**Introduction to Oxidation-Reduction (Redox) Reactions**

In Unit 4, you have seen examples of the following reaction types.

**precipitation reaction**: Two aqueous solutions of ionic compounds are combined. The cations of one solution combine with the anions of another solution to form an insoluble ionic compound, known as a precipitate.

**acid-base reaction**: A Bronsted-Lowry acid ($H^+$ donor) reacts with a Bronsted-Lowry base ($H^+$ acceptor). The $H^+$ is transferred from the acid to the base.

We will now study another type of reaction known as oxidation-reduction or redox.

Consider the following balanced chemical equation for the reaction between magnesium (Mg) and oxygen ($O_2$).

$$2\,Mg(s) + O_2(g) \rightarrow 2\,MgO(s)$$

This equation can be split up into two separate processes.

| An atom of Mg changes into a $Mg^{2+}$ ion | A molecule of $O_2$ changes into two $O^{2-}$ ions |
|---|---|
| $Mg \rightarrow Mg^{2+}$ | $O_2 \rightarrow 2\,O^{2-}$ |

You should know and understand each of the following definitions about oxidation and reduction.

**oxidation**: a substance loses electrons **reduction**: a substance gains electrons

You can use the mnemonic device **OIL RIG** to help you remember these definitions.

**OIL**: Oxidation is the Loss of electrons **RIG**: Reduction is the Gain of electrons

Equations that show either oxidation or reduction separately are known as **half-reactions**.

- In an oxidation half-reaction, electrons appear on the right side of the equation.
- In a reduction half-reaction, electrons appear on the left side of the equation.

The half-reactions for the reaction involving Mg and $O_2$ are written as follows. Notice that the overall charge is balanced on both sides of the equation.

| Oxidation Half-Reaction | Reduction Half-Reaction |
|---|---|
| $Mg \rightarrow Mg^{2+} + 2\,e^-$ | $4\,e^- + O_2 \rightarrow 2\,O^{2-}$ |

- The number of electrons lost in the oxidation half-reaction must be equal to the number of electrons gained in the reduction half-reaction.
- One (or both) of the half-reactions may need to be multiplied by a certain number so that the electrons can be cancelled out on both sides of the equation.
- When the two half-reactions are added together, you should get the overall balanced redox equation.

$$\begin{array}{l} 2\,Mg \rightarrow 2\,Mg^{2+} + 4\,e^{-} \\ 4\,e^{-} + O_2 \rightarrow 2\,O^{2-} \\ \hline 2\,Mg + O_2 \rightarrow 2\,MgO \end{array}$$

(Note that this half-reaction was multiplied by 2)

(Electrons do not appear in the overall redox equation.)

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In each of the following, you are given two separate half-reactions. You need to add the two half-reactions together to produce a balanced redox equation based on the rules listed on the previous page.

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1.

| Oxidation Half-Reaction | Reduction Half-Reaction |
|---|---|
| $Zn \rightarrow Zn^{2+} + 2\,e^{-}$ | $2\,e^{-} + 2\,H^{+} \rightarrow H_2$ |

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2.

| Oxidation Half-Reaction | Reduction Half-Reaction |
|---|---|
| $Al \rightarrow Al^{3+} + 3\,e^{-}$ | $2\,e^{-} + Cu^{2+} \rightarrow Cu$ |

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3.

| Oxidation Half-Reaction | Reduction Half-Reaction |
|---|---|
| $Fe^{2+} \rightarrow Fe^{3+} + e^{-}$ | $5\,e^{-} + MnO_4^{-} + 8\,H^{+} \rightarrow Mn^{2+} + 4\,H_2O$ |

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4.

| Oxidation Half-Reaction | Reduction Half-Reaction |
|---|---|
| $2\,Cl^{-} \rightarrow Cl_2 + 2\,e^{-}$ | $6\,e^{-} + 14\,H^{+} + Cr_2O_7^{2-} \rightarrow 2\,Cr^{3+} + 7\,H_2O$ |

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5.

| Oxidation Half-Reaction | Reduction Half-Reaction |
|---|---|
| $Ni \rightarrow Ni^{2+} + 2\,e^{-}$ | $3\,e^{-} + 4\,H^{+} + NO_3^{-} \rightarrow NO + 2\,H_2O$ |

Oxidation Numbers

In an oxidation-reduction (redox) reaction, electrons are transferred from the substance that is oxidized to the substance that is reduced. One of the ways to recognize that a reaction is a redox reaction is to look for changes in the oxidation numbers of the elements involved.

Oxidation numbers are a bookkeeping system that is used to keep track of any changes in electrons being gained or lost. The following rules are used when assigning oxidation numbers to the atoms in a molecule or an ion.

| Rule | Examples |
|---|---|
| #1: For an atom in its elemental form, the oxidation number is zero. | $H_2$ = zero  Li = zero  $N_2$ = zero |
| #2: For a monoatomic ion, the oxidation number is the charge on the ion. | $Na^+$ = +1  $Mg^{2+}$ = +2  $Cl^-$ = –1 |
| #3: The oxidation number of Group 1 metals is +1 in all compounds, and the oxidation number of Group 2 metals is +2 in all compounds. | Na = +1 in compounds such as NaCl, $NaNO_3$, and $Na_2O$.<br><br>Mg = +2 in compounds such as $MgBr_2$, $MgSO_4$, and MgS. |
| #4: The oxidation number of F is –1 in all compounds. | F = –1 in compounds such as NaF, $CH_3F$, and HF. |
| #5: The oxidation number of H is usually +1 when bonded to nonmetals and –1 when bonded to metals. | H = +1 in compounds such as $H_2O$, $NH_4Cl$, and $CH_4$.<br><br>H = –1 in compounds such as LiH, NaH, and $CaH_2$. |
| #6: The oxidation number of O is usually –2 in most compounds. Two important exceptions are $H_2O_2$ and $OF_2$. | O = –2 in compounds such as $CO_2$, $KNO_3$, and $H_2SO_4$.<br><br>O = –1 in $H_2O_2$<br><br>O = +2 in $OF_2$ |
| #7: The oxidation number of Cl, Br, and I is –1 in most binary compounds. However, the oxidation number of Cl, Br, and I has a positive value when bonded with O or F. | Cl = –1 in compounds such as KCl, $CCl_4$, and $SCl_2$.<br><br>Cl = +1 in NaClO and HClO<br><br>Cl = +3 in $NaClO_2$ and $HClO_2$<br><br>Cl = +5 in $NaClO_3$ and $HClO_3$<br><br>Cl = +7 in $NaClO_4$ and $HClO_4$ |

- The sum of the oxidation numbers of all atoms in a neutral compound is equal to zero.
- The sum of the oxidation numbers of all atoms in a polyatomic ion is equal to the overall charge on the polyatomic ion.

6. Based on the rules on the previous page, assign oxidation numbers to each atom in each substance.

| Substance | Oxidation Numbers | | |
|---|---|---|---|
| $Fe$ | Fe = | | |
| $Cl_2$ | Cl = | | |
| $P_4$ | P = | | |
| $NaBr$ | Na = | Br = | |
| $Ca_3N_2$ | Ca = | N = | |
| $Fe_2O_3$ | Fe = | O = | |
| $MgS$ | Mg = | S = | |
| $SO_2$ | S = | O = | |
| $CH_3F$ | C = | H = | F = |
| $CHF_3$ | C = | H = | F = |
| $CO$ | C = | O = | |
| $OF_2$ | O = | F = | |
| $H_2O_2$ | H = | O = | |
| $KH$ | K = | H = | |
| $CaH_2$ | Ca = | H = | |
| $HBr$ | H = | Br = | |
| $HBrO$ | H = | Br = | O = |
| $HBrO_2$ | H = | Br = | O = |
| $HBrO_3$ | H = | Br = | O = |
| $Na_2CO_3$ | Na = | C = | O = |
| $BaCrO_4$ | Ba = | Cr = | O = |
| $NH_4^+$ | N = | H = | |
| $OH^-$ | O = | H = | |
| $NO_3^-$ | N = | O = | |
| $ClO_4^-$ | Cl = | O = | |
| $SO_3^{2-}$ | S = | O = | |
| $SO_4^{2-}$ | S = | O = | |
| $Cr_2O_7^{2-}$ | Cr = | O = | |
| $PO_4^{3-}$ | P = | O = | |
| $H_3O^+$ | H = | O = | |

7. Each of the following equations represents an oxidation-reduction reaction.
   - Assign oxidation numbers to each element on each side of the equation.
   - Determine which element is oxidized and which element is reduced in the reaction.
   - If the oxidation number increases, the element is oxidized.
   - If the oxidation number decreases, the element is reduced.

| Balanced Equation | | Element that is Oxidized | Element that is Reduced |
|---|---|---|---|
| $N_2 + 3\ H_2 \rightarrow 2\ NH_3$ | | | |
| N = H = | N = H = | | |

| Balanced Equation | | Element that is Oxidized | Element that is Reduced |
|---|---|---|---|
| $3\ Fe(NO_3)_2 + 2\ Al \rightarrow 3\ Fe + 2\ Al(NO_3)_3$ | | | |
| Fe = N = O = Al = | Fe = N = O = Al = | | |

| Balanced Equation | | Element that is Oxidized | Element that is Reduced |
|---|---|---|---|
| $2\ BrO_3^- + 3\ N_2H_4 \rightarrow 2\ Br^- + 6\ H_2O + 3\ N_2$ | | | |
| Br = O = N = H = | Br = O = N = H = | | |

| Balanced Equation | | Element that is Oxidized | Element that is Reduced |
|---|---|---|---|
| $P_4 + 10\ HClO + 6\ H_2O \rightarrow 4\ H_3PO_4 + 10\ HCl$ | | | |
| P = H = Cl = O = | P = H = Cl = O = | | |

| Balanced Equation | | Element that is Oxidized | Element that is Reduced |
|---|---|---|---|
| $2\ C_2H_6 + 7\ O_2 \rightarrow 4\ CO_2 + 6\ H_2O$ | | | |
| C = H = O = | C = H = O = | | |

| Balanced Equation | | Element that is Oxidized | Element that is Reduced |
|---|---|---|---|
| $Mn + H_2SO_4 \rightarrow MnSO_4 + H_2$ | | | |
| Mn = H = S = O = | Mn = H = S = O = | | |

8. Classify each of the following reactions as precipitation, acid-base, or redox.

| Balanced Equation for the Reaction | Reaction Type |
|---|---|
| $Zn(s) + 2\,HCl(aq) \rightarrow ZnCl_2(aq) + H_2(g)$ | |
| $Ca(OH)_2(aq) + 2\,HCl(aq) \rightarrow CaCl_2(aq) + 2\,H_2O(l)$ | |
| $AgNO_3(aq) + HCl(aq) \rightarrow AgCl(s) + HNO_3(aq)$ | |
| $2\,KOH(aq) + H_2SO_4(aq) \rightarrow 2\,H_2O(l) + K_2SO_4(aq)$ | |
| $Pb(NO_3)_2(aq) + K_2SO_4(aq) \rightarrow 2\,KNO_3(aq) + PbSO_4(s)$ | |
| $PbS(s) + 4\,H_2O_2(aq) \rightarrow 4\,H_2O(l) + PbSO_4(s)$ | |

Net Ionic Equations

9. For each of the following redox reactions, write the balanced net ionic equation.

(a) $Ca(s) + 2\,HCl(aq) \rightarrow CaCl_2(aq) + H_2(g)$

(b) $2\,Fe(s) + 6\,HBr(aq) \rightarrow 2\,FeBr_3(aq) + 3\,H_2(g)$

(c) $Zn(s) + Ni(NO_3)_2(aq) \rightarrow Zn(NO_3)_2(aq) + Ni(s)$

(d) $Al(s) + 3\,AgNO_3(aq) \rightarrow Al(NO_3)_3(aq) + 3\,Ag(s)$

(e) $Cu(s) + 4\,HNO_3(aq) \rightarrow Cu(NO_3)_2(aq) + 2\,H_2O(l) + 2\,NO_2(g)$

(f) $Br_2(aq) + 2\,KI(aq) \rightarrow 2\,KBr(aq) + I_2(aq)$