Section 9.1: Electron Transfer Reactions

Tutorial 1 Practice, page 601

1. (a) Given: $Sn^{2+}(aq) + Co(s) \rightarrow Sn(s) + Co^{2+}(aq)$
Solution:
Step 1: Separate the equation into two half-reactions, one for each element.
$Co(s) \rightarrow Co^{2+}(aq)$
$Sn^{2+}(aq) \rightarrow Sn(s)$
Step 2: Add electrons to both equations so that the net charge on both sides of each equation is equal.
$Co(s) \rightarrow Co^{2+}(aq) + 2\ e^{-}$
$Sn^{2+}(aq) + 2\ e^{-} \rightarrow Sn(s)$
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 $Sn^{2+}(aq)$ ion gains electrons.
Statement: The oxidation and reduction half-reaction equations for the reaction are as follows:
Oxidation: $Co(s) \rightarrow Co^{2+}(aq) + 2\ e^{-}$
Reduction: $Sn^{2+}(aq) + 2\ e^{-} \rightarrow Sn(s)$
(b) Given: $2\ Ag^{+}(aq) + Pb(s) \rightarrow 2\ Ag(s) + Pb^{2+}(aq)$
Solution:
Step 1: Separate the equation into two half-reactions, one for each element.
$Pb(s) \rightarrow Pb^{2+}(aq)$
$2\ Ag^{+}(aq) \rightarrow 2\ Ag(s)$
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:
$Pb(s) \rightarrow Pb^{2+}(aq)$
$Ag^{+}(aq) \rightarrow Ag(s)$
Step 3: Add electrons to both equations so that the net charge on both sides of each equation is equal.
$Pb(s) \rightarrow Pb^{2+}(aq) + 2\ e^{-}$
$Ag^{+}(aq) + e^{-} \rightarrow Ag(s)$
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 $Ag^{+}(aq)$ ion gains an electron.
Statement: The oxidation and reduction half-reaction equations for the reaction are as follows:
Oxidation: $Pb(s) \rightarrow Pb^{2+}(aq) + 2\ e^{-}$
Reduction: $Ag^{+}(aq) + e^{-} \rightarrow Ag(s)$

(c) Given: $2\ Fe^{2+}(aq) + I_2(s) \rightarrow 2\ I^-(aq) + 2\ Fe^{3+}(aq)$
Solution:
Step 1: Separate the equation into two half-reactions, one for each element.
$2\ Fe^{2+}(aq) \rightarrow 2\ Fe^{3+}(aq)$
$I_2(s) \rightarrow 2\ I^-(aq)$
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:
$Fe^{2+}(aq) \rightarrow Fe^{3+}(aq)$
$I_2(s) \rightarrow 2\ I^-(aq)$
Step 3: Add electrons to both equations so that the net charge on both sides of each equation is equal.
$Fe^{2+}(aq) \rightarrow Fe^{3+}(aq) + e^-$
$I_2(s) + 2\ e^- \rightarrow 2\ I^-(aq)$
Step 4: Determine whether each half-reaction equation represents an oxidation or a reduction. The first equation represents an oxidation half-reaction since the $Fe^{2+}(aq)$ ion loses an electron. The second equation represents a reduction half-reaction since $I_2(s)$ gains electrons.
Statement: The oxidation and reduction half-reaction equations for the reaction are
Oxidation: $Fe^{2+}(aq) \rightarrow Fe^{3+}(aq) + e^-$
Reduction: $I_2(s) + 2\ e^- \rightarrow 2\ I^-(aq)$
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: $Ni(s) + CuCl_2(aq) \rightarrow NiCl_2(aq) + Cu(s)$
Solution:
Step 1: Write the equation with its ionic compounds dissociated.
$Ni(s) + Cu^{2+}(aq) + 2\ Cl^-(aq) \rightarrow Ni^{2+}(aq) + 2\ Cl^-(aq) + Cu(s)$
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:
$Ni(s) + Cu^{2+}(aq) \rightarrow Ni^{2+}(aq) + Cu(s)$
Step 3: Separate the equation into two half-reactions, one for each element.
$Ni(s) \rightarrow Ni^{2+}(aq)$
$Cu^{2+}(aq) \rightarrow Cu(s)$
Step 4: Add electrons to both equations so that the net charge on both sides of each equation is equal.
$Ni(s) \rightarrow Ni^{2+}(aq) + 2\ e^-$
$Cu^{2+}(aq) + 2\ e^- \rightarrow Cu(s)$
Step 5: Determine whether each half-reaction equation represents an oxidation or a reduction. The first equation represents an oxidation half-reaction since Ni(s) loses electrons. The second equation represents a reduction half-reaction since the $Cu^{2+}(aq)$ ion gains electrons.

Statement: The oxidation and reduction half-reaction equations for the reaction are:
Oxidation: $Ni(s) \rightarrow Ni^{2+}(aq) + 2\ e^-$
Reduction: $Cu^{2+}(aq) + 2\ e^- \rightarrow Cu(s)$
(b) Solution:
Step 1: Write a balanced chemical equation for the reaction:
$Sn(NO_3)_2(aq) + 2\ Cr(NO_3)_2(aq) \rightarrow Sn(s) + 2\ Cr(NO_3)_3(aq)$
Step 2: Write the equation with its ionic compounds dissociated:
$Sn^{2+}(aq) + 2\ NO_3^-(aq) + 2\ Cr^{2+}(aq) + 4\ NO_3^-(aq) \rightarrow Sn(s) + 2\ Cr^{3+}(aq) + 6\ NO_3^-(aq)$
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:
$Sn^{2+}(aq) + 2\ Cr^{2+}(aq) \rightarrow Sn(s) + 2\ Cr^{3+}(aq)$
Step 4: Separate the equation into two half-reactions, one for each element.
$2\ Cr^{2+}(aq) \rightarrow 2\ Cr^{3+}(aq)$
$Sn^{2+}(aq) \rightarrow Sn(s)$
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:
$Cr^{2+}(aq) \rightarrow Cr^{3+}(aq)$
$Sn^{2+}(aq) \rightarrow Sn(s)$
Step 6: Add electrons to both equations so that the net charge on both sides of each equation is equal.
$Cr^{2+}(aq) \rightarrow Cr^{3+}(aq) + e^-$
$Sn^{2+}(aq) + 2\ e^- \rightarrow Sn(s)$
Step 7: Determine whether each half-reaction equation represents an oxidation or a reduction. The first equation represents an oxidation half-reaction since $Cr^{2+}(aq)$ loses an electron. The second equation represents a reduction half-reaction since the $Sn^{2+}(aq)$ ion gains electrons.
Statement: The oxidation and reduction half-reaction equations for the reaction are:
Oxidation: $Cr^{2+}(aq) \rightarrow Cr^{3+}(aq) + e^-$
Reduction: $Sn^{2+}(aq) + 2\ e^- \rightarrow Sn(s)$
4. (a) The balanced chemical equation for this reaction is
$Cl_2(g) + 2\ KI(aq) \rightarrow I_2(s) + 2\ KCl(aq)$
(b) First, write the equation with its ionic compounds dissociated:
$Cl_2(g) + 2\ K^+(aq) + 2\ I^-(aq) \rightarrow I_2(s) + 2\ K^+(aq) + 2\ Cl^-(aq)$
Eliminate the spectator ions, the $K^+(aq)$ ions, from the above total ionic equation.
The net ionic equation for the reaction is
$Cl_2(g) + 2\ I^-(aq) \rightarrow I_2(s) + 2\ Cl^-(aq)$
(c) Solution:
Step 1: Separate the net ionic equation into two half-reactions, one for each element.
$2\ I^-(aq) \rightarrow I_2(s)$
$Cl_2(g) \rightarrow 2\ Cl^-(aq)$
Step 2: Add electrons to both equations so that the net charge on both sides of each equation is equal.
$2\ I^-(aq) \rightarrow I_2(s) + 2\ e^-$
$Cl_2(g) + 2\ e^- \rightarrow 2\ Cl^-(aq)$

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: $2\ I^-(aq) \rightarrow I_2(s) + 2\ e^-$
Reduction: $Cl_2(g) + 2\ e^- \rightarrow 2\ Cl^-(aq)$
(d) 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.
(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 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.

$$\overset{+4}{N}\overset{-2}{O_2} \quad \text{for each atom}$$

Step 3: Check that the sum of the oxidation numbers is equal to 0. Account for the number of each atom in the nitrogen dioxide molecule.

$$\underset{\substack{\uparrow \\ \text{number of N atoms}}}{1}(+4) + \underset{\substack{\uparrow \\ \text{number of O atoms}}}{2}(-2) = 0$$

Statement: In a nitrogen dioxide molecule, NO_2, the oxidation number of the nitrogen 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 zero-sum rule to assign oxidation numbers of other elements.
Since N_2O 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_2O, each nitrogen atom in a molecule of N_2O must have an oxidation number of +1 to balance the –2 of the oxygen atom.

$$\overset{+1}{N_2}\overset{-2}{O} \quad \text{for each atom}$$

Step 3: Check that the sum of the oxidation numbers is equal to 0. Account for the number of each atom in the N_2O molecule.

$$\underset{\substack{\uparrow \\ \text{number of N atoms}}}{2}(+1) + \underset{\substack{\uparrow \\ \text{number of O atoms}}}{1}(-2) = 0$$

Statement: In a nitrogen dioxide molecule, N_2O, 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.

$\overset{+1}{Na}\overset{+5}{N}\overset{-2}{O_3}$ for each atom

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

$1(+1) \quad + \quad 1(+5) \quad + \quad 3(-2) = 0$

↑ number of Na atoms ↑ number of N atoms ↑ number of O atoms

Statement: In a sodium nitrate molecule, $NaNO_3$, 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 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 in a molecule of NH_3 must have an oxidation number of –3 to balance the +3 of the hydrogen atoms.

$\overset{-3}{N}\overset{+1}{H_3}$ for each atom

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

$1(-3) \quad + \quad 3(+1) = 0$

↑ number of N atoms ↑ number of H atoms

Statement: In an ammonia molecule, NH_3, 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.

$\overset{+2}{C}\overset{-2}{O}$ for each atom

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 CH_4 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 CH_4, the carbon atom in a molecule of CH_4 must have an oxidation number of –4 to balance the +4 of the hydrogen atoms.

$\overset{-4}{C}\overset{+1}{H}_4$ for each atom

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, CH_4, 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.
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 of the carbon atom must be +4.

$\overset{+1}{Na}_2\overset{+4}{C}\overset{-2}{O}_3$ for each atom

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, Na_2CO_3, 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.

$\overset{0}{C}_6\overset{+1}{H}_{12}\overset{-2}{O}_6$ for each atom

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

$$\underset{\substack{\uparrow \\ \text{number of C atoms}}}{6(0)} + \underset{\substack{\uparrow \\ \text{number of H atoms}}}{12(+1)} + \underset{\substack{\uparrow \\ \text{number of O atoms}}}{6(-2)} = 0$$

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 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 SO_2, the sulfur atom in a molecule of SO_2 must have an oxidation number of +4 to balance the –4 of the oxygen atoms.

$\overset{+4}{S}\overset{-2}{O}_2$ for each atom

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

$$\underset{\substack{\uparrow \\ \text{number of S atoms}}}{1(+4)} + \underset{\substack{\uparrow \\ \text{number of O atoms}}}{2(-2)} = 0$$

Statement: In a sulfur dioxide molecule, 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 SO_3^{2-} ion has a net charge of –2, the sum of the oxidation numbers of all the atoms in the SO_3^{2-} ion must equal –2. Since each 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.

$\overset{+4}{S}\overset{-2}{O}_3{}^{2-}$ for each atom

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

$$\underset{\substack{\uparrow \\ \text{number of S atoms}}}{1(+4)} + \underset{\substack{\uparrow \\ \text{number of O atoms}}}{3(-2)} = -2$$

Statement: In a sulfite ion, 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 SO_4^{2-} ion has a net charge of –2, the sum of the oxidation numbers of all the atoms in the SO_4^{2-} ion must equal –2. Since each 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.

$$\overset{+6}{S}\overset{-2}{O}{}_4^{2-} \quad \text{for each atom}$$

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

$$\underset{\substack{\uparrow \\ \text{number of S atoms}}}{1}(+6) \quad + \quad \underset{\substack{\uparrow \\ \text{number of O atoms}}}{4}(-2) \quad = \quad -2$$

Statement: In the sulfate ion, 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 $S_2O_8^{2-}$ ion has a net charge of –2, the sum of the oxidation numbers of all the atoms in the $S_2O_8^{2-}$ ion must equal –2. Since each $S_2O_8^{2-}$ 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.

$$\overset{+7}{S}{}_2\overset{-2}{O}{}_8^{2-} \quad \text{for each atom}$$

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

$$\underset{\substack{\uparrow \\ \text{number of S atoms}}}{2}(+7) \quad + \quad \underset{\substack{\uparrow \\ \text{number of O atoms}}}{8}(-2) \quad = \quad -2$$

Statement: In the sulfur octroxide ion, $S_2O_8^{2-}$, 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.

$$\overset{+1}{Na}{}_2\overset{+4}{C}\overset{-2}{O}{}_3 \quad \text{for each atom}$$

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

$$\underset{\substack{\uparrow \\ \text{number of Na atoms}}}{2(+1)} \quad + \quad \underset{\substack{\uparrow \\ \text{number of C atoms}}}{1(+4)} \quad + \quad \underset{\substack{\uparrow \\ \text{number of O atoms}}}{3(-2)} = 0$$

Statement: In the sodium carbonate molecule, Na_2CO_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.

$$\overset{+1}{K_2}\overset{+6}{Cr_2}\overset{-2}{O_7} \quad \text{for each atom}$$

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

$$\underset{\substack{\uparrow \\ \text{number of K atoms}}}{2(+1)} \quad + \quad \underset{\substack{\uparrow \\ \text{number of Cr atoms}}}{2(+6)} \quad + \quad \underset{\substack{\uparrow \\ \text{number of O atoms}}}{7(-2)} = 0$$

Statement: In the potassium chromate molecule, $K_2Cr_2O_7$, 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.

$$\overset{+1}{H}\overset{+7}{Cl}\overset{-2}{O_4} \quad \text{for each atom}$$

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

$$\underset{\substack{\uparrow \\ \text{number of H atoms}}}{1(+1)} + \underset{\substack{\uparrow \\ \text{number of Cl atoms}}}{1(+7)} + \underset{\substack{\uparrow \\ \text{number of O atoms}}}{4(-2)} = 0$$

Statement: In perchloric acid, $HClO_4$, 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.

$$\overset{+2}{Cu}_3(\overset{+5}{P}\overset{-2}{O}_4)_2 \quad \text{for each atom}$$

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

$$\underset{\substack{\uparrow \\ \text{number of Cu atoms}}}{3(+2)} + 2\,[\underset{\substack{\uparrow \\ \text{number of P atoms}}}{1(+5)} + \underset{\substack{\uparrow \\ \text{number of O atoms}}}{4(-2)}] = 0$$

Statement: In a molecule of copper(II) phosphate, $Cu_3(PO_4)_2$, 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.

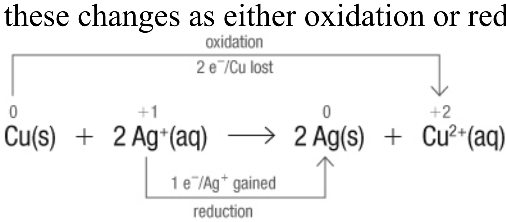
Tutorial 3 Practice, page 606

1. (a) Solution:

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

$$\overset{0}{Cu}(s) + 2\,\overset{+1}{Ag}{}^{+}(aq) \rightarrow 2\,\overset{0}{Ag}(s) + \overset{+2}{Cu}{}^{2+}(aq)$$

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

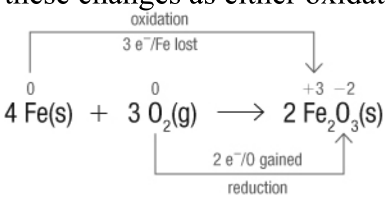


The oxidation number of copper increases from 0 to +2, so it is oxidized.
The oxidation number of silver decreases from +1 to 0, so it is reduced.
Statement: Copper is oxidized and silver is reduced.

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

$4\ \overset{0}{\text{Fe}}(s) + 3\ \overset{0}{\text{O}}_2(g) \rightarrow 2\ \overset{+3}{\text{Fe}}_2\overset{-2}{\text{O}}_3(s)$

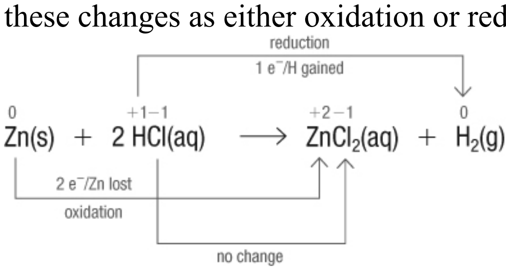
Step 2: Determine how the oxidation number on each element changes, and identify these changes as either oxidation or reduction.



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.

$\overset{0}{\text{Zn}}(s) + 2\ \overset{+1}{\text{H}}\overset{-1}{\text{Cl}}(aq) \rightarrow \overset{+2}{\text{Zn}}\overset{-1}{\text{Cl}}_2(aq) + \overset{0}{\text{H}}_2(g)$

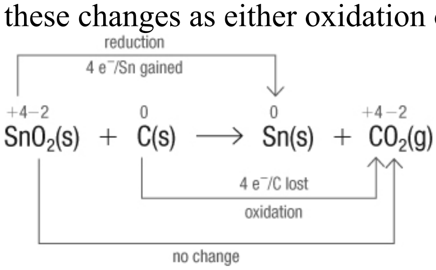
Step 2: Determine how the oxidation number on each element changes, and identify these changes as either oxidation or reduction.



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, H^+, gains electrons, the oxidizing agent is 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.

$\overset{+4}{\text{Sn}}\overset{-2}{\text{O}}_2(s) + \overset{0}{\text{C}}(s) \rightarrow \overset{0}{\text{Sn}}(s) + \overset{+4}{\text{C}}\overset{-2}{\text{O}}_2(g)$

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.

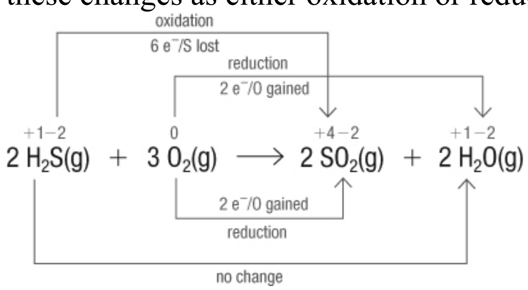
Statement: Since tin gains electrons, the oxidizing agent is SnO_2(s). Since carbon loses electrons, the reducing agent is carbon.

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_2S and the sulfur in SO_2.

$$2\ \overset{+1}{H_2}\overset{-2}{S}(g) + 3\ \overset{0}{O_2}(g) \rightarrow 2\ \overset{+4}{S}\overset{-2}{O_2}(g) + 2\ \overset{+1}{H_2}\overset{-2}{O}(g)$$

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



The oxidation number of 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.

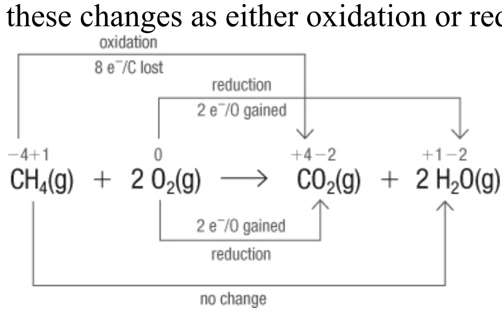
Statement: The oxidation numbers of the atoms in their order in the equation are hydrogen: +1; sulfur: –2; oxygen: 0; sulfur: +4; oxygen: –2; hydrogen: +1; oxygen: –2 Since oxygen gains electrons, the oxidizing agent is O_2(g). Since sulfur in H_2S loses electrons, the reducing agent is H_2S(g).

(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.

$$\overset{-4}{C}\overset{+1}{H_4}(g) + 2\ \overset{0}{O_2}(g) \rightarrow \overset{+4}{C}\overset{-2}{O_2}(g) + 2\ \overset{+1}{H_2}\overset{-2}{O}(g)$$

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 +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 $O_2(g)$. Since carbon in CH_4 loses electrons, the reducing agent is $CH_4(g)$.

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: $2\ Al(s) + Fe_2O_3(s) \rightarrow 2\ Fe(l) + Al_2O_3(s)$

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 half-reaction 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: $Al(s) \rightarrow Al^{3+}(s) + 3\ e^{-}$
Reduction: $Fe^{3+}(s) + 3\ e^{-} \rightarrow Fe(l)$
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 S_8 is 0.
(b) According to the rules for assigning oxidation numbers, the sum of the oxidation numbers of all atoms in the $Cr_2O_7^{2-}$ ion must equal the overall charge of the ion. Since each $Cr_2O_7^{2-}$ 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 $Cr_2O_7^{2-}$ 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 N_2H_4 must equal 0. Since each N_2H_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 N_2H_4 must have an oxidation number of –4 to balance the +4 of the hydrogen atoms. Therefore, the oxidation number of nitrogen in N_2H_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 MgI_2 must equal 0. The oxidation number of the magnesium ion, a monatomic ion, is +2. Since each MgI_2 molecule contains 1 magnesium ion, the 2 iodide ions in a molecule of MgI_2 must have an oxidation number of –2 to balance the +2 of the magnesium ion. Therefore, the oxidation number of iodine in 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 CO must equal 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. Therefore, the oxidation number of carbon in 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 NH_3 must equal 0. Since each hydrogen atom has an oxidation number of +1, and there are 3 hydrogen atoms in NH_3, the nitrogen atom in a molecule of NH_3 must have an oxidation number of –3 to balance the +3 of the hydrogen atoms. Therefore, the oxidation number of nitrogen in 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 P_4O_6 must equal 0. Since each P_4O_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 P_4O_6 must have an oxidation number of +12 to balance the –12 of the oxygen atoms. Therefore, the oxidation number of phosphorus in P_4O_6 is +3.

(h) According to the rules for assigning oxidation numbers, the sum of the oxidation numbers of all atoms in the MnO_4^- ion must equal the overall charge of the ion. Since each 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 MnO_4^- is +7.
(i) According to the rules for assigning oxidation numbers, the sum of the oxidation numbers of all atoms in the electrically neutral compound C_2H_5OH must equal 0. In a C_2H_5OH 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 C_2H_5OH is –2.
(j) According to the rules for assigning oxidation numbers, the sum of the oxidation numbers of all atoms in the electrically neutral compound $Al_2(SO_3)_3$ must equal 0. Since each aluminum ion has an oxidation number of +3 and there are 2 aluminum ions, the 3 SO_3^{2-} ions must have an overall charge of –6 to balance the +6 of the aluminum ions. Each 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 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 $Al_2(SO_3)_3$ 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 CH_4 and the carbon in CO. The oxidation number of carbon in CH_4 is –4 and the oxidation number of C in CO is +2.

$$\overset{-4}{C}\overset{+1}{H}_4(g) + \overset{+1}{H}_2\overset{-2}{O}(g) \rightarrow \overset{+2}{C}\overset{-2}{O}(g) + 3\ \overset{0}{H}_2(g)$$

Step 2: Determine how the oxidation number on each element changes, and identify these changes as either oxidation or reduction.
The oxidation number of 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 H^+ is +1, the oxidation number of iron in Fe^{2+} is +2, and the oxidation number of iron in Fe^{3+} is +3.
In the polyatomic ion 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 MnO_4^- is +7.

$$8\ \overset{+1}{H}{}^+(aq) + \overset{+7}{Mn}\overset{-2}{O}{}_4^-(aq) + \overset{+2}{Fe}{}^{2+}(aq) \rightarrow \overset{+2}{Mn}{}^{2+}(aq) + \overset{+3}{Fe}{}^{3+}(aq) + 4\ \overset{+1}{H}{}_2\overset{-2}{O}(l)$$

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 $Fe^{2+}(aq)$ is oxidized, and the manganese ion, Mn^{7+}, in $MnO_4^-(aq)$ is reduced.
(c) Solution:
Step 1: Write a balanced chemical equation for the reaction.
$Cu(s) + 2\ AgNO_3(aq) \rightarrow 2\ Ag(s) + Cu(NO_3)_2(aq)$
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 to the charge on the ion, the oxidation number of Cu^{2+} in $Cu(NO_3)_2(aq)$ is +2, and the oxidation number of Ag^+ in $AgNO_3(aq)$ is +1.
In the polyatomic nitrate ion, 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 NO_3^- is +5.
Step 3: Determine how the oxidation number on each element changes, and identify these changes as either oxidation or reduction.

$$\overset{0}{Cu}(s) + 2\ \overset{+1}{Ag}\overset{+5}{N}\overset{-2}{O}{}_3(aq) \rightarrow 2\ \overset{0}{Ag}(s) + \overset{+2}{Cu}(\overset{+5}{N}\overset{-2}{O}{}_3)_2(aq)$$

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 Cu(s) is oxidized and the silver ion, Ag^+, in $AgNO_3(aq)$ 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.

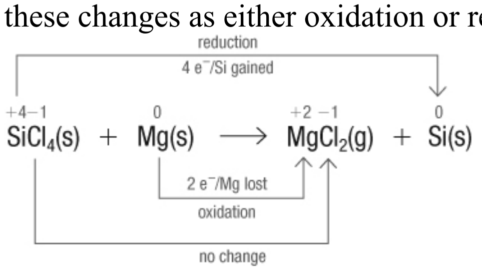
6. (a) Solution:
Step 1: Assign oxidation numbers to each atom and ion in the chemical equation.
The oxidation number of hydrogen in its compounds is +1.
Since the oxidation number of a monatomic ion is the same as its charge, the oxidation number of chlorine in HCl and in NH_4Cl is –1.
Since NH_3 is 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_4Cl 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_4Cl, the oxidation number of nitrogen must be –3.

$$\overset{+1}{H}\overset{-1}{Cl}(g) + 2\,\overset{-3}{N}\overset{+1}{H_3}(g) \rightarrow \overset{-3}{N}\overset{+1}{H_4}\overset{-1}{Cl}(s)$$

Step 2: Determine how the oxidation number on each element changes, and identify these changes as either oxidation or reduction.
There is no change in the oxidation numbers of the elements.
Statement: Since there is no change in oxidation of the elements, the reaction is not a redox reaction.
(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.

$$\overset{+4}{Si}\overset{-1}{Cl_4}(s) + 2\,\overset{0}{Mg}(s) \rightarrow 2\,\overset{+2}{Mg}\overset{-1}{Cl_2}(s) + \overset{0}{Si}(s)$$

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

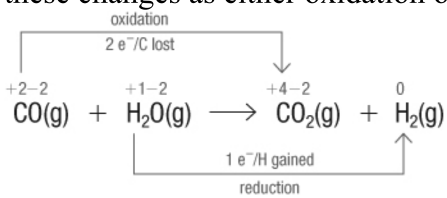


Statement: The oxidation number of magnesium increases from 0 to +2, so magnesium is oxidized. The oxidation number of silicon decreases from +4 to 0, so silicon is reduced. The oxidation number of chlorine does not change. Since Si^{4+} gains electrons, the oxidizing agent is $SiCl_2(s)$. Since magnesium loses electrons, the reducing agent is Mg(s).

(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 H_2 is 0.
The oxidation number of hydrogen in its compounds is +1, so the oxidation number of hydrogen in $H_2O(g)$ is +1. The oxidation number of oxygen in its compounds is –2, so the oxidation number of oxygen in H_2O 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 CO_2 is +4.

$$\overset{+2}{C}\overset{-2}{O}(g) + \overset{+1}{H_2}\overset{-2}{O}(g) \rightarrow \overset{+4}{C}\overset{-2}{O_2}(g) + \overset{0}{H_2}(g)$$

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 $H_2O(g)$. 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 O_2: Since the oxidation number of an atom in an element is 0, the oxidation number of oxygen in O_2 is 0.
H and O in H_2O: The oxidation number of hydrogen in its compounds is +1, so the oxidation number of hydrogen in H_2O is +1. The oxidation number of oxygen in its compounds is –2, so the oxidation number of oxygen in H_2O is –2.
P in PH_3: Since PH_3 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 PH_3, the phosphorus atom must have an oxidation number of –3 to balance the +3 of the hydrogen atoms.
O in P_4O_{10}: Since P_4O_{10} 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 P_4O_{10}, 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 $O_2(g)$ 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_2SO_4: The oxidation number of hydrogen in its compounds is +1, so the oxidation number of hydrogen in H_2SO_4 and H_2O is +1. The oxidation number of oxygen in its compounds is –2, so the oxidation number of oxygen in $H_2O(g)$ is –2.
S in H_2SO_4: Since there are 2 hydrogen atoms in H_2SO_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; oxygem: –2; hydrogen: +1; sulfur: +6
Products: lead: +2; sulfur: +6; oxygen: –2; hydrogen: +1
8. (a) From the previous question, for the reaction
$4\ PH_3(g) + 8\ O_2(g) \rightarrow P_4O_{10}(s) + 6\ H_2O(l)$,
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: $P^{3-}(g) \rightarrow P^{5+}(s) + 8\ e^-$
Reduction: $O_2(g) + 4\ e^- \rightarrow 2\ O^{2-}(s)$
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
$2\ KClO_3(s) \rightarrow 2\ KCl(s) + 3\ O_2(g)$,
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: $2\ O^{2-}(g) \rightarrow O_2(g) + 4\ e^-$
Reduction: $Cl^{5+}(s) + 6\ e^- \rightarrow Cl^-(s)$
Since the chlorine atom in $KClO_3$ loses electrons, the oxidizing agent is $KClO_3(s)$. Since oxygen ion, O^{2-}, in $KClO_3$ gains electrons, the reducing agent is also $KClO_3(s)$.
(c) From the previous question, for the reaction
$Pb(s) + PbO_2(s) + 2\ H_2SO_4(aq) \rightarrow 2\ PbSO_4(s) + 2\ H_2O(g)$,
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
Since the oxidation number of lead increases from 0 to +2, it is oxidized.
Since the oxidation number of lead in PbO_2 decreases from +4 to +2, it is reduced.
The reaction can be represented by the following oxidation–reduction half-reactions:
Oxidation: $Pb(s) \rightarrow Pb^{2+}(s) + 2\ e^-$
Reduction: $Pb^{4+}(s) + 2\ e^- \rightarrow Pb^{2+}(s)$
Since the lead ion, Pb^{4+}, gains electrons, the oxidizing agent is $PbO_2(s)$. 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 CO_2 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 CO_2, the total contribution of the oxygen atoms is –4. Therefore, the carbon atom in CO_2 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 $C_6H_{12}O_6$ 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 $C_6H_{12}O_6$, 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 CO_2 is +4, and the oxidation number of carbon in $C_6H_{12}O_6$ is 0.

(b) For the photosynthesis reaction
$6\ CO_2(g) + 6\ H_2O(l) \rightarrow C_6H_{12}O_6(s) + 6\ O_2(g)$,
the oxidation numbers of the elements are:
Reactants: carbon: +4; hydrogen: +1; oxygen: –2
Products: carbon: 0; hydrogen: +1; oxygen: –2; oxygen in O_2: 0
Since the oxidation number of oxygen increases from –2 to 0, it is oxidized.
Since the oxidation number of carbon decreases from +4 to 0, it is reduced.
The reaction can be represented by the following oxidation–reduction half-reactions:
Oxidation: $2\ O^{2-}(g) \rightarrow O_2(g) + 4\ e^-$
Reduction: $C^{4+}(g) + 4\ e^- \rightarrow C^0(s)$
Since the carbon ion, C^{4+}, gains electrons, the oxidizing agent is $CO_2(g)$. Since oxygen ion, O^{2-}, loses electrons and 24 electrons have to come from the reducing agent, the reducing agent is also $CO_2(g)$.

10. Solution: Use the zero-sum rule to assign oxidation numbers to the carbon in each compound: The sum of the oxidation numbers of all atoms in each of the electrically neutral compounds CO_2, CH_2O_2, CH_2O, CH_4O, and CH_4 must be 0.

CO_2: The oxidation number of oxygen in its compounds is –2. Since there are 2 oxygen atoms in CO_2, the total contribution of the oxygen atoms is –4. Therefore, the carbon atom in CO_2 must have an oxidation number of +4 to give a sum of 0.

CH_2O_2: The oxidation number of hydrogen in its compounds is +1 and the oxidation number of oxygen in its compounds is –2. Since there are 2 hydrogen atoms in CH_2O_2, the total contribution of the hydrogen atoms is +2. Since there are 2 oxygen atoms in CH_2O_2, the total contribution of the oxygen atoms is –4. Therefore, the carbon atom in CH_2O_2 must have an oxidation number of +2 to give a sum of 0.

CH_2O: The oxidation number of hydrogen in its compounds is +1 and the oxidation number of oxygen in its compounds is –2. Since there are 2 hydrogen atoms in CH_2O, the total contribution of the hydrogen atoms is +2. Since there is 1 oxygen atom in CH_2O, the contribution of the oxygen atom is –2. Therefore, the carbon atom in CH_2O must have an oxidation number of 0 to give a sum of 0.

CH_4O: The oxidation number of hydrogen in its compounds is +1 and the oxidation number of oxygen in its compounds is –2. Since there are 4 hydrogen atoms in CH_4O, the total contribution of the hydrogen atoms is +4. Since there is 1 oxygen atom in CH_4O, the contribution of the oxygen atom is –2. Therefore, the carbon atom in CH_4O must have an oxidation number of –2 to give a sum of 0.

CH_4: The oxidation number of hydrogen in its compounds is +1. Since there are 4 hydrogen atoms in CH_4, the total contribution of the hydrogen atoms is +4. Therefore, the carbon atom in CH_4 must have an oxidation number of –4 to give a sum of 0.

Statement: The oxidation number of carbon is +4 in CO_2, +2 in CH_2O_2, 0 in CH_2O, –2 in CH_4O, and –4 in CH_4.