Worksheet 1 - 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 or F.

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

   Na\textsubscript{2}CrO\textsubscript{4}  \quad Na =   \quad O =   \quad Cr =

   K\textsubscript{2}Cr\textsubscript{2}O\textsubscript{7}  \quad K =   \quad O =   \quad Cr =

   CO\textsubscript{2}  \quad O =   \quad C =

   CH\textsubscript{4}  \quad H =   \quad C =

   HClO\textsubscript{4}  \quad O =   \quad H =   \quad Cl =

   MnO\textsubscript{2}  \quad O =   \quad Mn =

   SO\textsubscript{3}^{2-}  \quad O =   \quad S =

   SF\textsubscript{4}  \quad F =   \quad 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>NH₃</td>
<td></td>
<td></td>
</tr>
<tr>
<td>nitrogen</td>
<td></td>
<td></td>
</tr>
<tr>
<td>nitrite</td>
<td>NO₃⁻</td>
<td></td>
</tr>
<tr>
<td>dinitrogen monoxide</td>
<td>NO₂</td>
<td></td>
</tr>
<tr>
<td>hydroxylamine</td>
<td>NH₂OH</td>
<td></td>
</tr>
<tr>
<td>nitrogen monoxide</td>
<td></td>
<td></td>
</tr>
<tr>
<td>hydrazine</td>
<td>N₂H₄</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.

\[
\begin{align*}
\text{N (in } \text{N}_2 & = \quad \text{N (in } \text{NH}_3 & = \\
\text{H (in } \text{H}_2 & = \quad \text{H (in } \text{NH}_3 & = \\
\end{align*}
\]

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. $4 \text{Fe} + 3 \text{O}_2 \rightarrow 2 \text{Fe}_2\text{O}_3$

change in oxidation number

**oxidizing agent**
**reducing agent**

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

change in oxidation number

**oxidizing agent**
**reducing agent**

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

change in oxidation number

**oxidizing agent**
**reducing agent**

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

change in oxidation number

**oxidizing agent**
**reducing agent**

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

change in oxidation number

**oxidizing agent**
**reducing agent**
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                  Manganese half-reaction
\[ \text{Br}^- \rightarrow \text{Br}_2 \] \[ \text{MnO}_4^- \rightarrow \text{Mn}^{2+} \]

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

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
      \[ \_\_\_ \text{MnO}_4^- \rightarrow \_\_\_ \text{Mn}^{2+} \]
   b. Next, balance oxygen by adding water molecules (H_2O)
      \[ \_\_\_ \text{MnO}_4^- \rightarrow \_\_\_ \text{Mn}^{2+} \]
c. Next, balance **hydrogen** by adding protons (H<sup>+</sup>)

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

d. Finally, balance **charge** by adding electrons (e<sup>-</sup>).

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

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.**

\[
\begin{align*}
2 \text{ Br}^- & \rightarrow \text{ Br}_2 + 2\text{e}^- \\
5\text{e}^- + 8\text{H}^+ + \text{MnO}_4^- & \rightarrow \text{Mn}^{2+} + 4\text{H}_2\text{O}
\end{align*}
\]

multiply this half-reaction by ____ multiply this half-reaction by ___

\[
\begin{align*}
\text{____ Br}^- & \rightarrow \text{____ Br}_2 + \text{____ e}^- \\
\text{____ e}^- + \text{____ H}^+ + \text{____ MnO}_4^- & \rightarrow \text{____ Mn}^{2+} + \text{____ H}_2\text{O}
\end{align*}
\]

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

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

Which compound is the **oxidizing agent**?

Which compound is the **reducing agent**?

Notice that there are protons (H<sup>+</sup>) 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<sup>-</sup>) 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:

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

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 **acidic** solution.

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

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