

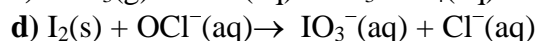
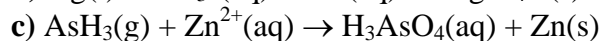
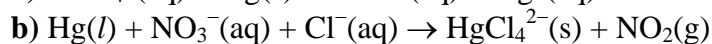
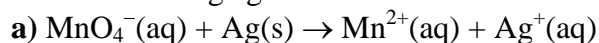
Chapter 12 Oxidation-Reduction Reactions

Solutions to Practice Problems

1.

Problem

Balance each of the following ionic equations for acidic conditions. Identify the oxidizing agent and the reducing agent in each case:



What Is Required?

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

What Is Given?

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

Plan Your Strategy

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

Step 1. Write unbalanced half-reactions that show the formulas of the given reactant(s) and product(s). Write only those compounds that contain the atom that is oxidized or the atom that is reduced.

Step 2. Balance any atoms other than oxygen and hydrogen first.

Step 3. Balance any oxygen atoms by adding water molecules.

Step 4. Balance any hydrogen atoms by adding hydrogen ions.

Step 5. Balance the charges by adding electrons.

Step 6. Determine the lowest common multiple (LCM) of the numbers of electrons in the oxidation and reduction half-reactions.

Step 7. Multiply one or both half-reactions by the number that will bring the number of electrons to the LCM and therefore an equal number of electrons will be lost and gained.

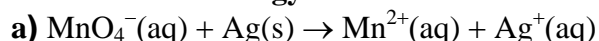
Step 8. Add the balanced half-reactions.

Step 9. Cancel the electrons and any other identical molecules or ions present on both sides of the equation.

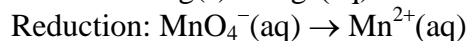
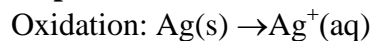
Step 10. If spectator ions were removed when forming half-reactions, add them back to the equation.

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

Act on Your Strategy

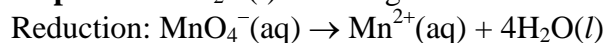


Step 1. The two unbalanced half-reactions are

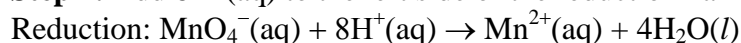


Step 2. Atoms other than oxygen and hydrogen are balanced.

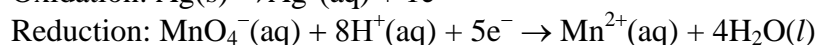
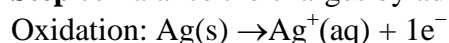
Step 3. Add $4\text{H}_2\text{O}(l)$ to the right side of the reduction half-reaction to balance the oxygen.



Step 4. Add $8\text{H}^+(aq)$ to the left side of the reduction half-reaction to balance the hydrogen.

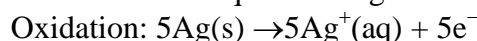


Step 5. Balance the charges by adding electrons.

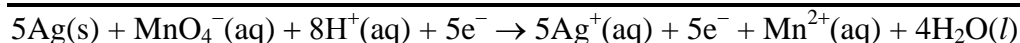
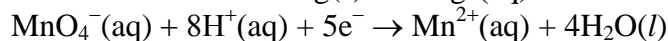


Step 6. The LCM of electrons is 5.

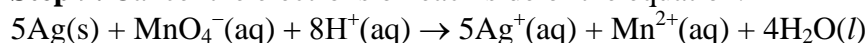
Step 7. Multiply the oxidation half-reaction by 5 so that the loss of electrons in the oxidation half-reaction is equal to the gain of electrons in the reduction half-reaction.



Step 8. Add the two half-reactions

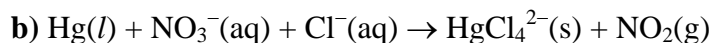


Step 9. Cancel the electrons on each side of the equation.

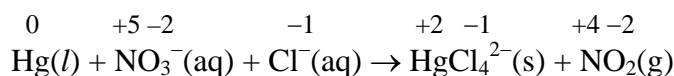


Since the MnO_4^- gains electrons, it is reduced and it is the oxidizing agent.

Since the Ag loses electrons, it is oxidized and it is the reducing agent.



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



The oxidation number of mercury increases (oxidation).

The oxidation number of nitrogen decreases (reduction).

Step 1. The two unbalanced half-reactions are



Reduction: $\text{NO}_3^-(\text{aq}) \rightarrow \text{NO}_2(\text{g})$

Step 2. Balance the chlorine in the oxidation half-reaction.

Oxidation: $\text{Hg}(\text{l}) + 4\text{Cl}^-(\text{aq}) \rightarrow \text{HgCl}_4^{2-}(\text{s})$

Step 3. Add $\text{H}_2\text{O}(\text{l})$ to the right side of the reduction half-reaction to balance the oxygen.

Reduction: $\text{NO}_3^-(\text{aq}) \rightarrow \text{NO}_2(\text{g}) + \text{H}_2\text{O}(\text{l})$

Step 4. Add $2\text{H}^+(\text{aq})$ to the left side of the reduction half-reaction to balance the hydrogen.

Reduction: $2\text{H}^+(\text{aq}) + \text{NO}_3^-(\text{aq}) \rightarrow \text{NO}_2(\text{g}) + \text{H}_2\text{O}(\text{l})$

Step 5. Balance the charges by adding electrons.

Oxidation: $\text{Hg}(\text{l}) + 4\text{Cl}^-(\text{aq}) \rightarrow \text{HgCl}_4^{2-}(\text{s}) + 2\text{e}^-$

Reduction: $2\text{H}^+(\text{aq}) + \text{NO}_3^-(\text{aq}) + 1\text{e}^- \rightarrow \text{NO}_2(\text{g}) + \text{H}_2\text{O}(\text{l})$

Step 6. The LCM of electrons is 2.

Step 7. Multiply the reduction half-reaction by 2 so that the loss of electrons in the oxidation half-reaction is equal to the gain of electrons in the reduction half-reaction.

Reduction: $4\text{H}^+(\text{aq}) + 2\text{NO}_3^-(\text{aq}) + 2\text{e}^- \rightarrow 2\text{NO}_2(\text{g}) + 2\text{H}_2\text{O}(\text{l})$

Step 8. Add the two half-reactions

$$\begin{array}{l} \text{Hg}(\text{l}) + 4\text{Cl}^-(\text{aq}) \rightarrow \text{HgCl}_4^{2-}(\text{s}) + 2\text{e}^- \\ 4\text{H}^+(\text{aq}) + 2\text{NO}_3^-(\text{aq}) + 2\text{e}^- \rightarrow 2\text{NO}_2(\text{g}) + 2\text{H}_2\text{O}(\text{l}) \end{array}$$

$\text{Hg}(\text{l}) + 4\text{Cl}^-(\text{aq}) + 4\text{H}^+(\text{aq}) + 2\text{NO}_3^-(\text{aq}) + 2\text{e}^- \rightarrow \text{HgCl}_4^{2-}(\text{s}) + 2\text{e}^- + 2\text{NO}_2(\text{g}) + 2\text{H}_2\text{O}(\text{l})$

Step 9. Cancel the electrons on each side of the equation.

$\text{Hg}(\text{l}) + 4\text{Cl}^-(\text{aq}) + 4\text{H}^+(\text{aq}) + 2\text{NO}_3^-(\text{aq}) \rightarrow \text{HgCl}_4^{2-}(\text{s}) + 2\text{NO}_2(\text{g}) + 2\text{H}_2\text{O}(\text{l})$

Since the NO_3^- is reduced it is the oxidizing agent.

Since the Hg is oxidized it is the reducing agent.

c) $\text{AsH}_3(\text{g}) + \text{Zn}^{2+}(\text{aq}) \rightarrow \text{H}_3\text{AsO}_4(\text{aq}) + \text{Zn}(\text{s})$

Step 1. The two unbalanced half-reactions are

Oxidation: $\text{AsH}_3(\text{g}) \rightarrow \text{H}_3\text{AsO}_4(\text{aq})$

Reduction: $\text{Zn}^{2+}(\text{aq}) \rightarrow \text{Zn}(\text{s})$

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

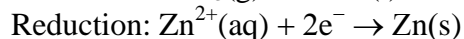
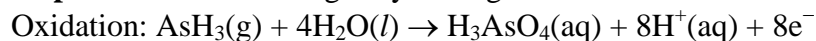
Step 3. Add $4\text{H}_2\text{O}(\text{l})$ to the left side of the oxidation half-reaction to balance the oxygen.

Oxidation: $\text{AsH}_3(\text{g}) + 4\text{H}_2\text{O}(\text{l}) \rightarrow \text{H}_3\text{AsO}_4(\text{aq})$

Step 4. Add $8\text{H}^+(\text{aq})$ to the right side of the oxidation half-reaction to balance the hydrogen.

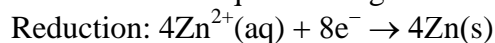
Oxidation: $\text{AsH}_3(\text{g}) + 4\text{H}_2\text{O}(\text{l}) \rightarrow \text{H}_3\text{AsO}_4(\text{aq}) + 8\text{H}^+(\text{aq})$

Step 5. Balance the charges by adding electrons.

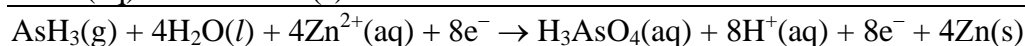
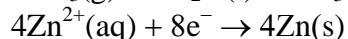


Step 6. The LCM of electrons is 8.

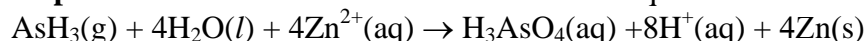
Step 7. Multiply the reduction half-reaction by 4 so that the loss of electrons in the oxidation half-reaction is equal to the gain of electrons in the reduction half-reaction.



Step 8. Add the two half-reactions

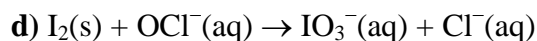


Step 9. Cancel the electrons on each side of the equation.

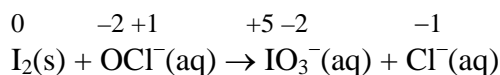


Since the $\text{Zn}^{2+}(\text{aq})$ gains electrons, it is reduced and it is the oxidizing agent.

Since the $\text{AsH}_3(\text{g})$ loses electrons, it is oxidized and it is the reducing agent.



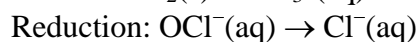
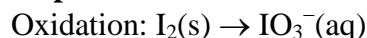
Assign an oxidation number to each element to identify the oxidation and reduction half-reactions.



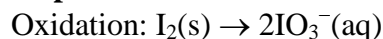
The oxidation number of iodine increases (oxidation).

The oxidation number of chlorine decreases (reduction).

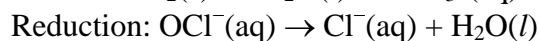
Step 1. The two unbalanced half-reactions are



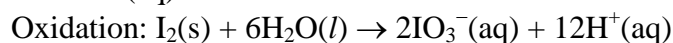
Step 2. Balance the iodine in the oxidation half-reaction.



Step 3. Add $6\text{H}_2\text{O}(\text{l})$ to the left side of the oxidation half-reaction to balance the oxygen. Add $\text{H}_2\text{O}(\text{l})$ to the right side of the reduction half-reaction to balance the oxygen.

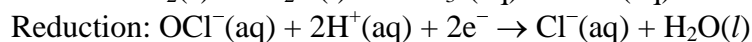
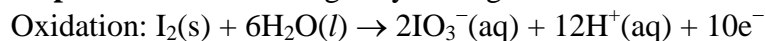


Step 4. Add $12\text{H}^+(\text{aq})$ to the right side of the reduction half-reaction to balance the hydrogen. Add $2\text{H}^+(\text{aq})$ to the left side of the reduction half-reaction to balance the hydrogen.



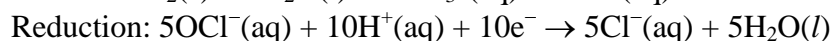
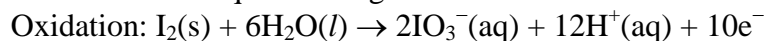


Step 5. Balance the charges by adding electrons.

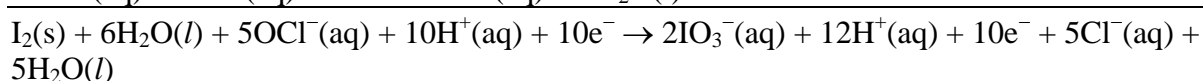
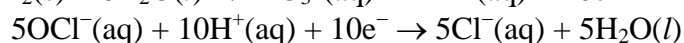
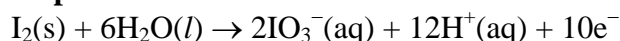


Step 6. The LCM of electrons is 10.

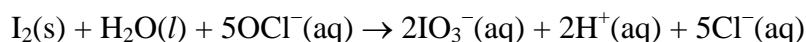
Step 7. Multiply the reduction half-reaction by 5 so that the loss of electrons in the oxidation half-reaction is equal to the gain of electrons in the reduction half-reaction.



Step 8. Add the two half-reactions.



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

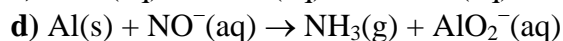
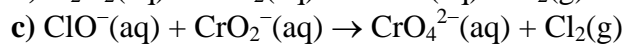
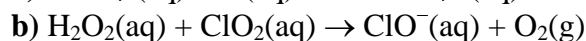
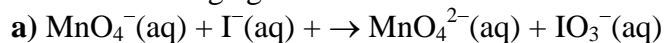


Since $\text{I}_2(\text{s})$ is oxidized, it is a reducing agent. Since $\text{OCl}^-(\text{aq})$ is reduced, it is an oxidizing agent.

2.

Problem

Balance each of the following ionic equations for basic conditions. Identify the oxidizing agent and the reducing agent in each case:



What Is Required?

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

What Is Given?

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

Plan Your Strategy

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

Step 1. Write unbalanced half-reactions that show the formulas of the given reactant(s) and product(s). Write only those compounds that contain the atom that is oxidized or the atom that is reduced.

Step 2. Balance any atoms other than oxygen and hydrogen first.

Step 3. Balance any oxygen atoms by adding water molecules.

Step 4. Balance any hydrogen atoms by adding hydrogen ions.

Step 5. Adjust for basic conditions by adding to both sides the same number of hydroxide ions as the number of hydrogen ions already present.

Step 6. Simplify the equation by combining hydrogen ions and hydroxide ions on the same side of the equation into water molecules.

Step 7. Cancel any water molecules present on both sides of the equation.

Step 8 Balance the charges by adding electrons.

Step 9. Determine the least common multiple of the numbers of electrons in the oxidation and reduction half-reactions.

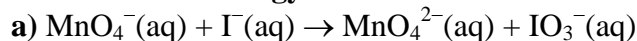
Step 10. Multiply one or both half-reactions by the number that will bring the number of electrons to the least common multiple and therefore an equal number of electrons are lost and gained.

Step 11. Add the balanced half-reactions.

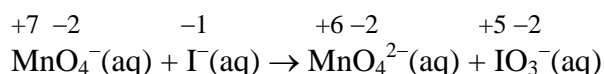
Step 12. Cancel the electrons on both sides of the equation and simplify by combining $\text{H}_2\text{O}(l)$ and $\text{OH}^-(aq)$ on each side of the equation.

Step 13. If spectator ions were removed when forming half-reactions, add them back to the equation.

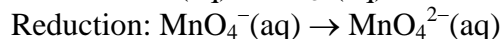
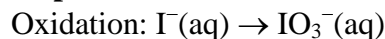
Act on Your Strategy



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

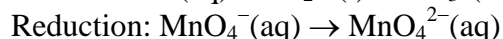
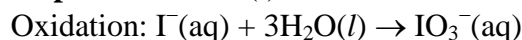


Step 1. The two unbalanced half-reactions are

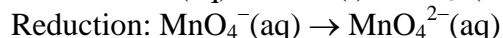
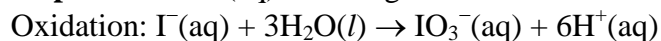


Step 2. Atoms other than oxygen and hydrogen are balanced.

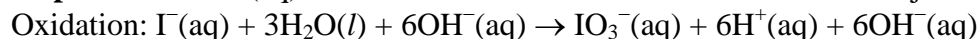
Step 3. Add $3\text{H}_2\text{O}(l)$ to the left side of the oxidation half-reaction to balance the oxygen.



Step 4. Add $6\text{H}^+(aq)$ to the right side of the oxidation half-reaction to balance the hydrogen.



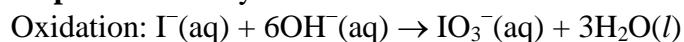
Step 5. Add $6\text{OH}^-(\text{aq})$ to each side of the oxidation half-reaction to adjust for basic conditions.



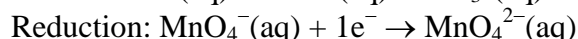
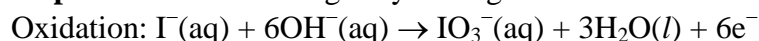
Step 6. Combine $\text{H}^+(\text{aq})$ and $\text{OH}^-(\text{aq})$ on the same side of each equation into water molecules.



Step 7. Cancel any water molecules common to both sides of the equation.

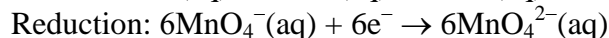
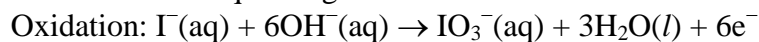


Step 8. Balance the charges by adding electrons.

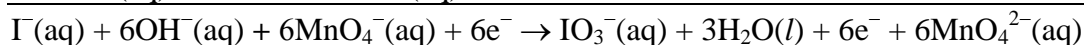
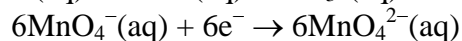


Step 9. The LCM of electrons is 6.

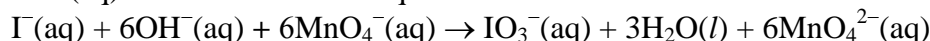
Step 10. Multiply the reduction half-reaction by 6 so that the loss of electrons in the oxidation half-reaction is equal to gain of electrons in the reduction half-reaction.



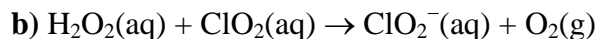
Step 11. Add the two half-reactions.



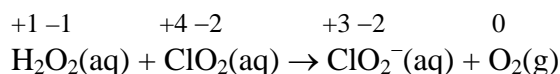
Step 12. Cancel the electrons on each side of the equation and simplify by combining $\text{H}_2\text{O}(\text{l})$ and $\text{OH}^-(\text{aq})$ on each side of the equation.



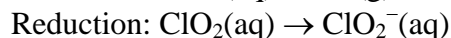
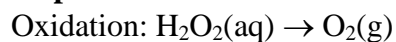
Since the $\text{MnO}_4^-(\text{aq})$ is reduced, it is the oxidizing agent. Since the $\text{I}^-(\text{aq})$ is oxidized it is the reducing agent.



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



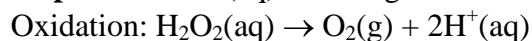
Step 1. The two unbalanced half-reactions are



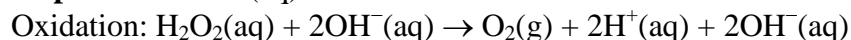
Step 2. Atoms other than oxygen and hydrogen are balanced.

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

Step 4. Add $2\text{H}^+(\text{aq})$ to the right side of the oxidation half-reaction to balance the hydrogen.



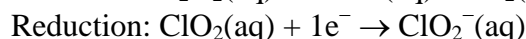
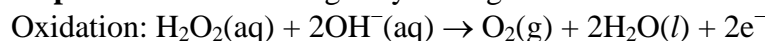
Step 5. Add $2\text{OH}^-(\text{aq})$ to each side of the oxidation half-reaction to adjust for basic conditions.



Step 6. Combine $\text{H}^+(\text{aq})$ and $\text{OH}^-(\text{aq})$ on the same side of each equation into water molecules.

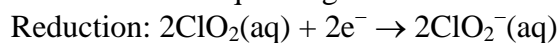


Step 7. Balance the charges by adding electrons.

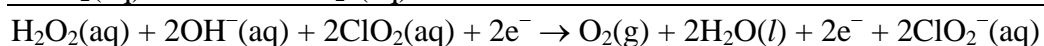
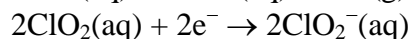
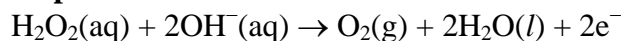


Step 8. The LCM of electrons is 2.

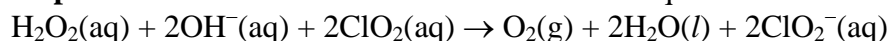
Step 9. Multiply the reduction half-reaction by 2 so that the loss of electrons in the oxidation half-reaction is equal to gain of electrons in the reduction half-reaction.



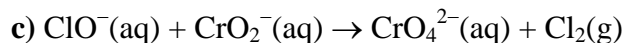
Step 10. Add the two half-reactions.



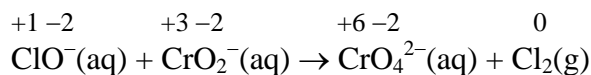
Step 11. Cancel the electrons on each side of the equation.



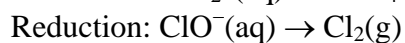
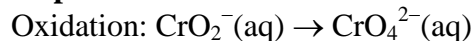
Since the $\text{ClO}_2(\text{aq})$ is reduced, it is the oxidizing agent. Since the $\text{H}_2\text{O}_2(\text{aq})$ is oxidized it is the reducing agent.



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



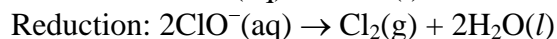
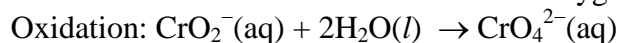
Step 1. The two unbalanced half-reactions are



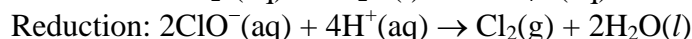
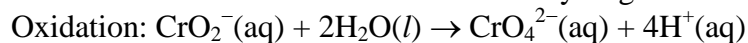
Step 2. Balance the chlorine in the reduction half-reaction.



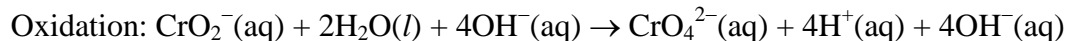
Step 3. Add $2\text{H}_2\text{O}(\text{l})$ to the left side of the oxidation half-reaction and $2\text{H}_2\text{O}(\text{l})$ to the right side of the reduction half-reaction to balance the oxygen.



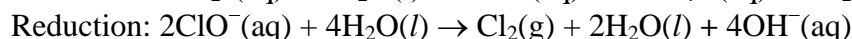
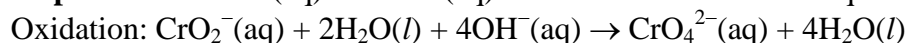
Step 4. Add $4\text{H}^+(\text{aq})$ to the right side of the oxidation half-reaction and $4\text{H}^+(\text{aq})$ to the left side of the reduction half-reaction to balance the hydrogen.



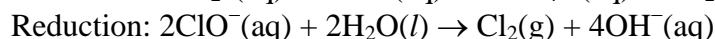
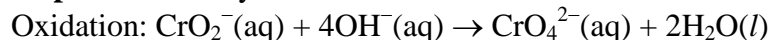
Step 5. Add $4\text{OH}^-(\text{aq})$ to each side of the oxidation and reduction half-reactions to adjust for basic conditions.



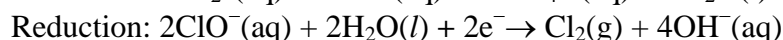
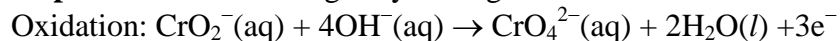
Step 6. Combine $\text{H}^+(\text{aq})$ and $\text{OH}^-(\text{aq})$ on the same side of each equation into water molecules.



Step 7. Cancel any water molecules common to both sides of the equation.

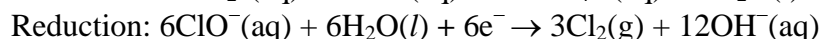
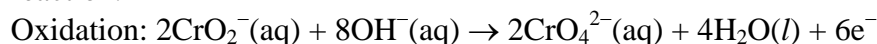


Step 8. Balance the charges by adding electrons.

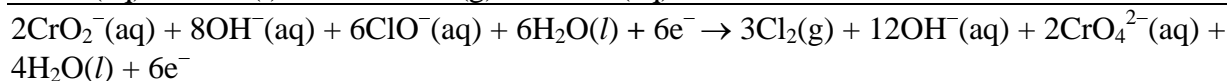
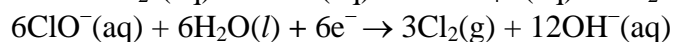
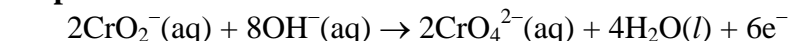


Step 9. The LCM of electrons is 6.

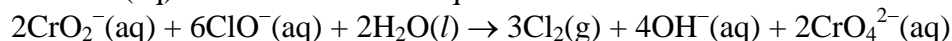
Step 10. Multiply the oxidation half-reaction by 2 and the reduction half-reaction by 3 so that the loss of electrons in the oxidation half-reaction is equal to gain of electrons in the reduction half-reaction.



Step 11. Add the two half-reactions

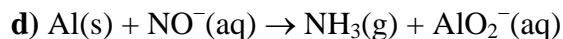


Step 12. Cancel the electrons on each side of the equation and simplify by combining $\text{H}_2\text{O}(l)$ and $\text{OH}^-(aq)$ on each side of the equation.

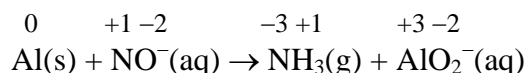


Since the $\text{ClO}^-(aq)$ is reduced, it is the oxidizing agent.

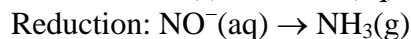
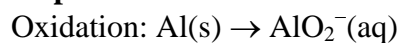
Since the $\text{CrO}_2^-(aq)$ is oxidized it is the reducing agent.



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

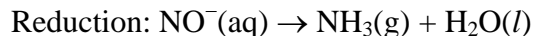
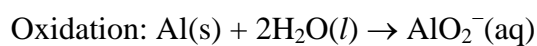


Step 1. The two unbalanced half-reactions are

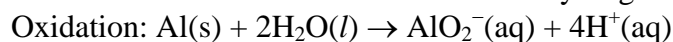


Step 2. Atoms other than oxygen and hydrogen are balanced.

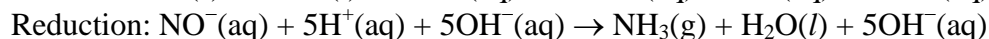
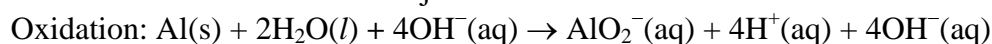
Step 3. Add $2\text{H}_2\text{O}(l)$ to the left side of the oxidation half-reaction and $\text{H}_2\text{O}(l)$ to the right side of the reduction half-reaction to balance the oxygen.



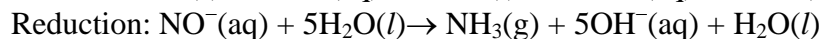
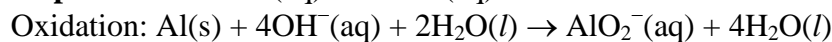
Step 4. Add $4\text{H}^+(aq)$ to the right side of the oxidation half-reaction and $5\text{H}^+(aq)$ to the left side of the reduction half-reaction to balance the hydrogen.



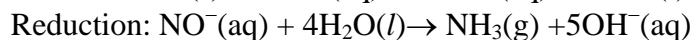
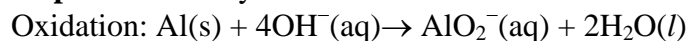
Step 5. Add $4\text{OH}^-(aq)$ to each side of the oxidation half-reaction and $5\text{OH}^-(aq)$ to each side of the reduction half-reaction to adjust for basic conditions.



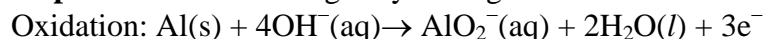
Step 6. Combine $\text{H}^+(aq)$ and $\text{OH}^-(aq)$ on the same side of each equation into water molecules.

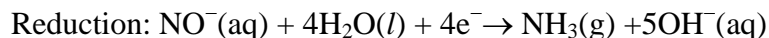


Step 7. Cancel any water molecules common to both sides of the equation.



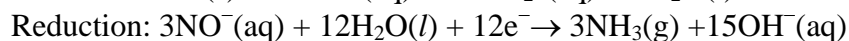
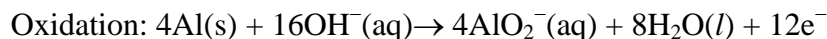
Step 8. Balance the charges by adding electrons.



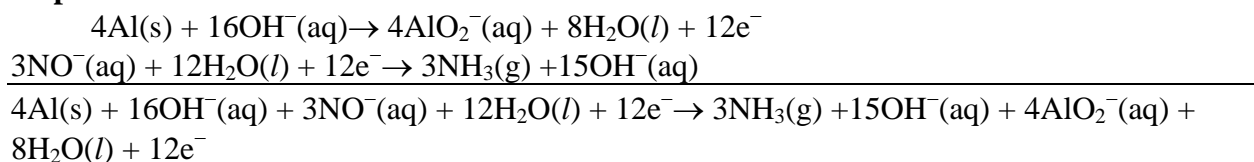


Step 9. The LCM of electrons is 12.

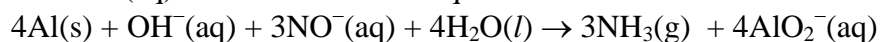
Step 10. Multiply the oxidation half-reaction by 4 and the reduction half-reaction by 3 so that the loss of electrons in the oxidation half-reaction is equal to gain of electrons in the reduction half-reaction.



Step 11. Add the two half-reactions



Step 12. Cancel the electrons on each side of the equation and simplify by combining $\text{H}_2\text{O}(\text{l})$ and $\text{OH}^-(\text{aq})$ on each side of the equation.



Since the $\text{NO}^-(\text{aq})$ is reduced, it is the oxidizing agent.

Since the $\text{Al}(\text{s})$ is oxidized it is the reducing agent.

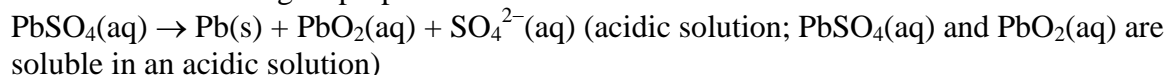
Check Your Solutions

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

3.

Problem

Balance the following disproportionation reaction:



What is Required?

You must balance a disproportionation reaction that occurs in acidic conditions.

What is Given?

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

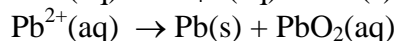
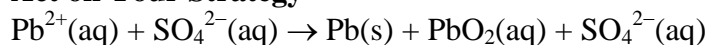
Plan Your Strategy

Write the equation as a net ionic equation.

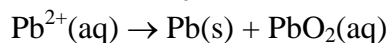
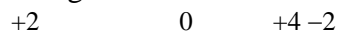
Cancel spectator ions common to both sides of the reaction.

Use the rules for balancing a redox reaction in acid conditions.

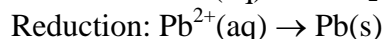
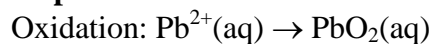
Act on Your Strategy



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

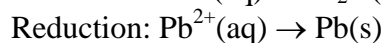
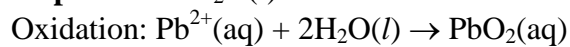


Step 1. The two unbalanced half-reactions are

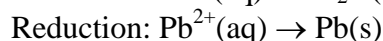


Step 2. Atoms other than oxygen and hydrogen are balanced

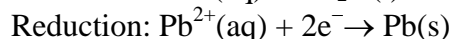
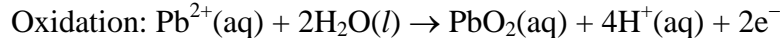
Step 3. Add $2\text{H}_2\text{O}(\text{l})$ to the left side of the oxidation half-reaction to balance the oxygen.



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



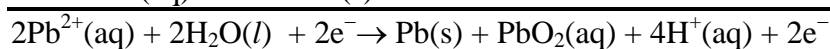
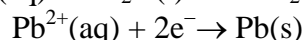
Step 5. Balance the charges by adding electrons.



Step 6. The LCM of electrons is 2.

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

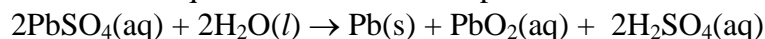
Step 8. Add the two half-reactions



Step 9. Cancel the electrons on each side of the equation.



Rewrite the equation to include the spectator ions.



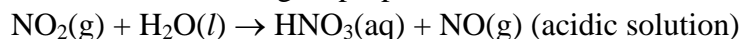
Check Your Solution

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

4.

Problem

Balance the following disproportionation reaction:



What is Required?

You must balance a disproportionation reaction that occurs in acidic conditions.

What is Given?

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

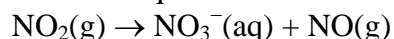
Plan Your Strategy

Write the equation in the form of a net ionic reaction.

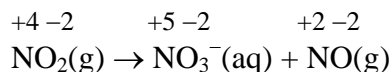
Use the rules for balancing a redox reaction in acid conditions.

Act on Your Strategy

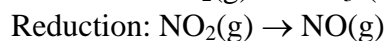
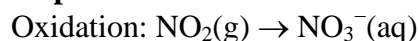
The ionic equation for this reaction is



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

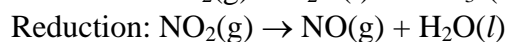
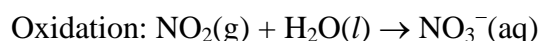


Step 1. The two unbalanced half-reactions are

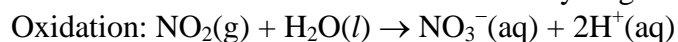


Step 2. Atoms other than oxygen and hydrogen are balanced

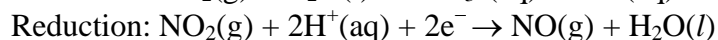
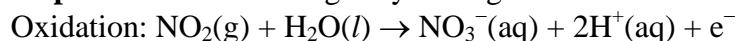
Step 3. Add $\text{H}_2\text{O}(\text{l})$ to the left side of the oxidation half-reaction and $\text{H}_2\text{O}(\text{l})$ to the right side of the reduction half-reaction to balance the oxygen.



Step 4. Add $2\text{H}^+(\text{aq})$ to the right side of the oxidation half-reaction and $2\text{H}^+(\text{aq})$ to the left side of the reduction half-reaction to balance the hydrogen.

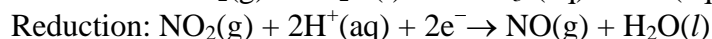
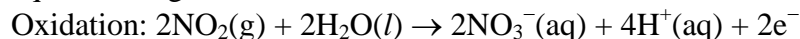


Step 5. Balance the charges by adding electrons.

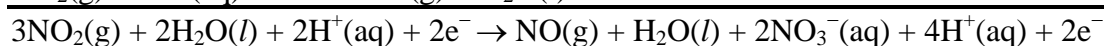
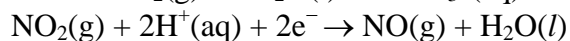
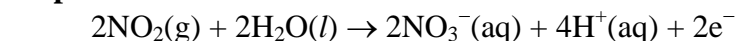


Step 6. The LCM of electrons is 2.

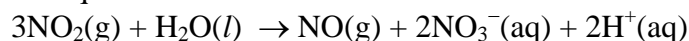
Step 7. Multiply the oxidation half-reaction by 2 so that the loss of electrons in the oxidation is equal to the gain of electrons in the reduction.



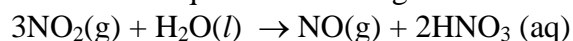
Step 8. Add the two half-reactions



Step 9. Cancel the electrons on each side of the equation and combine $\text{H}_2\text{O}(\text{l})$ from each side of the equation.



Rewrite the equation in its original non-ionic form.



Check Your Solution

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

5.

Problem

Balance the following disproportionation reaction:



What is Required?

You must balance a disproportionation reaction that occurs in basic conditions.

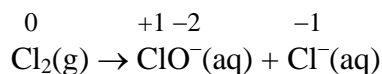
What is Given?

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

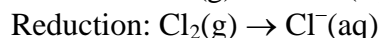
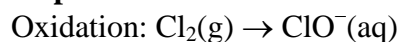
Plan Your Strategy

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

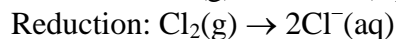
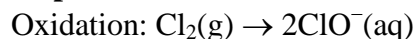
Assign an oxidation number to each element to determine the oxidation and reduction half-reactions.



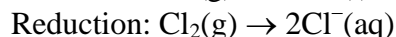
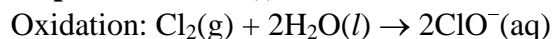
Step 1. The two unbalanced half-reactions are



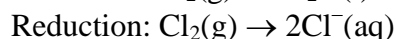
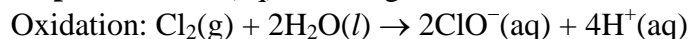
Step 2. Balance the chlorine atoms in each half-reaction.



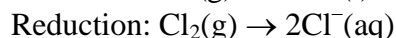
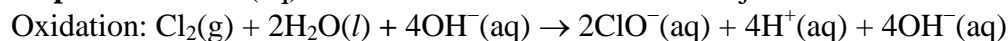
Step 3. Add $2\text{H}_2\text{O}(\text{l})$ to the left side of the oxidation half-reaction to balance the oxygen.



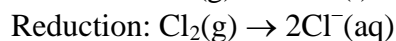
Step 4. Add $4\text{H}^+(\text{aq})$ to the right side of the oxidation half-reaction to balance the hydrogen.



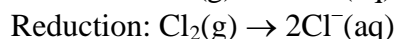
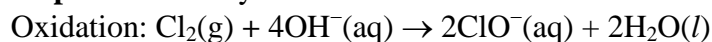
Step 5. Add $4\text{OH}^-(\text{aq})$ to each side of the oxidation to adjust for basic conditions.



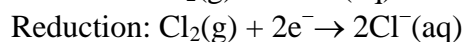
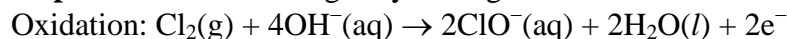
Step 6. Combine $\text{H}^+(\text{aq})$ and $\text{OH}^-(\text{aq})$ on the same side of each equation into water molecules.



Step 7. Cancel any water molecules common to both sides of the equation.



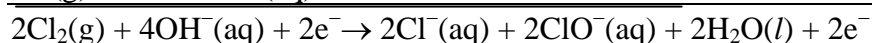
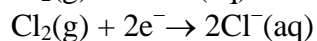
Step 8. Balance the charges by adding electrons.



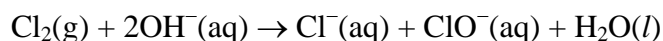
Step 9. The LCM of electrons is 2.

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

Step 11. Add the two half-reactions



Step 12. Cancel the electrons on each side of the equation and reduce the coefficients to the LCM.



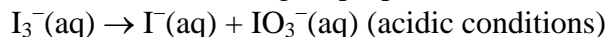
Check Your Solution

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

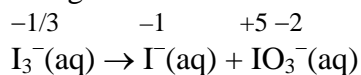
6.

Problem

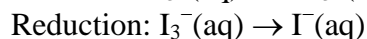
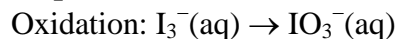
Balance the following disproportionation reaction:



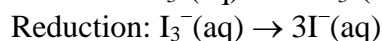
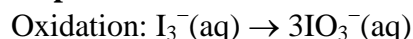
Assign an oxidation number to each element to identify oxidation and reduction half-reactions.



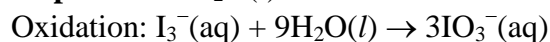
Step 1. The two unbalanced half-reactions are



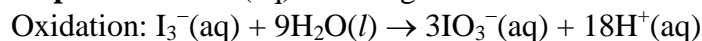
Step 2. Balance the iodine in the oxidation and reduction half-reactions.



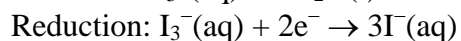
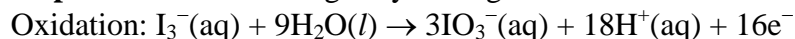
Step 3. Add $9\text{H}_2\text{O}(l)$ to the left side of the oxidation half-reaction to balance the oxygen.



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

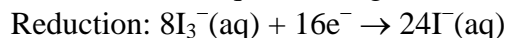


Step 5. Balance the charges by adding electrons.

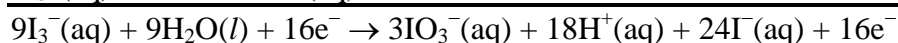
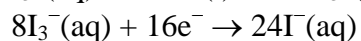
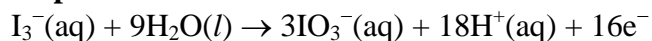


Step 6. The LCM of electrons is 16.

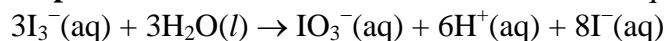
Step 7. Multiply the reduction half-reaction by 8 so that the loss of electrons in the oxidation half-reaction is equal to the gain of electrons in the reduction half-reaction.



Step 8. Add the two half-reactions



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



Check Your Solution

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

7.

Problem

Determine the oxidation number of the atoms of the specified element in each of the following:

- a) N in $\text{NF}_3(\text{g})$
- b) S in $\text{S}_8(\text{s})$
- c) Cr in $\text{CrO}_4^{2-}(\text{aq})$
- d) P in $\text{P}_2\text{O}_5(\text{s})$
- e) C in $\text{C}_{12}\text{H}_{22}\text{O}_{11}(\text{s})$
- f) C in $\text{CHCl}_3(\text{g})$

What is Required?

You must determine the oxidation number of the specified element.

What is Given?

The formula of the compound is known.

Plan Your Strategy

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

Act on Your Strategy

- a) N in $\text{NF}_3(\text{g})$

The sum of the O.N. = 0

$$(1 \text{ atom N} \times \text{O.N. of N}) + (3 \text{ atoms F} \times \text{O.N. of F}) = 0$$

$$(1 \text{ atom N} \times \text{O.N. of N}) + (3 \text{ atoms F} \times -1) = 0$$

$$\text{O.N. of N} = +3$$

- b) S in $\text{S}_8(\text{s})$ $S = 0$ since this is the elemental form of sulfur

- c) Cr in $\text{CrO}_4^{2-}(\text{aq})$

The sum of the O.N. = charge on the ion = 2-

$$(1 \text{ atom Cr} \times \text{O.N. of Cr}) + (4 \text{ atoms of O} \times \text{O.N. of O}) = 2-$$

$$(1 \text{ atom Cr} \times \text{O.N. of Cr}) + (4 \text{ atoms of O} \times -2) = 2-$$

$$\text{O.N. of Cr} = +6$$

- d) P in $\text{P}_2\text{O}_5(\text{s})$

The sum of the O.N. = 0

$$(2 \text{ atom P} \times \text{O.N. of P}) + (5 \text{ atoms of O} \times \text{O.N. of O}) = 0$$

$$(2 \text{ atom P} \times \text{O.N. of P}) + (5 \text{ atoms of O} \times -2) = 0$$

$$\text{O.N. of P} = +5$$

- e) C in $\text{C}_{12}\text{H}_{22}\text{O}_{11}(\text{s})$

The sum of the O.N. = 0

$$(12 \text{ atom C} \times \text{O.N. of C}) + (22 \text{ atoms of H} \times \text{O.N. of H}) + (11 \text{ atoms of O} \times \text{O.N. of O}) = 0$$

$$(12 \text{ atom C} \times \text{O.N. of C}) + (22 \text{ atoms of H} \times +1) + (11 \text{ atoms of O} \times -2) = 0$$

$$\text{O.N. of C} = 0$$

f) C in $\text{CHCl}_3(\text{g})$

The sum of the O.N. = 0

$$(1 \text{ atom C} \times \text{O.N. of C}) + (1 \text{ atom of H} \times \text{O.N. of H}) + (3 \text{ atoms of Cl} \times \text{O.N. of Cl}) = 0$$

$$(1 \text{ atom C} \times \text{O.N. of C}) + (1 \text{ atom of H} \times +1) + (3 \text{ atoms of Cl} \times -1) = 0$$

$$\text{O.N. of C} = +2$$

Check Your Solution

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

8.

Problem

Determine the oxidation number of each of the atoms in each of the following compounds:



What is Required?

You must determine the oxidation number of each element in the given compounds.

What is Given?

The formula of the compound is known.

Plan Your Strategy

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

Act on Your Strategy



The sum of the O.N. = 0

From the rules for assigning O.N., H = +1 and O = -2

$$(2 \text{ atoms H} \times \text{O.N. of H}) + (1 \text{ atom of S} \times \text{O.N. of S}) + (3 \text{ atoms of O} \times \text{O.N. of O}) = 0$$

$$(2 \text{ atoms H} \times +1) + (1 \text{ atom of S} \times \text{O.N. S}) + (3 \text{ atoms of O} \times -2) = 0$$

$$\text{O.N. of S} = +4$$



The sum of the O.N. = net charge on $\text{OH}^-(\text{aq}) = 1-$

From the rules for assigning O.N., H = +1 and O = -2



The sum of the O.N. = net charge on the $\text{HPO}_4^{2-}(\text{aq}) = 2-$

From the rules for assigning O.N. H = +1 and O = -2

$$(1 \text{ atom H} \times \text{O.N. of H}) + (1 \text{ atom of P} \times \text{O.N. of P}) + (4 \text{ atoms of O} \times \text{O.N. of O}) = 2 -$$

$$(1 \text{ atom H} \times +1) + (1 \text{ atom of P} \times \text{O.N. of P}) + (4 \text{ atoms of O} \times -2) = 2 -$$

$$\text{O.N. for P} = +5$$

Check Your Solution

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

9.

Problem

As stated in rule 4, oxygen does not always have its usual oxidation number of -2 . Determine the oxidation number of oxygen in each of the following:

- a) the compound oxygen difluoride, $\text{OF}_2(\text{g})$
- b) the peroxide ion, $\text{O}_2^{2-}(\text{aq})$

What is Required?

You must determine the oxidation number of oxygen in each compound.

What is Given?

The formula of the compound is known.

Plan Your Strategy

Follow the rules for assigning oxidation number (O.N.) except the O.N. for oxygen $\neq -2$.

Act on Your Strategy

- a) $\text{OF}_2(\text{g})$

The sum of the O.N. = 0

$$(1 \text{ atom O} \times \text{O.N. of O}) + (2 \text{ atoms F} \times \text{O.N. F}) = 0$$

$$(1 \text{ atom O} \times \text{O.N. of O}) + (2 \text{ atoms F} \times -1) = 0$$

$$\text{O.N. of O} = +2$$

- b) $\text{O}_2^{2-}(\text{aq})$

The sum of the O.N. = charge on $\text{O}_2^{2-}(\text{aq}) = 2-$

$$2 \text{ atoms O} \times \text{O.N. of O} = 2-$$

$$\text{O.N. of O} = -1$$

Check Your Solution

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

10.

Problem

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

- a) $\text{Al}(\text{HCO}_3)_3(\text{s})$
- b) $(\text{NH}_4)_3\text{PO}_4(\text{aq})$
- c) $\text{K}_2\text{H}_3\text{IO}_6(\text{aq})$

What is Required?

You must determine the oxidation number of each element in the given compounds.

What is Given?

The formula of the compound is known.

Plan Your Strategy

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

Act on Your Strategy



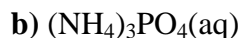
The sum of the O.N. = 0

From the rules for assigning O.N., H = +1 and O = -2, Al = +3

$$(1 \text{ atom Al} \times \text{O.N. of Al}) + (3 \text{ atoms of H} \times \text{O.N. of H}) + (3 \text{ atoms of C} \times \text{O.N. of C}) + (9 \text{ atoms O} \times \text{O.N. of O}) = 0$$

$$1 \text{ atom Al} \times \text{O.N. of +3} + (3 \text{ atoms of H} \times \text{O.N. of +1}) + (3 \text{ atoms of C} \times \text{O.N. of C}) + (9 \text{ atoms O} \times \text{O.N. of -2}) = 0$$

$$\text{O.N. of C} = +4$$



The sum of the O.N. = 0

From the rules for assigning O.N., H = +1 and O = -2

The sum of the O.N. for NH_4^+ = charge on NH_4^+ = 1+

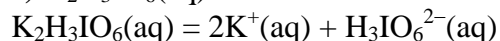
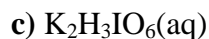
$$(1 \text{ atom N} \times \text{O.N. N}) + 4 \text{ atom H} \times \text{O.N. of H} = 1+$$

$$\text{O.N. of N} = -3$$

The sum of the O.N. for PO_4^{3-} = charge on PO_4^{3-} = 3-

$$(1 \text{ atom P} \times \text{O.N. of P}) + (4 \text{ atoms O} \times \text{O.N. of O}) = 3-$$

$$\text{O.N. of P} = +5$$



From the rules for assigning oxidation numbers, O.N., H = +1, O = -2, K = +1

The sum of the O.N. for $\text{H}_3\text{IO}_6^{2-}(\text{aq})$ = net charge on $\text{H}_3\text{IO}_6^{2-}(\text{aq})$ = 2-

$$(3 \text{ atoms of H} \times \text{O.N. of H}) + (1 \text{ atom of I} \times \text{O.N. of I}) + (6 \text{ atoms O} \times \text{O.N. of O}) = 2-$$

$$(3 \text{ atoms of H} \times +1) + (1 \text{ atom of I} \times \text{O.N. of I}) + (6 \text{ atoms O} \times \text{O.N. of -2}) = 2-$$

$$\text{O.N. of I} = +7$$

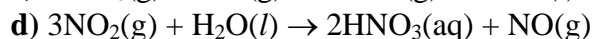
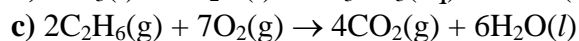
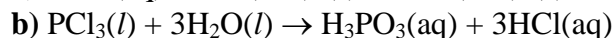
Check Your Solution

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

11.

Problem

Which of the following are redox reactions? Identify any disproportionation reactions:



What is Required?

You must determine if the reactions are oxidation-reduction and identify any disproportionation reactions.

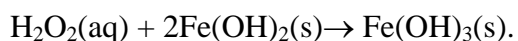
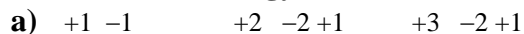
What is Given?

Balanced equations showing the formulas for the reactants and products are given.

Plan Your Strategy

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

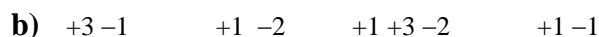
Act on Your Strategy



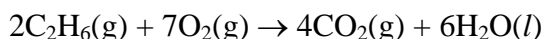
O.N. of Fe increases from +2 in $\text{Fe}(\text{OH})_2(\text{s})$ to +3 in $\text{Fe}(\text{OH})_3(\text{s})$ (oxidation)

O.N. of O decreases from -1 in $\text{H}_2\text{O}_2(\text{aq})$ to -2 in $\text{Fe}(\text{OH})_3(\text{s})$ (reduction)

This is a redox reaction.



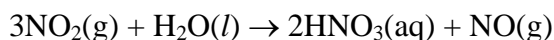
There is no change in the O.N. of any element. This is not a redox reaction.



The O.N. of C increases from -3 in $\text{C}_2\text{H}_6(\text{g})$ to +4 in $\text{CO}_2(\text{g})$ (oxidation)

The O.N. of oxygen in decreases from 0 in $\text{O}_2(\text{g})$ to -2 in $\text{CO}_2(\text{g})$ (reduction)

This is a redox reaction.



The O.N of N increases from +4 in $\text{NO}_2(\text{g})$ to +5 in $\text{HNO}_3(\text{aq})$ (oxidation)

The O.N of N decreases from +4 in $\text{NO}_2(\text{g})$ to +2 in $\text{NO}(\text{g})$ (reduction)

This is a redox reaction. Since N in $\text{NO}_2(\text{g})$ undergoes an increase and a decrease in O.N., this is a disproportionation reaction.

Check Your Solution

This answer corresponds to the definitions of oxidation and reduction.

12.

Problem

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

What is Required?

You must identify the oxidizing agent and the reducing agent in each reaction from question 10.

What is Given?

The answer from question 10 identified the species undergoing oxidation and reduction.

Plan Your Strategy

The reducing agent undergoes oxidation and the oxidizing agent undergoes reduction.

Act on Your Strategy

a) $\text{H}_2\text{O}_2(\text{aq})$ undergoes reduction. It is an oxidizing agent.

$\text{Fe}(\text{OH})_2(\text{s})$ undergoes oxidation. It is a reducing agent.

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

c) $\text{O}_2(\text{g})$ undergoes reduction. It is the oxidizing agent.

$\text{C}_2\text{H}_6(\text{g})$ undergoes oxidation. It is the reducing agent

d) $\text{NO}_2(\text{g})$ undergoes both reduction and oxidation. It is both an oxidizing agent and a reducing agent.

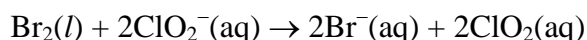
Check Your Solution

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

13.

Problem

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



What is Required?

You must identify the reactant that undergoes oxidation and the reactant that undergoes reduction.

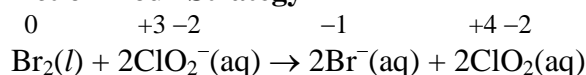
What is Given?

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

Plan Your Strategy

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

Act on Your Strategy



The O.N. of Br decreases from 0 in $\text{Br}_2(l)$ to -1 in $2\text{Br}^-(\text{aq})$. This is reduction. $\text{Br}_2(l)$ is an oxidizing agent.

The O.N. of Cl increases from +3 in $\text{ClO}_2^-(\text{aq})$ to +4 in $\text{ClO}_2(\text{aq})$. This is oxidation. $\text{ClO}_2^-(\text{aq})$ is a reducing agent.

Check Your Solution

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

14.

Problem

Nickel and copper ores usually contain the metals as sulfides, such as $\text{NiS}(\text{s})$ and $\text{Cu}_2\text{S}(\text{s})$. Does the extraction of these pure elemental metals from their ores involve redox reactions? Explain your reasoning.

What is Required?

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

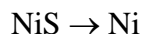
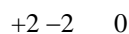
What is Given?

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

Plan Your Strategy

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

Act on Your Strategy



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

The same analysis can be used with CuS and Cu .

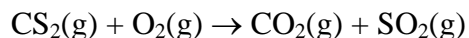
Check Your Solution

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

15.

Problem

Use the oxidation number method to balance the following equation for the combustion of carbon disulfide:



What is Required?

You must balance the redox reaction using oxidation numbers.

What is Given?

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

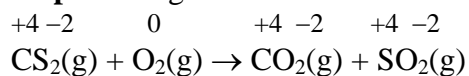
Plan Your Strategy

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

Act on Your Strategy

Step 1. The unbalanced equation is $\text{CS}_2(\text{g}) + \text{O}_2(\text{g}) \rightarrow \text{CO}_2(\text{g}) + \text{SO}_2(\text{g})$

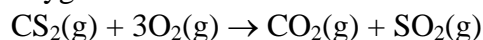
Step 2. Assign an O.N. to each element.



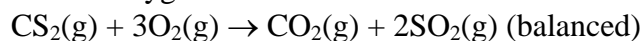
Step 3. The O.N. of sulfur undergoes an increase of 6 from -2 in $\text{CS}_2(\text{g})$ to $+4$ in $\text{SO}_2(\text{g})$. The O.N. of oxygen undergoes a decrease of 2 from 0 in $\text{O}_2(\text{g})$ to -2 in $\text{SO}_2(\text{g})$.

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

Step 5. Use the 3:1 ratio in **Step 4** to balance the oxygen and sulfur atoms. There must be 3 oxygen atoms for each sulfur atom.



Step 6. Complete the balancing by inspection. Balance the sulfur on the right side of the equation and the oxygen on the left side.



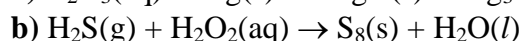
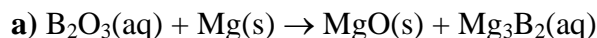
Check Your Solution

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

16.

Problem

Use the oxidation number method to balance the following equations:



What is Required?

You must balance the redox reactions using oxidation numbers.

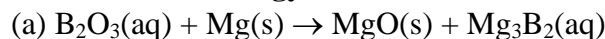
What is Given?

The unbalanced equations are given with the formula of the reactants and products.

Plan Your Strategy

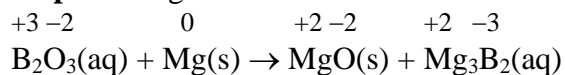
Assign an oxidation number (O.N.) to each element. Follow the rules for balancing an equation using O.N.

Act on Your Strategy



Step 1. The unbalanced equation is $\text{B}_2\text{O}_3(\text{aq}) + \text{Mg}(\text{s}) \rightarrow \text{MgO}(\text{s}) + \text{Mg}_3\text{B}_2(\text{aq})$

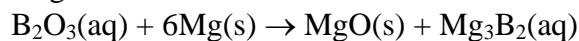
Step 2. Assign an O.N. to each element.



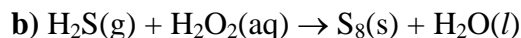
Step 3. The O.N. of magnesium undergoes an increase of 2 from 0 in $\text{Mg}(\text{s})$ to +2 in $\text{Mg}_3\text{B}_2(\text{aq})$. The O.N. of boron undergoes a decrease of 6 from +3 in $\text{B}_2\text{O}_3(\text{aq})$ to -3 in $\text{Mg}_3\text{B}_2(\text{aq})$.

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

Step 5. Use the 3:1 ratio in **Step 4** to balance the magnesium and boron atoms. There must be 3 magnesium atoms for each boron atom.

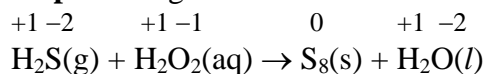


Step 6. Complete the balancing of the oxygen by inspection.



Step 1. The unbalanced equation is $\text{H}_2\text{S}(\text{g}) + \text{H}_2\text{O}_2(\text{aq}) \rightarrow \text{S}_8(\text{s}) + \text{H}_2\text{O}(\text{l})$

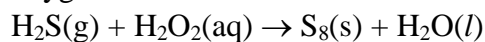
Step 2. Assign an O.N. to each element.



Step 3. The O.N. of sulfur undergoes an increase of 2 from -2 in $\text{H}_2\text{S}(\text{g})$ to 0 in $\text{S}_8(\text{s})$. The O.N. of oxygen undergoes a decrease of 1 from -1 in $\text{H}_2\text{O}_2(\text{aq})$ to -2 in $\text{H}_2\text{O}(\text{l})$.

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

Step 5. Use the 2:1 ratio in **Step 4** to balance the oxygen and sulfur atoms. There must be 2 oxygen atoms for each sulfur atom.



Step 6. Complete the balancing of hydrogen, sulfur and the oxygen by inspection.



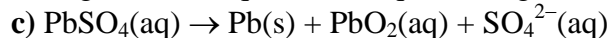
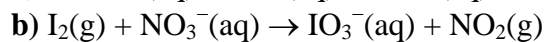
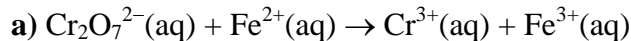
Check Your Solution

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

17.

Problem

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



What is Required?

You must balance the redox reactions using oxidation numbers.

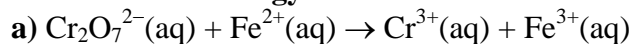
What is Given?

The unbalanced equations are given with the formula of the reactants and products.

Plan Your Strategy

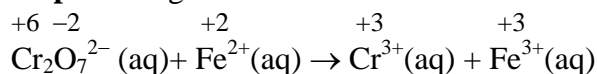
Assign an oxidation number (O.N.) to each element. Follow the rules for balancing an equation in acid conditions using O.N.

Act on Your Strategy



Step 1. The unbalanced equation is $\text{Cr}_2\text{O}_7^{2-}(\text{aq}) + \text{Fe}^{2+}(\text{aq}) \rightarrow \text{Cr}^{3+}(\text{aq}) + \text{Fe}^{3+}(\text{aq})$

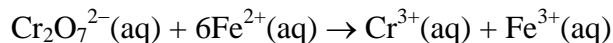
Step 2. Assign an O.N. to each element.



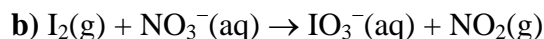
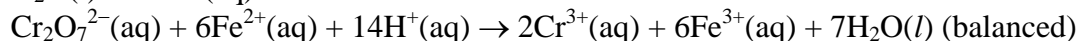
Step 3. The O.N. of iron undergoes an increase of 1 from +2 in $\text{Fe}^{2+}(\text{aq})$ to +3 in $\text{Fe}^{3+}(\text{aq})$. The O.N. of chromium undergoes a decrease of 3 from +6 in $\text{Cr}_2\text{O}_7^{2-}(\text{aq})$ to +3 in $\text{Cr}^{3+}(\text{aq})$.

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

Step 5. Use the 3:1 ratio in **Step 4** to balance the iron and chromium atoms. There must be 3 iron atoms for each chromium atom. Consider that there are 2 chromium atoms on the left.



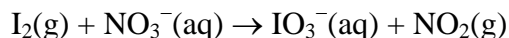
Step 6. Complete the balancing of iron by inspection. Balance for acidic conditions by adding $\text{H}_2\text{O}(\text{l})$ and $\text{H}^+(\text{aq})$.



Step 1. The unbalanced equation is $\text{I}_2(\text{g}) + \text{NO}_3^-(\text{aq}) \rightarrow \text{IO}_3^-(\text{aq}) + \text{NO}_2(\text{g})$

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

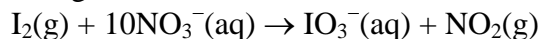




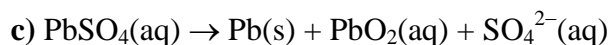
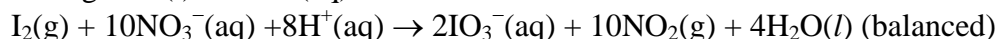
Step 3. The O.N. of iodine undergoes an increase of 5 from 0 in $\text{I}_2(\text{g})$ to +5 in $\text{IO}_3^-(\text{aq})$. The O.N. of nitrogen undergoes a decrease of 1 from +5 in $\text{NO}_3^-(\text{aq})$ to +4 in $\text{NO}_2(\text{g})$.

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

Step 5. Use the 5:1 ratio in **Step 4** to balance the nitrogen and iodine atoms. There must be 5 nitrogen atoms for each iodine atom.

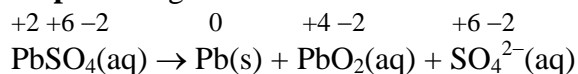


Step 6. Complete the balancing of nitrogen by inspection. Balance for acidic conditions by adding $\text{H}_2\text{O}(\text{l})$ and $\text{H}^+(\text{aq})$.



Step 1. The unbalanced equation is $\text{PbSO}_4(\text{aq}) \rightarrow \text{Pb}(\text{s}) + \text{PbO}_2(\text{aq}) + \text{SO}_4^{2-}(\text{aq})$

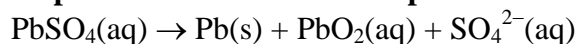
Step 2. Assign an O.N. to each element.



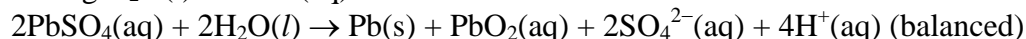
Step 3. The O.N. of lead undergoes an increase of 2 from +2 in $\text{PbSO}_4(\text{aq})$ to +4 in $\text{PbO}_2(\text{aq})$. The O.N. of lead undergoes a decrease of 2 from +2 in $\text{PbSO}_4(\text{aq})$ to 0 in $\text{Pb}(\text{s})$.

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

Step 5. Use the 1:1 ratio in **Step 4** to balance the lead in $\text{Pb}(\text{s})$ and $\text{PbO}_2(\text{aq})$.



Step 6. Complete the balancing of lead and sulfur by inspection. Balance for acidic conditions by adding $\text{H}_2\text{O}(\text{l})$ and $\text{H}^+(\text{aq})$.



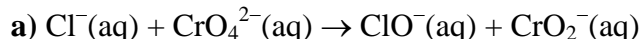
Check Your Solution

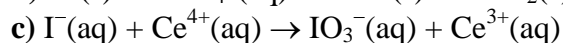
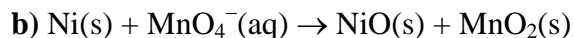
Since there is the same number of atoms of each element on both sides of the equations and the net charge is the same on both sides of the equations, the equation is balanced.

18.

Problem

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





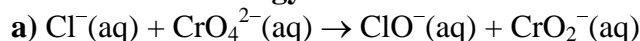
What is Given?

The unbalanced equations are given with the formula of the reactants and products.

Plan Your Strategy

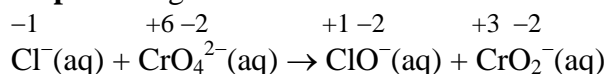
Assign an oxidation number (O.N.) to each element. Follow the rules for balancing an equation in basic conditions using O.N.

Act on Your Strategy



Step 1. The unbalanced equation is $\text{Cl}^-(\text{aq}) + \text{CrO}_4^{2-}(\text{aq}) \rightarrow \text{ClO}^-(\text{aq}) + \text{CrO}_2^-(\text{aq})$

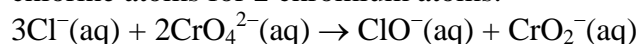
Step 2. Assign an O.N. to each element.



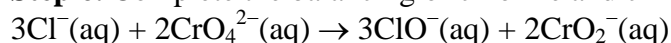
Step 3. The O.N. of chlorine undergoes an increase of 2 from -1 in $\text{Cl}^-(\text{aq})$ to $+1$ in $\text{ClO}^-(\text{aq})$. The O.N. of chromium undergoes a decrease of 3 from $+6$ in $\text{CrO}_4^{2-}(\text{aq})$ to $+3$ in $\text{CrO}_2^-(\text{aq})$.

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

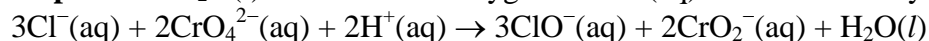
Step 5. Use the 3:2 ratio in **Step 4** to balance the chlorine and chromium atoms. There must be 3 chlorine atoms for 2 chromium atoms.



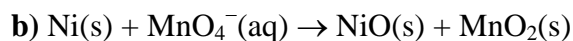
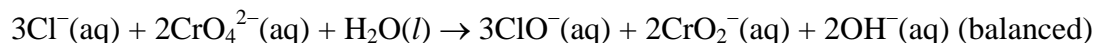
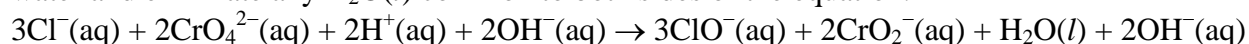
Step 6. Complete the balancing of chlorine and chromium by inspection.



Step 7. Add $\text{H}_2\text{O}(l)$ to balance the oxygen and $\text{H}^+(\text{aq})$ to balance the hydrogen.

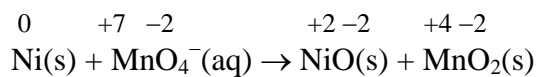


Step 8. Add $\text{OH}^-(\text{aq})$ to balance for basic conditions. Combine the $\text{H}^+(\text{aq})$ and $\text{OH}^-(\text{aq})$ to form water and eliminate any $\text{H}_2\text{O}(l)$ common to both sides of the equation.



Step 1. The unbalanced equation is $\text{Ni(s)} + \text{MnO}_4^-(\text{aq}) \rightarrow \text{NiO(s)} + \text{MnO}_2(\text{s})$

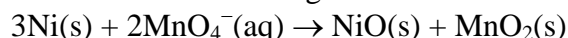
Step 2. Assign an O.N. to each element.



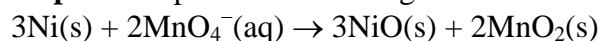
Step 3. The O.N. of nickel undergoes an increase of 2 from 0 in Ni(s) to +2 in NiO(s). The O.N. of manganese undergoes a decrease of 3 from +7 in MnO₄⁻(aq) to +4 in MnO₂(s).

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

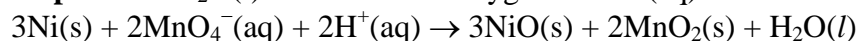
Step 5. Use the 3:2 ratio in **Step 4** to balance the nickel and manganese atoms. There must be 3 nickel atoms for 2 manganese atoms.



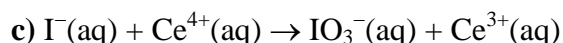
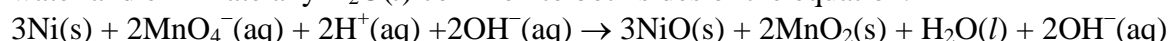
Step 6. Complete the balancing of chlorine and chromium by inspection.



Step 7. Add H₂O(l) to balance the oxygen and H⁺(aq) to balance the hydrogen.

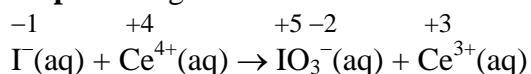


Step 8. Add OH⁻(aq) to balance for basic conditions. Combine the H⁺(aq) and OH⁻(aq) to form water and eliminate any H₂O(l) common to both sides of the equation.



Step 1. The unbalanced equation is $\text{I}^-(\text{aq}) + \text{Ce}^{4+}(\text{aq}) \rightarrow \text{IO}_3^-(\text{aq}) + \text{Ce}^{3+}(\text{aq})$

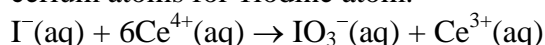
Step 2. Assign an O.N. to each element.



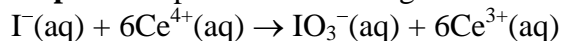
Step 3. The O.N. of iodine undergoes an increase of 6 from -1 in I⁻(aq) to +5 in IO₃⁻(aq). The O.N. of cerium undergoes a decrease of 1 from +4 in Ce⁴⁺(aq) to +3 in Ce³⁺(aq).

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

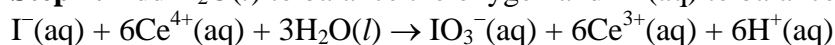
Step 5. Use the 6:1 ratio in **Step 4** to balance the cerium and iodine atoms. There must be 6 cerium atoms for 1 iodine atom.



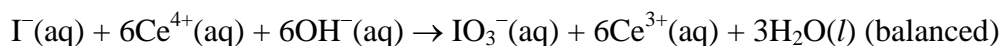
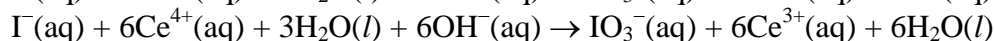
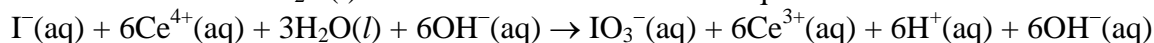
Step 6. Complete the balancing of cerium by inspection.



Step 7. Add $\text{H}_2\text{O}(l)$ to balance the oxygen and $\text{H}^+(\text{aq})$ to balance the hydrogen.



Step 8. Add $\text{OH}^-(\text{aq})$ to balance for basic conditions. Combine the $\text{H}^+(\text{aq})$ and $\text{OH}^-(\text{aq})$ to form water and eliminate $\text{H}_2\text{O}(l)$ common to both sides of the equation.



Check Your Solution

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

19.

Problem

An analyst prepares a $\text{H}_2\text{O}_2(\text{aq})$ sample solution by placing 1.284 g of $\text{H}_2\text{O}_2(\text{aq})$ solution in a flask then diluting it with water and adding sulfuric acid to acidify it. The analyst titrates the $\text{H}_2\text{O}_2(\text{aq})$ sample solution with 0.020 45 mol/L $\text{KMnO}_4(\text{aq})$ and determines that 38.95 mL of $\text{KMnO}_4(\text{aq})$ is required to reach the endpoint.

- What is the mass of pure $\text{H}_2\text{O}_2(\text{aq})$ that is present in the sample solution?
- What is the mass percent of pure $\text{H}_2\text{O}_2(\text{aq})$ in the sample solution?

What is Required?

You must calculate the mass of $\text{H}_2\text{O}_2(\text{aq})$ in a sample and express this result in terms of mass percent.

What is Given?

$$C_{\text{MnO}_4}(\text{aq}) = 0.020\ 45\ \text{mol/L}$$

$$V_{\text{MnO}_4}(\text{aq}) = 38.95\ \text{ml} = 0.038\ 95\ \text{L}$$

$$m_{\text{H}_2\text{O}_2}(\text{aq}) = 1.284\ \text{g}$$

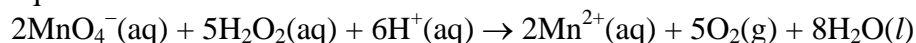
Plan Your Strategy

Write a balanced net ionic equation for the reaction between $\text{MnO}_4^-(\text{aq})$ and $\text{H}_2\text{O}_2(\text{aq})$ in acid solution. Calculate the number of moles (n) of $\text{MnO}_4^-(\text{aq})$ using $n = CV$. Use the mol ratio from the balanced equation to calculate the number of moles of $\text{H}_2\text{O}_2(\text{aq})$. Calculate the molar mass of $\text{H}_2\text{O}_2(\text{aq})$ and find the mass of $\text{H}_2\text{O}_2(\text{aq})$ using $m = n \times M$. The mass percent $\text{H}_2\text{O}_2(\text{aq})$ can be calculated using

$$\text{mass percent} = \frac{\text{mass H}_2\text{O}_2(\text{aq})}{\text{mass of solution}} \times 100\%$$

Act on Your Strategy

Following the rules for balancing a redox reaction in acid solution, the balanced net ionic equation is



$$\text{number of mol MnO}_4(\text{aq}) = C \times V = 0.020\ 45\ \text{mol/L} \times 0.038\ 95\ \text{L} = 0.000\ 7965\ \text{mol}$$

$$\text{mol H}_2\text{O}_2(\text{aq}) = 0.000\ 7965\ \text{mol MnO}_4(\text{aq}) \times \frac{5\ \text{mol H}_2\text{O}_2}{2\ \text{mol MnO}_4^-} = 0.001\ 991\ \text{mol H}_2\text{O}_2$$

$$M(\text{H}_2\text{O}_2) = 34.02\ \text{g/mol}$$

$$m_{\text{H}_2\text{O}_2} = 0.001\ 991 \times 34.02\ \text{g/mol} = 0.067\ 74\ \text{g}$$

$$\text{mass percent} = \frac{\text{mass H}_2\text{O}_2}{\text{mass of solution}} \times 100\ \% = \frac{0.06774\ \text{g}}{1.284\ \text{g}} \times 100\ \% = 5.276\%$$

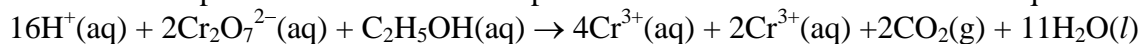
Check Your Solution

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

20.

Problem

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



If 32.35 mL of 0.050 23 mol/L $\text{Cr}_2\text{O}_7^{2-}(\text{aq})$ is required to titrate 27.00 g plasma, what is the mass percent of alcohol in the plasma?

What is Required?

You must calculate the mass percent of alcohol in the plasma.

What is Given?

The balanced net ionic equation for the redox reaction is given.

$$\text{mass plasma} = 27.99\ \text{g}$$

$$\text{volume Cr}_2\text{O}_7^{2-}(\text{aq}) = V = 32.35\ \text{L} = 0.032\ 35\ \text{L}$$

$$\text{concentration Cr}_2\text{O}_7^{2-}(\text{aq}) = C = 0.050\ 23\ \text{mol/L}$$

Plan Your Strategy

Calculate the number of moles (n) of $\text{Cr}_2\text{O}_7^{2-}(\text{aq})$ using $n = CV$. Use the mol ratio from the balanced equation to calculate the number of moles of $\text{C}_2\text{H}_5\text{OH}(\text{aq})$. Calculate the molar mass of $\text{C}_2\text{H}_5\text{OH}(\text{aq})$ and find the mass of $\text{C}_2\text{H}_5\text{OH}(\text{aq})$ using $m = n \times M$. The mass percent $\text{C}_2\text{H}_5\text{OH}(\text{aq})$ can be calculated using

$$\text{mass percent} = \frac{\text{mass C}_2\text{H}_5\text{OH}}{\text{mass of solution}} \times 100\%$$

Act on Your Strategy

$$\text{amount Cr}_2\text{O}_7^{2-}(\text{aq}) = CV = 0.050\ 23\ \text{mol/L} \times 0.032\ 35\ \text{L} = 0.001\ 625\ \text{mol}$$

$$\text{amount C}_2\text{H}_5\text{OH}(\text{aq}) = 0.001\ 625\ \text{mol Cr}_2\text{O}_7^{2-}(\text{aq}) \times \frac{1\ \text{mol C}_2\text{H}_5\text{OH}}{2\ \text{mol Cr}_2\text{O}_7^{2-}} = 0.000\ 8125\ \text{mol}$$

$$M(\text{C}_2\text{H}_5\text{OH}) = 46.08\ \text{g/mol}$$

$$\text{mass C}_2\text{H}_5\text{OH}(\text{aq}) = 0.000\ 8125\ \text{mol} \times 46.08\ \text{g/mol} = 0.037\ 44\ \text{g}$$

$$\text{mass percent} = \frac{\text{mass C}_2\text{H}_5\text{OH}}{\text{mass of solution}} \times 100\% = \frac{0.037\ 44\ \text{g}}{27.00\ \text{g}} \times 100\% = 0.1387\%$$

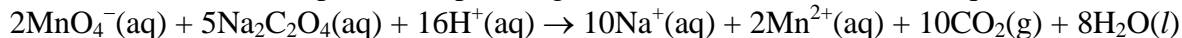
Check Your Solution

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

21.

Problem

An analyst titrates an acidified solution containing 0.153 g of purified sodium oxalate, $\text{Na}_2\text{C}_2\text{O}_4(\text{aq})$, with a potassium permanganate solution, $\text{KMnO}_4(\text{aq})$. The light purple endpoint is reached when the chemist has added 41.45 mL of potassium permanganate solution. What is the molar concentration of the potassium permanganate solution? The balanced equation is



What is Required?

You must calculate the molar concentration of a potassium permanganate solution.

What is Given?

$$\text{mass Na}_2\text{C}_2\text{O}_4(\text{aq}) = 0.153\ \text{g}$$

$$\text{volume KMnO}_4(\text{aq}) = V = 41.45\ \text{mL} = 0.041\ 45\ \text{L}$$

Plan Your Strategy

Determine the molar mass of $\text{Na}_2\text{C}_2\text{O}_4(\text{aq})$ and calculate the number of moles of this

reactant using $n = \frac{m}{M}$. Use the mol ratio in the balanced equation to calculate the amount

of $\text{MnO}_4^-(\text{aq})$. Calculate the concentration of $\text{KMnO}_4(\text{aq})$ using $C = \frac{n}{V}$.

Act on Your Strategy

$$M\text{Na}_2\text{C}_2\text{O}_4(\text{aq}) = 134.00\ \text{g/mol}$$

amount of $\text{Na}_2\text{C}_2\text{O}_4(\text{aq}) = \text{amount of C}_2\text{O}_4^{2-}(\text{aq})$, and amount of $\text{KMnO}_4(\text{aq}) = \text{amount of MnO}_4^-(\text{aq})$

$$\text{amount } \text{C}_2\text{O}_4^{2-}(\text{aq}) = \frac{0.153\text{g}}{134.00\text{g}} = 0.001\ 142\ \text{mol}$$

$$\text{amount } \text{MnO}_4^{-}(\text{aq}) = 0.001\ 142\ \text{mol } \text{C}_2\text{O}_4^{2-}(\text{aq}) \times \frac{2\ \text{mol } \text{MnO}_4^{-}}{5\ \text{mol } \text{C}_2\text{O}_4^{2-}} = 0.0004567\ \text{mol}$$

$$\text{concentration } \text{KMnO}_4(\text{aq}) = \frac{n}{V} = \frac{0.004567\ \text{mol}}{0.04145\ \text{L}} = 0.0110\ \text{mol/L}$$

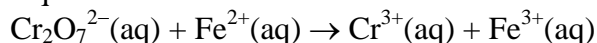
Check Your Solution

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

22.

Problem

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



What is Required?

You must calculate the final molar concentration of the $\text{Fe}^{2+}(\text{aq})$ in the original solution.

What is Given?

concentration $\text{Cr}_2\text{O}_7^{2-}(\text{aq}) = C = 0.020\ 43\ \text{mol/L}$

volume $\text{Cr}_2\text{O}_7^{2-}(\text{aq}) = V = 35.55\ \text{mL} = 0.035\ 55\ \text{L}$

original volume $\text{Cr}_2\text{O}_7^{2-}(\text{aq}) = 25.00\ \text{mL} = 0.025\ 00\ \text{L}$

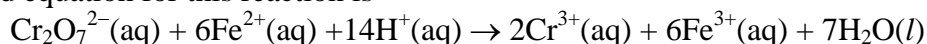
Plan Your Strategy

Follow the rules for balancing a redox reaction in acid solution to obtain the balanced equation for this reaction. Calculate the amount of $\text{Cr}_2\text{O}_7^{2-}(\text{aq})$ that react, $n = CV$. Use the mole ratio from the balanced equation to determine the amount of $\text{Fe}^{2+}(\text{aq})$ that reacted. Calculate the

concentration of $\text{Fe}^{2+}(\text{aq})$, $C = \frac{n}{V}$

Act on Your Strategy

The balanced equation for this reaction is



$$\text{amount } \text{Cr}_2\text{O}_7^{2-}(\text{aq}) = 0.020\ 43\ \text{mol/L} \times 0.035\ 55\ \text{L} = 0.000\ 7263\ \text{mol}$$

$$\text{amount Fe}^{2+}(\text{aq}) = 0.000\ 7263\ \text{mol Cr}_2\text{O}_7^{2-}(\text{aq}) \times \frac{6\ \text{mol Fe}^{2+}}{1\ \text{mol Cr}_2\text{O}_7^{2-}} = 0.004\ 358\ \text{mol}$$

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

Check Your Solution

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