Chemistry 12
Solutions Manual Part B

Unit 5 Electrochemistry
Solutions to Practice Problems in
Chapter 9 Oxidation-Reduction Reactions

Balancing an Equation for a Reaction That Occurs in an Acidic Solution
Balancing an Equation for a Reaction That Occurs in a Basic Solution
(Student textbook page 598)

1. Balance the following ionic equation for acidic conditions. Identify the oxidizing agent and the reducing agent.

$$MnO_4^{-}(aq) + Ag(s) \rightarrow Mn^{2+}(aq) + Ag^{+}(aq)$$

What Is Required?
You need to balance the equation for the redox reaction in acidic conditions and identify the oxidizing agent and reducing agent in the reaction.

What Is Given?
You are given the unbalanced equation: $MnO_4^{-}(aq) + Ag(s) \rightarrow Mn^{2+}(aq) + Ag^{+}(aq)$

Plan Your Strategy	Act on Your Strategy
Step 1 Write the unbalanced half-reactions. Include only those compounds that contain the atom that is oxidized or the atom that is reduced. (For this step, ignore the fact that one side of the equation might have oxygen atoms and the other side has none.)	*Oxidation half-reaction*: $Ag(s) \rightarrow Ag^{+}(aq)$ *Reduction half- reaction*: $MnO_4^{-}(aq) \rightarrow Mn^{2+}(aq)$
Step 2 Balance atoms other than oxygen and hydrogen.	$Ag(s) \rightarrow Ag^{+}(aq)$ (already balanced) $MnO_4^{-}(aq) \rightarrow Mn^{2+}(aq)$ (already balanced)
Step 3 Balance oxygen atoms by adding water molecules. *Oxidation half-reaction:* Already balanced. *Reduction half-reaction:* There are four oxygen atoms on the left side of the equation and none on the right. Add four water molecules to the right side.	$Ag(s) \rightarrow Ag^{+}(aq)$ (already balanced) $MnO_4^{-}(aq) \rightarrow Mn^{2+}(aq) + 4H_2O(\ell)$

Step 4 Balance hydrogen atoms by adding hydrogen ions. *Reduction half-reaction*: There are eight hydrogen atoms in the water molecules on the right side of the equation, so add eight hydrogen ions to the left side. Since the half-reaction is to be balanced in acidic conditions, these hydrogen ions will remain in the equation.	$Ag(s) \rightarrow Ag^{+}(aq)$ (already balanced) $8H^{+}(aq) + MnO_4^{-}(aq) \rightarrow Mn^{2+}(aq) + 4H_2O(\ell)$
Step 5 Balance the charges by adding electrons. *Oxidation half-reaction*: There are zero net charges on the left side of the equation and one positive charge on the right side. Therefore, give the right side a net charge of zero by adding one electron to the right side. *Reduction half-reaction*: There are eight positive charges and one negative charge on the left side of the equation, giving it a net charge of 7+. The net charge on the right side of the equation is 2+. Add five electrons to the left side to give it a net 2+ charge.	$Ag(s) \rightarrow Ag^{+}(aq) + 1e^{-}$ $8H^{+}(aq) + MnO_4^{-}(aq) + 5e^{-} \rightarrow Mn^{2+}(aq) + 4H_2O(\ell)$
Step 6 Multiply one or both half-reactions by the number that will bring the number of electrons to the lowest common multiple. *Oxidation half-reaction*: Multiply the reactants and products by 5. *Reduction half-reaction*: Multiply the reactants and products by 1.	$5Ag(s) \rightarrow 5Ag^{+}(aq) + 5e^{-}$ $8H^{+}(aq) + MnO_4^{-}(aq) + 5e^{-} \rightarrow Mn^{2+}(aq) + 4H_2O(\ell)$
Step 7 Add the balanced half-reactions.	$5Ag(s) + 8H^{+}(aq) + MnO_4^{-}(aq) + 5e^{-} \rightarrow Mn^{2+}(aq) + 4H_2O(\ell)\ 5Ag^{+}(aq) + 5e^{-}$
Step 8 Cancel the electrons present on both sides of the equation.	$5Ag(s) + 8H^{+}(aq) + MnO_4^{-}(aq) \rightarrow Mn^{2+}(aq) + 4H_2O(\ell) + 5Ag^{+}(aq)$
Step 9 Identify the oxidizing agent and reducing agent. The reducing agent gets oxidized and the oxidizing agent gets reduced.	Since Ag(s) gets oxidized, it is the reducing agent and since $MnO_4^{-}(aq)$ gets reduced, it is the oxidizing agent.

Check Your Solution

The number of atoms of each element is equal on each side of the equation and the total charge on each side is 7+, indicating that both mass and charge are balanced. The equation is balanced.

2. Balance the following ionic equation for acidic conditions. Identify the oxidizing agent and the reducing agent.

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

What Is Required?

You need to balance the equation for the redox reaction in acidic conditions and identify the oxidizing agent and reducing agent in the reaction.

What Is Given?

You are given the unbalanced equation: $Hg(\ell) + NO_3^-(aq) + Cl^-(aq) \rightarrow HgCl_4^{2-}(s) + NO_2(g)$

Plan Your Strategy	Act on Your Strategy
Step 1 Write the unbalanced half-reactions. Include only those compounds that contain the atom that is oxidized or the atom that is reduced. (For this step, ignore the fact that one side of the equation might have oxygen atoms and the other side has none.)	*Oxidation half-reaction*: $Hg(\ell) + Cl^-(aq) \rightarrow HgCl_4^{2-}(s)$ *Reduction half- reaction*: $NO_3^-(aq) \rightarrow NO_2(g)$
Step 2 Balance atoms other than oxygen and hydrogen.	$Hg(\ell) + 4Cl^-(aq) \rightarrow HgCl_4^{2-}(s)$ $NO_3^-(aq) \rightarrow NO_2(g)$ (already balanced for atoms other than oxygen and hydrogen)
Step 3 Balance oxygen atoms by adding water molecules. *Oxidation half-reaction:* Already balanced. *Reduction half-reaction:* There are three oxygen atoms on the left side of the equation and two on the right. Add one water molecule to the right side.	$Hg(\ell) + 4Cl^-(aq) \rightarrow HgCl_4^{2-}(s)$ (already balanced) $NO_3^-(aq) \rightarrow NO_2(g) + H_2O(\ell)$
Step 4 Balance hydrogen atoms by adding hydrogen ions. *Oxidation half-reaction:* Already balanced. *Reduction half-reaction*: There are two hydrogen atoms in the water molecules on the right side of the equation, so add two hydrogen ions to the left side. Since the half-reaction is to be balanced in acidic conditions, these hydrogen ions will remain in the equation.	$Hg(\ell) + 4Cl^-(aq) \rightarrow HgCl_4^{2-}(s)$ (already balanced) $2H^+(aq) + NO_3^-(aq) \rightarrow NO_2(g) + H_2O(\ell)$

Step 5 Balance the charges by adding electrons. *Oxidation half-reaction*: There is a net charge of 4− on the left side of the equation and 2− on the right side. Therefore, give the right side a net charge of 4− by adding two electrons to the right side. *Reduction half-reaction*: There is a net charge of 1+ on the left side of the equation. The net charge on the right side of the equation is zero. Add one electron to the left side to give it a net charge of zero.	$Hg(\ell) + 4Cl^{-}(aq) \rightarrow HgCl_4^{2-}(s) + 2e^{-}$ $2H^{+}(aq) + NO_3^{-}(aq) + 1e^{-} \rightarrow NO_2(g) + H_2O(\ell)$
Step 6 Multiply one or both half-reactions by the number that will bring the number of electrons to the lowest common multiple. *Oxidation half-reaction*: Multiply the reactants and products by 1. *Reduction half-reaction*: Multiply the reactants and products by 2.	$Hg(\ell) + 4Cl^{-}(aq) \rightarrow HgCl_4^{2-}(s) + 2e^{-}$ $4H^{+}(aq) + 2NO_3^{-}(aq) + 2e^{-} \rightarrow$ $2NO_2(g) + 2H_2O(\ell)$
Step 7 Add the balanced half-reactions.	$Hg(\ell) + 4Cl^{-}(aq) + 4H^{+}(aq) + 2NO_3^{-}(aq) + 2e^{-} \rightarrow$ $2NO_2(g) + 2H_2O(\ell) + HgCl_4^{2-}(s) + 2e^{-}$
Step 8 Cancel the electrons present on both sides of the equation.	$Hg(\ell) + 4Cl^{-}(aq) + 4H^{+}(aq) + 2NO_3^{-}(aq) \rightarrow$ $2NO_2(g) + 2H_2O(\ell) + HgCl_4^{2-}(s)$
Step 9 Identify the oxidizing agent and reducing agent. The reducing agent gets oxidized and the oxidizing agent gets reduced.	Since $Hg(\ell)$ gets oxidized, it is the reducing agent, and since $NO_3^{-}(aq)$ gets reduced, it is the oxidizing agent.

Check Your Solution

The number of atoms of each element is equal on each side of the equation and the total charge on each side is 2–, indicating that both mass and charge are balanced. The equation is balanced.

3. Balance the following ionic equation for acidic conditions. Identify the oxidizing agent and the reducing agent.

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

What Is Required?
You need to balance the equation for the redox reaction in acidic conditions and identify the oxidizing agent and reducing agent in the reaction.

What Is Given?
You are given the unbalanced equation: $AsH_3(s) + Zn^{2+}(aq) \rightarrow H_3AsO_4(aq) + Zn(s)$

Plan Your Strategy	Act on Your Strategy
Step 1 Write the unbalanced half-reactions. Include only those compounds that contain the atom that is oxidized or the atom that is reduced. (For this step, ignore the fact that one side of the equation might have oxygen atoms and the other side has none.)	*Oxidation half-reaction*: $AsH_3(s) \rightarrow H_3AsO_4(aq)$ *Reduction half- reaction*: $Zn^{2+}(aq) \rightarrow Zn(s)$
Step 2 Balance atoms other than oxygen and hydrogen.	$AsH_3(s) \rightarrow H_3AsO_4(aq)$ $Zn^{2+}(aq) \rightarrow Zn(s)$ (both are already balanced for atoms other than oxygen and hydrogen)
Step 3 Balance oxygen atoms by adding water molecules. *Oxidation half-reaction:* There are 4 oxygen atoms on the right side and none on the left side. Therefore, add four water molecules to the left side to balance the oxygen atoms. *Reduction half-reaction:* Already balanced.	$AsH_3(s) + 4H_2O(\ell) \rightarrow H_3AsO_4(aq)$ $Zn^{2+}(aq) \rightarrow Zn(s)$ (already balanced)
Step 4 Balance hydrogen atoms by adding hydrogen ions. *Oxidation half-reaction:* There are 3 hydrogen atoms in total on the right side of the equation, and 11 hydrogen atoms on the left side, so add 8 hydrogen ions to the right side. Since the half-reaction is to be balanced in acidic conditions, these hydrogen ions will remain in the equation. *Reduction half-reaction*: Already balanced.	$AsH_3(s) + 4H_2O(\ell) \rightarrow H_3AsO_4(aq) + 8H^+(aq)$ $Zn^{2+}(aq) \rightarrow Zn(s)$ (already balanced)

Step 5 Balance the charges by adding electrons. *Oxidation half-reaction*: There is a net charge of zero on the left side of the equation and 8+ on the right side. Therefore, give the right side a net charge of zero by adding 8 electrons to the right side. *Reduction half-reaction*: There is a net charge of 2+ on the left side of the equation. The net charge on the right side of the equation is zero. Add two electrons to the left side to give it a net charge of zero.	$AsH_3(s) + 4H_2O(\ell) \rightarrow H_3AsO_4(aq) + 8H^+(aq) + 8e^-$ $Zn^{2+}(aq) + 2e^- \rightarrow Zn(s)$
Step 6 Multiply one or both half-reactions by the number that will bring the number of electrons to the lowest common multiple. *Oxidation half-reaction*: Multiply the reactants and products by 1. *Reduction half-reaction*: Multiply the reactants and products by 4.	$AsH_3(s) + 4H_2O(\ell) \rightarrow H_3AsO_4(aq) + 8H^+(aq) + 8e^-$ $4Zn^{2+}(aq) + 8e^- \rightarrow 4Zn(s)$
Step 7 Add the balanced half-reactions.	$AsH_3(s) + 4H_2O(\ell) + 4Zn^{2+}(aq) + 8e^- \rightarrow H_3AsO_4(aq) + 8H^+(aq) + 8e^- + 4Zn(s)$
Step 8 Cancel the electrons present on both sides of the equation.	$AsH_3(s) + 4H_2O(\ell) + 4Zn^{2+}(aq) \rightarrow H_3AsO_4(aq) + 8H^+(aq) + 4Zn(s)$
Step 9 Identify the oxidizing agent and reducing agent. The reducing agent gets oxidized and the oxidizing agent gets reduced.	Since $AsH_3(s)$ gets oxidized, it is the reducing agent and since $Zn^{2+}(aq)$ gets reduced, it is the oxidizing agent.

Check Your Solution

The number of atoms of each element is equal on each side of the equation and the total charge on each side is 8+, indicating that both mass and charge are balanced. The equation is balanced.

4. Balance the following ionic equation for acidic conditions. Identify the oxidizing agent and the reducing agent.

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

What Is Required?
You need to balance the equation for the redox reaction in acidic conditions and identify the oxidizing agent and reducing agent in the reaction.

What Is Given?
You are given the unbalanced equation: $I_2(s) + ClO^-(aq) \rightarrow IO_3^-(aq) + Cl^-(aq)$

Plan Your Strategy	Act on Your Strategy
Step 1 Write the unbalanced half-reactions. Include only those compounds that contain the atom that is oxidized or the atom that is reduced. (For this step, ignore the fact that one side of the equation might have oxygen atoms and the other side has none.)	*Oxidation half-reaction*: $I_2(s) \rightarrow IO_3^-(aq)$ *Reduction half- reaction*: $ClO^-(aq) \rightarrow Cl^-(aq)$
Step 2 Balance atoms other than oxygen and hydrogen.	$I_2(s) \rightarrow 2IO_3^-(aq)$ $ClO^-(aq) \rightarrow Cl^-(aq)$ (already balanced for atoms other than oxygen and hydrogen)
Step 3 Balance oxygen atoms by adding water molecules. *Oxidation half-reaction:* There are 6 oxygen atoms on the right side and none on the left side. Therefore, add 6 water molecules to the left side to balance the oxygen atoms. *Reduction half-reaction:* There is one oxygen atom on the left side and none on the right, so add one water molecule to the right side.	$6H_2O(\ell) + I_2(s) \rightarrow 2IO_3^-(aq)$ $ClO^-(aq) \rightarrow Cl^-(aq) + H_2O(\ell)$
Step 4 Balance hydrogen atoms by adding hydrogen ions. *Oxidation half-reaction:* There are 12 hydrogen atoms in total on the left side of the equation, and none on the right side, so add 12 hydrogen ions to the right side. *Reduction half-reaction*: There are no hydrogen atoms on the left side and 2 on the right, so add 2 hydrogen ions to the left side.	$6H_2O(\ell) + I_2(s) \rightarrow 2IO_3^-(aq) + 12H^+(aq)$ $2H^+(aq) + ClO^-(aq) \rightarrow Cl^-(aq) + H_2O(\ell)$

Step 5 Balance the charges by adding electrons. *Oxidation half-reaction*: There is a net charge of zero on the left side of the equation and 10+ on the right side. Therefore, give the right side a net charge of zero by adding 10 electrons to the right side. *Reduction half-reaction*: There is a net charge of 1+ on the left side of the equation and 1– on the right. Add two electrons to the left side to give it a net charge of 1–.	$6H_2O(\ell) + I_2(s) \rightarrow 2IO_3^-(aq) + 12H^+(aq) + 10e^-$ $2e^- + 2H^+(aq) + ClO^-(aq) \rightarrow Cl^-(aq) + H_2O(\ell)$
Step 6 Multiply one or both half-reactions by the number that will bring the number of electrons to the lowest common multiple. *Oxidation half-reaction*: Multiply the reactants and products by 1. *Reduction half-reaction*: Multiply the reactants and products by 5.	$6H_2O(\ell) + I_2(s) \rightarrow 2IO_3^-(aq) + 12H^+(aq) + 10e^-$ $10e^- + 10H^+(aq) + 5ClO^-(aq) \rightarrow 5Cl^-(aq) + 5H_2O(\ell)$
Step 7 Add the balanced half-reactions.	$10e^- + 10H^+(aq) + 5ClO^-(aq) + 6H_2O(\ell) + I_2(s) \rightarrow 2IO_3^-(aq) + 12H^+(aq) + 10e^- + 5Cl^-(aq) + 5H_2O(\ell)$
Step 8 Cancel the electrons and common ions or atoms present on both sides of the equation.	$5ClO^-(aq) + H_2O(\ell) + I_2(s) \rightarrow 2IO_3^-(aq) + 2H^+(aq) + 5Cl^-(aq)$
Step 9 Identify the oxidizing agent and reducing agent. The reducing agent gets oxidized and the oxidizing agent gets reduced.	Since $I_2(s)$ gets oxidized, it is the reducing agent and since $ClO^-(aq)$ gets reduced, it is the oxidizing agent.

Check Your Solution

The number of atoms of each element is equal on each side of the equation and the total charge on each side is 5–, indicating that both mass and charge are balanced. The equation is balanced.

5. Balance the following ionic equation for basic conditions. Identify the oxidizing agent and the reducing agent.

$$MnO_4^-(aq) + I^-(aq) \rightarrow MnO_4^{2-}(aq) + IO_3^-(aq)$$

What Is Required?
You need to balance the equation for the redox reaction in basic conditions and identify the oxidizing agent and reducing agent in the reaction.

What Is Given?
You are given the unbalanced equation: $MnO_4^-(aq) + I^-(aq) \rightarrow MnO_4^{2-}(aq) + IO_3^-(aq)$

Plan Your Strategy	Act on Your Strategy
Step 1 Write the unbalanced half-reactions. Include only those compounds that contain the atom that is oxidized or the atom that is reduced. (For this step, ignore the fact that one side of the equation might have oxygen atoms and the other side has none.)	*Oxidation half-reaction*: $I^-(aq) \rightarrow IO_3^-(aq)$ *Reduction half- reaction*: $MnO_4^-(aq) \rightarrow MnO_4^{2-}(aq)$
Step 2 Balance atoms other than oxygen and hydrogen.	Equations for both half reactions are already balanced for atoms other than oxygen and hydrogen
Step 3 Balance oxygen atoms by adding water molecules. *Oxidation half-reaction:* There are 3 oxygen atoms on the right side and none on the left side. Therefore, add 3 water molecules to the left side to balance the oxygen atoms. *Reduction half-reaction:* There are 4 oxygen atoms on each side of the equation, therefore this is already balanced for oxygen atoms.	$3H_2O(\ell) + I^-(aq) \rightarrow IO_3^-(aq)$ $MnO_4^-(aq) \rightarrow MnO_4^{2-}(aq)$ (already balanced for oxygen atoms)
Step 4 Balance hydrogen atoms by adding hydrogen ions. *Oxidation half-reaction:* There are 6 hydrogen atoms on the left side of the equation, and none on the right side, so add 6 hydrogen ions to the right side. *Reduction half-reaction*: Already balanced for hydrogen atoms.	$3H_2O(\ell) + I^-(aq) \rightarrow IO_3^-(aq) + 6H^+(aq)$ $MnO_4^-(aq) \rightarrow MnO_4^{2-}(aq)$ (already balanced for hydrogen atoms)
Step 5 Adjust for basic conditions. *Oxidation half-reaction:* Add 6 hydroxide ions to each side of the	$6OH^-(aq) + 3H_2O(\ell) + I^-(aq) \rightarrow IO_3^-(aq) + 6H_2O(\ell)$ $MnO_4^-(aq) \rightarrow MnO_4^{2-}(aq)$

equation, forming 6 water molecules on the right side.	
Step 6 Cancel any water molecules present on both sides of the equation. *Oxidation half-reaction:* remove 3 water molecules from each side of the equation, leaving three water molecules on the right.	$6OH^-(aq) + I^-(aq) \rightarrow IO_3^-(aq) + 3H_2O(\ell)$ $MnO_4^-(aq) \rightarrow MnO_4^{2-}(aq)$
Step 7 Balance the charges by adding electrons. *Oxidation half-reaction*: There is a net charge of 7– on the left side of the equation and 1– on the right side. Therefore, give the right side a net charge of 7– by adding 6 electrons to the right side. *Reduction half-reaction*: There is a net charge of 1– on the left side of the equation and 2– on the right. Add one electron to the left side to give it a net charge of 2–.	$6OH^-(aq) + I^-(aq) \rightarrow IO_3^-(aq) + 3H_2O(\ell) + 6\,e^-$ $1e^- + MnO_4^-(aq) \rightarrow MnO_4^{2-}(aq)$
Step 8 Multiply one or both half-reactions by the number that will bring the number of electrons to the lowest common multiple. *Oxidation half-reaction*: Multiply the reactants and products by 1. *Reduction half-reaction*: Multiply the reactants and products by 6.	$6OH^-(aq) + I^-(aq) \rightarrow IO_3^-(aq) + 3H_2O(\ell) + 6\,e^-$ $6e^- + 6MnO_4^-(aq) \rightarrow 6MnO_4^{2-}(aq)$
Step 9 Add the balanced half-reactions.	$6OH^-(aq) + I^-(aq) + 6e^- + 6MnO_4^-(aq) \rightarrow$ $IO_3^-(aq) + 3H_2O(\ell) + 6e^- + 6MnO_4^{2-}(aq)$
Step 10 Cancel the electrons present on both sides of the equation.	$6OH^-(aq) + I^-(aq) + 6MnO_4^-(aq) \rightarrow$ $IO_3^-(aq) + 3H_2O(\ell) + 6MnO_4^{2-}(aq)$
Step 11 Identify the oxidizing agent and reducing agent. The reducing agent gets oxidized and the oxidizing agent gets reduced.	Since $I^-(aq)$ gets oxidized, it is the reducing agent; since $MnO_4^-(aq)$ gets reduced, it is the oxidizing agent.

Check Your Solution

The number of atoms of each element is equal on each side of the equation and the total charge on each side is 13–, indicating that both mass and charge are balanced. The equation is balanced.

6. Balance the following ionic equation for basic conditions. Identify the oxidizing agent and the reducing agent.

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

What Is Required?
You need to balance the equation for the redox reaction in basic conditions and identify the oxidizing agent and reducing agent in the reaction.

What Is Given?
You are given the unbalanced equation: $H_2O_2(aq) + ClO_2(aq) \rightarrow ClO^-(aq) + O_2(g)$

Plan Your Strategy	Act on Your Strategy
Step 1 Write the unbalanced half-reactions. Include only those compounds that contain the atom that is oxidized or the atom that is reduced. (For this step, ignore the fact that one side of the equation might have oxygen atoms and the other side has none.)	*Oxidation half-reaction*: $H_2O_2(aq) \rightarrow O_2(g)$ *Reduction half- reaction*: $ClO_2(aq) \rightarrow ClO^-(aq)$
Step 2 Balance atoms other than oxygen and hydrogen.	Both equations are already balanced for atoms other than oxygen and hydrogen
Step 3 Balance oxygen atoms by adding water molecules. *Oxidation half-reaction:* Already balanced for oxygen atoms. *Reduction half-reaction:* There are 2 oxygen atoms on the left side of the equation and only one on the right, so add a water molecule to the right side to balance the oxygen atoms.	$H_2O_2(aq) \rightarrow$ $O_2(g)$ (already balanced for oxygen atoms) $ClO_2(aq) \rightarrow ClO^-(aq) + H_2O(\ell)$
Step 4 Balance hydrogen atoms by adding hydrogen ions. *Oxidation half-reaction:* There are 2 hydrogen atoms on the left side of the equation, and none on the right side, so add 2 hydrogen ions to the right side. *Reduction half-reaction*: There are 2 hydrogen atoms on the right side of the equation, and none on the left side, so add 2 hydrogen ions to the left side.	$H_2O_2(aq) \rightarrow O_2(g) + 2H^+(aq)$ $2H^+(aq) + ClO_2(aq) \rightarrow ClO^-(aq) + H_2O(\ell)$

Step 5 Adjust for basic conditions. *Oxidation half-reaction:* Add 2 hydroxide ions to each side of the equation, forming 2 water molecules on the right side. *Reduction half-reaction:* Add 2 hydroxide ions to each side of the equation, forming 2 water molecules on the left side.	$2OH^-(aq) + H_2O_2(aq) \rightarrow O_2(g) + 2H_2O(\ell)$ $2H_2O(\ell) + ClO_2(aq) \rightarrow$ $ClO^-(aq) + H_2O(\ell) + 2OH^-(aq)$
Step 6 Cancel any water molecules present on both sides of the equation. *Reduction half-reaction:* remove 1 water molecule from each side of the equation, leaving one water molecule on the left.	$2OH^-(aq) + H_2O_2(aq) \rightarrow O_2(g) + 2H_2O(\ell)$ $H_2O(\ell) + ClO_2(aq) \rightarrow ClO^-(aq) + 2OH^-(aq)$
Step 7 Balance the charges by adding electrons. *Oxidation half-reaction*: There is a net charge of 2– on the left side of the equation and zero on the right side. Therefore, give the right side a net charge of 2– by adding 2 electrons to the right side. *Reduction half-reaction*: There is a net charge of zero on the left side of the equation and 3– on the right. Add three electrons to the left side to give it a net charge of 3–.	$2OH^-(aq) + H_2O_2(aq) \rightarrow O_2(g) + 2H_2O(\ell) + 2e^-$ $3e^- + H_2O(\ell) + ClO_2(aq) \rightarrow$ $ClO^-(aq) + 2OH^-(aq)$
Step 8 Multiply one or both half-reactions by the number that will bring the number of electrons to the lowest common multiple. *Oxidation half-reaction*: Multiply the reactants and products by 3. *Reduction half-reaction*: Multiply the reactants and products by 2.	$6OH^-(aq) + 3H_2O_2(aq) \rightarrow$ $3O_2(g) + 6H_2O(\ell) + 6e^-$ $6e^- + 2H_2O(\ell) + 2ClO_2(aq) \rightarrow$ $2ClO^-(aq) + 4OH^-(aq)$
Step 9 Add the balanced half-reactions.	$6OH^-(aq) + 3H_2O_2(aq) + 6e^- + 2H_2O(\ell) +$ $2ClO_2(aq) \rightarrow$ $2ClO^-(aq) + 4OH^-(aq) + 3O_2(g) + 6H_2O(\ell) + 6e^-$

Step 10 Cancel the electrons, water molecules and hydroxide ions present on both sides of the equation.	$2OH^-(aq) + 3H_2O_2(aq) + 2ClO_2(aq) \rightarrow 2ClO^-(aq) + 3O_2(g) + 4H_2O(\ell)$
Step 11 Identify the oxidizing agent and reducing agent. The reducing agent gets oxidized and the oxidizing agent gets reduced.	Since $H_2O_2(aq)$ gets oxidized, it is the reducing agent; since $ClO_2(aq)$ gets reduced, it is the oxidizing agent.

Check Your Solution

The number of atoms of each element is equal on each side of the equation and the total charge on each side is 2–, indicating that both mass and charge are balanced. The equation is balanced.

7. Balance the following ionic equation for basic conditions. Identify the oxidizing agent and the reducing agent.

$$ClO^{-}(aq) + CrO_2^{-}(aq) \rightarrow CrO_4^{2-}(aq) + Cl_2(g)$$

What Is Required?
You need to balance the equation for the redox reaction in basic conditions and identify the oxidizing agent and reducing agent in the reaction.

What Is Given?
You are given the unbalanced equation: $ClO^{-}(aq) + CrO_2^{-}(aq) \rightarrow CrO_4^{2-}(aq) + Cl_2(g)$

Plan Your Strategy	Act on Your Strategy
Step 1 Write the unbalanced half-reactions. Include only those compounds that contain the atom that is oxidized or the atom that is reduced. (For this step, ignore the fact that one side of the equation might have oxygen atoms and the other side has none.)	*Oxidation half-reaction*: $CrO_2^{-}(aq) \rightarrow CrO_4^{2-}(aq)$ *Reduction half- reaction*: $ClO^{-}(aq) \rightarrow Cl_2(g)$
Step 2 Balance atoms other than oxygen and hydrogen. *Oxidation half-reaction:* Already balanced. *Reduction half-reaction:* There are two atoms of chlorine on the right side and only one on the left side, so balance with a 2 in front of the $ClO^{-}(aq)$.	$CrO_2^{-}(aq) \rightarrow CrO_4^{2-}(aq)$ $2ClO^{-}(aq) \rightarrow Cl_2(g)$
Step 3 Balance oxygen atoms by adding water molecules. *Oxidation half-reaction:* There are two oxygen atoms on the left side and four on the right side, so add 2 water molecules to the left side. *Reduction half-reaction:* There are 2 oxygen atoms on the left side of the equation and none on the right, so add two water molecules to the right side to balance the oxygen atoms.	$2H_2O(\ell) + CrO_2^{-}(aq) \rightarrow CrO_4^{2-}(aq)$ $2ClO^{-}(aq) \rightarrow Cl_2(g) + 2H_2O(\ell)$

Step 4 Balance hydrogen atoms by adding hydrogen ions. *Oxidation half-reaction:* There are 4 hydrogen atoms on the left side of the equation, and none on the right side, so add 4 hydrogen ions to the right side. *Reduction half-reaction*: There are 4 hydrogen atoms on the right side of the equation and none on the left side, so add 4 hydrogen ions to the left side.	$2H_2O(\ell) + CrO_2^-(aq) \rightarrow CrO_4^{2-}(aq) + 4H^+(aq)$ $4H^+(aq) + 2ClO^-(aq) \rightarrow Cl_2(g) + 2H_2O(\ell)$
Step 5 Adjust for basic conditions. *Oxidation half-reaction:* Add 4 hydroxide ions to each side of the equation, forming 4 water molecules on the right side. *Reduction half-reaction:* Add 4 hydroxide ions to each side of the equation, forming 4 water molecules on the left side.	$4OH^-(aq) + 2H_2O(\ell) + CrO_2^-(aq) \rightarrow CrO_4^{2-}(aq) + 4H_2O(\ell)$ $4H_2O(\ell) + 2ClO^-(aq) \rightarrow Cl_2(g) + 2H_2O(\ell) + 4OH^-(aq)$
Step 6 Cancel any water molecules present on both sides of the equation. *Oxidation half-reaction:* Remove 2 water molecules from each side, leaving 2 on the right. *Reduction half-reaction:* Remove 2 water molecules from each side of the equation, leaving 2 water molecules on the left.	$4OH^-(aq) + CrO_2^-(aq) \rightarrow CrO_4^{2-}(aq) + 2H_2O(\ell)$ $2H_2O(\ell) + 2ClO^-(aq) \rightarrow Cl_2(g) + 4OH^-(aq)$
Step 7 Balance the charges by adding electrons. *Oxidation half-reaction*: There is a net charge of 5– on the left side of the equation and 2– on the right side. Therefore, give the right side a net charge of 5– by adding 3 electrons to the right side. *Reduction half-reaction*: There is a net charge of 2– on the left side of the equation and 4– on the right. Add 2 electrons to the left side to give it a net charge of 4–.	$4OH^-(aq) + CrO_2^-(aq) \rightarrow CrO_4^{2-}(aq) + 2H_2O(\ell) + 3e^-$ $2e^- + 2H_2O(\ell) + 2ClO^-(aq) \rightarrow Cl_2(g) + 4OH^-(aq)$

Step 8 Multiply one or both half-reactions by the number that will bring the number of electrons to the lowest common multiple. *Oxidation half-reaction*: Multiply the reactants and products by 2. *Reduction half-reaction*: Multiply the reactants and products by 3.	$8OH^{-}(aq) + 2CrO_2^{-}(aq) \rightarrow$ $2CrO_4^{2-}(aq) + 4H_2O(\ell) + 6e^{-}$ $6e^{-} + 6H_2O(\ell) + 6ClO^{-}(aq) \rightarrow$ $3Cl_2(g) + 12OH^{-}(aq)$
Step 9 Add the balanced half-reactions.	$6e^{-} + 6H_2O(\ell) + 6ClO^{-}(aq) + 8OH^{-}(aq)$ $+ 2CrO_2^{-}(aq) \rightarrow$ $2CrO_4^{2-}(aq) + 4H_2O(\ell) + 6e^{-} + 3Cl_2(g) +$ $12OH^{-}(aq)$
Step 10 Cancel the electrons, water molecules and hydroxide ions present on both sides of the equation.	$2H_2O(\ell) + 6ClO^{-}(aq) + 2CrO_2^{-}(aq) \rightarrow$ $2CrO_4^{2-}(aq) + 3Cl_2(g) + 4OH^{-}(aq)$
Step 11 Identify the oxidizing agent and reducing agent. The reducing agent gets oxidized and the oxidizing agent gets reduced.	Since $CrO_2^{-}(aq)$ gets oxidized, it is the reducing agent; since $ClO^{-}(aq)$ gets reduced, it is the oxidizing agent.

Check Your Solution

The number of atoms of each element is equal on each side of the equation and the total charge on each side is 8–, indicating that both mass and charge are balanced. The equation is balanced.

8. Balance the following ionic equation for basic conditions. Identify the oxidizing agent and the reducing agent.

$$Al(s) + NO^{-}(aq) \rightarrow NH_3(g) + AlO^{2-}(aq)$$

What Is Required?
You need to balance the equation for the redox reaction in basic conditions and identify the oxidizing agent and reducing agent in the reaction.

What Is Given?
You are given the unbalanced equation: $Al(s) + NO^{-}(aq) \rightarrow NH_3(g) + AlO_2^{-}(aq)$

Plan Your Strategy	Act on Your Strategy
Step 1 Write the unbalanced half-reactions. Include only those compounds that contain the atom that is oxidized or the atom that is reduced. (For this step, ignore the fact that one side of the equation might have oxygen atoms and the other side has none.)	*Oxidation half-reaction*: $Al(s) \rightarrow AlO_2^{-}(aq)$ *Reduction half- reaction*: $NO^{-}(aq) \rightarrow NH_3(g)$
Step 2 Balance atoms other than oxygen and hydrogen.	All atoms other than hydrogen and oxygen are balanced in the two half-reactions.
Step 3 Balance oxygen atoms by adding water molecules. *Oxidation half-reaction:* There are two oxygen atoms on the right side and none on the left side, so add 2 water molecules to the left side. *Reduction half-reaction:* There is 1 oxygen atom on the left side of the equation and none on the right, so add one water molecule to the right side to balance the oxygen atoms.	$2H_2O(\ell) + Al(s) \rightarrow AlO_2^{-}(aq)$ $NO^{-}(aq) \rightarrow NH_3(g) + H_2O(\ell)$
Step 4 Balance hydrogen atoms by adding hydrogen ions. *Oxidation half-reaction:* There are 4 hydrogen atoms on the left side of the equation, and none on the right side, so add 4 hydrogen ions to the right side. *Reduction half-reaction*: There are 5 hydrogen atoms on the right side of the equation and none on the left side, so add 5 hydrogen ions to the left side.	$2H_2O(\ell) + Al(s) \rightarrow AlO_2^{-}(aq) + 4H^{+}(aq)$ $5H^{+}(aq) + NO^{-}(aq) \rightarrow NH_3(g) + H_2O(\ell)$

Step 5 Adjust for basic conditions. *Oxidation half-reaction:* Add 4 hydroxide ions to each side of the equation, forming 4 water molecules on the right side. *Reduction half-reaction:* Add 5 hydroxide ions to each side of the equation, forming 5 water molecules on the left side.	$4OH^-(aq) + 2H_2O(\ell) + Al(s) \rightarrow AlO_2^-(aq) + 4H_2O(\ell)$ $5H_2O(\ell) + NO^-(aq) \rightarrow NH_3(g) + H_2O(\ell) + 5OH^-(aq)$
Step 6 Cancel any water molecules present on both sides of the equation. *Oxidation half-reaction:* Remove 2 water molecules from each side, leaving 2 on the right. *Reduction half-reaction:* remove 1water molecule from each side of the equation, leaving 4 water molecules on the left.	$4OH^-(aq) + Al(s) \rightarrow AlO_2^-(aq) + 2H_2O(\ell)$ $4H_2O(\ell) + NO^-(aq) \rightarrow NH_3(g) + 5OH^-(aq)$
Step 7 Balance the charges by adding electrons. *Oxidation half-reaction*: There is a net charge of 4– on the left side of the equation and 1– on the right side. Therefore, give the right side a net charge of 4– by adding 3 electrons to the right side. *Reduction half-reaction*: There is a net charge of 1– on the left side of the equation and 5– on the right. Add 4 electrons to the left side to give it a net charge of 5–.	$4OH^-(aq) + Al(s) \rightarrow AlO_2^-(aq) + 2H_2O(\ell) + 3e^-$ $4e^- + 4H_2O(\ell) + NO^-(aq) \rightarrow NH_3(g) + 5OH^-(aq)$
Step 8 Multiply one or both half-reactions by the number that will bring the number of electrons to the lowest common multiple. *Oxidation half-reaction*: Multiply the reactants and products by 4. *Reduction half-reaction*: Multiply the reactants and products by 3.	$16OH^-(aq) + 4Al(s) \rightarrow 4AlO_2^-(aq) + 8H_2O(\ell) + 12e^-$ $12e^- + 12H_2O(\ell) + 3NO^-(aq) \rightarrow 3NH_3(g) + 15OH^-(aq)$
Step 9 Add the balanced half-reactions.	$12e^- + 12H_2O(\ell) + 3NO^-(aq) + 16OH^-(aq) + 4Al(s) \rightarrow$ $4AlO_2^-(aq) + 8H_2O(\ell) + 12e^- + 3NH_3(g) + 15OH^-(aq)$

Step 10 Cancel the electrons, water molecules and hydroxide ions present on both sides of the equation.	$4H_2O(\ell) + 3NO^{-}(aq) + 4Al(s) + OH^{-}(aq) \rightarrow 4AlO_2^{-}(aq) + 3NH_3(g)$
Step 11 Identify the oxidizing agent and reducing agent. The reducing agent gets oxidized and the oxidizing agent gets reduced.	Since Al(s) gets oxidized, it is the reducing agent; since $NO^{-}(aq)$ gets reduced, it is the oxidizing agent.

Check Your Solution

The number of atoms of each element is equal on each side of the equation and the total charge on each side is 4–, indicating that both mass and charge are balanced. The equation is balanced.

9. Balance the following ionic equation for the conditions indicated. Identify the oxidizing agent and the reducing agent.

$$ClO_3^-(aq) + MnO_2(s) \rightarrow Cl^-(aq) + MnO_4^-(aq) \text{ (basic conditions)}$$

What Is Required?
You need to balance the equation for the redox reaction in basic conditions and identify the oxidizing agent and reducing agent in the reaction.

What Is Given?
You are given the unbalanced equation: $ClO_3^-(aq) + MnO_2(s) \rightarrow Cl^-(aq) + MnO_4^-(aq)$

Plan Your Strategy	Act on Your Strategy
Step 1 Write the unbalanced half-reactions. Include only those compounds that contain the atom that is oxidized or the atom that is reduced. (For this step, ignore the fact that one side of the equation might have oxygen atoms and the other side has none.)	*Oxidation half-reaction*: $MnO_2(s) \rightarrow MnO_4^-(aq)$ *Reduction half- reaction*: $ClO_3^-(aq) \rightarrow Cl^-(aq)$
Step 2 Balance atoms other than oxygen and hydrogen.	All atoms other than hydrogen and oxygen are balanced in the two half-reactions.
Step 3 Balance oxygen atoms by adding water molecules. *Oxidation half-reaction:* There are 4 oxygen atoms on the right side and 2 on the left side, so add 2 water molecules to the left side. *Reduction half-reaction:* There are 3 oxygen atoms on the left side of the equation and none on the right, so add 3 water molecules to the right side to balance the oxygen atoms.	$2H_2O(\ell) + MnO_2(s) \rightarrow MnO_4^-(aq)$ $ClO_3^-(aq) \rightarrow Cl^-(aq) + 3H_2O(\ell)$
Step 4 Balance hydrogen atoms by adding hydrogen ions. *Oxidation half-reaction:* There are 4 hydrogen atoms on the left side of the equation, and none on the right side, so add 4 hydrogen ions to the right side. *Reduction half-reaction*: There are 6 hydrogen atoms on the right side of the equation and none on the left side, so add 6 hydrogen ions to the left side.	$2H_2O(\ell) + MnO_2(s) \rightarrow MnO_4^-(aq) + 4H^+(aq)$ $6H^+(aq) + ClO_3^-(aq) \rightarrow Cl^-(aq) + 3H_2O(\ell)$

Step 5 Adjust for basic conditions. *Oxidation half-reaction:* Add 4 hydroxide ions to each side of the equation, forming 4 water molecules on the right side. *Reduction half-reaction:* Add 6 hydroxide ions to each side of the equation, forming 6 water molecules on the left side.	$4OH^-(aq) + 2H_2O(\ell) + MnO_2(s) \rightarrow MnO_4^-(aq) + 4H_2O(\ell)$ $6H_2O(\ell) + ClO_3^-(aq) \rightarrow Cl^-(aq) + 3H_2O(\ell) + 6OH^-(aq)$
Step 6 Cancel any water molecules present on both sides of the equation. *Oxidation half-reaction:* Remove 2 water molecules from each side, leaving 2 on the right. *Reduction half-reaction:* remove 3water molecules from each side of the equation, leaving 3 water molecules on the left.	$4OH^-(aq) + MnO_2(s) \rightarrow MnO_4^-(aq) + 2H_2O(\ell)$ $3H_2O(\ell) + ClO_3^-(aq) \rightarrow Cl^-(aq) + 6OH^-(aq)$
Step 7 Balance the charges by adding electrons. *Oxidation half-reaction*: There is a net charge of 4– on the left side of the equation and 1– on the right side. Therefore, give the right side a net charge of 4– by adding 3 electrons to the right side. *Reduction half-reaction*: There is a net charge of 1– on the left side of the equation and 7– on the right. Add 6 electrons to the left side to give it a net charge of 7–.	$4OH^-(aq) + MnO_2(s) \rightarrow MnO_4^-(aq) + 2H_2O(\ell) + 3e^-$ $6e^- + 3H_2O(\ell) + ClO_3^-(aq) \rightarrow Cl^-(aq) + 6OH^-(aq)$
Step 8 Multiply one or both half-reactions by the number that will bring the number of electrons to the lowest common multiple. *Oxidation half-reaction*: Multiply the reactants and products by 2. *Reduction half-reaction*: Multiply the reactants and products by 1.	$8OH^-(aq) + 2MnO_2(s) \rightarrow 2MnO_4^-(aq) + 4H_2O(\ell) + 6e^-$ $6e^- + 3H_2O(\ell) + ClO_3^-(aq) \rightarrow Cl^-(aq) + 6OH^-(aq)$
Step 9 Add the balanced half-reactions.	$6e^- + 3H_2O(\ell) + ClO_3^-(aq) + 8OH^-(aq) + 2MnO_2(s) \rightarrow$ $2MnO_4^-(aq) + 4\ H_2O(\ell) + 6e^- + Cl^-(aq) + 6OH^-(aq)$

Step 10 Cancel the electrons, water molecules and hydroxide ions present on both sides of the equation.	$ClO_3^-(aq) + 2OH^-(aq) + 2MnO_2(s) \rightarrow$ $2MnO_4^-(aq) + H_2O(\ell) + Cl^-(aq)$
Step 11 Identify the oxidizing agent and reducing agent. The reducing agent gets oxidized and the oxidizing agent gets reduced.	Since $MnO_2(s)$ gets oxidized, it is the reducing agent; since $ClO_3^-(aq)$ gets reduced, it is the oxidizing agent.

Check Your Solution

The number of atoms of each element is equal on each side of the equation and the total charge on each side is 3–, indicating that both mass and charge are balanced. The equation is balanced.

10. Balance the following ionic equation for the conditions indicated. Identify the oxidizing agent and the reducing agent.

$PbO_2(s) + I_2(s) \rightarrow Pb^{2+}(aq) + IO_3^-(aq)$ (acidic conditions)

What Is Required?
You need to balance the equation for the redox reaction in acidic conditions and identify the oxidizing agent and reducing agent in the reaction.

What Is Given?
You are given the unbalanced equation: $I_2(s) + ClO^-(aq) \rightarrow IO_3^-(aq) + Cl^-(aq)$

Plan Your Strategy	Act on Your Strategy
Step 1 Write the unbalanced half-reactions. Include only those compounds that contain the atom that is oxidized or the atom that is reduced. (For this step, ignore the fact that one side of the equation might have oxygen atoms and the other side has none.)	*Oxidation half-reaction*: $I_2(s) \rightarrow IO_3^-(aq)$ *Reduction half- reaction*: $PbO_2(s) \rightarrow Pb^{2+}(aq)$
Step 2 Balance atoms other than oxygen and hydrogen.	$I_2(s) \rightarrow 2IO_3^-(aq)$ $PbO_2(s) \rightarrow Pb^{2+}(aq)$ (already balanced for atoms other than oxygen and hydrogen)
Step 3 Balance oxygen atoms by adding water molecules. *Oxidation half-reaction:* There are 6 oxygen atoms on the right side and none on the left side. Therefore, add 6 water molecules to the left side to balance the oxygen atoms. *Reduction half-reaction:* There are two oxygen atoms on the left side and none on the right, so add 2 water molecules to the right side.	$6H_2O(\ell) + I_2(s) \rightarrow 2IO_3^-(aq)$ $PbO_2(s) \rightarrow Pb^{2+}(aq) + 2H_2O(\ell)$
Step 4 Balance hydrogen atoms by adding hydrogen ions. *Oxidation half-reaction:* There are 12 hydrogen atoms in total on the left side of the equation, and none on the right side, so add 12 hydrogen ions to the right side. *Reduction half-reaction*: There are no hydrogen atoms on the left side and 4 on the right, so add 4 hydrogen ions to the left side.	$6H_2O(\ell) + I_2(s) \rightarrow 2IO_3^-(aq) + 12H^+(aq)$ $4H^+(aq) + PbO_2(s) \rightarrow Pb^{2+}(aq) + 2H_2O(\ell)$

Step 5 Balance the charges by adding electrons. *Oxidation half-reaction*: There is a net charge of zero on the left side of the equation and 10+ on the right side. Therefore, give the right side a net charge of zero by adding 10 electrons to the right side. *Reduction half-reaction*: There is a net charge of 4+ on the left side of the equation and 2+ on the right. Add two electrons to the left side to give it a net charge of 2+.	$6H_2O(\ell) + I_2(s) \rightarrow 2IO_3^-(aq) + 12H^+(aq) + 10\ e^-$ $2e^- + 4H^+(aq) + PbO_2(s) \rightarrow Pb^{2+}(aq) + 2H_2O(\ell)$
Step 6 Multiply one or both half-reactions by the number that will bring the number of electrons to the lowest common multiple. *Oxidation half-reaction*: Multiply the reactants and products by 1. *Reduction half-reaction*: Multiply the reactants and products by 5.	$6H_2O(\ell) + I_2(s) \rightarrow 2IO_3^-(aq) + 12H^+(aq) + 10e^-$ $10e^- + 20H^+(aq) + 5PbO_2(s) \rightarrow 5Pb^{2+}(aq) + 10H_2O(\ell)$
Step 7 Add the balanced half-reactions.	$10e^- + 20H^+(aq) + 5PbO_2(s) + 6H_2O(\ell) + I_2(s) \rightarrow 2IO_3^-(aq) + 12H^+(aq) + 10e^- + 5Pb^{2+}(aq) + 10H_2O(\ell)$
Step 8 Cancel the electrons and common ions or atoms present on both sides of the equation.	$8H^+(aq) + 5PbO_2(s) + I_2(s) \rightarrow 2IO_3^-(aq) + 5Pb^{2+}(aq) + 4H_2O(\ell)$
Step 9 Identify the oxidizing agent and reducing agent. The reducing agent gets oxidized and the oxidizing agent gets reduced.	Since $I_2(s)$ gets oxidized, it is the reducing agent; since $PbO_2(s)$ gets reduced, it is the oxidizing agent.

Check Your Solution

The number of atoms of each element is equal on each side of the equation and the total charge on each side is 8+, indicating that both mass and charge are balanced. The equation is balanced.

Assigning Oxidation Numbers
(Student textbook page 606)

11. Determine the oxidation number of the atoms of the specified element in the following: N in $NF_3(g)$

What Is Required?
You must determine the oxidation number of an atom of nitrogen in a given compound.

What Is Given?
You are given the chemical formula $NF_3(g)$. You can refer to the rules in **Table 9.3** on page 604 of the student textbook.

Plan Your Strategy	Act on Your Strategy
Apply rule 5: the oxidation number of the more electronegative element is the same as the charge it would have as an ion. Determine the electronegativity of the two elements.	Since F is more electronegative, the fluorine ion is assigned a value of –1.
The oxidation number of N is unknown, but you know that the compound has a net charge of zero. Therefore, apply rule 6: the oxidation numbers must add to zero. Let *x* represent the oxidation number of N.	$1 \text{ N atom} + 3 \text{ F atoms} = 0$ $x + 3(-1) = 0$ $x - 3 = 0$ $x = +3$

The oxidation number of nitrogen is +3.

Check Your Solution
The rules were followed and the answer is logical.

12. Determine the oxidation number of the atoms of the specified element in the following: S in S_8(s)

What Is Required?
You must determine the oxidation number of an atom of sulfur in a given compound.

What Is Given?
You are given the chemical formula S_8(s) . You can refer to the rules in **Table 9.3** on page 604 of the student textbook.

Plan Your Strategy	Act on Your Strategy
Apply rule 1: A pure element has an oxidation number of zero.	Since S_8(s) is made up only of sulfur atoms, each atom has an oxidation number of zero.

Check Your Solution
The rules were followed and the answer is logical.

13. Determine the oxidation number of the atoms of the specified element in the following: Cr in CrO_4^{2-}(aq)

What Is Required?
You must determine the oxidation number of an atom of chromium in a given compound.

What Is Given?
You are given the chemical formula CrO_4^{2-}(aq). You can refer to the rules in **Table 9.3** on page 604 of the student textbook.

Plan Your Strategy	Act on Your Strategy
CrO_4^{2-}(aq) contains oxygen. Apply rule 4.	The oxidation number of oxygen is –2.
Apply rule 7: the sum of the oxidation numbers of all atoms in a polyatomic ion equals the charge on the ion. Let *x* represent the oxidation number of Cr.	$\begin{aligned} 1\text{ Cr atom} + 4\text{ O atoms} &= -2 \\ x + 4(-2) &= -2 \\ x - 8 &= -2 \\ x &= +6 \end{aligned}$

The oxidation number of Cr is +6.

Check Your Solution
The rules were followed and the answer is logical.

14. Determine the oxidation number of the atoms of the specified element in the following: P in P_2O_5(s)

What Is Required?
You must determine the oxidation number of an atom of phosphorus in a given compound.

What Is Given?
You are given the chemical formula P_2O_5(s). You can refer to the rules in **Table 9.3** on page 604 of the student textbook.

Plan Your Strategy	Act on Your Strategy
P_2O_5(s) contains oxygen. Apply rule 4.	The oxidation number of oxygen is –2.
The oxidation number of P is unknown, but you know that the compound has a net charge of zero. Therefore, apply rule 6: the oxidation numbers must add to zero. Let *x* represent the oxidation number of P.	$2\text{ P atoms} + 5\text{ O atoms} = 0$ $2x + 5(-2) = 0$ $2x - 10 = 0$ $2x = 10$ $x = +5$

The oxidation number of P is + 5.

Check Your Solution
The rules were followed and the answer is logical.

15. Determine the oxidation number of the atoms of the specified element in the following: C in $C_{12}H_{22}O_{11}(s)$

What Is Required?
You must determine the oxidation number of an atom of carbon in a given compound.

What Is Given?
You are given the chemical formula $C_{12}H_{22}O_{11}(s)$. You can refer to the rules in **Table 9.3** on page 604 of the student textbook.

Plan Your Strategy	Act on Your Strategy
$C_{12}H_{22}O_{11}(s)$ contains oxygen and hydrogen. Apply rules 3 and 4.	The oxidation number of oxygen is –2 and the oxidation number of hydrogen is +1.
The oxidation number of C is unknown, but you know that the compound has a net charge of zero. Therefore, apply rule 6: the oxidation numbers must add to zero. Let *x* represent the oxidation number of C.	$\begin{aligned} 12\text{ C atoms} + 22\text{ H atoms} + 11\text{ O atoms} &= 0 \\ 12x + 22(+1) + 11(-2) &= 0 \\ 12x + 22 - 22 &= 0 \\ 12x &= 0 \\ x &= 0 \end{aligned}$

The oxidation number of C is 0.

Check Your Solution
The rules were followed and the answer is logical. Because there are 12 carbon atoms in the molecule and they are not all bonded to atoms of other elements in the same way, the oxidation number, 0, is an average of the 12 atoms and not the exact oxidation number of each carbon atom.

16. Determine the oxidation number of the atoms of the specified element in the following: H in $CaH_2(s)$

What Is Required?
You must determine the oxidation number of an atom of hydrogen in a given compound.

What Is Given?
You are given the chemical formula $CaH_2(s)$. You can refer to the rules in **Table 9.3** on page 604 of the student textbook.

Plan Your Strategy	Act on Your Strategy
$CaH_2(s)$ contains hydrogen, so apply rule 3.	This is a metallic hydride, so the oxidation number of hydrogen is –1.

Check Your Solution
The rules were followed. The oxidation number is logical because the oxidation numbers of the two hydride ions add to –2 and the oxidation number of calcium is +2. The total is zero and the compound is neutral.

17. Determine the oxidation number of each of the atoms in the following compound: $H_2SO_3(aq)$

What Is Required?
You must determine the oxidation number of each of the atoms in a given compound.

What Is Given?
You are given the chemical formula $H_2SO_3(aq)$. You can refer to the rules in **Table 9.3** on page 604 of the student textbook.

Plan Your Strategy	Act on Your Strategy
$H_2SO_3(aq)$ contains oxygen and hydrogen. Apply rules 3 and 4.	The oxidation number of oxygen is –2, and the oxidation number of hydrogen is +1.
The oxidation number of C is unknown, but you know that the compound has a net charge of zero. Therefore, apply rule 6: the oxidation numbers must add to zero. Let *x* represent the oxidation number of C.	$\begin{aligned} 2\text{ H atoms} + 1\text{ S atom} + 3\text{ O atoms} &= 0 \\ 2(+1) + x + 3(-2) &= 0 \\ 2 + x - 6 &= 0 \\ x &= 6 - 2 \\ x &= +4 \end{aligned}$

The oxidation number of H is +1, the oxidation number of O is –2 and the oxidation number of S is +4.

Check Your Solution
The rules were followed and the answer is logical.

18. Determine the oxidation number of each of the atoms in the following compound: $OH^{-}(aq)$

What Is Required?
You must determine the oxidation number of each of the atoms in a given compound.

What Is Given?
You are given the chemical formula $OH^{-}(aq)$. You can refer to the rules in **Table 9.3** on page 604 of the student textbook.

Plan Your Strategy	Act on Your Strategy
$OH^{-}(aq)$ contains oxygen and hydrogen. Apply rules 3 and 4.	The oxidation number of oxygen is –2, and the oxidation number of hydrogen is +1.

Check Your Solution
The rules were followed and the answer is logical.

19. Determine the oxidation number of each of the atoms in the following compound: $HPO_4^{2-}(aq)$

What Is Required?
You must determine the oxidation number of each of the atoms in a given compound.

What Is Given?
You are given the chemical formula $HPO_4^{2-}(aq)$. You can refer to the rules in **Table 9.3** on page 604 of the student textbook.

Plan Your Strategy	Act on Your Strategy
$HPO_4^{2-}(aq)$ contains oxygen and hydrogen. Apply rules 3 and 4.	The oxidation number of oxygen is –2, and the oxidation number of hydrogen is +1.
Apply rule 7: the sum of the oxidation numbers of all atoms in a polyatomic ion equals the charge on the ion. Let x represent the oxidation number of P.	$\begin{aligned} 1\text{ H atom} + 1\text{ P atom} + 4\text{ O atoms} &= -2 \\ +1 + x + 4(-2) &= -2 \\ x - 7 &= -2 \\ x &= +5 \end{aligned}$

The oxidation number of H is +1, the oxidation number of O is –2, and the oxidation number of P is +5.

Check Your Solution
The rules were followed and the answer is logical.

20. Determine the oxidation number of oxygen in the following: $O_2(g)$

What Is Required?
You must determine the oxidation number of an atom of oxygen in a given compound.

What Is Given?
You are given the chemical formula $O_2(g)$. You can refer to the rules in **Table 9.3** on page 604 of the student textbook.

Plan Your Strategy	Act on Your Strategy
Apply rule 1: A pure element has an oxidation number of zero.	Since $O_2(g)$ is made up only of oxygen atoms, each atom has an oxidation number of zero.

Check Your Solution
The rules were followed and the answer is logical.

21. Determine the oxidation number of oxygen in the following: the peroxide ion, $O_2^{2-}(aq)$

What Is Required?
You must determine the oxidation number of an atom of oxygen in a given compound.

What Is Given?
You are given the chemical formula $O_2^{2-}(aq)$. You can refer to the rules in **Table 9.3** on page 604 of the student textbook.

Plan Your Strategy	Act on Your Strategy
Apply rule 7: the sum of the oxidation numbers of all atoms in a polyatomic ion equals the charge on the ion. Let x represent the oxidation number of O.	2 O atoms $= -2$ $2x = -2$ $x = -1$

The oxidation number of O is –1.

Check Your Solution
The rules were followed. Rule 4 also says that the oxidation number of oxygen in peroxides and superoxides is not –2.

22. Determine the oxidation number of each element in the following ionic compound by considering the ions separately: $Al(HCO_3)_3(s)$

What Is Required?
You must determine the oxidation number of each element in a given compound.

What Is Given?
You are given the chemical formula $Al(HCO_3)_3(s)$. You can refer to the rules in **Table 9.3** on page 604 of the student textbook.

Plan Your Strategy	Act on Your Strategy
Separate the ions in the compound.	$Al(HCO_3)_3$ is made up of Al^{3+} and HCO_3^-
Apply rule 2 to Al^{3+}: The oxidation number of an element in a monoatomic ion equals the charge on the ion.	The oxidation number of Al^{3+} is +3.
Apply rule 7 to HCO_3^-: the sum of the oxidation numbers of all atoms in a polyatomic ion equals the charge on the ion. Let *x* represent the oxidation number of C.	$\begin{aligned} 1\text{ H atom} + 1\text{ C atom} + 3\text{ O atoms} &= -1 \\ +1 + x + 3(-2) &= -1 \\ x - 5 &= -1 \\ x &= +4 \end{aligned}$

The oxidation number of Al is +3, the oxidation number of H is +1, the oxidation number of C is +4, and the oxidation number of O is –2.

Check Your Solution
The rules were followed and the answer is logical.

23. Determine the oxidation number of each element in the following ionic compound by considering the ions separately: $(NH_4)_4PO_4(aq)$

What Is Required?
You must determine the oxidation number of each element in a given compound.

What Is Given?
You are given the chemical formula $(NH_4)_4PO_4(aq)$. You can refer to the rules in **Table 9.3** on page 604 of the student textbook.

Plan Your Strategy	Act on Your Strategy
Separate the ions in the compound.	$(NH_4)_4PO_4(aq)$ is made up of NH_4^+ and PO_4^{3-}
Apply rule 7 to NH_4^+ and PO_4^{3-}: the sum of the oxidation numbers of all atoms in a polyatomic ion equals the charge on the ion. Let *x* represent the oxidation number of N and let *y* represent the oxidation number of P.	$\begin{aligned} 1\text{ N atom} + 4\text{ H atoms} &= +1 \\ x + 4(+1) &= +1 \\ x + 4 &= +1 \\ x &= -3 \end{aligned}$ $\begin{aligned} 1\text{ P atom} + 4\text{ O atoms} &= -3 \\ y + 4(-2) &= -3 \\ y - 8 &= -3 \\ y &= +5 \end{aligned}$

The oxidation number of N is –3, the oxidation number of H is +1, the oxidation number of P is +5, and the oxidation number of O is –2.

Check Your Solution
The rules were followed and the answer is logical.

24. Determine the oxidation number of each element in the following ionic compound by considering the ions separately. (**Hint:** One formula unit of the compound contains two identical monatomic ions and one polyatomic ion.) $K_2H_3IO_6(aq)$

What Is Required?
You must determine the oxidation number of each element in a given compound.

What Is Given?
You are given the chemical formula $K_2H_3IO_6(aq)$. You can refer to the rules in **Table 9.3** on page 604 of the student textbook.

Plan Your Strategy	Act on Your Strategy
Separate the ions in the compound.	$K_2H_3IO_6(aq)$ is made up of K^+ and $H_3IO_6^{2-}$.
Apply rule 2 to K^+: The oxidation number of an element in a monoatomic ion equals the charge on the ion.	The oxidation number of K^+ is +1.
Apply rule 7 to $H_3IO_6^{2-}$: the sum of the oxidation numbers of all atoms in a polyatomic ion equals the charge on the ion. Let *x* represent the oxidation number of I.	$3\text{ H atoms} + 1\text{ I atoms} + 6\text{ O atoms} = -2$ $3(+1) + x + 6(-2) = -2$ $x - 9 = -2$ $x = +7$

Therefore, the oxidation number of K is +1, the oxidation number of H is +1, the oxidation number of I is +7, and the oxidation number of O is –2.

Check Your Solution
The rules were followed and the answer is logical.

Identifying Redox Reactions
(Student textbook page 611)

25. Is the reaction below a redox reaction? Identify if it is a disproportionation reaction. (**Hint:** Some of the oxidation numbers are averages.)

$$2H_2(g) + O_2(g) \rightarrow 2H_2O(\ell)$$

What Is Required?
You need to determine whether the given chemical reaction is or is not a redox reaction.

What Is Given?
You are given the balanced equation $2H_2(g) + O_2(g) \rightarrow 2H_2O(\ell)$. You are given rules for assigning oxidation numbers in **Table 9.3** on page 604 of the student textbook.

Plan Your Strategy	Act on Your Strategy
Apply the rules in **Table 9.3** to assign oxidation numbers to each atom in the equation.	$\underset{0}{2H_2}(g) + \underset{0}{O_2}(g) \rightarrow 2\underset{+1}{H_2}\underset{-2}{O}(\ell)$
Compare the oxidation numbers of the atoms of each element on both sides of the equation.	The oxidation number of hydrogen is zero on the left and +1 on the right. The oxidation number of oxygen is zero on the left and –2 on the right.
Determine whether the oxidation numbers of at least two atoms change during the reaction. If so, the reaction is a redox reaction. If oxidation numbers do not change, the reaction is not a redox reaction.	The oxidation number of hydrogen changed from 0 to + 1. The oxidation number of oxygen changed from 0 to –2. Therefore, the reaction is a redox reaction.

Check Your Solution
The rules were followed. Oxidation numbers for two elements changed and the answer is logical.

26. Is the reaction below a redox reaction? Identify if it is a disproportionation reaction. (**Hint:** Some of the oxidation numbers are averages.)

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

What Is Required?

You need to determine whether the given chemical reaction is or is not a redox reaction.

What Is Given?

You are given the balanced equation $2H_2(g) + O_2(g) \rightarrow 2H_2O(\ell)$. You are given rules for assigning oxidation numbers in **Table 9.3** on page 604 of the student textbook.

Plan Your Strategy	Act on Your Strategy
Apply the rules in **Table 9.3** to assign oxidation numbers to each atom in the equation.	$\underset{+3\,-1}{PCl_3(\ell)} + 3\underset{+1\,-2}{H_2O(\ell)} \rightarrow \underset{+1\,+3\,-2}{H_3PO_3(aq)} + 3\underset{+1\,-1}{HCl(aq)}$
Compare the oxidation numbers of the atoms of each element on both sides of the equation.	The oxidation number of all hydrogen atoms is +1. The oxidation number of chlorine is –1 on both sides of the equation. The oxidation number of oxygen atoms on both sides of the equation is –2. The oxidation number of phosphorus on both sides of the equation is +3.
Determine whether the oxidation numbers of at least two atoms change during the reaction. If so, the reaction is a redox reaction. If oxidation numbers do not change, the reaction is not a redox reaction.	There was no change in any of the oxidation numbers. Therefore, the reaction is *not* a redox reaction.

Check Your Solution

The rules were followed and the answer is logical.

27. Is the reaction below a redox reaction? Identify if it is a disproportionation reaction. (**Hint:** Some of the oxidation numbers are averages.)

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

What Is Required?
You need to determine whether the given chemical reaction is or is not a redox reaction.

What Is Given?
You are given the balanced equation $2C_2H_6(g) + 7O_2(g) \rightarrow 4CO_2(g) + 6H_2O(\ell)$
You are given rules for assigning oxidation numbers in **Table 9.3** on page 604 of the student textbook.

Plan Your Strategy	Act on Your Strategy
Apply the rules in **Table 9.3** to assign oxidation numbers to each atom in the equation.	$\underset{-3\ +1}{2C_2H_6(g)} + \underset{0}{7O_2(g)} \rightarrow \underset{+4\ -2}{4CO_2(g)} + \underset{+1\ -2}{6H_2O(\ell)}$
Compare the oxidation numbers of the atoms of each element on both sides of the equation.	The oxidation number of all hydrogen atoms is +1. The oxidation number of carbon is –3 on the left side and + 4 on the right side of the equation. The oxidation number of oxygen atoms on the left side is 0 and on the right side of the equation is –2.
Determine whether the oxidation numbers of at least two atoms change during the reaction. If so, the reaction is a redox reaction. If oxidation numbers do not change, the reaction is not a redox reaction.	The oxidation number of carbon changed from –3 to +4. The oxidation number of oxygen changed from 0 to –2. Therefore, the reaction is a redox reaction.

Check Your Solution
The rules were followed. Oxidation numbers for two elements changed and in the opposite direction. The carbon is bonded to more oxygen atoms and fewer hydrogen atoms on the right side of the equation. The oxygen atoms on the right are bonded to fewer oxygen atoms on the right side.

28. Is the reaction below a redox reaction? Identify if it is a disproportionation reaction. (**Hint:** Some of the oxidation numbers are averages.)

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

What Is Required?
You need to determine whether the given chemical reaction is or is not a redox reaction.

What Is Given?
You are given the balanced equation $3NO_2(g) + H_2O(\ell) \rightarrow 2HNO_3(aq) + NO(g)$
You are given rules for assigning oxidation numbers in **Table 9.3** on page 604 of the student textbook.

Plan Your Strategy	Act on Your Strategy
Apply the rules in **Table 9.3** to assign oxidation numbers to each atom in the equation.	$\underset{+4\ -2}{3NO_2(g)} + \underset{+1\ -2}{H_2O(\ell)} \rightarrow \underset{+1+5-2}{2HNO_3(aq)} + \underset{+2\ -2}{NO(g)}$
Compare the oxidation numbers of the atoms of each element on both sides of the equation.	The oxidation number of all hydrogen atoms is +1. The oxidation number of oxygen atoms on both sides of the equation is –2. Some atoms of nitrogen go from an oxidation number of +4 to an oxidation number of +5 while some go from + 5 to + 2.
Determine whether the oxidation numbers of at least two atoms change during the reaction. If so, the reaction is a redox reaction. If oxidation numbers do not change, the reaction is not a redox reaction.	Since nitrogen atoms go from + 4 to +5 and from +4 to + 2, the reaction is a disproportionation reaction.

Check Your Solution
The rules were followed. The oxidation number of some of the nitrogen atoms increased while that of others decreased, making the reaction both a redox reaction and a disproporation reaction.

29. Is the reaction below a redox reaction? Identify if it is a disproportionation reaction. (**Hint:** Some of the oxidation numbers are averages.)

$$6HI(aq) + 2HMnO_4(aq) \rightarrow 3I_2(s) + 2MnO_2(s) + 4H_2O(\ell)$$

What Is Required?
You need to determine whether the given chemical reaction is or is not a redox reaction.

What Is Given?
You are given the balanced equation $6HI(aq) + 2HMnO_4(aq) \rightarrow 3I_2(s) + 2MnO_2(s) + 4H_2O(\ell)$
You are given rules for assigning oxidation numbers in **Table 9.3** on page 604 of the student textbook.

Plan Your Strategy	Act on Your Strategy
Apply the rules in **Table 9.3** to assign oxidation numbers to each atom in the equation.	$6HI(aq) + 2HMnO_4(aq) \rightarrow 3I_2(s) + 2MnO_2(s) + 4H_2O(\ell)$ +1–1 +1 +7 –2 0 +4 –2 +1 –2
Compare the oxidation numbers of the atoms of each element on both sides of the equation.	The oxidation number of all hydrogen atoms is +1. The oxidation number of oxygen atoms on both sides of the equation is –2. The oxidation number of iodine is –1 on the left side and 0 on the right side of the equation. The oxidation number of manganese atoms on the left side is +7 and on the right side of the equation is +4.
Determine whether the oxidation numbers of at least two atoms change during the reaction. If so, the reaction is a redox reaction. If oxidation numbers do not change, the reaction is not a redox reaction.	The oxidation number of iodine changed from –1 to 0. The oxidation number of manganese changed from +7 to –2. Therefore, the reaction is a redox reaction.

Check Your Solution
The rules were followed. Of the atoms that changed in oxidation state, one increased and the other decreased which is necessary to be a redox reaction.

30. Is the reaction below a redox reaction? Identify if it is a disproportionation reaction. (**Hint:** Some of the oxidation numbers are averages.)

$$CH_3COOH(aq) + CH_3OH(aq) \rightarrow CH_3COOCH_3(aq) + H_2O(\ell)$$

What Is Required?
You need to determine whether the given chemical reaction is or is not a redox reaction.

What Is Given?
You are given the balanced chemical equation:
$CH_3COOH(aq) + CH_3OH(aq) \rightarrow CH_3COOCH_3(aq) + H_2O(\ell)$
You are given rules for assigning oxidation numbers in **Table 9.3** on page 604 of the student textbook.

Plan Your Strategy	Act on Your Strategy
Apply the rules in **Table 9.3** to assign oxidation numbers to each atom in the equation.	$CH_3COOH(aq) + CH_3OH(aq) \rightarrow CH_3COOCH_3(aq) + H_2O(\ell)$ +3+1–3–2–2+1 –2 +1 –2+1 –3+1 +3–2–2+2+1 +1 –2
Compare the oxidation numbers of the atoms of each element on both sides of the equation.	The oxidation number of all hydrogen atoms is +1. The average oxidation number of carbon is $\frac{(+3-3-2)}{3}=-\frac{2}{3}$ on the left side and $\frac{\left(-3+3+2\right)}{3}=-\frac{2}{3}$ on the right side of the equation. The numbers were determined by using the Lewis structure method. The oxidation number of all oxygen atoms is –2.
Determine whether the oxidation numbers of at least two atoms change during the reaction. If so, the reaction is a redox reaction. If oxidation numbers do not change, the reaction is not a redox reaction.	There is no change in any of the oxidation numbers. Therefore, the reaction is *not* a redox reaction.

Check Your Solution
The rules were followed. The carbon atoms on the two sides of the equation, were bonded to the same number of hydrogen atoms and oxygen atoms, also indicating that their oxidation numbers did not change during the reaction.

31. Is the reaction below a redox reaction? Identify if it is a disproportionation reaction. (**Hint:** Some of the oxidation numbers are averages.)

$$2NO_2(g) + 7H_2(g) \rightarrow 2NH_3(g) + 4H_2O(\ell)$$

What Is Required?

You need to determine whether the given chemical reaction is or is not a redox reaction.

What Is Given?

You are given the balanced equation $2NO_2(g) + 7H_2(g) \rightarrow 2NH_3(g) + 4H_2O(\ell)$. You are given rules for assigning oxidation numbers in **Table 9.3** on page 604 of the student textbook.

Plan Your Strategy	Act on Your Strategy
Apply the rules in **Table 9.3** to assign oxidation numbers to each atom in the equation.	$2NO_2(g) + 7H_2(g) \rightarrow 2NH_3(g) + 4H_2O(\ell)$ +4 –2 0 –3 +1 +1 –2
Compare the oxidation numbers of the atoms of each element on both sides of the equation.	The oxidation number of hydrogen atoms on the left side is 0 and on the right side is +1. The oxidation number of nitrogen is +4 on the left side and –3 on the right side of the equation. The oxidation number of all oxygen atoms is –2.
Determine whether the oxidation numbers of at least two atoms change during the reaction. If so, the reaction is a redox reaction. If oxidation numbers do not change, the reaction is not a redox reaction.	The oxidation number of nitrogen changed from +4 to –3. The oxidation number of hydrogen changed from 0 to +1. Therefore, the reaction is a redox reaction.

Check Your Solution

The rules were followed. The oxidation numbers of the two elements changed in opposite directions, as they should for a redox reaction.

32. Identify the oxidizing agent and the reducing agent for the redox reaction(s) in questions 25 through 31.

What Is Required?
You need to identify the oxidizing agent and reducing agent in the redox reactions from questions 25 through 31.

What Is Given?
You are given the information you collected in using oxidation numbers to determine if the reactions were redox reactions.
[Note: Since the reactions in questions 26 and 30 were not redox reactions, there is no oxidizing agent or reducing agent in these reactions.]

Plan Your Strategy	Act on Your Strategy
Step 1 Review answer from question 25.	The oxidation number of hydrogen changed from 0 to + 1. The oxidation number of oxygen changed from 0 to –2. Therefore, the reaction is a redox reaction.
Step 2 The oxidizing agent is reduced, which causes a decrease in the oxidation number of the atom. The reducing agent is oxidized, which causes an increase in the oxidation number of the atom.	Hydrogen increases in oxidation number, so the reducing agent is $H_2(g)$, and oxygen decreases in oxidation number, so the oxidizing agent is $O_2(g)$

Plan Your Strategy	Act on Your Strategy
Step 1 Review answer from question 27.	The oxidation number of carbon changed from –3 to +4. The oxidation number of oxygen changed from 0 to –2. Therefore, the reaction is a redox reaction.
Step 2 The oxidizing agent is reduced, which causes a decrease in the oxidation number of the atom. The reducing agent is oxidized, which causes an increase in the oxidation number of the atom.	Carbon increases in oxidation number, so the reducing agent is $C_2H_6(g)$; oxygen decreases in oxidation number, so the oxidizing agent is $O_2(g)$

Plan Your Strategy	Act on Your Strategy
Step 1 Review answer from question 28.	Since nitrogen atoms go from + 4 to +5 and from +4 to + 2, the reaction is a disproportionation reaction.
Step 2 The oxidizing agent is reduced, which causes a decrease in the oxidation number of the atom. The reducing agent is oxidized, which causes an increase in the oxidation number of the atom.	Nitrogen increases and decreases in oxidation number, so NO_2 (g) is both the oxidizing and reducing agent.

Plan Your Strategy	Act on Your Strategy
Step 1 Review answer from question 29.	The oxidation number of iodine changed from –1 to 0. The oxidation number of manganese changed from +7 to +4. Therefore, the reaction is a redox reaction.
Step 2 The oxidizing agent is reduced, which causes a decrease in the oxidation number of the atom. The reducing agent is oxidized, which causes an increase in the oxidation number of the atom.	Iodine increases in oxidation number so the reducing agent is HI(aq), and manganese decreases in oxidation number so the oxidizing agent is $HMnO_4$(aq)

Plan Your Strategy	Act on Your Strategy
Step 1 Review answer from question 31.	The oxidation number of nitrogen changed from +4 to –3. The oxidation number of hydrogen changed from 0 to +1. Therefore, the reaction is a redox reaction.
Step 2 The oxidizing agent is reduced, which causes a decrease in the oxidation number of the atom. The reducing agent is oxidized, which causes an increase in the oxidation number of the atom.	Hydrogen increases in oxidation number, so the reducing agent is H_2(g), and nitrogen decreases in oxidation number, so the oxidizing agent is NO_2(g)

Check Your Solution

In each case except one, the oxidation number of atoms of one element increased while atoms of another element decreased, which is necessary in order to have an oxidizing agent and a reducing agent. In the case of the disproportionation reaction, some atoms of the same element, nitrogen, increases while that of other atoms decreased.

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

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

What Is Required?
You are to identify the reactant that undergoes oxidation and the reactant that undergoes reduction in a redox reaction.

What Is Given?
You are given the balanced equation $Br_2(\ell) + 2ClO_2^-(aq) \rightarrow 2Br^-(aq) + 2ClO_2(aq)$
You are given rules for assigning oxidation numbers in **Table 9.3** on page 604 of the student textbook.

Plan Your Strategy	Act on Your Strategy
Apply the rules in **Table 9.3** to assign oxidation numbers to each atom in the equation.	$Br_2(\ell) + 2ClO_2^-(aq) \rightarrow 2Br^-(aq) + 2ClO_2(aq)$ 0 +3 –2 –1 +4 –2
Compare the oxidation numbers of the atoms of each element on both sides of the equation.	The oxidation number of bromine atoms on the left side is 0 and on the right side is –1. The oxidation number of chlorine is +3 on the left side and +4 on the right side of the equation. The oxidation number of all oxygen atoms is –2.
Determine whether the oxidation numbers of at least two atoms change during the reaction. If so, the reaction is a redox reaction. If oxidation numbers do not change, the reaction is not a redox reaction.	The oxidation number of bromine changed from 0 to –1. The oxidation number of chlorine changed from +3 to +4. Therefore, the reaction is a redox reaction.
An atom that undergoes an increase in oxidation number is oxidized and an atom that undergoes a decrease in oxidation number is reduced.	Chlorine increases in oxidation number, so $ClO_2^-(aq)$ is oxidized. Bromine decreases in oxidation number, so bromine is reduced.

Check Your Solution
Following the rules led to logical answers.

34. 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 are to determine whether extraction of pure elemental metals from their ores involves redox reactions.

What Is Given?
You are given that nickel and copper ores usually contain metals as sulfides, such as NiS(s) and Cu_2S(s). You are also given the rules for assigning oxidation numbers in **Table 9.3** on page 604 of the student textbook.

Plan Your Strategy	Act on Your Strategy
Identify oxidation number of the metals in their ores and in their pure elemental metal form.	NiS: Sulfur has an oxidation number of –2, therefore nickel has an oxidation number of +2. In pure elemental form, the oxidation number of nickel is 0. Cu_2S(s): sulfur has an oxidation number of –2, therefore copper has an oxidation number of +1. In pure elemental form, the oxidation number of copper is 0.
Identify if the oxidation number of the metals change in the process of extraction.	In both cases, the oxidation number of the metal decreases from a positive value to 0. This is a reduction. A reducing agent must be used to achieve this change, so this is a redox reaction.

Check Your Solution
The conclusion that metal ions are reduced to produce the metal atoms agrees with the original observations that the mass of the ores is reduced when metals extracted. This is the origin of the term, reduction.

Balancing a Redox Equation for an Acidic Solution Using the Oxidation Number Method
(Student textbook page 615)

35. 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 an equation for the combustion of carbon disulfide, using the oxidation number method.

What Is Given?
You are given the reactants, carbon disulfide and oxygen, and the products, carbon dioxide and sulfur dioxide.
You are given rules for assigning oxidation numbers in **Table 9.3** on page 604 of the student textbook.

Plan Your Strategy	Act on Your Strategy
Step 1 Assign an oxidation number to each atom in the equation and determine whether it is a redox reaction.	$CS_2(g) + O_2(g) \longrightarrow CO_2(g) + SO_2(g)$ +4 -2 0 +4 -2 +4 -2 The oxidation numbers of sulfur and oxygen change during the reaction. It is a redox reaction.
Step 2 Identify the atom or atoms that undergo an increase in oxidation number and the atom or atoms that undergo a decrease in oxidation number.	Each of two sulfur atoms lost 6 electrons and each of two oxygen atoms gained 2 electrons. So the O_2 lost 12 electrons and the CS_2 gained 4 electrons. $1(-12e^-) = -12e^-$ $CS_2(g) + O_2(g) \longrightarrow CO_2(g) + SO_2(g)$ +4 -2 0 +4 -2 +4 -2 $3(+4e^-) = +12e^-$

Step 3 Balance the atoms that undergo a change in oxidation number. Place coefficients in front of the reactants that will balance the number of electrons that are lost by one reactant and gained by another.	A coefficient of 3 is needed for the O_2. A coefficient of 1 is already implied for the CS_2. This ratio of 1:3 must be maintained for the remainder of the balancing. $CS_2(g) + 3O_2(g) \rightarrow CO_2(g) + SO_2(g)$
Step 4 Determine the coefficients of the products that are needed to complete the balancing of the equation.	There are 2 S atoms on the left and 1 on the right. Place a coefficient of 2 in front of the SO_2 in the products. $CS_2(g) + 3O_2(g) \rightarrow CO_2(g) + 2SO_2(g)$ Atoms of all elements are now balanced.

Check Your Solution

Atoms of all of the elements are balanced. The total charge on the left side is 0 and the total charge on the right side is also 0. The charge is balanced.

36. Use the oxidation number method to balance the following equation:

$B_2O_3(aq) + Mg(s) \rightarrow MgO(s) + Mg_3B_2(aq)$

What Is Required?
You must balance an equation for the given redox reaction using the oxidation number method.

What Is Given?
You are given the reaction: $B_2O_3(aq) + Mg(s) \rightarrow MgO(s) + Mg_3B_2(aq)$
You are given rules for assigning oxidation numbers in **Table 9.3** on page 604 of the student textbook.

Plan Your Strategy	Act on Your Strategy
Step 1 Assign an oxidation number to each atom in the equation and determine whether it is a redox reaction.	$\underset{+3}{B_2}\underset{-2}{O_3}(aq) + \underset{0}{Mg}(s) \longrightarrow \underset{+2}{Mg}\underset{-2}{O}(s) + \underset{+2}{Mg_3}\underset{-3}{B_2}(aq)$ The oxidation numbers of boron and magnesium change during the reaction. It is a redox reaction.
Step 2 Identify the atom or atoms that undergo an increase in oxidation number and the atom or atoms that undergo a decrease in oxidation number.	Each of two B atoms loses 6 electrons and the Mg atom gained 2 electrons. $1(+12e^-) = +12e^-$ $\underset{+3}{B_2}\underset{-2}{O_3}(aq) + \underset{0}{Mg}(s) \longrightarrow \underset{+2}{Mg}\underset{-2}{O}(s) + \underset{+2}{Mg_3}\underset{-3}{B_2}(aq)$ $6(-2e^-) = -12e^-$
Step 3 Balance the atoms that undergo a change in oxidation number. Place coefficients in front of the reactants that will balance the number of electrons that are lost by one reactant and gained by another.	Place a coefficient of 6 in front of the Mg reactant. A coefficient of 1 is already implied for the B_2O_3. This ratio of 1:6 must be maintained for the remainder of the balancing. $B_2O_3(aq) + 6Mg(s) \rightarrow MgO(s) + Mg_3B_2(aq)$
Step 4 Determine the coefficients of the products that are needed to complete the balancing of the equation.	Place a 3 in front of the magnesium oxide to balance the oxygen atoms $B_2O_3(aq) + 6Mg(s) \rightarrow 3MgO(s) + Mg_3B_2(aq)$ Atoms of all elements are now balanced.

Check Your Solution
Atoms of all of the elements are balanced. The total charge on the left side is 0 and the total charge on the right side is also 0. The charge is balanced.

37. Use the oxidation number method to balance the following equation:

$H_2S(g) + H_2O_2(aq) \rightarrow S_8(s) + H_2O(\ell)$

What Is Required?
You must balance an equation for the given redox reaction using the oxidation number method.

What Is Given?
You are given the reaction: $H_2S(g) + H_2O_2(aq) \rightarrow S_8(s) + H_2O(\ell)$
You are given rules for assigning oxidation numbers in **Table 9.3** on page 604 of the student textbook.

Plan Your Strategy	Act on Your Strategy
Step 1 Assign an oxidation number to each atom in the equation and determine whether it is a redox reaction.	$\underset{+1}{H_2}\underset{-2}{S}(g) + \underset{+1}{H_2}\underset{-1}{O_2}(aq) \longrightarrow \underset{0}{S_8}(s) + \underset{+1}{H_2}\underset{-2}{O}(\ell)$ The oxidation numbers of sulfur and oxygen change during the reaction. It is a redox reaction.
Step 2 Identify the atom or atoms that undergo an increase in oxidation number and the atom or atoms that undergo a decrease in oxidation number.	One sulfur atom lost 2 electrons and each of two oxygen atoms gained 1 electron for a total of 2. $1(-2e^-) = -2e^-$ $\underset{+1}{H_2}\underset{-2}{S}(g) + \underset{+1}{H_2}\underset{-1}{O_2}(aq) \longrightarrow \underset{0}{S_8}(s) + \underset{+1}{H_2}\underset{-2}{O}(\ell)$ $1(+2e^-) = +2e^-$
Step 3 Balance the atoms that undergo a change in oxidation number. Place coefficients in front of the reactants that will balance the number of electrons that are lost by one reactant and gained by another.	Coefficients of 1 are already implied for H_2S and H_2O_2. This ratio of 1:1 must be maintained while balancing the products. $H_2S(g) + H_2O_2(aq) \rightarrow S_8(s) + H_2O(\ell)$

Step 4 Determine the coefficients of the products that are needed to complete the balancing of the equation.	There are 8 S atoms on the right so a coefficient of 8 is needed for the H_2S on the left. If the H_2S has a coefficient of 8, then H_2O_2 must also have a coefficient of 8 in order to maintain the 1:1 ratio. $8H_2S(g) + 8H_2O_2(aq) \rightarrow S_8(s) + H_2O(\ell)$ Now there are 32 H atoms on the left and 2 on the right. Give the H_2O on the right a coefficient of 16. $8H_2S(g) + 8H_2O_2(aq) \rightarrow S_8(s) + 16H_2O(\ell)$ Atoms of all elements are now balanced.

Check Your Solution
Atoms of all of the elements are balanced. The total charge on the left side is 0 and the total charge on the right side is also 0. The charge is balanced.

38. Use the oxidation number method to balance the following equation:

$$ClO_2(aq) + SbO_2^-(aq) \rightarrow ClO_2^-(aq) + Sb(OH)_6^-(aq)$$

What Is Required?
You must balance an equation for the given redox reaction using the oxidation number method.

What Is Given?
You are given the reaction: $ClO_2(aq) + SbO_2^-(aq) \rightarrow ClO_2^-(aq) + Sb(OH)_6^-(aq)$
You are given rules for assigning oxidation numbers in **Table 9.3** on page 604 of the student textbook.

Plan Your Strategy	Act on Your Strategy
Step 1 Assign an oxidation number to each atom in the equation and determine whether it is a redox reaction.	$\underset{+4\ -2}{ClO_2}(aq) + \underset{+3\ -2}{SbO_2^-}(aq) \longrightarrow \underset{+3\ -2}{ClO_2^-}(aq) + \underset{+5\ \ -2\ +1}{Sb(OH)_6^-}(aq)$ The oxidation numbers of chlorine and antimony change during the reaction. It is a redox reaction.
Step 2 Identify the atom or atoms that undergo an increase in oxidation number and the atom or atoms that undergo a decrease in oxidation number.	The chlorine atom gained 1 electron and the antimony atom lost 2 electrons. $2(+1e^-) = +2e^-$ $\underset{+4\ -2}{ClO_2}(aq) + \underset{+3\ -2}{SbO_2^-}(aq) \longrightarrow \underset{+3\ -2}{ClO_2^-}(aq) + \underset{+5\ \ -2\ +1}{Sb(OH)_6^-}(aq)$ $1(-2e^-) = -2e^-$
Step 3 Determine the numerical values of the total increase and decrease in oxidation numbers. Place coefficients in front of the reactants that will balance the number of electrons that are lost by one reactant and gained by another.	The ClO_2 needs a coefficient of 2 and a coefficient of 1 is already implied for the SbO_2^-. This ratio of 2:1 must be maintained for the remainder of the balancing. $2ClO_2(aq) + SbO_2^-(aq) \rightarrow$ $ClO_2^-(aq) + Sb(OH)_6^-(aq)$
Step 4 The presence of the $Sb(OH)_6^-(aq)$ on the right implies that the reaction was carried out in a basic solution. Continue to balance the equation according to basic conditions. Balance atoms of all elements except oxygen and hydrogen.	There are 2 Cl atoms on the left and one on the right. Give the ClO_2^- a coefficient of 2. $2ClO_2(aq) + SbO_2^-(aq) \rightarrow$ $2ClO_2^-(aq) + Sb(OH)_6^-(aq)$

Step 5 Add water molecules to balance oxygen atoms.	There are 6 O atoms on the left and 10 on the right. Add 4 water molecules to the left side of the equation. $2ClO_2(aq) + SbO_2^-(aq) + 4H_2O(\ell) \rightarrow$ $2ClO_2^-(aq) + Sb(OH)_6^-(aq)$
Step 6 Balance the hydrogen atoms by adding hydrogen ions.	There are 8 H atoms on the left and 6 on the right. Add 2 H ions to the right side. $2ClO_2(aq) + SbO_2^-(aq) + 4H_2O(\ell) \rightarrow$ $2ClO_2^-(aq) + Sb(OH)_6^-(aq) + 2H^+(aq)$
Step 7 Add as many OH^- ions to both sides as there are H^+ ions. On the right, combine the OH^- and H^+ ions to make water.	There are 2 H^+ ions on the right. Add 2 OH^- to both sides. $2ClO_2(aq) + SbO_2^-(aq) + 4H_2O(\ell) +$ $2(OH)^-(aq) \rightarrow$ $2ClO_2^-(aq) + Sb(OH)_6^-(aq) + 2H_2O(\ell)$
Step 8 Cancel the number of water molecules that are on both sides of the equation.	There are 2 water molecules on the right and 4 on the left. Delete 2 water molecules from each side of the equation. $2ClO_2(aq) + SbO_2^-(aq) + 2H_2O(\ell) +$ $2(OH)^-(aq) \rightarrow$ $2ClO_2^-(aq) + Sb(OH)_6^-(aq)$ Atoms of all elements are now balanced.

Check Your Solution

Atoms of all of the elements are balanced. The total charge on the left side is –3 and the total charge on the right side is also –3. The charge is balanced.

39. Use the oxidation number method to balance the ionic equation in an acidic solution:

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

What Is Required?
You must balance an equation for the given redox reaction using the oxidation number method in an acidic solution.

What Is Given?
You are given the reaction: $Cr_2O_7^{2-}(aq) + Fe^{2+}(aq) \rightarrow Cr^{3+}(aq) + Fe^{3+}(aq)$
You are given rules for assigning oxidation numbers in **Table 9.3** on page 604 of the student textbook.

Plan Your Strategy	Act on Your Strategy
Step 1 Assign an oxidation number to each atom in the equation and determine whether it is a redox reaction.	$\underset{+6}{Cr_2}\underset{-2}{O_7^{2}}(aq) + \underset{+2}{Fe^{2+}}(aq) \longrightarrow 2\underset{+3}{Cr^{3+}}(aq) + \underset{+3}{Fe^{3+}}(aq)$ The oxidation numbers of chromium and iron change during the reaction. It is a redox reaction.
Step 2 Identify the atom or atoms that undergo an increase in oxidation number and the atom or atoms that undergo a decrease in oxidation number.	Each of two chromium atoms gains 3 electrons for a total of 6. The iron ion loses 1 electron. $1(+6e^-) = +6e^-$ $\underset{+6}{Cr_2}\underset{-2}{O_7^{2-}}(aq) + \underset{+2}{Fe^{2+}}(aq) \longrightarrow 2\underset{+3}{Cr^{3+}}(aq) + \underset{+3}{Fe^{3+}}(aq)$ $6(-1e^-) = -6e^-$
Step 3 Balance the atoms that undergo a change in oxidation number. Place coefficients in front of the reactants that will balance the number of electrons that are lost by one reactant and gained by another.	Place a coefficient of 6 in front of the iron ion and a coefficient of 1 is already implied for the $Cr_2O_7^{2-}$ ion. This ratio of 1:6 must be maintained for the remainder of the balancing. $Cr_2O_7^{2-}(aq) + 6Fe^{2+}(aq) \rightarrow Cr^{3+}(aq) + Fe^{3+}(aq)$
Step 4 Balance atoms of all elements except oxygen and hydrogen.	There are 2 Cr atoms on the left and 1 on the right. There are 6 Fe^{2+} ions on the left and 1 on the right. Place a coefficient of 2 in front of the Cr^{3+} ion and a 6 in front of the Fe^{3+} ion in the products. $Cr_2O_7^{2-}(aq) + 6Fe^{2+}(aq) \rightarrow 2Cr^{3+}(aq) + 6Fe^{3+}(aq)$

Step 5 Balance the O atoms by adding H_2O.	There are 7 O atoms on the left and none on the right so add 7 water molecules to the right side of the equation. $Cr_2O_7^{2-}(aq) + 6Fe^{2+}(aq) \rightarrow$ $2Cr^{3+}(aq) + 6Fe^{3+}(aq) + 7H_2O(\ell)$
Step 6 Balance the hydrogen atoms by adding hydrogen ions.	There are 14 H atoms on the right and none on the left, so add 14 H^+ ions to the left side of the equation. $Cr_2O_7^{2-}(aq) + 6Fe^{2+}(aq) + 14H^+ \rightarrow$ $2Cr^{3+}(aq) + 6Fe^{3+}(aq) + 7H_2O(\ell)$ Atoms of all elements are now balanced.

Check Your Solution

Atoms of all of the elements are balanced. The total charge on the left side is +24 and the total charge on the right side is also +24. The charge is balanced.

40. Use the oxidation number method to balance the ionic equation in an acidic solution:

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

What Is Required?
You must balance an equation for the given redox reaction using the oxidation number method in an acidic solution.

What Is Given?
You are given the reaction: $I_2(g) + NO_3^-(aq) \rightarrow IO_3^-(aq) + NO_2(g)$
You are given rules for assigning oxidation numbers in **Table 9.3** on page 604 of the student textbook.

Plan Your Strategy	Act on Your Strategy
Step 1 Assign an oxidation number to each atom in the equation and determine whether it is a redox reaction.	$I_2(g) + NO_3^-(aq) \longrightarrow IO_3^-(aq) + NO_2(g)$ 0 +5 -2 +5 -2 +4 -2 The oxidation numbers of iodine and nitrogen change during the reaction. It is a redox reaction.
Step 2 Identify the atom or atoms that undergo an increase in oxidation number and the atom or atoms that undergo a decrease in oxidation number.	Each of 2 I atoms in I_2 loses 5 electrons for a total of 10 electrons. The nitrogen atom in NO_3^- gains one electron. $1(-10e^-) = -10e^-$ $I_2(g) + NO_3^-(aq) \longrightarrow IO_3^-(aq) + NO_2(g)$ 0 +5 -2 +5 -2 +4 -2 $10(+1e^-) = +10e^-$
Step 3 Balance the atoms that undergo a change in oxidation number. Place coefficients in front of the reactants that will balance the number of electrons that are lost by one reactant and gained by another.	Give a coefficient of 10 to the NO_3^- and a coefficient of 1 is already implied for the I_2 molecule. $I_2(g) + 10NO_3^-(aq) \rightarrow IO_3^-(aq) + NO_2(g)$ This ratio of 1:10 must be maintained for the remainder of balancing.

Step 4 Balance atoms of all elements except oxygen and hydrogen.	There are 2 iodine atoms on the left and 1 on the right so place a coefficient of 2 in front of the iodate ion on the right of the equation. There are 10 nitrogen atoms on the left and 1 on the right so place a coefficient of 10 in front of the NO_2 molecule on the right of the equation. $I_2(g) + 10NO_3^-(aq) \rightarrow 2IO_3^-(aq) + 10NO_2(g)$
Step 5 Balance the O atoms by adding H_2O.	There are 30 O atoms on the left and 26 on the right so add 4 water molecules to the right side of the equation. $I_2(g) + 10NO_3^-(aq) \rightarrow$ $2IO_3^-(aq) + 10NO_2(g) + 4H_2O(\ell)$
Step 6 Balance the hydrogen atoms by adding hydrogen ions	There are 8 hydrogen atoms on the right side and none on the left, so add 8 hydrogen ions to the left side of the equation. $8H^+(aq) + I_2(g) + 10NO_3^-(aq) \rightarrow$ $2IO_3^-(aq) + 10NO_2(g) + 4H_2O(\ell)$ Atoms of all elements are now balanced.

Check Your Solution

Atoms of all of the elements are balanced. The total charge on the left side is –2 and the total charge on the right side is also –2. The charge is balanced.

41. Use the oxidation number method to balance the ionic equation in an acidic solution:

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

What Is Required?
You must balance an equation for the given redox reaction using the oxidation number method in an acidic solution.

What Is Given?
You are given the reaction: $PbSO_4(aq) \rightarrow Pb(s) + PbO_2(aq) + SO_4^{2-}(aq)$
You are given rules for assigning oxidation numbers in **Table 9.3** on page 604 of the student textbook.

Plan Your Strategy	Act on Your Strategy
Step 1 Assign an oxidation number to each atom in the equation and determine whether it is a redox reaction.	$\underset{+2\ +6\ -2}{PbSO_4}(aq) \longrightarrow \underset{0}{Pb}(s) + \underset{+4\ -2}{PbO_2}(aq) + \underset{+6\ -2}{SO_4^{2-}}(aq)$ The oxidation numbers of lead changes during the reaction. It is a disproportionation reaction.
Step 2 Identify the atom or atoms that undergo an increase in oxidation number and the atom or atoms that undergo a decrease in oxidation number.	For every Pb atom that loses an electron, another atom of Pb gains an electron. $-2e^-$ $\underset{+2\ +6\ -2}{PbSO_4}(aq) \longrightarrow \underset{0}{Pb}(s) + \underset{+4\ -2}{PbO_2}(aq) + \underset{+6\ -2}{SO_4^{2-}}(aq)$ $+2e^-$
Step 3 Balance the atoms that undergo a change in oxidation number. Place coefficients in front of the reactants that will balance the number of electrons that are lost by one reactant and gained by another.	The same number of electrons that are gained by one Pb atom are lost by another. Therefore, there must be an even number of Pb atoms on the left so exactly half of them can lose electrons and half of them can gain electrons. Place a coefficient of 2 in front of the $PbSO_4$ on the left side of the equation. This coefficient must remain even for the remainder of the balancing. $2PbSO_4(aq) \rightarrow Pb(s) + PbO_2(aq) + SO_4^{2-}(aq)$

Step 4 Balance atoms of all elements except oxygen and hydrogen.	The Pb atoms are balanced. There are 2 S atoms on the left and 1 on the right. Place a coefficient of 2 in front of the SO_4^{2-} on the right side of the equation. $2PbSO_4(aq) \rightarrow Pb(s) + PbO_2(aq) + 2SO_4^{2-}(aq)$
Step 5 Balance the oxygen atoms by adding water molecules.	There are 8 atoms of oxygen on the left side and 10 on the right side, so add 2 water molecules to the left side of the equation. $2H_2O(\ell) + 2PbSO_4(aq) \rightarrow$ $Pb(s) + PbO_2(aq) + 2SO_4^{2-}(aq)$
=**Step 6** Balance the hydrogen atoms by adding hydrogen ions	There are 4 hydrogen atoms on the left side and none on the right, so add 4 hydrogen ions to the right side of the equation. $2H_2O(\ell) + 2PbSO_4(aq) \rightarrow$ $Pb(s) + PbO_2(aq) + 2SO_4^{2-}(aq) + 4H^+(aq)$ Atoms of all elements are now balanced.

Check Your Solution
Atoms of all of the elements are balanced. The total charge on the left side is 0 and the total charge on the right side is also 0. The charge is balanced.

42. Use the oxidation number method to balance the ionic equation in a basic solution:

$$Cl^{-}(aq) + CrO_4^{2-}(aq) \rightarrow ClO^{-}(aq) + CrO_2^{-}(aq)$$

What Is Required?
You must balance an equation for the given redox reaction using the oxidation number method in a basic solution.

What Is Given?
You are given the reaction: $Cl^{-}(aq) + CrO_4^{2-}(aq) \rightarrow ClO^{-}(aq) + CrO_2^{-}(aq)$
You are given rules for assigning oxidation numbers in **Table 9.3** on page 604 of the student textbook.

Plan Your Strategy	Act on Your Strategy
Step 1 Assign an oxidation number to each atom in the equation and determine whether it is a redox reaction.	$\underset{-1}{Cl}^{-}(aq) + \underset{+6}{Cr}\underset{-2}{O}_4^{2-}(aq) \longrightarrow \underset{+1}{Cl}\underset{-2}{O}^{-}(aq) + \underset{+3}{Cr}\underset{-2}{O}_2^{-}(aq)$ The oxidation numbers of chromium and chlorine change during the reaction. It is a redox reaction.
Step 2 Identify the atom or atoms that undergo an increase in oxidation number and the atom or atoms that undergo a decrease in oxidation number.	The chlorine atom lost 2 electrons and the chromium atom gained 3 electrons. $3(-2e^{-}) = -6e^{-}$ $\underset{-1}{Cl}^{-}(aq) + \underset{+6}{Cr}\underset{-2}{O}_4^{2-}(aq) \longrightarrow \underset{+1}{Cl}\underset{-2}{O}^{-}(aq) + \underset{+3}{Cr}\underset{-2}{O}_2^{-}(aq)$ $2(+3e^{-}) = +6e^{-}$
Step 3 Determine the numerical values of the total increase and decrease in oxidation numbers. Place coefficients in front of the reactants that will balance the number of electrons that are lost by one reactant and gained by another.	The Cl^{-} needs a coefficient of 3 and the CrO_4^{2-} needs a coefficient of 2. This ratio of 3:2 must be maintained for the remainder of the balancing. $3Cl^{-}(aq) + 2CrO_4^{2-}(aq) \rightarrow$ $ClO^{-}(aq) + CrO_2^{-}(aq)$
Step 4 Balance atoms of all elements except oxygen and hydrogen.	There are 3 Cl atoms on the left and 1 on the right. Place a coefficient of 3 in front of the ClO^{-} on the right side of the equation. There are 2 Cr atoms on the left and 1 on the right. Place a coefficient of 2 in front of the CrO_2^{-} on the right side of the equation. $3Cl^{-}(aq) + 2CrO_4^{2-}(aq) \rightarrow$ $3ClO^{-}(aq) + 2CrO_2^{-}(aq)$

Step 5 Balance the oxygen atoms by adding water molecules.	There are 8 atoms of oxygen on the left side and 7 on the right side, so add 1 water molecule to the right side. $3Cl^{-}(aq) + 2CrO_4^{2-}(aq) \rightarrow$ $3ClO^{-}(aq) + 2CrO_2^{-}(aq) + H_2O(\ell)$
Step 6 Balance the hydrogen atoms by adding hydrogen ions	There are 2 hydrogen atoms on the right side and none on the left, so add 2 hydrogen ions to the left side. $2H^{+}(aq) + 3Cl^{-}(aq) + 2CrO_4^{2-}(aq) \rightarrow$ $3ClO^{-}(aq) + 2CrO_2^{-}(aq) + H_2O(\ell)$
Step 7 Add as many OH^{-} ions to both sides as there are H^{+} ions. When H^{+} and OH^{-} ions appear on the same side of the equation, combine them to make water.	There are 2 hydrogen ions on the left side, so add 2 hydroxide ions to each side. $2H_2O(\ell) + 3Cl^{-}(aq) + 2CrO_4^{2-}(aq) \rightarrow$ $3ClO^{-}(aq) + 2CrO_2^{-}(aq) + H_2O(\ell) + 2OH^{-}(aq)$
Step 8 Cancel the number of water molecules that are on both sides of the equation.	There are 2 water molecules on the left and 1 on the right. Delete 1 water molecule from each side. $H_2O(\ell) + 3Cl^{-}(aq) + 2CrO_4^{2-}(aq) \rightarrow$ $3ClO^{-}(aq) + 2CrO_2^{-}(aq) + 2OH^{-}(aq)$ Atoms of all elements are now balanced.

Check Your Solution

Atoms of all of the elements are balanced. The total charge on the left side is –7 and the total charge on the right side is also –7. The charge is balanced.

43. Use the oxidation number method to balance the ionic equation in a basic solution:

$$Ni(s) + MnO_4^-(aq) \rightarrow NiO(s) + MnO_2(s)$$

What Is Required?
You must balance an equation for the given redox reaction using the oxidation number method in a basic solution.

What Is Given?
You are given the reaction: $Ni(s) + MnO_4^-(aq) \rightarrow NiO(s) + MnO_2(s)$
You are given rules for assigning oxidation numbers in **Table 9.3** on page 604 of the student textbook.

Plan Your Strategy	Act on Your Strategy
Step 1 Assign an oxidation number to each atom in the equation and determine whether it is a redox reaction.	$\underset{0}{Ni}(s) + \underset{+7}{Mn}\underset{-2}{O_4^-}(aq) \longrightarrow \underset{+2}{Ni}\underset{-2}{O}(s) + \underset{+4}{Mn}\underset{-2}{O_2}(s)$ The oxidation numbers of nickel and manganese change during the reaction. It is a redox reaction.
Step 2 Identify the atom or atoms that undergo an increase in oxidation number and the atom or atoms that undergo a decrease in oxidation number.	The nickel atom lost 2 electron and the manganese atom gained 3 electrons. $3(-2e^-) = -6e^-$ $\underset{0}{Ni}(s) + \underset{+7}{Mn}\underset{-2}{O_4^-}(aq) \longrightarrow \underset{+2}{Ni}\underset{-2}{O}(s) + \underset{+4}{Mn}\underset{-2}{O_2}(s)$ $2(+3e^-) = +6e^-$
Step 3 Place coefficients in front of the reactants that will balance the number of electrons that are lost by one reactant and gained by another.	The Ni needs a coefficient of 3 and the MnO_4^- needs a coefficient of 2. This ratio of 3:2 must be maintained for the remainder of the balancing. $3Ni(s) + 2MnO_4^-(aq) \rightarrow NiO(s) + MnO_2(s)$
Step 4 Balance atoms of all elements except oxygen and hydrogen.	There are three Ni atoms on the left and 1 on the right. Place a coefficient of 3 in front of the NiO on the right side of the equation. There are 2 Mn atoms on the left and 1 on the right. Place a coefficient of 2 in front of the MnO_2 on the right side of the equation. $3Ni(s) + 2MnO_4^-(aq) \rightarrow 3NiO(s) + 2MnO_2(s)$

Step 5 Balance the oxygen atoms by adding water molecules.	There are 8 atoms of oxygen on the left side and 7 on the right side, so add 1 water molecule to the right side. $3Ni(s) + 2MnO_4^-(aq) \rightarrow$ $3NiO(s) + 2MnO_2(s) + H_2O(\ell)$
Step 6 Balance the hydrogen atoms by adding hydrogen ions.	There are 2 hydrogen atoms on the right side and none on the left, so add 2 hydrogen ions to the left side. $2H^+(aq) + 3Ni(s) + 2MnO_4^-(aq) \rightarrow$ $3NiO(s) + 2MnO_2(s) + H_2O(\ell)$
Step 7 Add as many OH^- ions to both sides as there are H^+ ions. When H^+ and OH^- ions appear on the same side of the equation, combine them to make water.	There are 2 hydrogen ions on the left side, so add 2 hydroxide ions to each side. $3Ni(s) + 2MnO_4^-(aq) + 2H_2O(\ell) \rightarrow$ $3NiO(s) + 2MnO_2(s) + H_2O(\ell) + 2OH^-(aq)$
Step 8 Cancel the number of water molecules that are on both sides of the equation.	There are 2 water molecules of the left side and 1 on the right side of the equation. Delete 1 water molecule from each side. $3Ni(s) + 2MnO_4^-(aq) + H_2O(\ell) \rightarrow$ $3NiO(s) + 2MnO_2(s) + 2OH^-(aq)$ Atoms of all elements are now balanced.

Check Your Solution

Atoms of all of the elements are balanced. The total charge on the left side is –2 and the total charge on the right side is also –2. The charge is balanced.

44. Use the oxidation number method to balance the ionic equation in a basic solution:

$$I^-(aq) + Ce^{4+}(aq) \rightarrow IO_3^-(aq) + Ce^{3+}(aq)$$

What Is Required?
You must balance an equation for the given redox reaction using the oxidation number method in a basic solution.

What Is Given?
You are given the reaction: $I^-(aq) + Ce^{4+}(aq) \rightarrow IO_3^-(aq) + Ce^{3+}(aq)$
You are given rules for assigning oxidation numbers in **Table 9.3** on page 604 of the student textbook.

Plan Your Strategy	Act on Your Strategy
Step 1 Assign an oxidation number to each atom in the equation and determine whether it is a redox reaction.	$\underset{-1}{I^-}(aq) + \underset{+4}{Ce^{4+}}(aq) \longrightarrow \underset{+5\ -2}{IO_3^-}(aq) + \underset{+3}{Ce^{3+}}(aq)$ The oxidation numbers of cerium and iodine change during the reaction. It is a redox reaction.
Step 2 Identify the atom or atoms that undergo an increase in oxidation number and the atom or atoms that undergo a decrease in oxidation number.	The iodine ion lost 6 electrons and the cerium ion gained 1 electron. $1(-6e^-) = -6e^-$ $\underset{-1}{I^-}(aq) + \underset{+4}{Ce^{4+}}(aq) \longrightarrow \underset{+5\ -2}{IO_3^-}(aq) + \underset{+3}{Ce^{3+}}(aq)$ $6(+1e^-) = +6e^-$
Step 3 Determine the numerical values of the total increase and decrease in oxidation numbers. Place coefficients in front of the reactants that will balance the number of electrons that are lost by one reactant and gained by another.	The I^- needs a coefficient of 1 which is already implied. The Ce^{4+} needs a coefficient of 6. This ratio of 1:6 must be maintained for the remainder of the balancing. $I^-(aq) + 6Ce^{4+}(aq) \rightarrow IO_3^-(aq) + Ce^{3+}(aq)$
Step 4 Balance atoms of all elements except oxygen and hydrogen.	The I is balanced. There are 6 Ce atoms on the left and 1 on the right side of the equation. Place a coefficient of 6 in front of the Ce^{3+} on the right side of the equation. $I^-(aq) + 6Ce^{4+}(aq) \rightarrow IO_3^-(aq) + 6Ce^{3+}(aq)$ This ratio of 1:6 must be maintained for the remainder of balancing.
Step 5 Balance the oxygen atoms by	There are 3 oxygen atoms on the right and none

adding water molecules.	on the left side of the equation. Add 3 water molecules to the left side of the equation. $I^-(aq) + 6Ce^{4+}(aq) + 3H_2O(\ell) \rightarrow$ $IO_3^-(aq) + 6Ce^{3+}(aq)$
Step 6 Balance the hydrogen atoms by adding hydrogen ions.	There are 6 H atoms on the left and none on the right side of the equation. Add 6 H^+ ions to the right side of the equation. $I^-(aq) + 6Ce^{4+}(aq) + 3H_2O(\ell) \rightarrow$ $IO_3^-(aq) + 6Ce^{3+}(aq) + 6H^+(aq)$
Step 7 Add as many OH^- ions to both sides as there are H^+ ions. When H^+ and OH^- ions appear on the same side of the equation, combine them to make water.	There are 6 H^+ ions on the right side so add 6 OH^- to both sides of the equation. $I^-(aq) + 6Ce^{4+}(aq) + 3H_2O(\ell) + 6OH^-(aq) \rightarrow$ $IO_3^-(aq) + 6Ce^{3+}(aq) + 6H_2O(\ell)$
Step 8 Cancel the number of water molecules that are on both sides of the equation.	There are 3 water molecules on the left and 6 on the right so delete 3 water molecules from both sides of the equation. $I^-(aq) + 6Ce^{4+}(aq) + 6OH^-(aq) \rightarrow$ $IO_3^-(aq) + 6Ce^{3+}(aq) + 3H_2O(\ell)$ Atoms of all elements are now balanced.

Check Your Solution

Atoms of all of the elements are balanced. The total charge on the left side is +17 and the total charge on the right side is also +17. The charge is balanced.

Solutions to Practice Problems in Chapter 10 Electrochemical Cells

Using Half-Reactions to Sketch a Galvanic Cell
(Student textbook page 641)

1. In a galvanic cell involving zinc and magnesium, which electrode will be the anode and which will be the cathode?

What Is Required?
You are asked to identify the anode and cathode of a galvanic cell.

What Is Given?
You are given that one electrode is zinc and the second electrode is magnesium. You are also given Relative Strengths of Oxidizing and Reducing Agents, **Table 9.2** on page 587 of the student textbook.

Plan Your Strategy	Act on Your Strategy
Identify the stronger reducing agent from Table 9.2.	Between zinc and magnesium, magnesium is the stronger reducing agent.
Since the stronger reducing agent is always oxidized, and oxidation occurs at the anode, the anode can be identified.	Since magnesium is the stronger reducing agent, magnesium is oxidized and thus is the anode in the galvanic cell.
Once the anode is identified, the other electrode will be the cathode.	Given that the magnesium electrode is the anode, the zinc electrode must be the cathode.

Check Your Solution
The stronger reducing agent, according to Table 9.2, is the anode, and the weaker reducing agent is the cathode. The solution is correct.

2. Explain where the oxidation and the reduction are occurring in a galvanic cell involving zinc and magnesium.

What Is Required?
You are to identify where the oxidation and the reduction are occurring in a galvanic cell.

What Is Given?
You are given that one electrode is zinc and the second electrode is magnesium. You are also given Relative Strengths of Oxidizing and Reducing Agents, **Table 9.2** on page 587 of the student textbook.

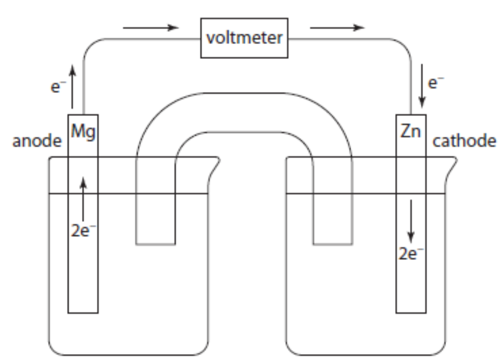
Plan Your Strategy	Act on Your Strategy
Identify the stronger reducing agent from Table 9.2.	Between zinc and magnesium, magnesium is the stronger reducing agent.
Since the stronger reducing agent is always oxidized, oxidation occurs at the anode.	Oxidation occurs at the magnesium electrode.
The weaker reducing agent is always reduced and reduction occurs at the cathode.	Reduction occurs at the zinc electrode.

Check Your Solution
Oxidation occurs at the electrode that is the stronger reducing agent, and reduction occurs at the electrode that is the weaker reducing agent. The solution is correct.

3. Draw a diagram of a galvanic cell involving zinc and magnesium, and indicate the direction of electron flow.

What Is Required?
You are to draw a galvanic cell, labelling the anode, the cathode, and the direction of electron flow.

What Is Given?
You are given the electrodes of magnesium and zinc for this galvanic cell.

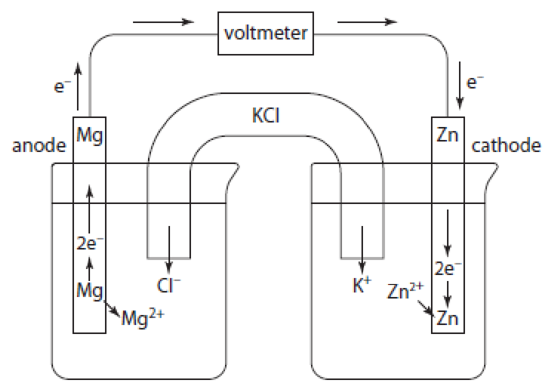
Plan Your Strategy	Act on Your Strategy
Identify the pieces of the galvanic cell.	The magnesium is the anode—the site of oxidation, where magnesium ions flow into solution and electrons leave the electrode and flow through the voltmeter as they move to the zinc cathode. The zinc is the cathode—the site of reduction, where zinc ions flow onto the electrode as they pick up electrons.
Identify the direction of electron flow.	In a galvanic cell, electrons flow from the anode to the cathode, in this case, from the magnesium electrode to the zinc electrode.
Identify the ion flow in the half-cells.	Ions of magnesium will flow into the oxidation half-reaction as the atoms of magnesium lose electrons. Ions of zinc will flow onto the zinc electrode as they pick up electrons during the reduction.
Identify the ion flow in the salt bridge.	Negative ions in the salt bridge will migrate towards the oxidation half-reaction and positive charge builds up in the cell. Positive ions will migrate towards the reduction half-cell as positive charge is removed from the half-cell, leaving a net negative charge in this half-cell that will attract the positive ions in the salt bridge.
Sketch the apparatus, including the beakers, electrodes, conducting wires, voltmeter, salt bridge, and electron flow.	

Check Your Solution
Electrons flow from the site of oxidation at the magnesium anode to the site of reduction at the zinc cathode. The drawing is correct.

4. Indicate the direction of the ionic movement in a galvanic cell involving zinc and magnesium.

What Is Required?
You are asked to determine the direction of the ionic movement in a galvanic cell.

What Is Given?
You are given that the electrodes of the galvanic cell are magnesium and zinc. You have also identified that the magnesium is the anode and the zinc is the cathode in the galvanic cell.

Plan Your Strategy	Act on Your Strategy
Identify the oxidation half-cell where electrons will be lost as positive ions flow into solution.	The magnesium atoms lose electrons and thus magnesium ions will flow into the magnesium half-cell.
Identify the reduction half-cell where the electrons will be gained as positive ions flow to the electrode.	The zinc ions in solution will flow onto the zinc electrode in the zinc half-cell.
Add the ion flow to your diagram.	

Check Your Solution
Positive magnesium ions flow from the site of oxidation at the magnesium anode into the solution. Positive zinc ions flow onto the site of reduction at the zinc cathode. The drawing is correct.

5. Write a balanced ionic equation that represents the reaction in a galvanic cell involving zinc and magnesium.

What Is Required?
You are asked to write a balanced ionic equation to represent a galvanic reaction.

What Is Given?
You are given the galvanic cell involving zinc and magnesium.

Plan Your Strategy	Act on Your Strategy
Write the oxidation half-reaction.	$Mg(s) \rightarrow Mg^{2+}(aq) + 2e^{-}$
Write the reduction half-reaction.	$Zn^{2+}(aq) + 2e^{-} \rightarrow Zn(s)$
Balance the electrons lost and gained in the two half-reactions.	The number of electrons lost is 2 and the number gained is 2, therefore the gain and loss are already balanced.
Combine the two half-reactions, cancelling the electrons lost and gained.	$Mg(s) + Zn^{2+}(aq) \rightarrow Zn(s) + Mg^{2+}(aq)$

Check Your Solution
The two sides are balanced by mass and charge. Therefore, the reaction is balanced.

6. In a galvanic cell involving nickel and silver, which electrode will be the anode and which will be the cathode?

What Is Required?
You are asked to identify the anode and cathode of a galvanic cell.

What Is Given?
You are given that one electrode is nickel and the second electrode is silver. You are also given Relative Strengths of Oxidizing and Reducing Agents, **Table 9.2** on page 587 of the student textbook.

Plan Your Strategy	Act on Your Strategy
Identify the stronger reducing agent from Table 9.2	Between nickel and silver, nickel is the stronger reducing agent.
Since the stronger reducing agent is always oxidized, and oxidation occurs at the anode, the anode can be identified.	Since nickel is the stronger reducing agent, nickel is oxidized and thus is the anode in the galvanic cell.
Once the anode is identified, the other electrode will be the cathode.	Given that the nickel electrode is the anode, the silver electrode must be the cathode.

Check Your Solution
The stronger reducing agent, according to Table 9.2, is the anode, and the weaker reducing agent is the cathode. The solution is correct.

7. Explain where the oxidation and the reduction are occurring in a galvanic cell involving silver and nickel.

What Is Required?
You are to identify where the oxidation and the reduction are occurring in a galvanic cell.

What Is Given?
You are given that one electrode is silver and the second electrode is nickel. You are also given Relative Strengths of Oxidizing and Reducing Agents, **Table 9.2** on page 587 of the student textbook.

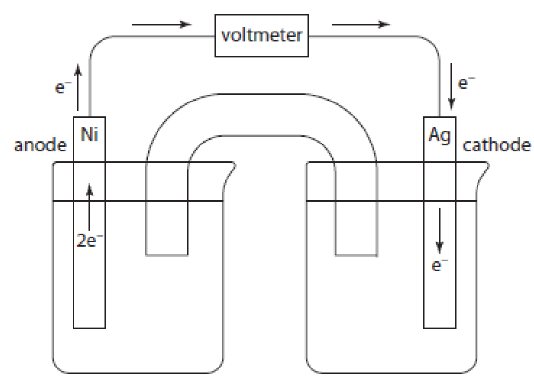
Plan Your Strategy	Act on Your Strategy
Identify the stronger reducing agent from Table 9.2.	Between silver and nickel, nickel is the stronger reducing agent.
Since the stronger reducing agent is always oxidized, oxidation occurs at the anode.	Oxidation occurs at the nickel electrode.
The weaker reducing agent is always reduced and reduction occurs at the cathode.	Reduction occurs at the silver electrode.

Check Your Solution
Oxidation occurs at the electrode that is the stronger reducing agent, and reduction occurs at the electrode that is the weaker reducing agent. The solution is correct.

8. Draw a diagram of a galvanic cell involving silver and nickel, and indicate the direction of electron flow.

What Is Required?
You are to draw a galvanic cell, labelling the anode, the cathode, and the direction of electron flow.

What Is Given?
You are given the electrodes of nickel and silver for this galvanic cell.

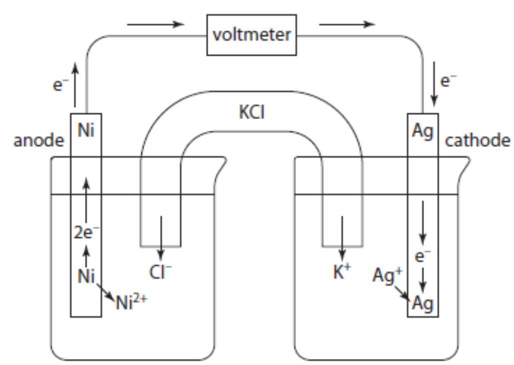
Plan Your Strategy	Act on Your Strategy
Identify the pieces of the galvanic cell.	The nickel is the anode—the site of oxidation where nickel ions flow into solution and electrons leave the electrode and flow through the voltmeter as they move to the silver cathode. The silver is the cathode—the site of reduction where silver ions flow onto the electrode as they pick up electrons.
Identify the direction of electron flow.	In a galvanic cell, electrons flow from the anode to the cathode, in this case, from the nickel electrode to the silver electrode.
Identify the ion flow in the half-cells.	Ions of nickel will flow into the oxidation half-reaction as the atoms of nickel lose electrons. Ions of silver will flow onto the silver electrode as they pick up electrons during the reduction.
Identify the ion flow in the salt bridge.	Negative ions in the salt bridge will migrate towards the oxidation half-reaction and positive charge builds up in the cell. Positive ions will migrate towards the reduction half-cell as positive charge is removed from the half-cell, leaving a net negative charge in this half-cell that will attract the positive ions in the salt bridge.
Sketch the apparatus, including the beakers, electrodes, conducting wires, voltmeter, salt bridge, and electron flow.	

Check Your Solution
Electrons flow from the site of oxidation at the nickel anode to the site of reduction at the silver cathode. The drawing is correct.

9. Indicate the direction of the ionic movement in a galvanic cell involving silver and nickel.

What Is Required?
You are asked to determine the direction of the ionic movement in a galvanic cell.

What Is Given?
You are given that the electrodes of the galvanic cell are nickel and silver. You have also identified that the nickel is the anode and the silver is the cathode in the galvanic cell.

Plan Your Strategy	Act on Your Strategy
Identify the oxidation half-cell where electrons will be lost as positive ions flow into solution.	The nickel atoms lose electrons and thus nickel ions will flow into the nickel half-cell.
Identify the reduction half-cell where the electrons will be gained as positive ions flow to the electrode.	The silver ions in solution will flow onto the silver electrode in the silver half-cell.
Add the ion flow to your diagram.	

Check Your Solution
Positive nickel ions flow from the site of oxidation at the nickel anode into the solution. Positive silver ions flow onto the site of reduction at the silver cathode. The drawing is correct.

10. Write a balanced ionic equation that represents the reaction in a galvanic cell involving silver and nickel.

What Is Required?
You are asked to write a balanced ionic equation to represent a galvanic reaction.

What Is Given?
You are given the galvanic cell involving silver and nickel.

Plan Your Strategy	Act on Your Strategy
Write the oxidation half-reaction.	$Ni(s) \rightarrow Ni^{2+}(aq) + 2e^-$
Write the reduction half-reaction.	$Ag^+(aq) + e^- \rightarrow Ag(s)$
Balance the electrons lost and gained in the two half-reactions.	The number of electrons lost is 2 and the number gained is 1. Therefore, the gain needs to be multiplied by 2. $2Ag^+(aq) + 2e^- \rightarrow 2Ag(s)$
Combine the two half-reactions, cancelling the electrons lost and gained.	$Ni(s) + 2Ag^+(aq) \rightarrow 2Ag(s) + Ni^{2+}(aq)$

Check Your Solution
The two sides are balanced by mass and charge. Therefore, the reaction is balanced.

Calculating a Standard Cell Potential, Given a Net Ionic Equation
Calculating a Standard Cell Potential, Given a Chemical Reaction
(Student textbook page 647)

11. For the reaction $Cl_2(g) + 2Br^-(aq) \rightarrow 2Cl^-(aq) + Br_2(\ell)$:
a. write the oxidation and reduction half-reactions.
b. determine the standard cell potentials for galvanic cells in which these reactions occur.
c. predict whether the reaction will proceed spontaneously as written.

What Is Required?
a. You need to write the oxidation and reduction half reactions for a given redox reaction.
b. You need to determine the standard cell potential for the galvanic cell given.
c. You need to determine whether the reaction will proceed spontaneously as written.

What Is Given?
You are given the balanced net ionic equation.
You have a table of standard reduction potentials.

Plan Your Strategy	Act on Your Strategy
a. Break the redox reaction into the two half-reactions that make up the overall reaction.	Oxidation: $2Br^-(aq) \rightarrow 2Br^-(aq) + 2e^-$ Reduction: $Cl_2(g) + 2e^- \rightarrow 2Cl^-(aq)$
b. Calculate the standard cell-potential for the galvanic cell by locating the relevant reduction potentials in a table of standard reduction potentials.	$Br_2(\ell) + 2e^- \rightarrow 2Br^-(aq)$ $E^{o}_{anode} = +1.07\text{ V}$ $Cl_2(g) + 2e^- \rightarrow 2Cl^-(aq)$ $E^{o}_{cathode} = +1.36\text{ V}$
Subtract the reduction potentials to determine the cell potential by using the formula $E^{o}_{cell} = E^{o}_{cathode} - E^{o}_{anode}$.	$E^{o}_{cell} = E^{o}_{cathode} - E^{o}_{anode}$ $= 1.36\text{ V} - (1.07\text{ V})$ $= +0.29\text{ V}$
c. Predict whether the reaction will proceed spontaneously.	Since the cell potential is positive, the reaction is spontaneous in the direction written.

Check Your Solution
The half-reactions are correct, and the reduction potentials are correct according to the table of standard reduction potentials on page 748 of the Chemistry 12 student textbook. The prediction that a positive cell potential will result in a spontaneous reaction is correct as well.

12. For the reaction $Mg(s) + 2AgNO_3(aq) \rightarrow Mg(NO_3)_2(aq) + 2Ag(s)$:

a. write the oxidation and reduction half-reactions.
b. determine the standard cell potentials for galvanic cells in which these reactions occur.
c. predict whether the reaction will proceed spontaneously as written.

What Is Required?

a. You need to write the oxidation and reduction half reactions for a given redox reaction.
b. You need to determine the standard cell potential for the galvanic cell given.
c. You need to determine whether the reaction will proceed spontaneously as written.

What Is Given?

You are given the balanced net ionic equation.
You have a table of standard reduction potentials.

Plan Your Strategy	Act on Your Strategy
a. Break the redox reaction into the two half-reactions that make up the overall reaction.	Oxidation: $Mg(s) \rightarrow Mg^{2+}(aq) + 2e^-$ Reduction: $Ag^+(aq) + 1e^- \rightarrow Ag(s)$
b. Calculate the standard cell-potential for the galvanic cell by locating the relevant reduction potentials in a table of standard reduction potentials.	$Mg^{2+}(aq) + 2e^- \rightarrow Mg(s)$ $E^\circ_{\text{anode}} = -2.37\text{ V}$ $Ag^+(aq) + 1e^- \rightarrow Ag(s)$ $E^\circ_{\text{cathode}} = +0.80\text{ V}$
Subtract the reduction potentials to determine the cell potential by using the formula $E^\circ_{\text{cell}} = E^\circ_{\text{cathode}} - E^\circ_{\text{anode}}$.	$E^\circ_{\text{cell}} = E^\circ_{\text{cathode}} - E^\circ_{\text{anode}}$ $= 0.80\text{ V} - (-2.37\text{ V})$ $= +3.17\text{ V}$
c. Predict whether the reaction will proceed spontaneously.	Since the cell potential is positive, the reaction is spontaneous in the direction written.

Check Your Solution

The half-reactions are correct, and the reduction potentials are correct according to the table of standard reduction potentials on page 748 of the Chemistry 12 student textbook. The prediction that a positive cell potential will result in a spontaneous reaction is correct as well.

13. For the reaction $Sn(s) + 2HBr(aq) \rightarrow SnBr_2(aq) + H_2(g)$:

a. write the oxidation and reduction half-reactions.
b. determine the standard cell potentials for galvanic cells in which these reactions occur.
c. predict whether the reaction will proceed spontaneously as written.

What Is Required?

a. You need to write the oxidation and reduction half reactions for a given redox reaction.
b. You need to determine the standard cell potential for the galvanic cell given.
c. You need to determine whether the reaction will proceed spontaneously as written.

What Is Given?

You are given the balanced net ionic equation.
You have a table of standard reduction potentials.

Plan Your Strategy	Act on Your Strategy
a. Break the redox reaction into the two half-reactions that make up the overall reaction.	Oxidation: $Sn(s) \rightarrow Sn^{2+}(aq) + 2e^-$ Reduction: $2H^+ + 2e^- \rightarrow H_2(g)$
b. Calculate the standard cell-potential for the galvanic cell by locating the relevant reduction potentials in a table of standard reduction potentials.	$Sn^{2+}(aq) + 2e^- \rightarrow Sn(s)$ $E^\circ_{\text{anode}} = -0.14\text{ V}$ $2H^+(aq) + 2e^- \rightarrow H_2(g)$ $E^\circ_{\text{cathode}} = 0.00\text{ V}$
Subtract the reduction potentials to determine the cell potential by using the formula $E^\circ_{\text{cell}} = E^\circ_{\text{cathode}} - E^\circ_{\text{anode}}$.	$E^\circ_{\text{cell}} = E^\circ_{\text{cathode}} - E^\circ_{\text{anode}}$ $= 0.00\text{ V} - (-0.14\text{ V})$ $= +0.14\text{ V}$
c. Predict whether the reaction will proceed spontaneously.	Since the cell potential is positive, the reaction is spontaneous in the direction written.

Check Your Solution

The half-reactions are correct, and the reduction potentials are correct according to the table of standard reduction potentials on page 748 of the Chemistry 12 student textbook. The prediction that a positive cell potential will result in a spontaneous reaction is correct as well.

14. For the reaction $Cr(s) + 3AgCl(s) \rightarrow CrCl_3(aq) + 3Ag(s)$:

a. write the oxidation and reduction half-reactions.
b. determine the standard cell potentials for galvanic cells in which these reactions occur.
c. predict whether the reaction will proceed spontaneously as written.

What Is Required?

a. You need to write the oxidation and reduction half reactions for a given redox reaction.
b. You need to determine the standard cell potential for the galvanic cell given.
c. You need to determine whether the reaction will proceed spontaneously as written.

What Is Given?

You are given the balanced net ionic equation.
You have a table of standard reduction potentials.

Plan Your Strategy	Act on Your Strategy
a. Break the redox reaction into the two half-reactions that make up the overall reaction.	Oxidation: $Cr(s) \rightarrow Cr^{3+}(aq) + 3e^-$ Reduction: $Ag^+(aq) + 1e^- \rightarrow Ag(s)$
b. Calculate the standard cell-potential for the galvanic cell by locating the relevant reduction potentials in a table of standard reduction potentials.	$Cr^{3+}(aq) + 3e^- \rightarrow Cr(s) \quad E^o_{anode} = -0.74\ V$ $Ag^+(aq) + 1e^- \rightarrow Ag(s) \quad E^o_{cathode} = +0.80\ V$
Subtract the reduction potentials to determine the cell potential by using the formula $E^o_{cell} = E^o_{cathode} - E^o_{anode}$.	$E^o_{cell} = E^o_{cathode} - E^o_{anode}$ $= 0.80\ V - (-0.74\ V)$ $= +1.54\ V$
c. Predict whether the reaction will proceed spontaneously.	Since the cell potential is positive, the reaction is spontaneous in the direction written.

Check Your Solution

The half-reactions are correct, and the reduction potentials are correct according to the table of standard reduction potentials on page 748 of the Chemistry 12 student textbook. The prediction that a positive cell potential will result in a spontaneous reaction is correct as well.

15. For the reaction $3Fe(s) + 2Cr(NO_3)_3(aq) \rightarrow 2Cr(s) + 3Fe(NO_3)_2(aq)$:

a. write the oxidation and reduction half-reactions.
b. determine the standard cell potentials for galvanic cells in which these reactions occur.
c. predict whether the reaction will proceed spontaneously as written.

What Is Required?

a. You need to write the oxidation and reduction half reactions for a given redox reaction.
b. You need to determine the standard cell potential for the galvanic cell given.
c. You need to determine whether the reaction will proceed spontaneously as written.

What Is Given?

You are given the balanced net ionic equation.
You have a table of standard reduction potentials.

Plan Your Strategy	Act on Your Strategy
a. Break the redox reaction into the two half-reactions that make up the overall reaction.	Oxidation: $Fe(s) \rightarrow Fe^{2+}(aq) + 2e^-$ Reduction: $Cr^{3+}(aq) + 3e^- \rightarrow Cr(s)$
b. Calculate the standard cell-potential for the galvanic cell by locating the relevant reduction potentials in a table of standard reduction potentials.	$Fe^{2+}(aq) + 2e^- \rightarrow Fe(s)$ $\quad E^\circ_{\text{anode}} = -0.45\ \text{V}$ $Cr^{3+}(aq) + 3e^- \rightarrow Cr(s)$ $\quad E^\circ_{\text{cathode}} = -0.74\ \text{V}$
Subtract the reduction potentials to determine the cell potential by using the formula $E^\circ_{\text{cell}} = E^\circ_{\text{cathode}} - E^\circ_{\text{anode}}$.	$E^\circ_{\text{cell}} = E^\circ_{\text{cathode}} - E^\circ_{\text{anode}}$ $= -0.74\ \text{V} - (-0.45\ \text{V})$ $= -0.29\ \text{V}$
c. Predict whether the reaction will proceed spontaneously.	Since the cell potential is negative, the reaction is not spontaneous in the direction written.

Check Your Solution

The half-reactions are correct, and the reduction potentials are correct according to the table of standard reduction potentials on page 748 of the Chemistry 12 student textbook. The prediction that a negative cell potential will result in a non-spontaneous reaction is correct as well.

16. For the reaction $2Al(s) + 3ZnCl_2(aq) \rightarrow 3Zn(s) + 2AlCl_3(aq)$:

a. write the oxidation and reduction half-reactions.
b. determine the standard cell potentials for galvanic cells in which these reactions occur.
c. predict whether the reaction will proceed spontaneously as written.

What Is Required?

a. You need to write the oxidation and reduction half reactions for a given redox reaction.
b. You need to determine the standard cell potential for the galvanic cell given.
c. You need to determine whether the reaction will proceed spontaneously as written.

What Is Given?

You are given the balanced net ionic equation.
You have a table of standard reduction potentials.

Plan Your Strategy	**Act on Your Strategy**
a. Break the redox reaction into the two half-reactions that make up the overall reaction.	Oxidation: $Al(s) \rightarrow Al^{3+}(aq) + 3e^-$ Reduction: $Zn^{2+}(aq) + 2e^- \rightarrow Zn(s)$
b. Calculate the standard cell-potential for the galvanic cell by locating the relevant reduction potentials in a table of standard reduction potentials.	$Al^{3+}(aq) + 3e^- \rightarrow Al(s)$ $E^o_{anode} = -1.66$ V $Zn^{2+}(aq) + 2e^- \rightarrow Zn(s)$ $E^o_{cathode} = -0.76$ V
Subtract the reduction potentials to determine the cell potential by using the formula $E^o_{cell} = E^o_{cathode} - E^o_{anode}$.	$E^o_{cell} = E^o_{cathode} - E^o_{anode}$ $= -0.76\text{ V} - (-1.66\text{ V})$ $= +0.90\text{ V}$
c. Predict whether the reaction will proceed spontaneously.	Since the cell potential is positive, the reaction is spontaneous in the direction written.

Check Your Solution

The half-reactions are correct, and the reduction potentials are correct according to the table of standard reduction potentials on page 748 of the Chemistry 12 student textbook. The prediction that a positive cell potential will result in a spontaneous reaction is correct as well.

17. For the reaction $2AgNO_3(aq) + Zn(s) \rightarrow 2Ag(s) + Zn(NO_3)_2(aq)$:

a. write the oxidation and reduction half-reactions.

b. determine the standard cell potentials for galvanic cells in which these reactions occur.

c. predict whether the reaction will proceed spontaneously as written.

What Is Required?

a. You need to write the oxidation and reduction half reactions for a given redox reaction.

b. You need to determine the standard cell potential for the galvanic cell given.

c. You need to determine whether the reaction will proceed spontaneously as written.

What Is Given?

You are given the balanced net ionic equation.

You have a table of standard reduction potentials.

Plan Your Strategy	Act on Your Strategy
a. Break the redox reaction into the two half-reactions that make up the overall reaction.	Oxidation: $Zn(s) \rightarrow Zn^{2+}(aq) + 2e^-$ Reduction: $Ag^+(aq) + 1e^- \rightarrow Ag(s)$
b. Calculate the standard cell-potential for the galvanic cell by locating the relevant reduction potentials in a table of standard reduction potentials.	$Zn^{2+}(aq) + 2e^- \rightarrow Zn(s)$ $E^o_{anode} = -0.76$ V $Ag^+(aq) + 1e^- \rightarrow Ag(s)$ $E^o_{cathode} = +0.80$ V
Subtract the reduction potentials to determine the cell potential by using the formula $E^o_{cell} = E^o_{cathode} - E^o_{anode}$.	$E^o_{cell} = E^o_{cathode} - E^o_{anode}$ $= 0.80 \text{ V} - (-0.76 \text{ V})$ $= +1.56 \text{ V}$
c. Predict whether the reaction will proceed spontaneously.	Since the cell potential is positive, the reaction is spontaneous in the direction written.

Check Your Solution

The half-reactions are correct, and the reduction potentials are correct according to the table of standard reduction potentials on page 748 of the Chemistry 12 student textbook. The prediction that a positive cell potential will result in a spontaneous reaction is correct as well.

18. For the reaction $3Cu(NO_3)_2(aq) + 2Al(s) \rightarrow 2Al(NO_3)_3(aq) + 3Cu(s)$:

a. write the oxidation and reduction half-reactions.
b. determine the standard cell potentials for galvanic cells in which these reactions occur.
c. predict whether the reaction will proceed spontaneously as written.

What Is Required?
a. You need to write the oxidation and reduction half reactions for a given redox reaction.
b. You need to determine the standard cell potential for the galvanic cell given.
c. You need to determine whether the reaction will proceed spontaneously as written.

What Is Given?
You are given the balanced net ionic equation.
You have a table of standard reduction potentials.

Plan Your Strategy	Act on Your Strategy
a. Break the redox reaction into the two half-reactions that make up the overall reaction.	Oxidation: $Al(s) \rightarrow Al^{3+}(aq) + 3e^-$ Reduction: $Cu^{2+}(aq) + 2e^- \rightarrow Cu(s)$
b. Calculate the standard cell-potential for the galvanic cell by locating the relevant reduction potentials in a table of standard reduction potentials.	$Al^{3+}(aq) + 3e^- \rightarrow Al(s)$ $\quad E^\circ_{\text{anode}} = -1.66\text{ V}$ $Cu^{2+}(aq) + 2e^- \rightarrow Cu(s)$ $\quad E^\circ_{\text{cathode}} = +0.34\text{ V}$
Subtract the reduction potentials to determine the cell potential by using the formula $E^\circ_{\text{cell}} = E^\circ_{\text{cathode}} - E^\circ_{\text{anode}}$.	$E^\circ_{\text{cell}} = E^\circ_{\text{cathode}} - E^\circ_{\text{anode}}$ $= 0.34\text{ V} - (-1.66\text{ V})$ $= +2.00\text{ V}$
c. Predict whether the reaction will proceed spontaneously.	Since the cell potential is positive, the reaction is spontaneous in the direction written.

Check Your Solution
The half-reactions are correct, and the reduction potentials are correct according to the table of standard reduction potentials on page 748 of the Chemistry 12 student textbook. The prediction that a positive cell potential will result in a spontaneous reaction is correct as well.

19. For the reaction $CuSO_4(aq) + 2Ag(s) \rightarrow Ag_2SO_4(aq) + Cu(s)$:

a. write the oxidation and reduction half-reactions.
b. determine the standard cell potentials for galvanic cells in which these reactions occur.
c. predict whether the reaction will proceed spontaneously as written.

What Is Required?

a. You need to write the oxidation and reduction half reactions for a given redox reaction.
b. You need to determine the standard cell potential for the galvanic cell given.
c. You need to determine whether the reaction will proceed spontaneously as written.

What Is Given?

You are given the balanced net ionic equation.
You have a table of standard reduction potentials.

Plan Your Strategy	Act on Your Strategy
a. Break the redox reaction into the two half-reactions that make up the overall reaction.	Oxidation: $Ag(s) \rightarrow Ag^{+(}aq) + 1e^-$ Reduction: $Cu^{2+}(aq) + 2e^- \rightarrow Cu(s)$
b. Calculate the standard cell-potential for the galvanic cell by locating the relevant reduction potentials in a table of standard reduction potentials.	$Ag^+(aq) + 1e^- \rightarrow Ag(s) \quad E^\circ_{anode} = +0.80\text{ V}$ $Cu^{2+}(aq) + 2e^- \rightarrow Cu(s) \quad E^\circ_{cathode} = +0.34\text{ V}$
Subtract the reduction potentials to determine the cell potential by using the formula $E^\circ_{cell} = E^\circ_{cathode} - E^\circ_{anode}$.	$E^\circ_{cell} = E^\circ_{cathode} - E^\circ_{anode}$ $= 0.34\text{V} - (0.80\text{ V})$ $= -0.46\text{ V}$
c. Predict whether the reaction will proceed spontaneously.	Since the cell potential is negative, the reaction is not spontaneous in the direction written.

Check Your Solution

The half-reactions are correct, and the reduction potentials are correct according to the table of standard reduction potentials on page 748 of the Chemistry 12 student textbook. The prediction that a negative cell potential will result in a non-spontaneous reaction is correct as well.

20. For the reaction $Br_2(\ell) + 2NaI(aq) \rightarrow 2NaBr(aq) + I_2(s)$:

a. write the oxidation and reduction half-reactions.

b. determine the standard cell potentials for galvanic cells in which these reactions occur.

c. predict whether the reaction will proceed spontaneously as written.

What Is Required?

a. You need to write the oxidation and reduction half reactions for a given redox reaction.

b. You need to determine the standard cell potential for the galvanic cell given.

c. You need to determine whether the reaction will proceed spontaneously as written.

What Is Given?

You are given the balanced net ionic equation.

You have a table of standard reduction potentials.

Plan Your Strategy	Act on Your Strategy
a. Break the redox reaction into the two half-reactions that make up the overall reaction.	Oxidation: $2I^-(aq) \rightarrow I_2(s) + 2e^-$ Reduction: $Br_2(\ell) + 2e^- \rightarrow 2Br^-(aq)$
b. Calculate the standard cell-potential for the galvanic cell by locating the relevant reduction potentials in a table of standard reduction potentials.	$I_2(s) + 2e^- \rightarrow 2I^-(aq)$ $E^o_{anode} = +0.54\text{ V}$ $Br_2(\ell) + 2e^- \rightarrow 2Br^-(aq)$ $E^o_{cathode} = +1.07\text{ V}$
Subtract the reduction potentials to determine the cell potential by using the formula $E^o_{cell} = E^o_{cathode} - E^o_{anode}$.	$E^o_{cell} = E^o_{cathode} - E^o_{anode}$ $= 1.07\text{ V} - (0.54\text{ V})$ $= +0.53\text{ V}$
c. Predict whether the reaction will proceed spontaneously.	Since the cell potential is positive, the reaction is spontaneous in the direction written.

Check Your Solution

The half-reactions are correct, and the reduction potentials are correct according to the table of standard reduction potentials on page 748 of the Chemistry 12 student textbook. The prediction that a positive cell potential will result in a spontaneous reaction is correct as well.