Identifying Redox Reactions
(Student textbook page 611)

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

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

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

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

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

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

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

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

What Is Required?

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

What Is Given?

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

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

Check Your Solution

The rules were followed and the answer is logical.

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

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

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

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

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

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

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

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

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

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

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

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

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

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

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

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

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

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

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

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

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

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

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

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

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

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

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

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

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

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

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

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

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

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

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

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

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

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

Check Your Solution

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

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

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

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

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

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

Check Your Solution
Following the rules led to logical answers.