Worksheet 7 - Oxidation/Reduction Reactions

Oxidation number rules:

**Elements** have an oxidation number of 0

**Group I** and **II** – In addition to the elemental oxidation state of 0, Group I has an oxidation state of +1 and Group II has an oxidation state of +2.

**Hydrogen** – usually +1, except when bonded to Group I or Group II, when it forms hydrides, -1.

**Oxygen** – usually -2, except when it forms a O-O single bond, a peroxide, when it is -1.

**Fluorine** is always -1. Other halogens are usually -1, except when bonded to O.

1. Assign **oxidation numbers** to each of the atoms in the following compounds:

   - $\text{Na}_2\text{CrO}_4$: Na =   O =   Cr =
   - $\text{K}_2\text{Cr}_2\text{O}_7$: K =   O =   Cr =
   - CO$_2$: O =   C =
   - CH$_4$: H =   C =
   - HClO$_4$: O =   H =   Cl =
   - MnO$_2$: O =   Mn =
   - SO$_3^{2-}$: O =   S =
   - SF$_4$: F =   S =

   a. What is the range of oxidation states for carbon?
   b. Which compound has C in a +4 state?
   c. Which compound has C in a -4 state?
2. Nitrogen has 5 valence electrons (Group V). It can gain up to 3 electrons (-3), or lose up to 5 (+5) electrons. Fill in the missing names or formulas and assign an oxidation state to each of the following nitrogen containing compounds:

<table>
<thead>
<tr>
<th>name</th>
<th>formula</th>
<th>oxidation state of N</th>
</tr>
</thead>
<tbody>
<tr>
<td>nitrogen</td>
<td>NH(_3)</td>
<td></td>
</tr>
<tr>
<td>nitrite</td>
<td>NO(_3^–)</td>
<td></td>
</tr>
<tr>
<td>dinitrogen monoxide</td>
<td>NO(_2)</td>
<td></td>
</tr>
<tr>
<td>hydroxylamine</td>
<td>NH(_2)OH</td>
<td></td>
</tr>
<tr>
<td>nitrogen monoxide</td>
<td></td>
<td></td>
</tr>
<tr>
<td>hydrazine</td>
<td>N(_2)H(_4)</td>
<td></td>
</tr>
</tbody>
</table>

During chemical reactions, the **oxidation state** of atoms can change. This occurs when compounds gain or lose electrons, or when the bonds to an atom change. This is illustrated by the reaction between nitrogen and hydrogen to make ammonia:

\[
\text{N}_2(\text{g}) + 3 \text{H}_2(\text{g}) \rightarrow 2 \text{NH}_3(\text{g})
\]

a. Assign oxidation numbers to each of the atoms in this reaction.

- N (in N\(_2\)) =
- H (in H\(_2\)) =
- N (in NH\(_3\)) =
- H (in NH\(_3\)) =

When an oxidation number **increases**, that species has been oxidized.

b. Which reactant undergoes an increase in its oxidation number?

When an oxidation number **decreases**, that species has been reduced.

c. Which reactant undergoes a decrease in its oxidation number?
The species that is oxidized is called the reducing agent because it gives up an electron, so that another species can gain an electron (be reduced).

d. What is the reducing agent in this reaction?

The species that is reduced is called the oxidizing agent because it takes an electron away from another group, raising that group’s oxidation number.

e. What is the oxidizing agent in this reaction?

3. In each of the following reactions, assign oxidation numbers to all of the elements and identify the oxidizing and reducing agents and the change in oxidation number.

a. \( 4 \text{Fe} + 3 \text{O}_2 \rightarrow 2 \text{Fe}_2\text{O}_3 \)

   oxidizing agent
   reducing agent
   change in oxidation number

b. \( \text{Cr}_2\text{O}_7^{2-} + 2\text{OH}^- \rightarrow 2 \text{CrO}_4^{2-} + \text{H}_2\text{O} \)

   oxidizing agent
   reducing agent
   change in oxidation number

c. \( \text{NH}_4\text{NO}_2 \rightarrow \text{N}_2 + 2 \text{H}_2\text{O} \)

   oxidizing agent
   reducing agent
   change in oxidation number

d. \( \text{P}_4 + 10 \text{Cl}_2 \rightarrow 4 \text{PCl}_5 \)

   oxidizing agent
   reducing agent
   change in oxidation number

e. \( 2 \text{Cr}^{3+} + \text{H}_2\text{O} + 6 \text{ClO}_3^- \rightarrow \text{Cr}_2\text{O}_7^{2-} + 6\text{ClO}_2 + 2 \text{H}^+ \)

   oxidizing agent
   reducing agent
   change in oxidation number
Balancing Redox Reactions

Oxidation/Reduction (Redox) reactions can be balanced using the oxidation state changes, as seen in the previous example. However, there is an easier method, which involves breaking a redox reaction into two half-reactions. This is best shown by working an example.

Hydrobromic acid will react with permanganate to form elemental bromine and the manganese(II) ion. The unbalanced, net reaction is shown below,

\[ \text{Br}^- + \text{MnO}_4^- \rightarrow \text{Br}_2 + \text{Mn}^{2+} \]

1. Break this into two half-reactions, one involving bromine and the other involving manganese.

   **Bromine half-reaction**
   
   \[ \text{Br}^- \rightarrow \text{Br}_2 \]

   **Manganese half-reaction**
   
   \[ \text{MnO}_4^- \rightarrow \text{Mn}^{2+} \]

2. First balance the bromine half-reaction first.
   a. Balance the bromine atoms of the reaction
      \[ \underline{ \quad \text{Br}^- \rightarrow \quad \text{Br}_2 } \]
   b. Now balance charge by adding electrons (\(e^-\))
      \[ \underline{ \quad \text{Br}^- \rightarrow \quad \text{Br}_2 } \]

   This half-reaction is producing/consuming electrons. This is an oxidation/reduction half-reaction. Confirm this by assigning oxidation numbers to the bromine species.

3. Next, balance the manganese half-reaction.
   a. Balance the manganese atoms of the half-reaction
      \[ \underline{ \quad \text{MnO}_4^- \rightarrow \quad \text{Mn}^{2+} } \]
   b. Next, balance oxygen by adding water molecules (\(\text{H}_2\text{O}\))
      \[ \underline{ \quad \text{MnO}_4^- \rightarrow \quad \text{Mn}^{2+} } \]
c. Next, balance **hydrogen** by adding protons (H⁺)

\[ \underline{\text{____ MnO}_4^-} \to \underline{\text{____ Mn}^{2+}} \]

d. Finally, balance **charge** by adding electrons (e⁻).

\[ \underline{\text{____ MnO}_4^-} \to \underline{\text{____ Mn}^{2+}} \]

This half-reaction is **producing/consuming electrons**. This is a **oxidation/reduction** half-reaction. Confirm this by assigning oxidation numbers to the manganese atoms.

Notice that the number of electrons equals the change in oxidation number.

4. Now put the two **half-reactions** together. The number of electrons produced must equal the number of electrons consumed.

\[
2 \text{ Br}^- \to \text{ Br}_2 + 2\text{e}^- \\
5\text{e}^- + 8\text{H}^+ + \text{MnO}_4^- \to \text{Mn}^{2+} + 4\text{H}_2\text{O}
\]

Multiply this half-reaction by _____

\[
\underline{\text{____ Br}^-} \to \underline{\text{____ Br}_2 + \underline{\text{____ e}^-}} \\
\underline{\text{____ e}^- + \underline{\text{____ H}^+ + \underline{\text{____ MnO}_4^-}}} \to \underline{\text{____ Mn}^{2+} + \underline{\text{____ H}_2\text{O}}}
\]

Add the two half-reactions, canceling out species that appear on both sides (including electrons)

\[
\underline{\text{____ Br}^- + \text{____ H}^+ + \text{____ MnO}_4^-} \to \underline{\text{____ Br}_2 + \text{____ Mn}^{2+} + \text{____ H}_2\text{O}}
\]

Which compound is the **oxidizing agent**?

Which compound is the **reducing agent**?

Notice that there are protons (H⁺) present in the reactants. This indicates that the reaction is carried out in an **acidic solution**. To carry this out in a **basic solution**, simply add enough hydroxide ions (OH⁻) to each side of the equation to neutralize the protons. The product of the neutralization reaction will be water.
The overall balanced reaction under basic conditions is:

\[ \underline{\text{Br}^-} + \underline{\text{MnO}_4^-} + \underline{\text{H}_2\text{O}} \rightarrow \underline{\text{Br}_2} + \underline{\text{Mn}^{2+}} + \underline{\text{OH}^-} \]

Now, balance the redox reaction between methanol and dichromate, which produces methanal and chromium (III), as shown below:

\[ \text{CH}_3\text{OH} + \text{Cr}_2\text{O}_7^{2-} \rightarrow \text{CH}_2\text{O} + \text{Cr}^{3+} \]

First, separate this into two half-reactions
Then, balance the redox active species.
Then, balance oxygens with \( \text{H}_2\text{O} \)
Balance hydrogen with \( \text{H}^+ \)
Balance charge with electrons.
Equalize the number of electrons lost and gained

This indicates that the reaction must be carried out in an \textbf{acidic} solution.

To carry it out in a \textbf{basic} solution, just add enough \( \text{OH}^- \) to neutralize the acid, \( \text{H}^+ \)
Balance the following redox-reaction which takes place in basic solutions.

\[
\text{Zn (s) + NO}_2^- \rightarrow \text{NH}_3 + \text{Zn(OH)}_4^{2-}
\]