Section 9.1: Electron Transfer Reactions

Tutorial 1 Practice, page 601

1. (a) Given: \( \text{Sn}^{2+}(aq) + \text{Co(s)} \rightarrow \text{Sn}(s) + \text{Co}^{2+}(aq) \)

Solution:

Step 1: Separate the equation into two half-reactions, one for each element.
\( \text{Co(s)} \rightarrow \text{Co}^{2+}(aq) \)
\( \text{Sn}^{2+}(aq) \rightarrow \text{Sn}(s) \)

Step 2: Add electrons to both equations so that the net charge on both sides of each equation is equal.
\( \text{Co(s)} \rightarrow \text{Co}^{2+}(aq) + 2 \text{ e}^- \)
\( \text{Sn}^{2+}(aq) + 2 \text{ e}^- \rightarrow \text{Sn}(s) \)

Step 3: Determine whether each half-reaction equation represents an oxidation or a reduction. The first equation represents an oxidation half-reaction since Co(s) loses electrons. The second equation represents a reduction half-reaction since the \( \text{Sn}^{2+}(aq) \) ion gains electrons.

Statement: The oxidation and reduction half-reaction equations for the reaction are as follows:
Oxidation: \( \text{Co(s)} \rightarrow \text{Co}^{2+}(aq) + 2 \text{ e}^- \)
Reduction: \( \text{Sn}^{2+}(aq) + 2 \text{ e}^- \rightarrow \text{Sn}(s) \)

(b) Given: \( 2 \text{ Ag}^+(aq) + \text{Pb(s)} \rightarrow 2 \text{ Ag(s)} + \text{Pb}^{2+}(aq) \)

Solution:

Step 1: Separate the equation into two half-reactions, one for each element.
\( \text{Pb(s)} \rightarrow \text{Pb}^{2+}(aq) \)
\( 2 \text{ Ag}^+(aq) \rightarrow 2 \text{ Ag(s)} \)

Step 2: Divide each equation by a whole number so that the coefficients in the equation are in the simplest whole-number ratio. In this case, simplify the silver half-reaction by dividing both sides by 2:
\( \text{Pb(s)} \rightarrow \text{Pb}^{2+}(aq) \)
\( \text{Ag}^+(aq) \rightarrow \text{Ag(s)} \)

Step 3: Add electrons to both equations so that the net charge on both sides of each equation is equal.
\( \text{Pb(s)} \rightarrow \text{Pb}^{2+}(aq) + 2 \text{ e}^- \)
\( \text{Ag}^+(aq) + \text{ e}^- \rightarrow \text{Ag(s)} \)

Step 4: Determine whether each half-reaction equation represents an oxidation or a reduction. The first equation represents an oxidation half-reaction since Pb(s) loses electrons. The second equation represents a reduction half-reaction since the \( \text{Ag}^+(aq) \) ion gains an electron.

Statement: The oxidation and reduction half-reaction equations for the reaction are as follows:
Oxidation: \( \text{Pb(s)} \rightarrow \text{Pb}^{2+}(aq) + 2 \text{ e}^- \)
Reduction: \( \text{Ag}^+(aq) + \text{ e}^- \rightarrow \text{Ag(s)} \)
(c) Given: $2 \text{Fe}^{2+}(aq) + \text{I}_2(s) \rightarrow 2 \text{I}^-(aq) + 2 \text{Fe}^{3+}(aq)$

Solution:
Step 1: Separate the equation into two half-reactions, one for each element.
$2 \text{Fe}^{2+}(aq) \rightarrow 2 \text{Fe}^{3+}(aq)$
$\text{I}_2(s) \rightarrow 2 \text{I}^-(aq)$

Step 2: Divide each equation by a whole number so that the coefficients in the equation are in the simplest whole-number ratio. In this case, simplify the iron half-reaction by dividing both sides by 2:
$\text{Fe}^{2+}(aq) \rightarrow \text{Fe}^{3+}(aq)$
$\text{I}_2(s) \rightarrow 2 \text{I}^-(aq)$

Step 3: Add electrons to both equations so that the net charge on both sides of each equation is equal.
$\text{Fe}^{2+}(aq) \rightarrow \text{Fe}^{3+}(aq) + e^-$
$\text{I}_2(s) + 2 e^- \rightarrow 2 \text{I}^-(aq)$

Step 4: Determine whether each half-reaction equation represents an oxidation or a reduction. The first equation represents an oxidation half-reaction since $\text{Fe}^{2+}(aq)$ ion loses an electron.

The second equation represents a reduction half-reaction since $\text{I}_2(s)$ gains electrons.

Statement: The oxidation and reduction half-reaction equations for the reaction are

Oxidation: $\text{Fe}^{2+}(aq) \rightarrow \text{Fe}^{3+}(aq) + e^-$

Reduction: $\text{I}_2(s) + 2 e^- \rightarrow 2 \text{I}^-(aq)$

2. (a) Since Co(s) loses 2 electrons, it is oxidized.
Since Sn$^{2+}$(aq) gains 2 electrons, it is reduced.
(b) Since Pb(s) loses 2 electrons, it is oxidized.
Since Ag$^+$ (aq) gains 1 electron, it is reduced.
(c) Since Fe$^{2+}$(aq) loses 1 electron, it is oxidized.
Since I$_2$(s) gains 2 electrons, it is reduced.

3. (a) Given: $\text{Ni}(s) + \text{CuCl}_2(aq) \rightarrow \text{NiCl}_2(aq) + \text{Cu}(s)$

Solution:
Step 1: Write the equation with its ionic compounds dissociated.
$\text{Ni}(s) + \text{Cu}^{2+}(aq) + 2 \text{Cl}^-(aq) \rightarrow \text{Ni}^{2+}(aq) + 2 \text{Cl}^-(aq) + \text{Cu}(s)$
The chloride ions appear on both sides of the equation, so they are the spectator ions.

Step 2: Write the net ionic equation for the reaction:
$\text{Ni}(s) + \text{Cu}^{2+}(aq) \rightarrow \text{Ni}^{2+}(aq) + \text{Cu}(s)$

Step 3: Separate the equation into two half-reactions, one for each element.
$\text{Ni}(s) \rightarrow \text{Ni}^{2+}(aq)$
$\text{Cu}^{2+}(aq) \rightarrow \text{Cu}(s)$

Step 4: Add electrons to both equations so that the net charge on both sides of each equation is equal.
$\text{Ni}(s) \rightarrow \text{Ni}^{2+}(aq) + 2 e^-$
$\text{Cu}^{2+}(aq) + 2 e^- \rightarrow \text{Cu}(s)$

Step 5: Determine whether each half-reaction equation represents an oxidation or a reduction. The first equation represents an oxidation half-reaction since $\text{Ni}(s)$ loses electrons. The second equation represents a reduction half-reaction since the $\text{Cu}^{2+}(aq)$ ion gains electrons.
Statement: The oxidation and reduction half-reaction equations for the reaction are:
Oxidation: \( \text{Ni}(s) \rightarrow \text{Ni}^{2+}(aq) + 2 \text{e}^- \)
Reduction: \( \text{Cu}^{2+}(aq) + 2 \text{e}^- \rightarrow \text{Cu}(s) \)

(b) Solution:

Step 1: Write a balanced chemical equation for the reaction:
\( \text{Sn(NO}_3)_2(aq) + 2 \text{Cr(NO}_3)_2(aq) \rightarrow \text{Sn}(s) + 2 \text{Cr(NO}_3)_3(aq) \)

Step 2: Write the equation with its ionic compounds dissociated:
\( \text{Sn}^{2+}(aq) + 2 \text{NO}_3^-(aq) + 2 \text{Cr}^{2+}(aq) + 4 \text{NO}_3^-(aq) \rightarrow \text{Sn}(s) + 2 \text{Cr}^{3+}(aq) + 6 \text{NO}_3^-(aq) \)
The nitrate ions appear on both sides of the equation, so they are the spectator ions.

Step 3: Write the net ionic equation for the reaction:
\( \text{Sn}^{2+}(aq) + 2 \text{Cr}^{2+}(aq) \rightarrow \text{Sn}(s) + 2 \text{Cr}^{3+}(aq) \)

Step 4: Separate the equation into two half-reactions, one for each element.
\( 2 \text{Cr}^{2+}(aq) \rightarrow 2 \text{Cr}^{3+}(aq) \)
\( \text{Sn}^{2+}(aq) \rightarrow \text{Sn}(s) \)

Step 5: Divide each equation by a whole number so that the coefficients in the equation are in the simplest whole-number ratio. In this case, simplify the chromium half-reaction by dividing both sides by 2:
\( \text{Cr}^{2+}(aq) \rightarrow \text{Cr}^{3+}(aq) \)
\( \text{Sn}^{2+}(aq) \rightarrow \text{Sn}(s) \)

Step 6: Add electrons to both equations so that the net charge on both sides of each equation is equal.
\( \text{Cr}^{2+}(aq) \rightarrow \text{Cr}^{3+}(aq) + \text{e}^- \)
\( \text{Sn}^{2+}(aq) + 2 \text{e}^- \rightarrow \text{Sn}(s) \)

Step 7: Determine whether each half-reaction equation represents an oxidation or a reduction. The first equation represents an oxidation half-reaction since \( \text{Cr}^{2+}(aq) \) loses an electron. The second equation represents a reduction half-reaction since the \( \text{Sn}^{2+}(aq) \) ion gains electrons.

Statement: The oxidation and reduction half-reaction equations for the reaction are:
Oxidation: \( \text{Cr}^{2+}(aq) \rightarrow \text{Cr}^{3+}(aq) + \text{e}^- \)
Reduction: \( \text{Sn}^{2+}(aq) + 2 \text{e}^- \rightarrow \text{Sn}(s) \)

4. (a) The balanced chemical equation for this reaction is
\( \text{Cl}_2(g) + 2 \text{Kl}(aq) \rightarrow \text{I}_2(s) + 2 \text{KCl}(aq) \)
(b) First, write the equation with its ionic compounds dissociated:
\( \text{Cl}_2(g) + 2 \text{K}^+(aq) + 2 \text{I}^-(aq) \rightarrow \text{I}_2(s) + 2 \text{K}^+(aq) + 2 \text{Cl}^-(aq) \)
Eliminate the spectator ions, the \( \text{K}^+ \) (aq) ions, from the above total ionic equation.
The net ionic equation for the reaction is
\( \text{Cl}_2(g) + 2 \text{I}^-(aq) \rightarrow \text{I}_2(s) + 2 \text{Cl}^-(aq) \)
(c) Solution:

Step 1: Separate the net ionic equation into two half-reactions, one for each element.
\( 2 \text{I}^-(aq) \rightarrow \text{I}_2(s) \)
\( \text{Cl}_2(g) \rightarrow 2 \text{Cl}^-(aq) \)

Step 2: Add electrons to both equations so that the net charge on both sides of each equation is equal.
\( 2 \text{I}^-(aq) \rightarrow \text{I}_2(s) + 2 \text{e}^- \)
\( \text{Cl}_2(g) + 2 \text{e}^- \rightarrow 2 \text{Cl}^-(aq) \)
Step 3: The first equation represents an oxidation half-reaction since I\(^{-}\) (aq) loses electrons. The second equation represents a reduction half-reaction since Cl\(_2\) (g) gains electrons.

**Statement:** The oxidation and reduction half-reaction equations for the reaction are

**Oxidation:** \(2 \text{I}^{-} \rightarrow \text{I}_2 + 2 \text{e}^{-}\)

**Reduction:** \(\text{Cl}_2 + 2 \text{e}^{-} \rightarrow 2 \text{Cl}^{-}\)

**Statement:** Since I\(^{-}\) (aq) loses 1 electron, it is oxidized. Since Cl\(_2\) (g) gains 2 electrons, it is reduced.

**Tutorial 2 Practice, page 604**

1. **(a)** According to the rules for assigning oxidation numbers, the oxidation number of an atom in an element is 0. Therefore, the oxidation number of each nitrogen atom in N\(_2\) is 0.

**Solution:**

**Step 1:** Assign oxidation numbers to elements as listed in Table 1. The oxidation number of oxygen in its compounds is −2.

**Step 2:** Use the zero-sum rule to assign oxidation numbers of other elements.

Since NO\(_2\) is an electrically neutral compound, the sum of the oxidation numbers of oxygen and nitrogen must be 0. Since each oxygen atom has an oxidation number of −2, and there are 2 oxygen atoms in nitrogen dioxide, the nitrogen atom in a molecule of nitrogen dioxide must have an oxidation number of +4 to balance the −4 of the oxygen atoms.

\[
\begin{align*}
\text{NO}_2 & \quad \text{Step 3:} & \text{Check that the sum of the oxidation numbers is equal to 0. Account for the number of each atom in the nitrogen dioxide molecule.} \\
1(+4) & + & 2(−2) = 0 \\
\text{↑} & & \text{↑}
\end{align*}
\]

**Statement:** In a nitrogen dioxide molecule, NO\(_2\), the oxidation number of the nitrogen atom is +4.

**Solution:**

**Step 1:** Assign oxidation numbers to elements as listed in Table 1. The oxidation number of oxygen in its compounds is −2.

**Step 2:** Use the zero-sum rule to assign oxidation numbers of other elements.

Since N\(_2\)O is an electrically neutral compound, the sum of the oxidation numbers of oxygen and nitrogen must be 0. Since each oxygen atom has an oxidation number of −2, and there are 2 nitrogen atoms in N\(_2\)O, each nitrogen atom in a molecule of N\(_2\)O must have an oxidation number of +1 to balance the −2 of the oxygen atom.

\[
\begin{align*}
\text{N}_2\text{O} & \quad \text{Step 3:} & \text{Check that the sum of the oxidation numbers is equal to 0. Account for the number of each atom in the N}_2\text{O molecule.} \\
2(+1) & + & 1(−2) = 0 \\
\text{↑} & & \text{↑}
\end{align*}
\]

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**Statement:** In a nitrogen dioxide molecule, N₂O, the oxidation number of each nitrogen atom is +1.

(d) **Solution:**

**Step 1:** Assign oxidation numbers to elements as listed in Table 1. The oxidation number of the sodium ion, a monatomic ion, is +1, and the oxidation number of oxygen in its compounds is −2.

**Step 2:** Use the zero-sum rule to assign oxidation numbers of other elements. The sum of the oxidation numbers of the atoms in the compound must be 0. Since each sodium ion has an oxidation number of +1 and there is 1 sodium ion, the polyatomic nitrate ion must have an overall charge of −1. Since each oxygen atom has an oxidation number of −2 and there are 3 oxygen atoms, the total contribution of the oxygen atoms is −6. Therefore, the oxidation number of the nitrogen atom must be +5.

\[
\begin{array}{c}
+1 & +5 & -2 \\
\hline
\text{NaNO}_3
\end{array}
\]

**Step 3:** Check that the sum of the oxidation numbers is equal to 0.

\[1(+1) + 1(+5) + 3(-2) = 0\]

\[\uparrow \quad \uparrow \quad \uparrow\]

number of Na atoms \quad number of N atoms \quad number of O atoms

**Statement:** In a sodium nitrate molecule, NaNO₃, the oxidation number of the nitrogen atom is +5.

(e) **Solution:**

**Step 1:** Assign oxidation numbers to elements as listed in Table 1. The oxidation number of hydrogen in its compounds is +1.

**Step 2:** Use the zero-sum rule to assign oxidation numbers of other elements. Since NH₃ is an electrically neutral compound, the sum of the oxidation numbers of nitrogen and hydrogen must be 0. Since each hydrogen atom has an oxidation number of +1, and there are 3 hydrogen atoms in NH₃, the nitrogen atom in a molecule of NH₃ must have an oxidation number of −3 to balance the +3 of the hydrogen atoms.

\[-3 \quad +1 \quad \text{for each atom}\]

\[
\begin{array}{c}
+1 \\
\hline
\text{NH}_3
\end{array}
\]

**Step 3:** Check that the sum of the oxidation numbers is equal to 0.

\[1(-3) + 3(+1) = 0\]

\[\uparrow \quad \uparrow\]

number of N atoms \quad number of H atoms

**Statement:** In an ammonia molecule, NH₃, the oxidation number of the nitrogen atom is −3.

2. (a) **Solution:**

**Step 1:** Assign oxidation numbers to elements as listed in Table 1. The oxidation number of oxygen in its compounds is −2.

**Step 2:** Use the zero-sum rule to assign oxidation numbers of other elements. Since CO is an electrically neutral compound, the sum of the oxidation numbers of carbon and oxygen must be 0. Since each oxygen atom has an oxidation number of −2, and there is 1 oxygen atom in CO, the carbon atom in a molecule of CO must have an oxidation number of +2 to balance the −2 of the oxygen atom.

\[+2 \quad -2 \quad \text{for each atom}\]

\[
\begin{array}{c}
+2 & -2 \\
\hline
\text{CO}
\end{array}
\]
Step 3: Check that the sum of the oxidation numbers is equal to 0.
1(+2) + 1(–2) = 0
↑ number of C atoms  ↑ number of O atoms

Statement: In a carbon monoxide molecule, CO, the oxidation number of the carbon atom is +2.

(b) Solution:
Step 1: Assign oxidation numbers to elements as listed in Table 1. The oxidation number of hydrogen in its compounds is +1.

Step 2: Use the zero-sum rule to assign oxidation numbers of other elements. Since CH₄ is an electrically neutral compound, the sum of the oxidation numbers of carbon and hydrogen must be 0. Since each hydrogen atom has an oxidation number of +1, and there are 4 hydrogen atoms in CH₄, the carbon atom in a molecule of CH₄ must have an oxidation number of –4 to balance the +4 of the hydrogen atoms.

–4 +1  for each atom
CH₄

Step 3: Check that the sum of the oxidation numbers is equal to 0.
1(–4) + 4(+1) = 0
↑ number of C atoms  ↑ number of H atoms

Statement: In a carbon tetrachloride molecule, CH₄, the oxidation number of the carbon atom is –4.

(c) Solution:
Step 1: Assign oxidation numbers to elements as listed in Table 1. The oxidation number of the sodium ion, a monatomic ion, is +1, and the oxidation number of oxygen in its compounds is –2.

Step 2: Use the zero-sum rule to assign oxidation numbers of other elements. Since each sodium ion has an oxidation number of +1 and there are 2 sodium ions, the polyatomic carbonate ion must have an overall charge of –2. Since each oxygen atom has an oxidation number of –2 and there are 3 oxygen atoms, the total contribution of the oxygen atoms is –6. Therefore, the oxidation number of the carbon atom must be +4.

+1  +4  –2  for each atom
Na₂CO₃

Step 3: Check that the sum of the oxidation numbers is equal to 0.
2(+1) + 1(+4) + 3(–2) = 0
↑ number of Na atoms  ↑ number of C atoms  ↑ number of O atoms

Statement: In a sodium carbonate molecule, Na₂CO₃, the oxidation number of the carbon atom is +4.

(d) Solution:
Step 1: Assign oxidation numbers to elements as listed in Table 1. The oxidation number of hydrogen in its compounds is +1 and the oxidation number of oxygen in its compounds is –2.
**Step 2:** Use the zero-sum rule to assign oxidation numbers of other elements. The sum of the oxidation numbers of the atoms in the compound must be 0. Since each hydrogen atom has an oxidation number of +1 and there are 12 hydrogen atoms, the total contribution of the hydrogen atoms is +12. Since each oxygen atom has an oxidation number of −2 and there are 6 oxygen atoms, the total contribution of the oxygen atoms is −12. Therefore, the total contribution of the 6 carbon atoms must be 0. That is, the oxidation number of each carbon atom is 0.

\[ C_6H_{12}O_6 \]

**Step 3:** Check that the sum of the oxidation numbers is equal to 0.

\[
6(0) + 12(+1) + 6(-2) = 0
\]

↑ number of C atoms  ↑ number of H atoms  ↑ number of O atoms

**Statement:** In a glucose molecule, \( C_6H_{12}O_6 \), the oxidation number of each carbon atom is 0.

3. (a) **Solution:**

**Step 1:** Assign oxidation numbers to elements as listed in Table 1. The oxidation number of oxygen in its compounds is −2.

**Step 2:** Use the zero-sum rule to assign oxidation numbers of other elements. Since \( \text{SO}_2 \) is an electrically neutral compound, the sum of the oxidation numbers of sulfur and oxygen must be 0. Since each oxygen atom has an oxidation number of −2, and there are 2 oxygen atoms in \( \text{SO}_2 \), the sulfur atom in a molecule of \( \text{SO}_2 \) must have an oxidation number of +4 to balance the −4 of the oxygen atoms.

\[ \text{SO}_2 \]

**Step 3:** Check that the sum of the oxidation numbers is equal to 0.

\[
1(+4) + 2(-2) = 0
\]

↑ number of S atoms  ↑ number of O atoms

**Statement:** In a sulfur dioxide molecule, \( \text{SO}_2 \), the oxidation number of the sulfur atom is +4.

(b) **Solution:**

**Step 1:** Assign oxidation numbers to elements as listed in Table 1. The oxidation number of oxygen in its compounds is −2.

**Step 2:** Use the ion-charge rule to assign oxidation numbers of other elements. Since the \( \text{SO}_3^{2-} \) ion has a net charge of −2, the sum of the oxidation numbers of all the atoms in the \( \text{SO}_3^{2-} \) ion must equal −2. Since each \( \text{SO}_3^{2-} \) ion contains 3 oxygen atoms, each of which has an oxidation number of −2, the total charge due to oxygen is −6. The net charge of the ion is −2. Therefore, the oxidation number of the sulfur atom must be +4.

\[ \text{SO}_3^{2-} \]

**Step 3:** Check that the sum of the oxidation numbers is equal to −2.

\[
1(+4) + 3(-2) = -2
\]

↑ number of S atoms  ↑ number of O atoms
Statement: In a sulfite ion, $\text{SO}_3^{2-}$, the oxidation number of the sulfur atom is +4.

(c) Solution:

Step 1: Assign oxidation numbers to elements as listed in Table 1. The oxidation number of oxygen in its compounds is −2.

Step 2: Use the ion-charge rule to assign oxidation numbers of other elements. Since the $\text{SO}_4^{2-}$ ion has a net charge of −2, the sum of the oxidation numbers of all the atoms in the $\text{SO}_4^{2-}$ ion must equal −2. Since each $\text{SO}_4^{2-}$ ion contains 4 oxygen atoms, each of which has an oxidation number of −2, the total charge due to oxygen is −8. The net charge of the ion is −2. Therefore, the oxidation number of the sulfur atom must be +6.

\[
\begin{align*}
\text{SO}_4^{2-} & \quad \text{for each atom} \\
+6 & \quad -2
\end{align*}
\]

Step 3: Check that the sum of the oxidation numbers is equal to 0.

\[
\begin{align*}
1(+6) & \quad + \quad 4(-2) \quad = \quad -2
\end{align*}
\]

\[
\begin{array}{c@{\quad}c@{\quad}c}
\uparrow & \uparrow & \\
\text{number of S atoms} & \text{number of O atoms}
\end{array}
\]

Statement: In the sulfate ion, $\text{SO}_4^{2-}$, the oxidation number of the sulfur atom is +6.

(d) Solution:

Step 1: Assign oxidation numbers to elements as listed in Table 1. The oxidation number of oxygen in its compounds is −2.

Step 2: Use the ion-charge rule to assign oxidation numbers of other elements. Since the $\text{S}_2\text{O}_8^{2-}$ ion has a net charge of −2, the sum of the oxidation numbers of all the atoms in the $\text{S}_2\text{O}_8^{2-}$ ion must equal −2. Since each $\text{S}_2\text{O}_8^{2-}$ ion contains 8 oxygen atoms, each of which has an oxidation number of −2, the total charge due to oxygen is −16. Since the total charge of the ion is −2, the total contribution from the 2 sulfur atoms is +14. Therefore, the oxidation number of each sulfur atom is +7.

\[
\begin{align*}
\text{S}_2\text{O}_8^{2-} & \quad \text{for each atom} \\
+7 & \quad -2
\end{align*}
\]

Step 3: Check that the sum of the oxidation numbers is equal to 0.

\[
\begin{align*}
2(+7) & \quad + \quad 8(-2) \quad = \quad -2
\end{align*}
\]

\[
\begin{array}{c@{\quad}c@{\quad}c}
\uparrow & \uparrow & \\
\text{number of S atoms} & \text{number of O atoms}
\end{array}
\]

Statement: In the sulfur octroxide ion, $\text{S}_2\text{O}_8^{2-}$, the oxidation number of each sulfur atom is +7.

4. (a) Solution:

Step 1: Assign oxidation numbers to elements as listed in Table 1. The oxidation number of the sodium ion, a monatomic ion, is +1, and the oxidation number of oxygen in its compounds is −2.

Step 2: Use the zero-sum rule to assign oxidation numbers of other elements. The sum of the oxidation numbers of the atoms in the compound must be 0. Since each sodium ion has an oxidation number of +1 and there are 2 sodium ions, the polyatomic carbonate ion must have an overall charge of −2. Since each oxygen atom has an oxidation number of −2 and there are 3 oxygen atoms, the total contribution of the oxygen atoms is −6. Therefore, the oxidation number for the carbon atom must be +4.

\[
\begin{align*}
\text{Na}_2\text{CO}_3 & \quad \text{for each atom} \\
+1 & \quad +4 & \quad -2
\end{align*}
\]
Step 3: Check that the sum of the oxidation numbers is equal to 0.

\[ 2(+1) + 1(+4) + 3(-2) = 0 \]

\[ \uparrow \quad \uparrow \quad \uparrow \]

number of Na atoms number of C atoms number of O atoms

**Statement:** In the sodium carbonate molecule, Na\(_2\)CO\(_3\), the oxidation number of each sodium ion is +1, the oxidation number of each carbon atom is +4, and the oxidation number of each oxygen atom is −2.

(b) **Solution:**

**Step 1:** Assign oxidation numbers to elements as listed in Table 1. The oxidation number of the potassium ion, a monatomic ion, is +1, and the oxidation number of oxygen in its compounds is −2.

**Step 2:** Use the zero-sum rule to assign oxidation numbers of other elements. The sum of the oxidation numbers of the atoms in the compound must be 0. Since each potassium ion has an oxidation number of +1 and there are 2 potassium ions, the polyatomic dichromate ion must have an overall charge of −2. Since each oxygen atom has an oxidation number of −2 and there are 7 oxygen atoms, the total contribution of the oxygen atoms is −14. Since the total charge of the dichromate ion is −2, the total contribution from the 2 chromium atoms must be +12. Therefore, the oxidation number of each chromium atom is +6.

\[ +1 \quad +6 \quad -2 \quad \text{for each atom} \]

K\(_2\)Cr\(_2\)O\(_7\)

**Step 3:** Check that the sum of the oxidation numbers is equal to 0.

\[ 2(+1) + 2(+6) + 7(-2) = 0 \]

\[ \uparrow \quad \uparrow \quad \uparrow \]

number of K atoms number of Cr atoms number of O atoms

**Statement:** In the potassium chromate molecule, K\(_2\)Cr\(_2\)O\(_7\), the oxidation number of each potassium ion is +1; the oxidation number of each chromium atom is +6; and the oxidation number of each oxygen atom is −2.

(c) **Solution:**

**Step 1:** Assign oxidation numbers to elements as listed in Table 1. The oxidation number of hydrogen in its compounds is +1, and the oxidation number of oxygen in its compounds is −2.

**Step 2:** Use the zero-sum rule to assign oxidation numbers of other elements. The sum of the oxidation numbers of the atoms in the compound must be 0. Since each hydrogen atom has an oxidation number of +1 and there is 1 hydrogen atom, the contribution of the hydrogen atoms is +1. Since each oxygen atom has an oxidation number of −2 and there are 4 oxygen atoms, the total contribution of the oxygen atoms is −8. Therefore, the oxidation number of the chlorine atom must be +7.

\[ +1 \quad +7 \quad -2 \quad \text{for each atom} \]

HClO\(_4\)
**Step 3:** Check that the sum of the oxidation numbers is equal to 0.

\[
1(+1) + 1(+7) + 4(-2) = 0
\]

<table>
<thead>
<tr>
<th>number of H atoms</th>
<th>number of Cl atoms</th>
<th>number of O atoms</th>
</tr>
</thead>
<tbody>
<tr>
<td>↑</td>
<td>↑</td>
<td>↑</td>
</tr>
</tbody>
</table>

**Statement:** In perchloric acid, HClO₄, the oxidation number of the hydrogen atom is +1; the oxidation number of the chlorine atom is +7; and the oxidation number of each oxygen atom is −2.

**d) Solution:**

**Step 1:** Assign oxidation numbers to elements as listed in Table 1. Since the charge on the copper(II) ion is \(2^+\), the oxidation number of copper(II) ion is +2, and the oxidation number of oxygen in its compounds is −2.

**Step 2:** Use the zero-sum rule to assign oxidation numbers of other elements. The sum of the oxidation numbers of the atoms in the compound must be 0.

Since each copper(II) ion has an oxidation number of +2 and there are 3 copper(II) ions, the two polyatomic phosphate ions must have an overall charge of −6. That means, each polyatomic phosphate ion has an overall charge of −3.

Since each oxygen atom has an oxidation number of −2 and there are 4 oxygen atoms in a polyatomic phosphate ion, the total contribution of the oxygen atoms in each is −8. There is 1 phosphorus atom in each polyatomic phosphate ion. Therefore, the oxidation number of the phosphorus atom must be +5 to balance the −8 of the oxygen atoms.

\[
\begin{align*}
Cu_{3}(PO_{4})_{2} & \quad +5 \quad -2 \\
\text{for each atom} & \quad & \\
\end{align*}
\]

**Step 3:** Check that the sum of the oxidation numbers is equal to 0.

\[
3(+2) + 2 \left[1(+5) + 4(-2)\right] = 0
\]

<table>
<thead>
<tr>
<th>number of Cu atoms</th>
<th>number of P atoms</th>
<th>number of O atoms</th>
</tr>
</thead>
<tbody>
<tr>
<td>↑</td>
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**Statement:** In a molecule of copper(II) phosphate, Cu₃(PO₄)₂, the oxidation number of each copper atom is +2; the oxidation number of each phosphorus atom is +5; and the oxidation number of each oxygen atom is −2.

**Tutorial 3 Practice, page 606**

1. (a) Solution:

**Step 1:** Assign oxidation numbers to each atom/ion in the chemical equation.

\[
Cu(s) + 2 \text{Ag}^{+}(aq) \rightarrow 2 \text{Ag}(s) + Cu^{2+}(aq)
\]

**Step 2:** Determine how the oxidation number on each element changes, and identify these changes as either oxidation or reduction.

\[
\begin{align*}
\text{Cu(s)} & \quad +1 \quad +1 \\
\text{2 e}^- & \quad /Cu \text{ lost} \\
\text{Cu(aq)} & \quad +1 \quad -1 \\
\end{align*}
\]

The oxidation number of copper increases from 0 to +2, so it is oxidized.

The oxidation number of silver decreases from +1 to 0, so it is reduced.

**Statement:** Copper is oxidized and silver is reduced.
(b) Solution:
Step 1: Assign oxidation numbers to each atom/ion in the chemical equation.

\[ 4 \text{Fe}(s) + 3 \text{O}_2(g) \rightarrow 2 \text{Fe}_2\text{O}_3(s) \]

Step 2: Determine how the oxidation number on each element changes, and identify these changes as either oxidation or reduction.

The oxidation number of iron increases from 0 to +3, so it is oxidized.
The oxidation number of oxygen decreases from 0 to –2, so it is reduced.

Statement: Iron is oxidized and oxygen is reduced.

2. (a) Solution:
Step 1: Assign oxidation numbers to each atom/ion in the chemical equation.

\[ \text{Zn}(s) + 2 \text{HCl}(aq) \rightarrow \text{ZnCl}_2(aq) + \text{H}_2(g) \]

Step 2: Determine how the oxidation number on each element changes, and identify these changes as either oxidation or reduction.

The oxidation number of zinc increases from 0 to +2, so it is oxidized.
The oxidation number of hydrogen decreases from +1 to 0, so hydrogen is reduced.
The oxidation number of chlorine does not change.

Statement: Since hydride ions, \( \text{H}^+ \), gains electrons, the oxidizing agent is \( \text{HCl}(aq) \). Since zinc loses electrons, the reducing agent is zinc.

(b) Solution:
Step 1: Assign oxidation numbers to each atom/ion in the chemical equation.

\[ \text{SnO}_2(s) + \text{C}(s) \rightarrow \text{Sn}(s) + \text{CO}_2(g) \]

Step 2: Determine how the oxidation number on each element changes, and identify these changes as either oxidation or reduction.

The oxidation number of carbon increases from 0 to +4, so carbon is oxidized.
The oxidation number of tin decreases from +4 to 0, so tin is reduced.
The oxidation number of oxygen does not change.
Statement: Since tin gains electrons, the oxidizing agent is SnO$_2$(s). Since carbon loses electrons, the reducing agent is carbon.

3. (a) Solution:
Step 1: Assign oxidation numbers to elements as listed in Table 1. The oxidation number of hydrogen in its compounds is +1 and the oxidation number of oxygen in the compound SO$_2$ is −2. Use the zero-sum rule to assign oxidation numbers to the sulfur in H$_2$S and the sulfur in SO$_2$.

\[
2 \text{H}_2\text{S}(g) + 3 \text{O}_2(g) \rightarrow 2 \text{SO}_2(g) + 2 \text{H}_2\text{O}(g)
\]

Step 2: Determine how the oxidation number on each element changes, and identify these changes as either oxidation or reduction.

\[
\begin{align*}
\text{oxidation} & \quad \text{reduction} \\
\text{0} & \quad \text{−2} \\
\text{0} & \quad \text{+4} \\
\text{−2} & \quad \text{0}
\end{align*}
\]

The oxidation number of sulfur increases from −2 to +4, so sulfur is oxidized. The oxidation number of oxygen decreases from 0 to −2, so oxygen is reduced. The oxidation number of hydrogen does not change.

Statement: The oxidation numbers of the atoms in their order in the equation are hydrogen: +1; sulfur: −2; oxygen: 0; sulfur: +4; oxygen: −2; hydrogen: +1; oxygen: −2.

Since oxygen gains electrons, the oxidizing agent is O$_2$(g). Since sulfur in H$_2$S loses electrons, the reducing agent is H$_2$S(g).

(b) Solution:
Step 1: Assign oxidation numbers to elements as listed in Table 1. The oxidation number of hydrogen in its compounds is +1 and the oxidation number of oxygen in the compound CO$_2$ is −2. Use the zero-sum rule to assign oxidation numbers to the carbon in CH$_4$ and the carbon in CO$_2$.

\[
\text{CH}_4(g) + 2 \text{O}_2(g) \rightarrow \text{CO}_2(g) + 2 \text{H}_2\text{O}(g)
\]

Step 2: Determine how the oxidation number on each element changes, and identify these changes as either oxidation or reduction.

\[
\begin{align*}
\text{oxidation} & \quad \text{reduction} \\
\text{0} & \quad \text{+4} \\
\text{0} & \quad \text{−2} \\
\text{−2} & \quad \text{0}
\end{align*}
\]

The oxidation number of carbon increases from −4 to +4, so carbon is oxidized. The oxidation number of oxygen decreases from 0 to −2, so oxygen is reduced. The oxidation number of hydrogen does not change.
**Statement:** The oxidation numbers of the atoms in their order in the equation are carbon: -4; hydrogen: +1; oxygen: 0; carbon: +4; oxygen: -2; hydrogen: +1; O: -2. Since oxygen gains electrons, the oxidizing agent is O\(_2\)\(\text{(g)}\). Since carbon in CH\(_4\) loses electrons, the reducing agent is CH\(_4\)\(\text{(g)}\).

**Section 9.1 Questions, page 607**

1. Answer may vary. Sample answer: A redox reaction is also known as an oxidation–reduction reaction. During the reaction, electrons are transferred from one entity to another. The element that loses electrons is said to be oxidized; its oxidation number increases by the number of electrons lost. The element that gains electrons is said to be reduced; its oxidation number decreases by the number of electrons gained.

2. (a) Given: Mg(s) + 2 H\(^+\)(aq) \(\rightarrow\) Mg\(^{2+}\)(aq) + H\(_2\)(g)

**Solution:**

**Step 1:** Separate the equation into two half-reactions, one for each element.

Mg(s) \(\rightarrow\) Mg\(^{2+}\)(aq)
2 H\(^+\)(aq) \(\rightarrow\) H\(_2\)(g)

**Step 2:** Add electrons to both equations so that the net charge on both sides of each equation is equal.

Mg(s) \(\rightarrow\) Mg\(^{2+}\)(aq) + 2 e\(^-\)
2 H\(^+\)(aq) + 2 e\(^-\) \(\rightarrow\) H\(_2\)(g)

**Step 3:** Determine whether each half-reaction equation represents an oxidation or a reduction. The first equation represents an oxidation half-reaction since Mg(s) loses electrons. The second equation represents a reduction half-reaction since each H\(^+\)(aq) ion gains an electron.

**Statement:** The oxidation and reduction half-reaction equations for the reaction are

**Oxidation:** Mg(s) \(\rightarrow\) Mg\(^{2+}\)(aq) + 2 e\(^-\)

**Reduction:** 2 H\(^+\)(aq) + 2 e\(^-\) \(\rightarrow\) H\(_2\)(g)

(b) Given: 2 Al(s) + Fe\(_2\)O\(_3\)(s) \(\rightarrow\) 2 Fe(l) + Al\(_2\)O\(_3\)(s)

**Solution:**

**Step 1:** Write the equation with its ionic compounds dissociated.

2 Al(s) + 2 Fe\(^{3+}\)(s) + 3 O\(^2-\)(s) \(\rightarrow\) 2 Fe(l) + 2 Al\(^{3+}\)(s) + 3 O\(^2-\)(s)

The oxide ions appear on both sides of the equation, so they are spectator ions.

**Step 2:** Write the net ionic equation for the reaction.

2 Al(s) + 2 Fe\(^{3+}\)(s) \(\rightarrow\) 2 Fe(l) + 2 Al\(^{3+}\)(s)

**Step 3:** Divide the equation by a whole number so that the coefficients in the equation are in the simplest whole-number ratio. In this case, simplify the equation by dividing both sides by 2:

Al(s) + Fe\(^{3+}\)(s) \(\rightarrow\) Fe(l) + Al\(^{3+}\)(s)

**Step 4:** Separate the equation into two half-reactions, one for each element.

Al(s) \(\rightarrow\) Al\(^{3+}\)(s)
Fe\(^{3+}\)(s) \(\rightarrow\) Fe(l)

**Step 5:** Add electrons to both equations so that the net charge on both sides of each equation is equal.

Al(s) \(\rightarrow\) Al\(^{3+}\)(s) + 3 e\(^-\)
Fe\(^{3+}\)(s) + 3 e\(^-\) \(\rightarrow\) Fe(l)
**Step 6:** Determine whether each half-reaction equation represents an oxidation or a reduction. The first equation represents an oxidation halfreaction since Al(s) loses electrons. The second equation represents a reduction half-reaction since each Fe^{3+}(s) ion gains electrons.

**Statement:** The oxidation and reduction half-reaction equations for the reaction are

Oxidation: \( \text{Al}(s) \rightarrow \text{Al}^{3+}(s) + 3 \text{e}^- \)

Reduction: \( \text{Fe}^{3+}(s) + 3 \text{e}^- \rightarrow \text{Fe}(l) \)

3. (a) According to the rules for assigning oxidation numbers, the oxidation number of an atom in an element is 0. Therefore, the oxidation number of sulfur in \( \text{S}_8 \) is 0.

(b) According to the rules for assigning oxidation numbers, the sum of the oxidation numbers of all atoms in the \( \text{Cr}_2\text{O}_7^{2-} \) ion must equal the overall charge of the ion. Since each \( \text{Cr}_2\text{O}_7^{2-} \) ion contains 7 oxygen atoms, each of which has an oxidation number of \(-2\), the total charge due to oxygen is \(-14\). Since the overall charge of the ion is \(-2\), the total contribution from the 2 chromium atoms is \(+12\). Therefore, the oxidation number of chromium in \( \text{Cr}_2\text{O}_7^{2-} \) is \(+6\).

(c) According to the rules for assigning oxidation numbers, the sum of the oxidation numbers of all atoms in the electrically neutral compound \( \text{N}_2\text{H}_4 \) must equal 0. Since each \( \text{N}_2\text{H}_4 \) molecule contains 4 hydrogen atoms, each of which has an oxidation number of \(+1\), the total charge due to hydrogen is \(+4\). The 2 nitrogen atoms in a molecule of \( \text{N}_2\text{H}_4 \) must have an oxidation number of \(-4\) to balance the \(+4\) of the hydrogen atoms. Therefore, the oxidation number of nitrogen in \( \text{N}_2\text{H}_4 \) is \(-2\).

(d) According to the rules for assigning oxidation numbers, the sum of the oxidation numbers of all atoms in the electrically neutral compound \( \text{MgI}_2 \) must equal 0. The oxidation number of the magnesium ion, a monatomic ion, is \(+2\). Since each \( \text{MgI}_2 \) molecule contains 1 magnesium ion, the 2 iodide ions in a molecule of \( \text{MgI}_2 \) must have an oxidation number of \(-2\) to balance the \(+2\) of the magnesium ion. Therefore, the oxidation number of iodine in \( \text{MgI}_2 \) is \(-1\).

(e) According to the rules for assigning oxidation numbers, the sum of the oxidation numbers of all atoms in the electrically neutral compound \( \text{CO} \) must equal 0. Since each oxygen atom has an oxidation number of \(-2\), and there is 1 oxygen atom in \( \text{CO} \), the carbon atom in a molecule of \( \text{CO} \) must have an oxidation number of \(+2\) to balance the \(-2\) of the oxygen atom. Therefore, the oxidation number of carbon in \( \text{CO} \) is \(+2\).

(f) According to the rules for assigning oxidation numbers, the sum of the oxidation numbers of all atoms in the electrically neutral compound \( \text{NH}_3 \) must equal 0. Since each hydrogen atom has an oxidation number of \(+1\), and there are 3 hydrogen atoms in \( \text{NH}_3 \), the nitrogen atom in a molecule of \( \text{NH}_3 \) must have an oxidation number of \(-3\) to balance the \(+3\) of the hydrogen atoms. Therefore, the oxidation number of nitrogen in \( \text{NH}_3 \) is \(-3\).

(g) According to the rules for assigning oxidation numbers, the sum of the oxidation numbers of all atoms in the electrically neutral compound \( \text{P}_4\text{O}_6 \) must equal 0. Since each \( \text{P}_4\text{O}_6 \) molecule contains 6 oxygen atoms, each of which has an oxidation number of \(-2\), the total charge due to oxygen is \(-12\). The 4 phosphorus atoms in a molecule of \( \text{P}_4\text{O}_6 \) must have an oxidation number of \(+12\) to balance the \(-12\) of the oxygen atoms. Therefore, the oxidation number of phosphorus in \( \text{P}_4\text{O}_6 \) is \(+3\).
According to the rules for assigning oxidation numbers, the sum of the oxidation numbers of all atoms in the $\text{MnO}_4^−$ ion must equal the overall charge of the ion. Since each $\text{MnO}_4^−$ ion contains 4 oxygen atoms, each of which has an oxidation number of $−2$, the total charge due to oxygen is $−8$. Since the overall charge of the ion is $−1$, the total contribution from the Mn atom must be $+7$. Therefore, the oxidation number of manganese in $\text{MnO}_4^−$ is $+7$.

According to the rules for assigning oxidation numbers, the sum of the oxidation numbers of all atoms in the electrically neutral compound $\text{C}_2\text{H}_5\text{OH}$ must equal 0. In a $\text{C}_2\text{H}_5\text{OH}$ molecule, there are 6 hydrogen atoms, each of which has an oxidation number of $+1$, so the total charge due to hydrogen is $+6$. Since each oxygen atom has an oxidation number of $−2$ and there is 1 oxygen atom, the total charge due to the oxygen atom is $−2$. So, the total charge due to the 2 carbon atoms must be $−4$. Therefore, the oxidation number of carbon in $\text{C}_2\text{H}_5\text{OH}$ is $−2$.

According to the rules for assigning oxidation numbers, the sum of the oxidation numbers of all atoms in the electrically neutral compound $\text{Al}_2(\text{SO}_3)_3$ must equal 0. Since each aluminum ion has an oxidation number of $+3$ and there are 2 aluminum ions, the $3 \text{SO}_3^{2−}$ ions must have an overall charge of $−6$ to balance the $+6$ of the aluminum ions. Each $\text{SO}_3^{2−}$ ion has an overall charge of $−2$. Since each oxygen atom has an oxidation number of $−2$ and there are 3 oxygen atoms in a $\text{SO}_3^{2−}$ ion, the total contribution of the oxygen atoms in each is $−6$. Since the overall charge is $−2$, the sulfur atom must have an oxidation number of $+4$ to balance the $−6$ of the oxygen atoms. Therefore, the oxidation number of sulfur in $\text{Al}_2(\text{SO}_3)_3$ is $+4$.

4. (a) Solution:

**Step 1:** Assign oxidation numbers to each atom and ion in the chemical equation. The oxidation number of hydrogen in its compounds is $+1$, and the oxidation number of oxygen in its compounds is $−2$. Use the zero-sum rule to assign oxidation numbers to the carbon in $\text{CH}_4$ and the carbon in $\text{CO}$. The oxidation number of carbon in $\text{CH}_4$ is $−4$ and the oxidation number of C in CO is $+2$.

$$\text{CH}_4(\text{g}) + \text{H}_2\text{O}(\text{g}) \rightarrow \text{CO}(\text{g}) + 3 \text{H}_2(\text{g})$$

**Step 2:** Determine how the oxidation number on each element changes, and identify these changes as either oxidation or reduction. The oxidation number of carbon increases from $−4$ to $+2$, so it is oxidized. The oxidation number of hydrogen decreases from $+1$ to $0$, so it is reduced. The oxidation number of oxygen does not change.

**Statement:** The carbon atom in CH$_4$(g) is oxidized, and the hydrogen atom in CH$_4$(g) is reduced.
(b) Solution:
Step 1: Assign oxidation numbers to each atom/ion in the chemical equation.
The oxidation number of hydrogen in its compounds is +1, and the oxidation number of each monatomic ion equal the charge on the ion. Therefore, the oxidation number of $\text{H}^+$ is +1, the oxidation number of iron in $\text{Fe}^{2+}$ is +2, and the oxidation number of iron in $\text{Fe}^{3+}$ is +3.
In the polyatomic ion $\text{MnO}_4^-$ of overall charge −1, there are 4 oxygen atoms, each of which has an oxidation number of −2. So the total charge due to oxygen is −8. Since the overall charge of the ion is −1, the total contribution from the manganese atom must be +7. Therefore, the oxidation number of manganese in $\text{MnO}_4^-$ is +7.

$$8 \text{H}^+(aq) + \text{MnO}_4^-(aq) + \text{Fe}^{2+}(aq) \rightarrow \text{Mn}^{2+}(aq) + \text{Fe}^{3+}(aq) + 4 \text{H}_2\text{O}(l)$$

Step 2: Determine how the oxidation number on each element changes, and identify these changes as either oxidation or reduction.
The oxidation number of iron increases from +2 to +3, so it is oxidized.
The oxidation number of manganese decreases from +7 to +2, so it is reduced.
Statement: The iron atom in $\text{Fe}^{2+}(aq)$ is oxidized, and the manganese ion, $\text{Mn}^{7+}$, in $\text{MnO}_4^-(aq)$ is reduced.

(c) Solution:
Step 1: Write a balanced chemical equation for the reaction.
$\text{Cu(s)} + 2 \text{AgNO}_3(aq) \rightarrow 2 \text{Ag(s)} + \text{Cu(NO}_3)_2(aq)$
Step 2: Assign oxidation numbers to each atom and ion in the chemical equation.
Since the oxidation number of an atom in an element is 0, the oxidation number of copper metal is 0 and the oxidation number of silver metal is 0.
Since the oxidation number of each monatomic ion equal the charge on the ion, the oxidation number of $\text{Cu}^{2+}$ in $\text{Cu(NO}_3)_2(aq)$ is +2, and the oxidation number of $\text{Ag}^+$ in $\text{AgNO}_3(aq)$ is +1.
In the polyatomic nitrate ion, $\text{NO}_3^-$, of overall charge −1, there are 3 oxygen atoms, each of which has an oxidation number of −2, so the total charge due to oxygen is −6. Since the overall charge of the ion is −1, the total contribution from the nitrogen atom must be +5. Therefore, the oxidation number of nitrogen in $\text{NO}_3^-$ is +5.
Step 3: Determine how the oxidation number on each element changes, and identify these changes as either oxidation or reduction.

$$^0\text{Cu(s)} + 2^0\text{AgNO}_3(aq) \rightarrow 2^0\text{Ag(s)} + ^0\text{Cu(NO}_3)_2(aq)$$
The oxidation number of copper increases from 0 to +2, so it is oxidized.
The oxidation number of silver decreases from +1 to 0, so it is reduced.
Statement: The copper atom in $\text{Cu(s)}$ is oxidized and the silver ion, $\text{Ag}^+$, in $\text{AgNO}_3(aq)$ is reduced.

5. (a) The oxidation number of an entity increases when it becomes oxidized.
(b) The oxidation number of an entity decreases when it is reduced.
(c) During a reaction in which an entity is oxidized, the entity loses electrons. This change is represented by an oxidation half-reaction. The electrons that were lost are transferred to another entity, which gains the electrons and becomes reduced. This change is represented by a reduction half-reaction. Since the electrons are exchanged from one half reaction to the other, both the oxidation and reduction half-reactions have to occur in the same reaction.
6. (a) Solution:
Step 1: Assign oxidation numbers to each atom and ion in the chemical equation.
The oxidation number of hydrogen in its compounds is +1.
Since the oxidation number of a monatomic ion is the same as its charge, the oxidation
number of chlorine in HCl and in NH$_4$Cl is –1.
Since NH$_3$ is an electrically neutral compound, the sum of the oxidation numbers of
nitrogen and hydrogen must be 0. Since each hydrogen atom has an oxidation number of
+1, and there are 3 hydrogen atoms in NH$_3$, the nitrogen atom must have an oxidation
number of –3 to balance the +3 of the hydrogen atoms.
In the NH$_4$Cl molecule, the oxidation number of the hydrogen is +1 and the oxidation
number of chlorine is –1. Since there are 4 hydrogen atoms in NH$_4$Cl, the oxidation
number of nitrogen must be –3.

\[
\begin{align*}
\text{HCl(g)} & \quad + \quad 2 \text{NH}_3(g) \quad \rightarrow \quad \text{NH}_4\text{Cl(s)} \\
\text{Step 2:} & \quad \text{Determine how the oxidation number on each element changes, and identify}
\text{these changes as either oxidation or reduction.}
\text{Statement:} & \quad \text{Since there is no change in oxidation of the elements, the reaction is not a}
\text{redox reaction.}
\end{align*}
\]

(b) Solution:
Step 1: Assign oxidation numbers to each atom/ion in the chemical equation.
Since the oxidation number of an atom in an element is 0, the oxidation number of
magnesium is 0 and the oxidation number of silicon is 0.
Since the oxidation number of each monatomic ion equals the charge on the ion, the
oxidation number of Si$^{4+}$ is +4, the oxidation number of Mg$^{2+}$ is +2, and the oxidation
number of Cl$^-$ is –1.

\[
\begin{align*}
\text{SiCl}_4(s) & \quad + \quad 2 \text{Mg(s)} \quad \rightarrow \quad 2 \text{MgCl}_2(s) \quad + \quad \text{Si(s)} \\
\text{Step 2:} & \quad \text{Determine how the oxidation number on each element changes, and identify}
\text{these changes as either oxidation or reduction.}
\text{Statement:} & \quad \text{The oxidation number of magnesium increases from 0 to +2, so magnesium}
\text{is oxidized. The oxidation number of silicon decreases from +4 to 0, so silicon is}
\text{reduced. The oxidation number of chlorine does not change. Since Si$^{4+}$ gains electrons,}
\text{the oxidizing agent is SiCl}_2(s). \text{Since magnesium loses electrons, the reducing agent}
\text{is Mg(s).}
\end{align*}
\]
(c) Solution:

Step 1: Assign oxidation numbers to each atom/ion in the chemical equation.

Since the oxidation number of an atom in an element is 0, the oxidation number of hydrogen in H₂ is 0.

The oxidation number of hydrogen in its compounds is +1, so the oxidation number of hydrogen in H₂O(g) is +1. The oxidation number of oxygen in its compounds is −2, so the oxidation number of oxygen in H₂O is −2. To balance the charge due to oxygen, the oxidation number of carbon in CO is +2, and the oxidation number of carbon in CO₂ is +4.

\[
\text{CO}(g) + \text{H}_2\text{O}(g) \rightarrow \text{CO}_2(g) + \text{H}_2(g)
\]

Step 2: Determine how the oxidation number on each element changes, and identify these changes as either oxidation or reduction.

Statement: The oxidation number of carbon increases from +2 to +4, so carbon is oxidized. The oxidation number of hydrogen decreases from +1 to 0, so hydrogen is reduced. The oxidation number of oxygen does not change. Since the hydride ion, H⁺, gains electrons, the oxidizing agent is H₂O(g). Since the carbon atom in CO loses electrons, the reducing agent is CO(g).

7. (a) Solution: Use the zero-sum rule to assign oxidation numbers.

O in O₂: Since the oxidation number of an atom in an element is 0, the oxidation number of oxygen in O₂ is 0.

H and O in H₂O: The oxidation number of hydrogen in its compounds is +1, so the oxidation number of hydrogen in H₂O is +1. The oxidation number of oxygen in its compounds is −2, so the oxidation number of oxygen in H₂O is −2.

P in PH₃: Since PH₃ is an electrically neutral compound, the sum of the oxidation numbers of phosphorus and hydrogen must be 0. Since each hydrogen atom has an oxidation number of +1, and there are 3 hydrogen atoms in PH₃, the phosphorus atom must have an oxidation number of −3 to balance the +3 of the hydrogen atoms.

O in P₄O₁₀: Since P₄O₁₀ is an electrically neutral compound, the sum of the oxidation numbers of phosphorus and oxygen must be 0. Since each oxygen atom has an oxidation number of −2, and there are 10 oxygen atoms in P₄O₁₀, the total contribution of the oxygen atoms is −20. The 4 phosphorus atoms must have an oxidation number of +20 to balance the −20 of the oxygen atoms. Therefore, the oxidation number of each phosphorus atom is +5.

Statement: The oxidation numbers of the elements are

Reactants: phosphorus: −3; hydrogen: +1; oxygen: 0

Products: phosphorus: +5; hydrogen: +1; oxygen: −2
(b) Solution: Use the zero-sum rule to assign oxidation numbers.
O in O\(_2\): Since the oxidation number of an atom in an element is 0, the oxidation number of oxygen in O\(_2\) is 0.
Cl in KCl: The oxidation number of the potassium ion, a monatomic ion, is +1, and the oxidation number of the chloride ion in KCl is –1.
Cl in KClO\(_3\): Since each potassium ion has an oxidation number of +1 and there is 1 potassium ion in KClO\(_3\), the polyatomic ClO\(_3\)\(^–\) ion must have an overall charge of –1. Since each oxygen atom has an oxidation number of –2 and there are 3 oxygen atoms, the total contribution of the oxygen atoms is –6. Therefore, the oxidation number of the chlorine atom must be +5.

Statement: The oxidation numbers of the elements are
Reactants: potassium: +1; chlorine: +5; oxygen: –2
Products: potassium: +1; chlorine: –1; oxygen: 0
(c) Solution: Use the zero-sum rule to assign oxidation numbers.
Pb: Since the oxidation number of an atom in an element is 0, the oxidation number of lead in Pb(s) is 0.
O in H\(_2\)SO\(_4\): The oxidation number of hydrogen in its compounds is +1, so the oxidation number of hydrogen in H\(_2\)SO\(_4\) and H\(_2\)O is +1. The oxidation number of oxygen in its compounds is –2, so the oxidation number of oxygen in H\(_2\)O(g) is –2.
S in H\(_2\)SO\(_4\): Since there are 2 hydrogen atoms in H\(_2\)SO\(_4\), the total contribution of the hydrogen atoms is +2. The polyatomic ion SO\(_4\)\(^2–\) must have an overall charge of –2. Since each SO\(_4\)\(^2–\) ion contains 4 oxygen atoms, each of which has an oxidation number of –2, the total contribution due to the oxygen atoms is –8. Therefore, the oxidation number of the sulfur atom must be +6.
Pb in PbO\(_2\): Since PbO\(_2\) is an electrically neutral compound, the sum of the oxidation numbers of lead and oxygen must be 0. Since each oxygen atom has an oxidation number of –2, and there are 2 oxygen atoms, the total contribution of the oxygen atoms is –4. Therefore, the oxidation number of the lead atom is +4.
Pb in PbSO\(_4\): Since PbSO\(_4\) is an electrically neutral compound, the sum of the oxidation numbers of lead and the SO\(_4\)\(^2–\) ion must be 0. Since each SO\(_4\)\(^2–\) ion has an overall charge of –2, the oxidation number of the lead atom is +2.

Statement: The oxidation numbers of the elements are:
Reactants: lead: 0; lead in PbO\(_2\): +4; oxygen: –2; hydrogen: +1; sulfur: +6
Products: lead: +2; sulfur: +6; oxygen: –2; hydrogen: +1

8. From the previous question, for the reaction
4 PH\(_3\)(g) + 8 O\(_2\)(g) → P\(_4\)O\(_{10}\)(s) + 6 H\(_2\)O(l),
the oxidation numbers of the elements are:
Reactants: phosphorus: –3; hydrogen: +1; oxygen: 0
Products: phosphorus: +5; hydrogen: +1; oxygen: –2
Since the oxidation number of P increases from –3 to +5, it is oxidized. Since the oxidation number of O decreases from 0 to –2, it is reduced.
The reaction can be represented by the following oxidation–reduction half-reactions:
Oxidation: P\(^3–\)(g) → P\(^5+\)(s) + 8 e\(^–\)
Reduction: O\(_2\)(g) + 4 e\(^–\) → 2 O\(^2–\)(s)
Since oxygen gains electrons, the oxidizing agent is O\(_2\)(g). Since the phosphorus atom in PH\(_3\) loses electrons, the reducing agent is PH\(_3\)(g).
(b) From the previous question, for the reaction
2 KClO₃(s) → 2 KCl(s) + 3 O₂(g),
the oxidation numbers of the elements are:
Reactants: potassium: +1; chlorine: +5; oxygen: −2
Products: potassium: +1; chlorine: −1; oxygen: 0
Since the oxidation number of oxygen increases from −2 to 0, it is oxidized.
Since the oxidation number of chlorine decreases from +5 to −1, it is reduced.
The reaction can be represented by the following oxidation–reduction half-reactions:
Oxidation: 2 O₂⁻(g) → O₂(g) + 4 e⁻
Reduction: Cl⁺⁺(s) + 6 e⁻ → Cl⁻(s)
Since the chlorine atom in KClO₃ loses electrons, the oxidizing agent is KClO₃(s). Since oxygen ion, O²⁻, in KClO₃ gains electrons, the reducing agent is also KClO₃(s).

(c) From the previous question, for the reaction
Pb(s) + PbO₂(s) + 2 H₂SO₄(aq) → 2 PbSO₄(s) + 2 H₂O(g),
the oxidation numbers of the elements are
Reactants: lead: 0; lead in PbO₂: +4; oxygen: −2; hydrogen: +1; sulfur: +6
Products: lead: +2; sulfur: +6; oxygen: −2; hydrogen: +1
Since the oxidation number of lead increases from 0 to +2, it is oxidized.
Since the oxidation number of lead in PbO₂ decreases from +4 to +2, it is reduced.
The reaction can be represented by the following oxidation–reduction half-reactions:
Oxidation: Pb(s) → Pb⁺⁺(s) + 2 e⁻
Reduction: Pb⁺⁺(s) + 2 e⁻ → Pb²⁺(s)
Since the lead ion, Pb⁺⁺, gains electrons, the oxidizing agent is PbO₂(s). Since lead loses electrons, the reducing agent is Pb(s).

9. (a) **Solution:** Use the zero-sum rule to assign oxidation numbers.

**Step 1:** The oxidation number of oxygen in its compounds is −2. Since CO₂ is an electrically neutral compound, the sum of the oxidation numbers of carbon and oxygen must be 0. Since each oxygen atom has an oxidation number of −2 and there are 2 oxygen atoms in CO₂, the total contribution of the oxygen atoms is −4. Therefore, the carbon atom in CO₂ must have an oxidation number of +4 to balance the −4 of the oxygen atoms.

**Step 2:** The oxidation number of hydrogen in its compounds is +1 and the oxidation number of oxygen in its compounds is −2. Since C₆H₁₂O₆ is an electrically neutral compound, the sum of the oxidation numbers of all the atoms must be 0. Since each hydrogen atom has an oxidation number of +1 and there are 12 hydrogen atoms in C₆H₁₂O₆, the total contribution of the hydrogen atoms is +12. Since each oxygen atom has an oxidation number of −2 and there are 6 oxygen atoms, the total contribution of the oxygen atoms is −12. Therefore, the total contribution of the 6 carbon atoms must be 0. That is, the oxidation number of each carbon atom is 0.

**Statement:** The oxidation number of carbon in CO₂ is +4, and the oxidation number of carbon in C₆H₁₂O₆ is 0.
(b) For the photosynthesis reaction

\[ 6 \text{CO}_2(g) + 6 \text{H}_2\text{O}(l) \rightarrow \text{C}_6\text{H}_{12}\text{O}_6(s) + 6 \text{O}_2(g) \]

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