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.