Section 9.2: Balancing Redox Reaction Equations

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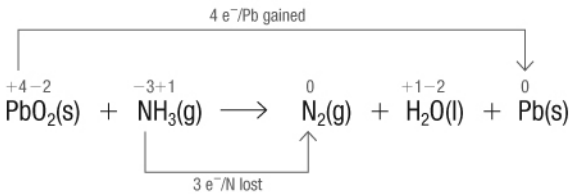
1. Solution:

Step 1: Write the unbalanced equation. Determine the oxidation numbers for each element in the equation, and identify the elements for which the oxidation numbers change.



The oxidation number of lead changes from +4 to 0, so lead gains 4 electrons.
The oxidation number of nitrogen changes from –3 to 0, so nitrogen loses 3 electrons.

Step 2: To balance electrons, adjust the coefficients by multiplying electrons gained by 3 and electrons lost by 4.

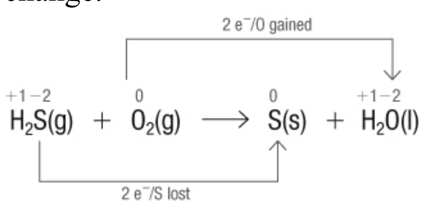
3 PbO_2(s) + 4 NH_3(g) → 2 N_2(g) + H_2O(l) + 3 Pb(s)

Step 3: Balance the remaining entities, oxygen and hydrogen, by inspection. Check the answer. The balanced equation is

3 PbO_2(s) + 4 NH_3(g) → 2 N_2(g) + 6 H_2O(l) + 3 Pb(s)

2. Solution:

Step 1: Use the unbalanced equation to determine the oxidation numbers for each element in the equation, and identify the elements for which the oxidation numbers change.



The oxidation number of sulfur changes from –2 to 0, so sulfur loses 2 electrons.
The oxidation number of oxygen changes from 0 to –2, so each oxygen molecule gains 4 electrons.

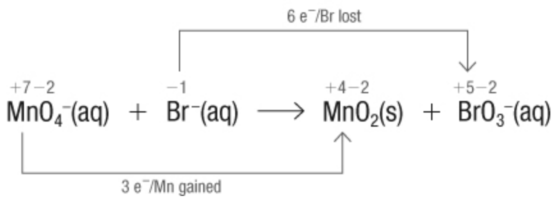
Step 2: To balance electrons, adjust the coefficients by multiplying electrons lost by 2.

2 H_2S(g) + O_2(g) → 2 S(s) + H_2O(l)

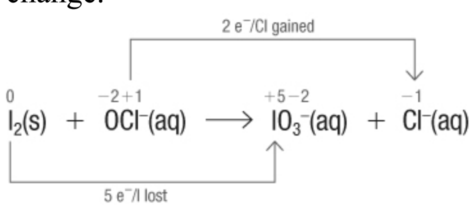
Step 3: Balance the remaining entities, oxygen and hydrogen, by inspection. Check the answer. The balanced equation is:

2 H_2S(g) + O_2(g) → 2 S(s) + 2 H_2O(l)

3. (a) Solution:
Step 1: Use the unbalanced equation to determine the oxidation numbers for each element in the equation, and identify the elements for which the oxidation numbers change.

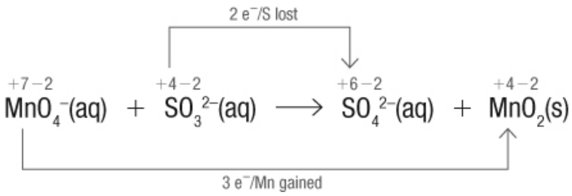


The oxidation number of manganese changes from +7 to +4, so manganese gains 3 electrons.
The oxidation number of bromine changes from –1 to +5, so bromine loses 6 electrons.
Step 2: To balance electrons, adjust the coefficients by multiplying electrons gained by 2.
$2\ MnO_4^-(aq) + Br^-(aq) \rightarrow 2\ MnO_2(s) + BrO_3^-(aq)$
Step 3: Balance the rest of the equation by inspection.
Balance oxygen by adding water.
$2\ MnO_4^-(aq) + Br^-(aq) \rightarrow 2\ MnO_2(s) + BrO_3^-(aq) + H_2O(l)$
Balance hydrogen by adding hydrogen ions.
$2\ MnO_4^-(aq) + Br^-(aq) + 2\ H^+(aq) \rightarrow 2\ MnO_2(s) + BrO_3^-(aq) + H_2O(l)$
Step 4: Check the answer. The balanced equation is:
$2\ MnO_4^-(aq) + Br^-(aq) + 2\ H^+(aq) \rightarrow 2\ MnO_2(s) + BrO_3^-(aq) + H_2O(l)$
(b) Solution:
Step 1: Write the unbalanced equation. Determine the oxidation numbers for each element in the equation, and identify the elements for which the oxidation numbers change.

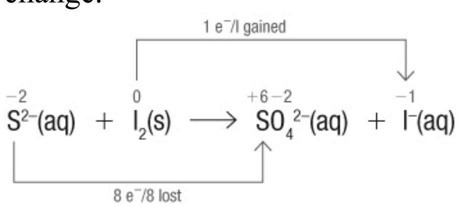


The oxidation number of iodine changes from 0 to +5, so each iodine molecule loses 10 electrons.
The oxidation number of chlorine changes from +1 to –1, so chlorine gains 2 electrons.
Step 2: To balance electrons, adjust the coefficients by multiplying electrons gained by 5.
Since each iodine molecule has 2 atoms, the coefficient for $IO_3^-(aq)$ must be 2.
$I_2(s) + 5\ OCl^-(aq) \rightarrow 2\ IO_3^-(aq) + 5\ Cl^-(aq)$
Step 3: Balance the rest of the equation by inspection.
Balance oxygen by adding water.
$I_2(s) + 5\ OCl^-(aq) + H_2O(l) \rightarrow 2\ IO_3^-(aq) + 5\ Cl^-(aq)$
Balance hydrogen by adding hydrogen ions.
$I_2(s) + 5\ OCl^-(aq) + H_2O(l) \rightarrow 2\ IO_3^-(aq) + 5\ Cl^-(aq) + 2\ H^+(aq)$
Step 4: Check the answer. The balanced equation is
$I_2(s) + 5\ OCl^-(aq) + H_2O(l) \rightarrow 2\ IO_3^-(aq) + 5\ Cl^-(aq) + 2\ H^+(aq)$

4. (a) Solution:
Step 1: Use the unbalanced equation to determine the oxidation numbers for each element in the equation, and identify the elements for which the oxidation numbers change.

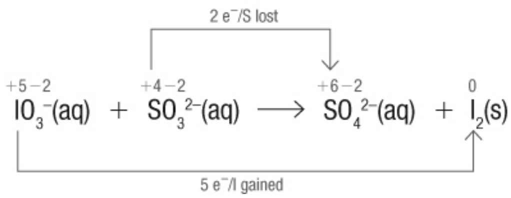


The oxidation number of manganese changes from +7 to +4, so manganese gains 3 electrons.
The oxidation number of sulfur changes from +4 to +6, so sulfur loses 2 electrons.
Step 2: To balance electrons, adjust the coefficients by multiplying electrons gained by 2 and electrons lost by 3.
$2\ MnO_4^-(aq) + 3\ SO_3^{2-}(aq) \rightarrow 3\ SO_4^{2-}(aq) + 2\ MnO_2(s)$
Step 3: Balance the rest of the equation by inspection. Balance oxygen by adding water.
$2\ MnO_4^-(aq) + 3\ SO_3^{2-}(aq) \rightarrow 3\ SO_4^{2-}(aq) + 2\ MnO_2(s) + H_2O(l)$
Step 4: Balance hydrogen by adding $H^+(aq)$ and $OH^-(aq)$.
Add $H^+(aq)$ to the reactant side of the equation.
$2\ MnO_4^-(aq) + 3\ SO_3^{2-}(aq) + 2\ H^+(aq) \rightarrow 3\ SO_4^{2-}(aq) + 2\ MnO_2(s) + H_2O(l)$
Since the reaction takes place in basic solution, eliminate $H^+(aq)$ by adding an equal number of $OH^-(aq)$ to both sides of the equation.
$2\ MnO_4^-(aq) + 3\ SO_3^{2-}(aq) + 2\ H^+(aq) + 2\ OH^-(aq) \rightarrow$
$3\ SO_4^{2-}(aq) + 2\ MnO_2(s) + H_2O(l) + 2\ OH^-(aq)$
Subtract 1 water molecule from each side to eliminate redundant water molecules.
$2\ MnO_4^-(aq) + 3\ SO_3^{2-}(aq) + H_2O(l) \rightarrow 3\ SO_4^{2-}(aq) + 2\ MnO_2(s) + 2\ OH^-(aq)$
Step 5: Check the answer. The balanced equation is:
$2\ MnO_4^-(aq) + 3\ SO_3^{2-}(aq) + H_2O(l) \rightarrow 3\ SO_4^{2-}(aq) + 2\ MnO_2(s) + 2\ OH^-(aq)$
(b) Solution:
Step 1: Use the unbalanced equation to determine the oxidation numbers for each element in the equation, and identify the elements for which the oxidation numbers change.

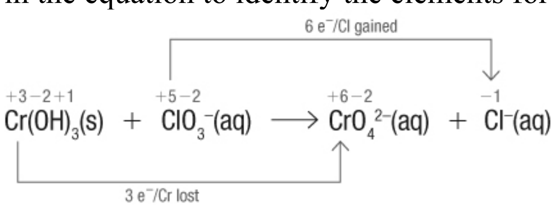


The oxidation number of sulfur changes from –2 to +6, so sulfur loses 8 electrons.
The oxidation number of iodine changes from 0 to –1, so each iodine molecule gains 2 electrons.
Step 2: To balance electrons, adjust the coefficients by multiplying electrons gained by 4. Since each iodine molecule has 2 atoms, the coefficient for $I^-(aq)$ must be 8.
$S^{2-}(aq) + 4\ I_2(s) \rightarrow SO_4^{2-}(aq) + 8\ I^-(aq)$
Step 3: Balance the rest of the equation by inspection. Balance oxygen by adding water.
$S^{2-}(aq) + 4\ I_2(s) + 4\ H_2O(l) \rightarrow SO_4^{2-}(aq) + 8\ I^-(aq)$

Step 4: Balance hydrogen by adding $H^+(aq)$ and $OH^-(aq)$.
Add $H^+(aq)$ to the product side of the equation.
$S^{2-}(aq) + 4\ I_2(s) + 4\ H_2O(l) \rightarrow SO_4^{2-}(aq) + 8\ I^-(aq) + 8\ H^+(aq)$
Since the reaction takes place in basic solution, eliminate $H^+(aq)$ by adding an equal number of $OH^-(aq)$ to both sides of the equation.
$S^{2-}(aq) + 4\ I_2(s) + 4\ H_2O(l) + 8\ OH^-(aq) \rightarrow SO_4^{2-}(aq) + 8\ I^-(aq) + 8\ H^+(aq) + 8\ OH^-(aq)$
Subtract 4 water molecules from each side to eliminate redundant water molecules.
$S^{2-}(aq) + 4\ I_2(s) + 8\ OH^-(aq) \rightarrow SO_4^{2-}(aq) + 8\ I^-(aq) + 4\ H_2O(l)$
Step 5: Check the answer. The balanced equation is
$S^{2-}(aq) + 4\ I_2(s) + 8\ OH^-(aq) \rightarrow SO_4^{2-}(aq) + 8\ I^-(aq) + 4\ H_2O(l)$
5. (a) Use the unbalanced equation to determine the oxidation numbers for each element in the equation to identify the elements for which the oxidation numbers change.



The oxidation number of iodine changes from +5 to 0, so iodine gains 5 electrons.
The oxidation number of sulfur changes from +4 to +6, so sulfur loses 2 electrons.
(b) To balance electrons, adjust the coefficients by multiplying electrons gained by 2 and electrons lost by 5.
$2\ IO_3^-(aq) + 5\ SO_3^-(aq) \rightarrow 5\ SO_4^{2-}(aq) + I_2(s)$
Balance the rest of the equation by inspection. Balance oxygen by adding water.
$2\ IO_3^-(aq) + 5\ SO_3^-(aq) \rightarrow 5\ SO_4^{2-}(aq) + I_2(s) + H_2O(l)$
Balance hydrogen by adding $H^+(aq)$.
$2\ IO_3^-(aq) + 5\ SO_3^{2-}(aq) + 2\ H^+(aq) \rightarrow 5\ SO_4^{2-}(aq) + I_2(s) + H_2O(l)$
Therefore, the balanced equation for the reaction in an acidic solution is
$2\ IO_3^-(aq) + 5\ SO_3^{2-}(aq) + 2\ H^+(aq) \rightarrow 5\ SO_4^{2-}(aq) + I_2(s) + H_2O(l)$
6. (a) Use the unbalanced equation to determine the oxidation numbers for each element in the equation to identify the elements for which the oxidation numbers change.



The oxidation number of chromium changes from +3 to +6, so chromium loses 3 electrons.
The oxidation number of chlorine changes from +5 to −1, so chlorine gains 6 electrons.
(b) To balance electrons, adjust the coefficients by multiplying electrons lost by 2.
$2\ Cr(OH)_3(s) + ClO_3^-(aq) \rightarrow 2\ CrO_4^{2-}(aq) + Cl^-(aq)$
Balance the rest of the equation by inspection. Balance oxygen by adding water.
$2\ Cr(OH)_3(s) + ClO_3^-(aq) \rightarrow 2\ CrO_4^{2-}(aq) + Cl^-(aq) + H_2O(l)$
Balance hydrogen by adding $H^+(aq)$.
$2\ Cr(OH)_3(s) + ClO_3^-(aq) \rightarrow 2\ CrO_4^{2-}(aq) + Cl^-(aq) + H_2O(l) + 4\ H^+(aq)$

Since the reaction takes place in a basic solution, eliminate $H^+(aq)$ by adding an equal number of $OH^-(aq)$ to both sides of the equation.

$2\ Cr(OH)_3(s) + ClO_3^-(aq) + 4\ OH^-(aq) \rightarrow$
$2\ CrO_4^{2-}(aq) + Cl^-(aq) + H_2O(l) + 4\ H^+(aq) + 4\ OH^-(aq)$

Therefore, the balanced equation for the reaction in a basic solution is:

$2\ Cr(OH)_3(s) + ClO_3^-(aq) + 4\ OH^-(aq) \rightarrow 2\ CrO_4^{2-}(aq) + Cl^-(aq) + 5\ H_2O(l)$

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1. (a) Solution:

Step 1: Assign oxidation numbers to all entities.

$\overset{0}{Zn}(s) + \overset{+1}{H^+}(aq) \rightarrow \overset{+2}{Zn^{2+}}(aq) + \overset{0}{H_2}(g)$

Step 2: Write unbalanced equations for oxidation and reduction half-reactions.

$\overset{0}{Zn}(s) \rightarrow \overset{+2}{Zn^{2+}}(aq)$ (oxidation)

$\overset{+1}{H^+}(aq) \rightarrow \overset{0}{H_2}(g)$ (reduction)

Step 3: Balance each half-reaction equation. Equations are already balanced.

$Zn(s) \rightarrow Zn^{2+}(aq)$
$2\ H^+(aq) \rightarrow H_2(g)$

Step 4: Use electrons to balance the charge in each half-reaction equation.

$Zn(s) \rightarrow Zn^{2+}(aq) + 2\ e^-$
$2\ H^+(aq) + 2\ e^- \rightarrow H_2(g)$

Step 5: The number of electrons transferred is equalized. So, add the half-reaction equations.

$Zn(s) + 2\ H^+(aq) \rightarrow Zn^{2+}(aq) + H_2(g)$

Step 6: Check the answer. The balanced equation is

$Zn(s) + 2\ H^+(aq) \rightarrow Zn^{2+}(aq) + H_2(g)$

(b) Solution:

Step 1: Assign oxidation numbers to all entities.

$\overset{+1}{H}\overset{+5}{N}\overset{-2}{O_3}(aq) + \overset{0}{Cu}(s) \rightarrow \overset{+4}{N}\overset{-2}{O_2}(g) + \overset{+2}{Cu^{2+}}(aq)$

Step 2: Write unbalanced equations for oxidation and reduction half-reactions.

$\overset{+1}{H}\overset{+5}{N}\overset{-2}{O_3}(aq) \rightarrow \overset{+4}{N}\overset{-2}{O_2}(g)$ (reduction)

$\overset{0}{Cu}(s) \rightarrow \overset{+2}{Cu^{2+}}(aq)$ (oxidation)

Step 3: Balance each half-reaction equation. Use $H_2O(l)$ to balance oxygen and $H^+(aq)$ to balance hydrogen.

$HNO_3(aq) + H^+(aq) \rightarrow NO_2(g) + H_2O(l)$
$Cu(s) \rightarrow Cu^{2+}(aq)$

Step 4: Use electrons to balance the charge in each half-reaction equation.

$HNO_3(aq) + H^+(aq) + e^- \rightarrow NO_2(g) + H_2O(l)$
$Cu(s) \rightarrow Cu^{2+}(aq) + 2\ e^-$

Step 5: Equalize the electron transfer in the two half-reaction equations. Multiply the reduction half-reaction by 2.

$2\ HNO_3(aq) + 2\ H^+(aq) + 2\ e^- \rightarrow 2\ NO_2(g) + 2\ H_2O(l)$

$Cu(s) \rightarrow Cu^{2+}(aq) + 2\ e^-$

Step 6: Add the half-reaction equations.

$2\ HNO_3(aq) + Cu(s) + 2\ H^+(aq) \rightarrow 2\ NO_2(g) + Cu^{2+}(aq) + 2\ H_2O(l)$

Step 7: Check the answer. The balanced equation is

$2\ HNO_3(aq) + Cu(s) + 2\ H^+(aq) \rightarrow 2\ NO_2(g) + Cu^{2+}(aq) + 2\ H_2O(l)$

2. (a) Solution:

Step 1: Assign oxidation numbers to all entities.

–2 +1 –2 +1   +7 –2   +4 –2   +6 –2

$CH_3OH(aq) + MnO_4^-(aq) \rightarrow CO_3^{2-}(aq) + MnO_4^{2-}(aq)$

Step 2: Write unbalanced equations for oxidation and reduction half-reactions.

–2 +1 –2 +1   +4 –2

$CH_3OH(aq) \rightarrow CO_3^{2-}(aq)$ (oxidation)

+7 –2   +6 –2

$MnO_4^-(aq) \rightarrow MnO_4^{2-}(aq)$ (reduction)

Step 3: Balance each half-reaction equation. Use $H_2O(l)$ to balance oxygen and $H^+(aq)$ to balance hydrogen.

$CH_3OH(aq) + 2\ H_2O(l) \rightarrow CO_3^{2-}(aq) + 8\ H^+(aq)$

$MnO_4^-(aq) \rightarrow MnO_4^{2-}(aq)$

Since the reaction takes place in basic solution, eliminate $H^+(aq)$ by adding an equal number of $OH^-(aq)$ to both sides of the equation.

$CH_3OH(aq) + 2\ H_2O(l) + 8\ OH^-(aq) \rightarrow CO_3^{2-}(aq) + 8\ H^+(aq) + 8\ OH^-(aq)$

$MnO_4^-(aq) \rightarrow MnO_4^{2-}(aq)$

Subtract 2 water molecules from each side to eliminate redundant water molecules.

$CH_3OH(aq) + 8\ OH^-(aq) \rightarrow CO_3^{2-}(aq) + 6\ H_2O(l)$

$MnO_4^-(aq) \rightarrow MnO_4^{2-}(aq)$

Step 4: Use electrons to balance the charge in each half-reaction equation.

$CH_3OH(aq) + 8\ OH^-(aq) \rightarrow CO_3^{2-}(aq) + 6\ H_2O(l) + 6\ e^-$

$MnO_4^-(aq) + e^- \rightarrow MnO_4^{2-}(aq)$

Statement: The balanced half-reaction equations for this reaction are

Oxidation: $CH_3OH(aq) + 8\ OH^-(aq) \rightarrow CO_3^{2-}(aq) + 6\ H_2O(l) + 6\ e^-$

Reduction: $MnO_4^-(aq) + e^- \rightarrow MnO_4^{2-}(aq)$

(b) Solution:

Step 1: Equalize the electron transfer in the two half-reaction equations. Multiply the reduction half-reaction by 6.

$CH_3OH(aq) + 8\ OH^-(aq) \rightarrow CO_3^{2-}(aq) + 6\ H_2O(l) + 6\ e^-$

$6\ MnO_4^-(aq) + 6\ e^- \rightarrow 6\ MnO_4^{2-}(aq)$

Step 2: Add the half-reaction equations to obtain the balanced equation.

$CH_3OH(aq) + 6\ MnO_4^-(aq) + 8\ OH^-(aq) \rightarrow CO_3^{2-}(aq) + 6\ MnO_4^{2-}(aq) + 6\ H_2O(l)$

Step 3: Check the answer. The balanced equation is

$CH_3OH(aq) + 6\ MnO_4^-(aq) + 8\ OH^-(aq) \rightarrow CO_3^{2-}(aq) + 6\ MnO_4^{2-}(aq) + 6\ H_2O(l)$

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1. (a) Solution:

Step 1: Write the unbalanced equation for the reaction.

$Mg(s) + HCl(aq) \rightarrow H_2(g) + MgCl_2(aq)$

Step 2: Write the unbalanced net ionic equation, without the spectator ion, $Cl^-(aq)$.

$Mg(s) + H^+(aq) \rightarrow H_2(g) + Mg^{2+}(aq)$

Step 3: Assign oxidation numbers to all entities.

$\overset{0}{Mg}(s) + \overset{+1}{H^+}(aq) \rightarrow \overset{0}{H_2}(g) + \overset{+2}{Mg^{2+}}(aq)$

Step 4: Write unbalanced equations for oxidation and reduction half-reactions.

$\overset{0}{Mg}(s) \rightarrow \overset{+2}{Mg^{2+}}(aq)$ (oxidation)

$\overset{+1}{H^+}(aq) \rightarrow \overset{0}{H_2}(g)$ (reduction)

Step 5: Balance each half-reaction equation.

$Mg(s) \rightarrow Mg^{2+}(aq)$

$2\ H^+(aq) \rightarrow H_2(g)$

Step 6: Use electrons to balance the charge in each half-reaction equation.

$Mg(s) \rightarrow Mg^{2+}(aq) + 2\ e^-$

$2\ H^+(aq) + 2\ e^- \rightarrow H_2(g)$

Statement: The balanced half-reaction equations for this reaction are

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

Reduction: $2\ H^+(aq) + 2\ e^- \rightarrow H_2(g)$

(b) Add the two half-reactions to obtain the balanced equation. Include the spectator ion, $Cl^-(aq)$.

The balanced equation for the reaction is

$Mg(s) + 2\ HCl(aq) \rightarrow H_2(g) + MgCl_2(aq)$

(c) The number of electrons transferred in the balanced equation is 2.

2. (a) Solution:

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

$Ag^+(aq) + NO_3^-(aq) + Cu(s) \rightarrow Cu^{2+}(aq) + 2\ NO_3^-(aq) + Ag(s)$

Step 2: Write the unbalanced net ionic equation, without the spectator ion, $NO_3^-(aq)$.

$Ag^+(aq) + Cu(s) \rightarrow Cu^{2+}(aq) + Ag(s)$

Step 3: Assign oxidation numbers to all entities.

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

Step 4: Write unbalanced equations for oxidation and reduction half-reactions.

$\overset{0}{Cu}(s) \rightarrow \overset{+2}{Cu^{2+}}(aq)$ (oxidation)

$\overset{+1}{Ag^+}(aq) \rightarrow \overset{0}{Ag}(s)$ (reduction)

Step 5: The half-reaction equations are each balanced. Use electrons to balance the charge in each half-reaction equation.

$Cu(s) \rightarrow Cu^{2+}(aq) + 2\ e^-$

$Ag^+(aq) + e^- \rightarrow Ag(s)$

Statement: The balanced half-reaction equations for this reaction are
Oxidation: $Cu(s) \rightarrow Cu^{2+}(aq) + 2\ e^{-}$
Reduction: $Ag^{+}(aq) + e^{-} \rightarrow Ag(s)$
(b) Solution:
Step 1: Assign oxidation numbers to all entities.
$$\overset{+3}{Cr}{}^{3+}(aq) + \overset{0}{Cl_2}(g) \rightarrow \overset{+6}{Cr_2}\overset{-2}{O_7}{}^{2-}(aq) + \overset{-1}{Cl}{}^{-}(aq)$$
Step 2: Write unbalanced equations for oxidation and reduction half-reactions.
$$\overset{+3}{Cr}{}^{3+}(aq) \rightarrow \overset{+6}{Cr_2}\overset{-2}{O_7}{}^{2-}(aq) \quad \text{(oxidation)}$$
$$\overset{0}{Cl_2}(g) \rightarrow \overset{-1}{Cl}{}^{-}(aq) \quad \text{(reduction)}$$
Step 3: Balance each half-reaction equation. Use $H_2O(l)$ to balance oxygen and $H^{+}(aq)$ to balance hydrogen.
$2\ Cr^{3+}(aq) + 7\ H_2O(l) \rightarrow Cr_2O_7{}^{2-}(aq) + 14\ H^{+}(aq)$
$Cl_2(g) \rightarrow 2\ Cl^{-}(aq)$
Step 4: Use electrons to balance the charge in each half-reaction equation.
$2\ Cr^{3+}(aq) + 7\ H_2O(l) \rightarrow Cr_2O_7{}^{2-}(aq) + 14\ H^{+}(aq) + 6\ e^{-}$
$Cl_2(g) + 2\ e^{-} \rightarrow 2\ Cl^{-}(aq)$
Statement: The balanced half-reaction equations are:
Oxidation: $2\ Cr^{3+}(aq) + 7\ H_2O(l) \rightarrow Cr_2O_7{}^{2-}(aq) + 14\ H^{+}(aq) + 6\ e^{-}$
Reduction: $Cl_2(g) + 2\ e^{-} \rightarrow 2\ Cl^{-}(aq)$
(c) Solution:
Step 1: Write the unbalanced equation with its ionic compounds dissociated.
$2\ H^{+}(aq) + SO_4{}^{2-}(aq) + Ca(s) \rightarrow Ca^{2+}(aq) + SO_4{}^{2-}(aq) + H_2(g)$
Step 2: Write the unbalanced net ionic equation, without the spectator ion, $SO_4{}^{2-}(aq)$.
$2\ H^{+}(aq) + Ca(s) \rightarrow Ca^{2+}(aq) + H_2(g)$
Step 3: Assign oxidation numbers to all entities.
$$2\ \overset{+1}{H}{}^{+}(aq) + \overset{0}{Ca}(s) \rightarrow \overset{+2}{Ca}{}^{2+}(aq) + \overset{0}{H_2}(g)$$
Step 4: Write unbalanced equations for oxidation and reduction half-reactions.
$$\overset{0}{Ca}(s) \rightarrow \overset{+2}{Ca}{}^{2+}(aq) \quad \text{(oxidation)}$$
$$2\ \overset{+1}{H}{}^{+}(aq) \rightarrow \overset{0}{H_2}(g) \quad \text{(reduction)}$$
Step 5: The half-reaction equations are each balanced. Use electrons to balance the charge in each half-reaction equation.
$Ca(s) \rightarrow Ca^{2+}(aq) + 2\ e^{-}$
$2\ H^{+}(aq) + 2\ e^{-} \rightarrow H_2(g)$
Statement: The balanced half-reaction equations are
Oxidation: $Ca(s) \rightarrow Ca^{2+}(aq) + 2\ e^{-}$
Reduction: $2\ H^{+}(aq) + 2\ e^{-} \rightarrow H_2(g)$

3. (a) Solution:
Step 1: Assign oxidation numbers to all entities.

$\overset{+5}{Cl}\overset{-2}{O}_3^-(aq) + \overset{0}{I}_2(aq) \rightarrow \overset{-1}{Cl}^-(aq) + \overset{+5}{I}\overset{-2}{O}_3^-(aq)$

The oxidation number of chlorine changes from +5 to –1, so chlorine gains 6 electrons.
The oxidation number of iodine changes from 0 to +5, so iodine loses 5 electrons.
Step 2: To balance electrons, adjust the coefficients by multiplying electrons gained by 5 and electrons lost by 6.
Since each iodine molecule has 2 atoms, the coefficient for $I_2(aq)$ would be 3.
$5\ ClO_3^-(aq) + 3\ I_2(aq) \rightarrow 5\ Cl^-(aq) + 6\ IO_3^-(aq)$
Step 3: Balance the rest of the equation by inspection.
Balance oxygen by adding water.
$5\ ClO_3^-(aq) + 3\ I_2(aq) + 3\ H_2O(l) \rightarrow 5\ Cl^-(aq) + 6\ IO_3^-(aq)$
Balance hydrogen by adding hydrogen ions.
$5\ ClO_3^-(aq) + 3\ I_2(aq) + 3\ H_2O(l) \rightarrow 5\ Cl^-(aq) + 6\ IO_3^-(aq) + 6\ H^+(aq)$
Statement: The balanced equation is:
$5\ ClO_3^-(aq) + 3\ I_2(aq) + 3\ H_2O(l) \rightarrow 5\ Cl^-(aq) + 6\ IO_3^-(aq) + 6\ H^+(aq)$
(b) Solution:
Step 1: Assign oxidation numbers to all entities.

$\overset{+6}{Cr}_2\overset{-2}{O}_7^{2-}(aq) + \overset{-1}{Cl}^-(aq) \rightarrow \overset{+3}{Cr}^{3+}(aq) + \overset{0}{Cl}_2(aq)$

The oxidation number of chromium changes from +6 to +3, so each chromium atom gains 3 electrons. Therefore, each $Cr_2O_7^{2-}(aq)$ gains 6 electrons.
The oxidation number of chlorine changes from –1 to 0, so each chlorine atom loses 1 electron.
Step 2: To balance electrons, adjust the coefficients by multiplying electrons lost by 6.
Since each chlorine molecule has 2 atoms, the coefficient for $Cl_2(aq)$ is 3.
$Cr_2O_7^{2-}(aq) + 6\ Cl^-(aq) \rightarrow Cr^{3+}(aq) + 3\ Cl_2(aq)$
Step 3: Balance the rest of the equation by inspection.
Balance chromium.
$Cr_2O_7^{2-}(aq) + 6\ Cl^-(aq) \rightarrow 2\ Cr^{3+}(aq) + 3\ Cl_2(aq)$
Balance oxygen by adding water.
$Cr_2O_7^{2-}(aq) + 6\ Cl^-(aq) \rightarrow 2\ Cr^{3+}(aq) + 3\ Cl_2(aq) + 7\ H_2O(l)$
Balance hydrogen by adding hydrogen ions.
$Cr_2O_7^{2-}(aq) + 6\ Cl^-(aq) + 14\ H^+(aq) \rightarrow 2\ Cr^{3+}(aq) + 3\ Cl_2(aq) + 7\ H_2O(l)$
Statement: The balanced equation is
$Cr_2O_7^{2-}(aq) + 6\ Cl^-(aq) + 14\ H^+(aq) \rightarrow 2\ Cr^{3+}(aq) + 3\ Cl_2(aq) + 7\ H_2O(l)$

4. (a) Solution:
Step 1: Assign oxidation numbers to all entities.

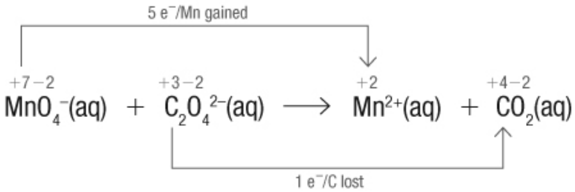
$\overset{+2\ -2\ +1}{Pb(OH)_4^{2-}}(aq) + \overset{+1\ -2}{ClO^-}(aq) \rightarrow \overset{+4\ -2}{PbO_2}(s) + \overset{-1}{Cl^-}(aq)$

The oxidation number of lead changes from +2 to +4, so lead loses 2 electrons.
The oxidation number of chlorine changes from +1 to –1, so chlorine gains 2 electrons.
Step 2: Electrons are balanced. Balance the rest of the equation by inspection.
Balance oxygen by adding water.
$Pb(OH)_4^{2-}(aq) + ClO^-(aq) \rightarrow PbO_2(s) + Cl^-(aq) + 3\ H_2O(l)$
Balance hydrogen by adding hydrogen ions.
$Pb(OH)_4^{2-}(aq) + ClO^-(aq) + 2\ H^+(aq) \rightarrow PbO_2(s) + Cl^-(aq) + 3\ H_2O(l)$
Since the reaction takes place in basic solution, eliminate $H^+(aq)$ by adding an equal number of $OH^-(aq)$ to both sides of the equation.
$Pb(OH)_4^{2-}(aq) + ClO^-(aq) + 2\ H^+(aq) + 2\ OH^-(aq) \rightarrow$
$PbO_2(s) + Cl^-(aq) + 3\ H_2O(l) + 2\ OH^-(aq)$
Subtract 2 water molecules from each side to eliminate redundant water molecules.
$Pb(OH)_4^{2-}(aq) + ClO^-(aq) \rightarrow PbO_2(s) + Cl^-(aq) + H_2O(l) + 2\ OH^-(aq)$
Statement: The balanced equation for the reaction in a basic solution is
$Pb(OH)_4^{2-}(aq) + ClO^-(aq) \rightarrow PbO_2(s) + Cl^-(aq) + H_2O(l) + 2\ OH^-(aq)$
(b) Solution:
Step 1: Assign oxidation numbers to all entities.

$\overset{+3\ -2}{NO_2^-}(aq) + \overset{0}{Al}(s) \rightarrow \overset{-3\ +1}{NH_3}(aq) + \overset{+3\ -2\ +1}{Al(OH)_4^-}(aq)$

The oxidation number of nitrogen changes from +3 to –3, so nitrogen gains 6 electrons.
The oxidation number of aluminum changes from 0 to +3, so aluminum loses 3 electrons.
Step 2: To balance electrons, adjust the coefficients by multiplying electrons lost by 2.
$NO_2^-(aq) + 2\ Al(s) \rightarrow NH_3(aq) + 2\ Al(OH)_4^-(aq)$
Step 3: Balance the rest of the equation by inspection.
Balance oxygen by adding water.
$NO_2^-(aq) + 2\ Al(s) + 6\ H_2O(l) \rightarrow NH_3(aq) + 2\ Al(OH)_4^-(aq)$
Balance hydrogen by adding hydrogen ions.
$NO_2^-(aq) + 2\ Al(s) + 6\ H_2O(l) \rightarrow NH_3(aq) + 2\ Al(OH)_4^-(aq) + H^+(aq)$
Since the reaction takes place in a basic solution, eliminate $H^+(aq)$ by adding an equal number of $OH^-(aq)$ to both sides of the equation.
$NO_2^-(aq) + 2\ Al(s) + 6\ H_2O(l) + OH^-(aq) \rightarrow$
$NH_3(aq) + 2\ Al(OH)_4^-(aq) + H^+(aq) + OH^-(aq)$
Subtract 1 water molecule from each side to eliminate redundant water molecules.
$NO_2^-(aq) + 2\ Al(s) + 5\ H_2O(l) + OH^-(aq) \rightarrow NH_3(aq) + 2\ Al(OH)_4^-(aq)$
Statement: The balanced equation for the reaction in a basic solution is
$NO_2^-(aq) + 2\ Al(s) + 5\ H_2O(l) + OH^-(aq) \rightarrow NH_3(aq) + 2\ Al(OH)_4^-(aq)$

5. (a) Determine the oxidation numbers for each element in the equation and identify the elements for which the oxidation numbers change.



The oxidation number of manganese changes from +7 to +2, so 5 electrons are transferred for each MnO_4^-(aq) ion.

(b) The oxidation number of carbon changes from +3 to +2, and there are two carbon atoms in each $C_2O_4^{2-}$(aq) ion, so 2 electrons are transferred for each $C_2O_4^{2-}$(aq) ion.

(c) Solution:

Step 1: Manganese gains 5 electrons, and each carbon atom loses 1 electron. To balance electrons, adjust the coefficients by multiplying electrons gained by 2 and electrons lost by 5.

$2\ MnO_4^-(aq) + 5\ C_2O_4^{2-}(aq) \rightarrow 2\ Mn^{2+}(aq) + 10\ CO_2(aq)$

Step 2: Balance the remaining entities by inspection.

Balance oxygen by adding water.

$2\ MnO_4^-(aq) + 5\ C_2O_4^{2-}(aq) \rightarrow 2\ Mn^{2+}(aq) + 10\ CO_2(aq) + 8\ H_2O(l)$

Balance hydrogen by adding hydrogen ions.

$2\ MnO_4^-(aq) + 5\ C_2O_4^{2-}(aq) + 16\ H^+(aq) \rightarrow 2\ Mn^{2+}(aq) + 10\ CO_2(aq) + 8\ H_2O(l)$

Statement: The balanced equation is

$2\ MnO_4^-(aq) + 5\ C_2O_4^{2-}(aq) + 16\ H^+(aq) \rightarrow 2\ Mn^{2+}(aq) + 10\ CO_2(aq) + 8\ H_2O(l)$

6. (a) Solution:

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

$Ag(s) + CN^-(aq) + O_2(g) \rightarrow Ag^+(aq) + 2\ CN^-(aq)$

Step 2: Write the unbalanced net ionic equation, without the spectator ion, CN^-(aq).

$Ag(s) + O_2(g) \rightarrow Ag^+(aq)$

Step 3: Write unbalanced equations for oxidation and reduction half-reactions.

The oxidation half-reaction equation is:

$Ag(s) \rightarrow Ag^+(aq)$

The reduction half-reaction involves oxygen.

Balance oxygen by adding water.

$O_2(g) \rightarrow 2\ H_2O(l)$

Step 4: Balance each half-reaction equation. The oxidation half-reaction is balanced.

For the reduction half-reaction, balance hydrogen by adding hydrogen ions.

$O_2(g) + 4\ H^+(aq) \rightarrow 2\ H_2O(l)$

Since the reaction takes place in a basic solution, eliminate H^+(aq) by adding an equal number of OH^-(aq) to both sides of the equation.

$O_2(g) + 4\ H^+(aq) + 4\ OH^-(aq) \rightarrow 2\ H_2O(l) + 4\ OH^-(aq)$

Subtract 2 water molecules from each side to eliminate redundant water molecules.

$O_2(g) + 2\ H_2O(l) \rightarrow 4\ OH^-(aq)$

Step 5: Use electrons to balance the charge in each half-reaction equation.

$Ag(s) \rightarrow Ag^+(aq) + e^-$

$O_2(g) + 2\ H_2O(l) + 4\ e^- \rightarrow 4\ OH^-(aq)$

Step 6: Equalize the electron transfer in the two half-reaction equations.
Multiply the oxidation half-reaction equation by 4.
$4\ Ag(s) \rightarrow 4\ Ag^{+}(aq) + 4\ e^{-}$
$O_2(g) + 2\ H_2O(l) + 4\ e^{-} \rightarrow 4\ OH^{-}(aq)$
Step 7: Add the two half-reaction equations. Include the spectator ion, $CN^{-}(aq)$.
$4\ Ag(s) + O_2(g) + 2\ H_2O(l) \rightarrow 4\ Ag^{+}(aq) + 4\ OH^{-}(aq)$
$4\ Ag(s) + 8\ CN^{-}(aq) + O_2(g) + 2\ H_2O(l) \rightarrow 4\ Ag(CN)_2^{-}(aq) + 4\ OH^{-}(aq)$
Statement: The balanced equation is
$4\ Ag(s) + 8\ CN^{-}(aq) + O_2(g) + 2\ H_2O(l) \rightarrow 4\ Ag(CN)_2^{-}(aq) + 4\ OH^{-}(aq)$
(b) Answers may vary. Sample answer: The second product of the reaction is hydroxide ions. A buildup of hydroxide ions in the environment makes the pH of the soil and water more basic, which can impact the types of plants that grow and the wildlife that eats them.
7. Answers may vary. Sample answer:
(a) Solution: $Cr_2O_7^{2-}(aq) + C_2H_5OH(aq) \rightarrow Cr^{3+}(aq) + CO_2(g)$
Step 1: Write unbalanced equations for oxidation and reduction half-reactions.
The reduction half-reaction equation is:
$Cr_2O_7^{2-}(aq) \rightarrow Cr^{3+}(aq)$
The oxidation half-reaction equation is
$C_2H_5OH(aq) \rightarrow CO_2(g)$
Step 2: Balance each half-reaction equation.
For the reduction half-reaction, balance chromium.
$Cr_2O_7^{2-}(aq) \rightarrow 2\ Cr^{3+}(aq)$
Balance oxygen by adding water.
$Cr_2O_7^{2-}(aq) \rightarrow 2\ Cr^{3+}(aq) + 7\ H_2O(l)$
Balance hydrogen by adding hydrogen ions.
$Cr_2O_7^{2-}(aq) + 14\ H^{+}(aq) \rightarrow 2\ Cr^{3+}(aq) + 7\ H_2O(l)$
For the oxidation half-reaction, balance carbon.
$C_2H_5OH(aq) \rightarrow 2\ CO_2(g)$
Balance oxygen by adding water.
$C_2H_5OH(aq) + 3\ H_2O(l) \rightarrow 2\ CO_2(g)$
Balance hydrogen by adding hydrogen ions.
$C_2H_5OH(aq) + 3\ H_2O(l) \rightarrow 2\ CO_2(g) + 12\ H^{+}(aq)$
Step 3: Use electrons to balance the charge in each half-reaction equation.
$Cr_2O_7^{2-}(aq) + 14\ H^{+}(aq) + 6\ e^{-} \rightarrow 2\ Cr^{3+}(aq) + 7\ H_2O(l)$
$C_2H_5OH(aq) + 3\ H_2O(l) \rightarrow 2\ CO_2(g) + 12\ H^{+}(aq) + 12\ e^{-}$
Step 4: Equalize the electron transfer in the two half-reaction equations.
Multiply the reduction half-reaction equation by 2.
$2\ Cr_2O_7^{2-}(aq) + 28\ H^{+}(aq) + 12\ e^{-} \rightarrow 4\ Cr^{3+}(aq) + 14\ H_2O(l)$
$C_2H_5OH(aq) + 3\ H_2O(l) \rightarrow 2\ CO_2(g) + 12\ H^{+}(aq) + 12\ e^{-}$
Step 5: Add the two half-reaction equations.
$2\ Cr_2O_7^{2-}(aq) + C_2H_5OH(aq) + 16\ H^{+}(aq) \rightarrow 4\ Cr^{3+}(aq) + 2\ CO_2(g) + 11\ H_2O(l)$
Statement: The balanced equation is
$2\ Cr_2O_7^{2-}(aq) + C_2H_5OH(aq) + 16\ H^{+}(aq) \rightarrow 4\ Cr^{3+}(aq) + 2\ CO_2(g) + 11\ H_2O(l)$

(b) Answers may vary. Sample answer: I prefer to use the half-reaction method because it easier to balance the charges.

8. (a) Solution:

Step 1: Write unbalanced equations for oxidation and reduction half-reactions.

The reduction half-reaction equation is

$MnO_4^-(aq) \rightarrow Mn^{2+}(aq)$

The oxidation half-reaction equation is

$H_2O_2(aq) \rightarrow O_2(g)$

Step 2: Balance each half-reaction equation.

For the reduction half-reaction, balance oxygen by adding water.

$MnO_4^-(aq) \rightarrow Mn^{2+}(aq) + 4\ H_2O(l)$

Balance hydrogen by adding hydrogen ions.

$MnO_4^-(aq) + 8\ H^+(aq) \rightarrow Mn^{2+}(aq) + 4\ H_2O(l)$

Add 5 electrons to the reactant side of the equation to balance the charge.

$MnO_4^-(aq) + 8\ H^+(aq) + 5\ e^- \rightarrow Mn^{2+}(aq) + 4\ H_2O(l)$

For the oxidation half-reaction, balance hydrogen by adding hydrogen ions.

$H_2O_2(aq) \rightarrow O_2(g) + 2\ H^+(aq)$

Add 2 electrons to the product side of the equation to balance the charge.

$H_2O_2(aq) \rightarrow O_2(g) + 2\ H^+(aq) + 2\ e^-$

Statement: The balanced half-reaction equations

Oxidation: $H_2O_2(aq) \rightarrow O_2(g) + 2\ H^+(aq) + 2\ e^-$

Reduction: $MnO_4^-(aq) + 8\ H^+(aq) + 5\ e^- \rightarrow Mn^{2+}(aq) + 4\ H_2O(l)$

(b) Solution:

Step 1: Equalize the electron transfer in the two half-reaction equations. Multiply the reduction half-reaction by 2 and the oxidation half-reaction by 5.

$2\ MnO_4^-(aq) + 16\ H^+(aq) + 10\ e^- \rightarrow 2\ Mn^{2+}(aq) + 8\ H_2O(l)$

$5\ H_2O_2(aq) \rightarrow 5\ O_2(g) + 10\ H^+(aq) + 10\ e^-$

Step 2: Add the two-half-reaction equations.

$2\ MnO_4^-(aq) + 5\ H_2O_2(aq) + 6\ H^+(aq) \rightarrow 2\ Mn^{2+}(aq) + 5\ O_2(g) + 8\ H_2O(l)$

Statement: The balanced equation for the redox reaction in an acidic solution is the following:

$2\ MnO_4^-(aq) + 5\ H_2O_2(aq) + 6\ H^+(aq) \rightarrow 2\ Mn^{2+}(aq) + 5\ O_2(g) + 8\ H_2O(l)$

9. Answers may vary. Sample answer:

Solution:

Step 1: Write an unbalanced equation for the reaction.

$SO_3^{2-}(aq) + NO_3^-(aq) \rightarrow SO_4^{2-}(aq) + NO_2(g)$

Step 2: Assign oxidation numbers to all entities.

$\overset{+4\,-2}{SO_3^{2-}}(aq) + \overset{+5\,-2}{NO_3^-}(aq) \rightarrow \overset{+6\,-2}{SO_4^{2-}}(aq) + \overset{+4\,-2}{NO_2}(g)$

The oxidation number of sulfur changes from +4 to +6, so sulfur loses 2 electrons.

The oxidation number of nitrogen changes from +5 to +4, so nitrogen gains 1 electron.

Step 3: To balance electrons, adjust the coefficients by multiplying electrons gained by 2.

$SO_3^{2-}(aq) + 2\ NO_3^-(aq) \rightarrow SO_4^{2-}(aq) + 2\ NO_2(g)$

Step 4: Balance the rest of the equation by inspection.
Balance oxygen by adding water.
$SO_3^{2-}(aq) + 2\ NO_3^-(aq) \rightarrow SO_4^{2-}(aq) + 2\ NO_2(g) + H_2O(l)$
Balance hydrogen by adding hydrogen ions.
$SO_3^{2-}(aq) + 2\ NO_3^-(aq) + 2\ H^+(aq) \rightarrow SO_4^{2-}(aq) + 2\ NO_2(g) + H_2O(l)$
Statement: The balanced equation for the reaction is
$SO_3^{2-}(aq) + 2\ NO_3^-(aq) + 2\ H^+(aq) \rightarrow SO_4^{2-}(aq) + 2\ NO_2(g) + H_2O(l)$
10. (a) Solution:
Step 1: Write unbalanced equations for oxidation and reduction half-reactions.
The reduction half-reaction equation is
$Cl_2(g) \rightarrow Cl^-(aq)$
The oxidation half-reaction equation is
$Cl_2(g) \rightarrow ClO_3^-(aq)$
Step 2: Balance each half-reaction equation.
For the reduction half-reaction, balance chlorine.
$Cl_2(g) \rightarrow 2\ Cl^-(aq)$
For the oxidation half-reaction, balance chlorine.
$Cl_2(g) \rightarrow 2\ ClO_3^-(aq)$
Balance oxygen by adding water.
$Cl_2(g) + 6\ H_2O(l) \rightarrow 2\ ClO_3^-(aq)$
Balance hydrogen by adding hydrogen ions.
$Cl_2(g) + 6\ H_2O(l) \rightarrow 2\ ClO_3^-(aq) + 12\ H^+(aq)$
Since the reaction takes place in a basic solution, eliminate $H^+(aq)$ by adding an equal number of $OH^-(aq)$ to both sides of the equation.
$Cl_2(g) + 6\ H_2O(l) + 12\ OH^-(aq) \rightarrow 2\ ClO_3^-(aq) + 12\ H^+(aq) + 12\ OH^-(aq)$
Subtract 6 water molecules from each side to eliminate redundant water molecules.
$Cl_2(g) + 12\ OH^-(aq) \rightarrow 2\ ClO_3^-(aq) + 6\ H_2O(l)$
Step 3: Use electrons to balance the charge in each half-reaction equation.
$Cl_2(g) + 2\ e^- \rightarrow 2\ Cl^-(aq)$
$Cl_2(g) + 12\ OH^-(aq) \rightarrow 2\ ClO_3^-(aq) + 6\ H_2O(l) + 10\ e^-$
Step 4: Equalize the electron transfer in the two half-reaction equations.
Multiply the reduction half-reaction equation by 5.
$5\ Cl_2(g) + 10\ e^- \rightarrow 10\ Cl^-(aq)$
$Cl_2(g) + 12\ OH^-(aq) \rightarrow 2\ ClO_3^-(aq) + 6\ H_2O(l) + 10\ e^-$
Step 5: Add the two half-reaction equations.
$6\ Cl_2(g) + 12\ OH^-(aq) \rightarrow 2\ ClO_3^-(aq) + 10\ Cl^-(aq) + 6\ H_2O(l)$
Divide all the coefficients by 2 to obtain the lowest possible coefficients.
$3\ Cl_2(g) + 6\ OH^-(aq) \rightarrow ClO_3^-(aq) + 5\ Cl^-(aq) + 3\ H_2O(l)$
Statement: The balanced equation for the reaction is:
$3\ Cl_2(g) + 6\ OH^-(aq) \rightarrow ClO_3^-(aq) + 5\ Cl^-(aq) + 3\ H_2O(l)$
(b) The oxidation numbers of all the elements in this reaction are:
Reactants: chlorine: 0; oxygen: –2; hydrogen: +1
Products: chlorine in ClO_3^-: +5; oxygen: –2; chlorine in Cl^-: –1; hydrogen: +1; O: –2

(c) Since the chlorine atom in Cl_2 gains electrons, the oxidizing agent is $Cl_2(g)$. Since chlorine in ClO_3^- loses electrons, the reducing agent is also $Cl_2(g)$. Elemental chlorine is both an oxidizing agent and a reducing agent in this reaction.