check Your Solution
This answer corresponds to the definitions of oxidation and reduction.

12.
Problem
Identify the oxidizing agent and the reducing agent for the redox reaction(s) in the previous question.

What is Required?
You must identify the oxidizing agent and the reducing agent in each reaction from question 10.

What is Given?
The answer from question 10 identified the species undergoing oxidation and reduction.

Plan Your Strategy
The reducing agent undergoes oxidation and the oxidizing agent undergoes reduction.

Act on Your Strategy
a) $H_2O_2(aq)$ undergoes reduction. It is an oxidizing agent.
$Fe(OH)_2(s)$ undergoes oxidation. It is a reducing agent.

b) This is not a redox reaction. There is no oxidizing agent or reducing agent.

c) $O_2(g)$ undergoes reduction. It is the oxidizing agent.
$C_2H_6(g)$ undergoes oxidation. It is the reducing agent

d) $NO_2(g)$ undergoes both reduction and oxidation. It is both an oxidizing agent and a reducing agent.

Check Your Solution
The answers correspond with the definitions of oxidizing agent and reducing agent.

13.
Problem
For the following balanced net ionic equation, identify the reactant that undergoes oxidation and the reactant that undergoes reduction:

$$Br_2(l) + 2ClO_2^-(aq) \rightarrow 2Br^-(aq) + 2ClO_2(aq)$$

What is Required?
You must identify the reactant that undergoes oxidation and the reactant that undergoes reduction.

What is Given?
The balanced equation showing the formulas of the reactants and products is given.

Plan Your Strategy
Assign an oxidation number (O.N.) to each element. The reducing agent will undergo oxidation and its O.N. will increase. The oxidizing agent will undergo reduction and its O.N. will decrease.

Act on Your Strategy

$$\overset{0}{Br_2}(l) + 2\overset{+3}{Cl}\overset{-2}{O_2}{}^-(aq) \rightarrow 2\overset{-1}{Br}{}^-(aq) + 2\overset{+4}{Cl}\overset{-2}{O_2}(aq)$$

The O.N. of Br decreases from 0 in $Br_2(l)$ to −1 in $2Br^-(aq)$. This is reduction. $Br_2(l)$ is an oxidizing agent.

The O.N. of Cl increases from +3 in ClO_2^-(aq) to +4 in ClO_2(aq). This is oxidation. ClO_2^-(aq) is a reducing agent.

Check Your Solution
This answer corresponds to the definition of oxidizing agent and reducing agent.

14.
Problem
Nickel and copper ores usually contain the metals as sulfides, such as NiS(s) and Cu_2S(s). Does the extraction of these pure elemental metals from their ores involve redox reactions? Explain your reasoning.

What is Required?
You must determine if the reaction in which metals are recovered from their sulfide ores is a redox reaction.

What is Given?
The formula of the sulfide ore is known and one product must be the metal in its elemental form.

Plan Your Strategy
Write the formula of the sulfide ore as a reactant and the corresponding metal as the product. Assign an oxidation number to each element and determine if there is a change in oxidation number from reactants to products. If a reduction occurs, a corresponding oxidation must also occur and the reaction is a redox reaction.

Act on Your Strategy
$\overset{+2}{Ni}\overset{-2}{S} \rightarrow \overset{0}{Ni}$

The O.N. of the nickel decreases from +2 in the sulfide to 0 in its elemental form. This is a reduction. A reducing agent must be used to achieve this change. This is a redox reaction.
The same analysis can be used with CuS and Cu.

Check Your Solution
The answer is consistent with the definition of a redox reaction.

15.
Problem
Use the oxidation number method to balance the following equation for the combustion of carbon disulfide:
$CS_2(g) + O_2(g) \rightarrow CO_2(g) + SO_2(g)$

What is Required?
You must balance the redox reaction using oxidation numbers.

What is Given?
The unbalanced equation is given with the formula of the reactants and products.

Plan Your Strategy
Assign an oxidation number (O.N.) to each element. Follow the rules for balancing an equation using O.N.

Act on Your Strategy
Step 1. The unbalanced equation is $CS_2(g) + O_2(g) \rightarrow CO_2(g) + SO_2(g)$

Step 2. Assign an O.N. to each element.

$$\overset{+4\ -2}{CS_2}(g) + \overset{0}{O_2}(g) \rightarrow \overset{+4\ -2}{CO_2}(g) + \overset{+4\ -2}{SO_2}(g)$$

Step 3. The O.N. of sulfur undergoes an increase of 6 from −2 in $CS_2(g)$ to +4 in $SO_2(g)$.
The O.N. of oxygen undergoes a decrease of 2 from 0 in $O_2(g)$ to −2 in $SO_2(g)$.

Step 4. For the total decrease in oxidation number to equal the total increase in oxidation number, the smallest whole number ratio of oxygen:sulfur atoms is 3:1.

Step 5. Use the 3:1 ratio in **Step 4** to balance the oxygen and sulfur atoms. There must be 3 oxygen atoms for each sulfur atom.
$CS_2(g) + 3O_2(g) \rightarrow CO_2(g) + SO_2(g)$

Step 6. Complete the balancing by inspection. Balance the sulfur on the right side of the equation and the oxygen on the left side.
$CS_2(g) + 3O_2(g) \rightarrow CO_2(g) + 2SO_2(g)$ (balanced)

Check Your Solution
Since there is the same number of atoms of each element on both sides of the equation, the equation is balanced.

16.
Problem
Use the oxidation number method to balance the following equations:
a) $B_2O_3(aq) + Mg(s) \rightarrow MgO(s) + Mg_3B_2(aq)$
b) $H_2S(g) + H_2O_2(aq) \rightarrow S_8(s) + H_2O(l)$

What is Required?
You must balance the redox reactions using oxidation numbers.

What is Given?
The unbalanced equations are given with the formula of the reactants and products.

Plan Your Strategy
Assign an oxidation number (O.N.) to each element. Follow the rules for balancing an equation using O.N.

Act on Your Strategy
(a) $B_2O_3(aq) + Mg(s) \rightarrow MgO(s) + Mg_3B_2(aq)$

Step 1. The unbalanced equation is $B_2O_3(aq) + Mg(s) \rightarrow MgO(s) + Mg_3B_2(aq)$

Step 2. Assign an O.N. to each element.
+3 –2 0 +2 –2 +2 –3
$B_2O_3(aq) + Mg(s) \rightarrow MgO(s) + Mg_3B_2(aq)$
Step 3. The O.N. of magnesium undergoes an increase of 2 from 0 in Mg(s) to +2 in $Mg_3B_2(aq)$.
The O.N. of boron undergoes a decrease of 6 from +3 in $B_2O_3(aq)$ to –3 in $Mg_3B_2(aq)$.

Step 4. For the total increase in oxidation number to equal the total decrease in oxidation number, the smallest whole number ratio of magnesium:boron atoms is 3:1.

Step 5. Use the 3:1 ratio in **Step 4** to balance the magnesium and boron atoms. There must be 3 magnesium atoms for each boron atom.
$B_2O_3(aq) + 6Mg(s) \rightarrow MgO(s) + Mg_3B_2(aq)$

Step 6. Complete the balancing of the oxygen by inspection.
$B_2O_3(aq) + 6Mg(s) \rightarrow 3MgO(s) + Mg_3B_2(aq)$ (balanced)

b) $H_2S(g) + H_2O_2(aq) \rightarrow S_8(s) + H_2O(l)$
Step 1. The unbalanced equation is $H_2S(g) + H_2O_2(aq) \rightarrow S_8(s) + H_2O(l)$

Step 2. Assign an O.N. to each element.
+1 –2 +1 –1 0 +1 –2
$H_2S(g) + H_2O_2(aq) \rightarrow S_8(s) + H_2O(l)$

Step 3. The O.N. of sulfur undergoes an increase of 2 from –2 in $H_2S(g)$ to 0 in$S_8(s)$.
The O.N. of oxygen undergoes a decrease of 1 from –1 in $H_2O_2(aq)$ to –2 in $H_2O(l)$.

Step 4. For the total increase in oxidation number to equal the total decrease in oxidation number, the smallest whole number ratio of oxygen:sulfur atoms is 2:1. There must be 2 oxygen atoms for each sulfur atom.

Step 5. Use the 2:1 ratio in **Step 4** to balance the oxygen and sulfur atoms. There must be 2 oxygen atoms for each sulfur atom.
$H_2S(g) + H_2O_2(aq) \rightarrow S_8(s) + H_2O(l)$

Step 6. Complete the balancing of hydrogen, sulfur and the oxygen by inspection.
$8H_2S(g) + 8H_2O_2(aq) \rightarrow S_8(s) + 16H_2O(l)$ (balanced)

Check Your Solution
Since there is the same number of atoms of each element on both sides of the equations, the equations are balanced.

17.
Problem
Use the oxidation number method to balance each ionic equation in an acidic solution:
a) $Cr_2O_7^{2-}(aq) + Fe^{2+}(aq) \rightarrow Cr^{3+}(aq) + Fe^{3+}(aq)$
b) $I_2(g) + NO_3^{-}(aq) \rightarrow IO_3^{-}(aq) + NO_2(g)$
c) $PbSO_4(aq) \rightarrow Pb(s) + PbO_2(aq) + SO_4^{2-}(aq)$

What is Required?
You must balance the redox reactions using oxidation numbers.

What is Given?
The unbalanced equations are given with the formula of the reactants and products.

Plan Your Strategy
Assign an oxidation number (O.N.) to each element. Follow the rules for balancing an equation in acid conditions using O.N.

Act on Your Strategy
a) $Cr_2O_7^{2-}(aq) + Fe^{2+}(aq) \rightarrow Cr^{3+}(aq) + Fe^{3+}(aq)$

Step 1. The unbalanced equation is $Cr_2O_7^{2-}(aq) + Fe^{2+}(aq) \rightarrow Cr^{3+}(aq) + Fe^{3+}(aq)$

Step 2. Assign an O.N. to each element.
$$\overset{+6\ -2}{Cr_2O_7^{2-}}(aq) + \overset{+2}{Fe^{2+}}(aq) \rightarrow \overset{+3}{Cr^{3+}}(aq) + \overset{+3}{Fe^{3+}}(aq)$$

Step 3. The O.N. of iron undergoes an increase of 1 from +2 in $Fe^{2+}(aq)$ to +3 in $Fe^{3+}(aq)$.
The O.N. of chromium undergoes a decrease of 3 from +6 in $Cr_2O_7^{2-}(aq)$ to +3 in $Cr^{3+}(aq)$.

Step 4. For the total decrease in oxidation number to equal the total increase in oxidation number, the smallest whole number ratio of iron:chromium atoms is 3:1.

Step 5. Use the 3:1 ratio in **Step 4** to balance the iron and chromium atoms. There must be 3 iron atoms for each chromium atom. Consider that there are 2 chromium atoms on the left.
$Cr_2O_7^{2-}(aq) + 6Fe^{2+}(aq) \rightarrow Cr^{3+}(aq) + Fe^{3+}(aq)$

Step 6. Complete the balancing of iron by inspection. Balance for acidic conditions by adding $H_2O(l)$ and $H^+(aq)$.
$Cr_2O_7^{2-}(aq) + 6Fe^{2+}(aq) + 14H^+(aq) \rightarrow 2Cr^{3+}(aq) + 6Fe^{3+}(aq) + 7H_2O(l)$ (balanced)

b) $I_2(g) + NO_3^{-}(aq) \rightarrow IO_3^{-}(aq) + NO_2(g)$

Step 1. The unbalanced equation is $I_2(g) + NO_3^{-}(aq) \rightarrow IO_3^{-}(aq) + NO_2(g)$

Step 2. Assign an O.N. to each element.
0 +5 –2 +5 –2 +4 –2

$I_2(g) + NO_3^-(aq) \rightarrow IO_3^-(aq) + NO_2(g)$

Step 3. The O.N. of iodine undergoes an increase of 5 from 0 in $I_2(g)$ to +5 in $IO_3^-(aq)$. The O.N. of nitrogen undergoes a decrease of 1 from +5 in $NO_3^-(aq)$ to +4 in $NO_2(g)$.

Step 4. For the total decrease in oxidation number to equal the total increase in oxidation number, the smallest whole number ratio of nitrogen:iodine atoms is 5:1.

Step 5. Use the 5:1 ratio in **Step 4** to balance the nitrogen and iodine atoms. There must be 5 nitrogen atoms for each iodine atom.
$I_2(g) + 10NO_3^-(aq) \rightarrow IO_3^-(aq) + NO_2(g)$

Step 6. Complete the balancing of nitrogen by inspection. Balance for acidic conditions by adding $H_2O(l)$ and $H^+(aq)$.
$I_2(g) + 10NO_3^-(aq) + 8H^+(aq) \rightarrow 2IO_3^-(aq) + 10NO_2(g) + 4H_2O(l)$ (balanced)

c) $PbSO_4(aq) \rightarrow Pb(s) + PbO_2(aq) + SO_4^{2-}(aq)$

Step 1. The unbalanced equation is $PbSO_4(aq) \rightarrow Pb(s) + PbO_2(aq) + SO_4^{2-}(aq)$

Step 2. Assign an O.N. to each element.

$$\overset{+2\ +6\ -2}{PbSO_4}(aq) \rightarrow \overset{0}{Pb}(s) + \overset{+4\ -2}{PbO_2}(aq) + \overset{+6\ -2}{SO_4^{2-}}(aq)$$

Step 3. The O.N. of lead undergoes an increase of 2 from +2 in $PbSO_4(aq)$ to +4 in $PbO_2(aq)$. The O.N. of lead undergoes a decrease of 2 from +2 in $PbSO_4(aq)$ to 0 in Pb(s).

Step 4. Lead undergoes both an increase and a decrease in oxidation number (disproportionation). The whole number ratio of 1:1 ensures that the total decrease in oxidation number to equal the total increase in oxidation number.

Step 5. Use the 1:1 ratio in **Step 4** to balance the lead in Pb(s) and $PbO_2(aq)$.
$PbSO_4(aq) \rightarrow Pb(s) + PbO_2(aq) + SO_4^{2-}(aq)$

Step 6. Complete the balancing of lead and sulfur by inspection. Balance for acidic conditions by adding $H_2O(l)$ and $H^+(aq)$.
$2PbSO_4(aq) + 2H_2O(l) \rightarrow Pb(s) + PbO_2(aq) + 2SO_4^{2-}(aq) + 4H^+(aq)$ (balanced)

Check Your Solution
Since there is the same number of atoms of each element on both sides of the equations and the net charge is the same on both sides of the equations, the equation is balanced.

18.
Problem
Use the oxidation number method to balance each ionic equation in a basic solution:
a) $Cl^-(aq) + CrO_4^{2-}(aq) \rightarrow ClO^-(aq) + CrO_2^-(aq)$

b) $Ni(s) + MnO_4^-(aq) \rightarrow NiO(s) + MnO_2(s)$
c) $I^-(aq) + Ce^{4+}(aq) \rightarrow IO_3^-(aq) + Ce^{3+}(aq)$

What is Given?
The unbalanced equations are given with the formula of the reactants and products.

Plan Your Strategy
Assign an oxidation number (O.N.) to each element. Follow the rules for balancing an equation in basic conditions using O.N.

Act on Your Strategy
a) $Cl^-(aq) + CrO_4^{2-}(aq) \rightarrow ClO^-(aq) + CrO_2^-(aq)$

Step 1. The unbalanced equation is $Cl^-(aq) + CrO_4^{2-}(aq) \rightarrow ClO^-(aq) + CrO_2^-(aq)$

Step 2. Assign an O.N. to each element.
−1 +6 −2 +1 −2 +3 −2
$Cl^-(aq) + CrO_4^{2-}(aq) \rightarrow ClO^-(aq) + CrO_2^-(aq)$

Step 3. The O.N. of chlorine undergoes an increase of 2 from −1 in $Cl^-(aq)$ to +1 in $ClO^-(aq)$. The O.N. of chromium undergoes a decrease of 3 from +6 in $CrO_4^{2-}(aq)$ to +3 in $CrO_2^-(aq)$.

Step 4. For the total decrease in oxidation number to equal the total increase in oxidation number, the smallest whole number ratio of chlorine:chromium atoms is 3:2.

Step 5. Use the 3:2 ratio in **Step 4** to balance the chlorine and chromium atoms. There must be 3 chlorine atoms for 2 chromium atoms.
$3Cl^-(aq) + 2CrO_4^{2-}(aq) \rightarrow ClO^-(aq) + CrO_2^-(aq)$

Step 6. Complete the balancing of chlorine and chromium by inspection.
$3Cl^-(aq) + 2CrO_4^{2-}(aq) \rightarrow 3ClO^-(aq) + 2CrO_2^-(aq)$

Step 7. Add $H_2O(l)$ to balance the oxygen and $H^+(aq)$ to balance the hydrogen.
$3Cl^-(aq) + 2CrO_4^{2-}(aq) + 2H^+(aq) \rightarrow 3ClO^-(aq) + 2CrO_2^-(aq) + H_2O(l)$

Step 8. Add $OH^-(aq)$ to balance for basic conditions. Combine the $H^+(aq)$ and $OH^-(aq)$ to form water and eliminate any $H_2O(l)$ common to both sides of the equation.
$3Cl^-(aq) + 2CrO_4^{2-}(aq) + 2H^+(aq) + 2OH^-(aq) \rightarrow 3ClO^-(aq) + 2CrO_2^-(aq) + H_2O(l) + 2OH^-(aq)$
$3Cl^-(aq) + 2CrO_4^{2-}(aq) + 2H_2O(l) \rightarrow 3ClO^-(aq) + 2CrO_2^-(aq) + H_2O(l) + 2OH^-(aq)$

$3Cl^-(aq) + 2CrO_4^{2-}(aq) + H_2O(l) \rightarrow 3ClO^-(aq) + 2CrO_2^-(aq) + 2OH^-(aq)$ (balanced)

b) $Ni(s) + MnO_4^-(aq) \rightarrow NiO(s) + MnO_2(s)$
Step 1. The unbalanced equation is $Ni(s) + MnO_4^-(aq) \rightarrow NiO(s) + MnO_2(s)$

Step 2. Assign an O.N. to each element.

$\overset{0}{Ni}(s) + \overset{+7}{Mn}\overset{-2}{O_4}^-(aq) \rightarrow \overset{+2}{Ni}\overset{-2}{O}(s) + \overset{+4}{Mn}\overset{-2}{O_2}(s)$

Step 3. The O.N. of nickel undergoes an increase of 2 from 0 in Ni(s) to +2 in NiO(s).
The O.N. of manganese undergoes a decrease of 3 from +7 in MnO_4^-(aq) to +4 in MnO_2(s).

Step 4. For the total decrease in oxidation number to equal the total increase in oxidation number, the smallest whole number ratio of nickel:manganese atoms is 3:2.

Step 5. Use the 3:2 ratio in **Step 4** to balance the nickel and manganese atoms. There must be 3 nickel atoms for 2 manganese atoms.
$3Ni(s) + 2MnO_4^-(aq) \rightarrow NiO(s) + MnO_2(s)$

Step 6. Complete the balancing of chlorine and chromium by inspection.
$3Ni(s) + 2MnO_4^-(aq) \rightarrow 3NiO(s) + 2MnO_2(s)$

Step 7. Add $H_2O(l)$ to balance the oxygen and H^+(aq) to balance the hydrogen.
$3Ni(s) + 2MnO_4^-(aq) + 2H^+(aq) \rightarrow 3NiO(s) + 2MnO_2(s) + H_2O(l)$

Step 8. Add OH^-(aq) to balance for basic conditions. Combine the H^+(aq) and OH^-(aq) to form water and eliminate any $H_2O(l)$ common to both sides of the equation.
$3Ni(s) + 2MnO_4^-(aq) + 2H^+(aq) +2OH^-(aq) \rightarrow 3NiO(s) + 2MnO_2(s) + H_2O(l) + 2OH^-(aq)$
$3Ni(s) + 2MnO_4^-(aq) + 2H_2O(l) \rightarrow 3NiO(s) + 2MnO_2(s) + H_2O(l) + 2OH^-(aq)$

$3Ni(s) + 2MnO_4^-(aq) + H_2O(l) \rightarrow 3NiO(s) + 2MnO_2(s) + 2OH^-(aq)$ (balanced)

c) $I^-(aq) + Ce^{4+}(aq) \rightarrow IO_3^-(aq) + Ce^{3+}(aq)$
Step 1. The unbalanced equation is $I^-(aq) + Ce^{4+}(aq) \rightarrow IO_3^-(aq) + Ce^{3+}(aq)$

Step 2. Assign an O.N. to each element.
$\overset{-1}{I^-}(aq) + \overset{+4}{Ce^{4+}}(aq) \rightarrow \overset{+5}{I}\overset{-2}{O_3^-}(aq) + \overset{+3}{Ce^{3+}}(aq)$

Step 3. The O.N. of iodine undergoes an increase of 6 from −1 in I^-(aq) to +5 in IO_3^-(aq).
The O.N. of cerium undergoes a decrease of 1 from +4 in Ce^{4+}(aq) to +3 in Ce^{3+}(aq).

Step 4. For the total decrease in oxidation number to equal the total increase in oxidation number, the smallest whole number ratio of cerium:iodine atoms is 6:1.

Step 5. Use the 6:1 ratio in **Step 4** to balance the cerium and iodine atoms. There must be 6 cerium atoms for 1iodine atom.
$I^-(aq) + 6Ce^{4+}(aq) \rightarrow IO_3^-(aq) + Ce^{3+}(aq)$

Step 6. Complete the balancing of cerium by inspection.
$I^-(aq) + 6Ce^{4+}(aq) \rightarrow IO_3^-(aq) + 6Ce^{3+}(aq)$

Step 7. Add $H_2O(l)$ to balance the oxygen and $H^+(aq)$ to balance the hydrogen.
$I^-(aq) + 6Ce^{4+}(aq) + 3H_2O(l) \rightarrow IO_3^-(aq) + 6Ce^{3+}(aq) + 6H^+(aq)$

Step 8. Add $OH^-(aq)$ to balance for basic conditions. Combine the $H^+(aq)$ and $OH^-(aq)$ to form water and eliminate $H_2O(l)$ common to both sides of the equation.
$I^-(aq) + 6Ce^{4+}(aq) + 3H_2O(l) + 6OH^-(aq) \rightarrow IO_3^-(aq) + 6Ce^{3+}(aq) + 6H^+(aq) + 6OH^-(aq)$
$I^-(aq) + 6Ce^{4+}(aq) + 3H_2O(l) + 6OH^-(aq) \rightarrow IO_3^-(aq) + 6Ce^{3+}(aq) + 6H_2O(l)$

$I^-(aq) + 6Ce^{4+}(aq) + 6OH^-(aq) \rightarrow IO_3^-(aq) + 6Ce^{3+}(aq) + 3H_2O(l)$ (balanced)

Check Your Solution
Since there is the same number of atoms of each element on both sides of the equations and the net charge is the same on both sides of the equations, the equation is balanced.

19.
Problem
An analyst prepares a $H_2O_2(aq)$ sample solution by placing 1.284 g of $H_2O_2(aq)$ solution in a flask then diluting it with water and adding sulfuric acid to acidify it. The analyst titrates the $H_2O_2(aq)$ sample solution with 0.020 45 mol/L $KMnO_4(aq)$ and determines that 38.95 mL of $KMnO_4(aq)$ is required to reach the endpoint.
a) What is the mass of pure $H_2O_2(aq)$ that is present in the sample solution?
b) What is the mass percent of pure $H_2O_2(aq)$ in the sample solution?

What is Required?
You must calculate the mass of $H_2O_2(aq)$ in a sample and express this result in terms of mass percent.

What is Given?
$C\text{MnO}_4(aq)$ = 0.020 45 mol/L
$V\text{MnO}_4(aq)$ = 38.95 ml = 0.038 95 L
$m\text{H}_2\text{O}_2(aq)$ = 1.284 g

Plan Your Strategy
Write a balanced net ionic equation for the reaction between $MnO_4^-(aq)$ and $H_2O_2(aq)$ in acid solution. Calculate the number of moles (n) of $MnO_4^-(aq)$ using $n = CV$. Use the mol ratio from the balanced equation to calculate the number of moles of $H_2O_2(aq)$. Calculate the molar mass of $H_2O_2(aq)$ and find the mass of $H_2O_2(aq)$ using $m = n \times M$. The mass percent $H_2O_2(aq)$ can be calculated using

$$\text{mass percent} = \frac{\text{mass}\,H_2O_2(aq)}{\text{mass of solution}} \times 100\%$$

Act on Your Strategy
Following the rules for balancing a redox reaction in acid solution, the balanced net ionic equation is
$2MnO_4^-(aq) + 5H_2O_2(aq) + 6H^+(aq) \rightarrow 2Mn^{2+}(aq) + 5O_2(g) + 8H_2O(l)$

number of mol MnO_4(aq) = $C \times V = 0.020\ 45\ \text{mol/L} \times 0.038\ 95\ \text{L} = 0.000\ 7965$ mol

mol H_2O_2(aq) = 0.000 7965 mol MnO_4(aq) $\times \dfrac{5\,\text{mol}\,H_2O_2}{2\,\text{mol}MnO_4{}^-}$ = 0.001 991 mol H_2O_2

$M(H_2O_2)$ = 34.02 g/mol
mH_2O_2 = 0.001 991 × 34.02 g/mol = 0.067 74 g

mass percent = $\dfrac{\text{mass}\,H_2O_2}{\text{mass of solution}} \times 100\ \% = \dfrac{0.06774\,\text{g}}{1.284\,\text{g}} \times 100\% = 5.276\%$

Check Your Solution
The answer has the correct units (g and %), the correct number of significant digits (4) and the answer is close to the expected 6%.

20.
Problem
A forensic chemist wants to determine the level of alcohol in a sample of blood plasma. The chemist titrates the plasma with a solution of potassium dichromate. The balanced equation is

$$16H^+(aq) + 2Cr_2O_7{}^{2-}(aq) + C_2H_5OH(aq) \rightarrow 4Cr^{3+}(aq) + 2Cr^{3+}(aq) + 2CO_2(g) + 11H_2O(l)$$

If 32.35 mL of 0.050 23 mol/L $Cr_2O_7{}^{2-}$(aq) is required to titrate 27.00 g plasma, what is the mass percent of alcohol in the plasma?

What is Required?
You must calculate the mass percent of alcohol in the plasma.

What is Given?
The balanced net ionic equation for the redox reaction is given.
mass plasma = 27.99 g
volume $Cr_2O_7{}^{2-}$(aq) = V =32.35 L = 0.032 35 L
concentration $Cr_2O_7{}^{2-}$(aq) = C = 0.050 23 mol/L

Plan Your Strategy
Calculate the number of moles (n) of $Cr_2O_7{}^{2-}$(aq) using $n = CV$. Use the mol ratio from the balanced equation to calculate the number of moles of C_2H_5OH(aq). Calculate the molar mass of C_2H_5OH(aq) and find the mass of C_2H_5OH(aq) using $m = n \times M$. The mass percent C_2H_5OH(aq) can be calculated using

mass percent = $\dfrac{\text{mass}\ C_2H_5OH}{\text{mass of solution}} \times 100\%$

Act on Your Strategy
amount $Cr_2O_7^{2-}$(aq) $= CV$ = 0.050 23 mol/L × 0.032 35 L = 0.001 625 mol

amount C_2H_5OH(aq) = 0.001 625 mol $Cr_2O_7^{2-}$(aq) × $\dfrac{1\,\text{mol}\,C_2H_5OH}{2\,\text{mol}\,Cr_2O_7^{\,2-}}$ = 0.000 8125 mol

$M(C_2H_5OH)$ = 46.08 g/mol
mass C_2H_5OH(aq) = 0.000 8125 mol × 46.08 g/mol = 0.037 44 g

mass percent = $\dfrac{\text{mass}\ C_2H_5OH}{\text{mass of solution}}$ × 100% = $\dfrac{0.037\,44\,\text{g}}{27.00\,\text{g}}$ × 100% = 0.1387%

Check Your Solution
The answer has the correct unit (%), and the correct number of significant digits (4).

21.
Problem
An analyst titrates an acidified solution containing 0.153 g of purified sodium oxalate, $Na_2C_2O_4$(aq), with a potassium permanganate solution, $KMnO_4$(aq). The light purple endpoint is reached when the chemist has added 41.45 mL of potassium permanganate solution. What is the molar concentration of the potassium permanganate solution? The balanced equation is

$$2MnO_4^-(aq) + 5Na_2C_2O_4(aq) + 16H^+(aq) \rightarrow 10Na^+(aq) + 2Mn^{2+}(aq) + 10CO_2(g) + 8H_2O(l)$$

What is Required?
You must calculate the molar concentration of a potassium permanganate solution.

What is Given?
mass $Na_2C_2O_4$(aq) = 0.153 g
volume $KMnO_4$(aq) = V = 41.45 mL = 0.041 45 L

Plan Your Strategy
Determine the molar mass of $Na_2C_2O_4$(aq) and calculate the number of moles of this reactant using $n = \dfrac{m}{M}$. Use the mol ratio in the balanced equation to calculate the amount of MnO_4^-(aq). Calculate the concentration of $KMnO_4$(aq) using $C = \dfrac{n}{V}$.

Act on Your Strategy
$M Na_2C_2O_4$(aq) = 134.00 g/mol
amount of $Na_2C_2O_4$(aq) = amount of $C_2O_4^{2-}$(aq), and amount of $KMnO_4$(aq) = amount of MnO_4^-(aq)

amount $C_2O_4^{2-}$(aq) = $\dfrac{0.153\,\text{g}}{134.00\,\text{g}}$ = 0.001 142 mol

amount MnO_4^{-}(aq) = 0.001 142 mol $C_2O_4^{2-}$(aq) × $\dfrac{2\,\text{mol}\,MnO_4^{-}}{5\,\text{mol}\,C_2O_4^{2-}}$ = 0.0004567 mol

concentration $KMnO_4$ (aq) = $\dfrac{n}{V}$ = $\dfrac{0.004567\,\text{mol}}{0.04145\,\text{L}}$ = 0.0110 mol/L

Check Your Solution
The answer has the correct unit (mol/L) and the correct number of significant digits (3).

22.
Problem
25.00 mL of a solution containing iron(II) ions was titrated with a 0.020 43 mol/L potassium dichromate solution. The endpoint was reached when 35.55 mL of potassium dichromate solution had been added. What was the molar concentration of iron(II) ions in the original, acidic solution? The *unbalanced* equation is

$$Cr_2O_7^{2-}(aq) + Fe^{2+}(aq) \rightarrow Cr^{3+}(aq) + Fe^{3+}(aq)$$

What is Required?
You must calculate the final molar concentration of the Fe^{2+}(aq) in the original solution.

What is Given?
concentration $Cr_2O_7^{2-}$(aq) = C = 0.020 43 mol/L
volume $Cr_2O_7^{2-}$(aq) = V = 35.55 mL = 0.035 55 L
original volume $Cr_2O_7^{2-}$(aq) = 25.00 mL = 0.025 00 L

Plan Your Strategy
Follow the rules for balancing a redox reaction in acid solution to obtain the balanced equation for this reaction. Calculate the amount of $Cr_2O_7^{2-}$(aq) that react, $n = CV$. Use the mole ratio from the balanced equation to determine the amount of Fe^{2+}(aq) that reacted. Calculate the concentration of Fe^{2+}(aq), $C = \dfrac{n}{V}$

Act on Your Strategy
The balanced equation for this reaction is

$$Cr_2O_7^{2-}(aq) + 6Fe^{2+}(aq) + 14H^{+}(aq) \rightarrow 2Cr^{3+}(aq) + 6Fe^{3+}(aq) + 7H_2O(l)$$

amount $Cr_2O_7^{2-}$(aq) = 0.020 43 mol/L × 0.035 55 L = 0.000 7263 mol amount Fe^{2+}(aq) = 0.000 7263 mol $Cr_2O_7^{2-}$(aq) $\times \dfrac{6\,\text{mol}\,Fe^{2+}}{1\,\text{mol}\,Cr_2O_7^{\;2-}}$ = 0.004 358 mol

concentration of Fe^{2+}(aq) = $\dfrac{0.004\,358\,\text{mol}}{0.025\,00\,\text{L}}$ = 0.1743 mol/L

Check Your Solution
The answer has the correct unit (mol/L) and the correct number of significant digits (4).