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:

Na_2CrO_4	Na =	O =	Cr =
$K_2Cr_2O_7$	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:

name	**formula**	**oxidation state of N**
	NH_3	
nitrogen		
nitrite		
	NO_3^-	
dinitrogen monoxide		
	NO_2	
hydroxylamine	NH_2OH	
nitrogen monoxide		
hydrazine	N_2H_4	

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:

$$\mathbf{N_2(g) + 3\ H_2(g) \rightarrow 2\ NH_3(g)}$$

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

N (in N_2) = N (in NH_3) =

H (in H_2) = 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. $4\ Fe + 3\ O_2 \rightarrow 2\ Fe_2O_3$

change in oxidation number

oxidizing agent

reducing agent

b. $Cr_2O_7^{2-} + 2OH^- \rightarrow 2\ CrO_4^{2-} + H_2O$

change in oxidation number

oxidizing agent

reducing agent

c. $NH_4NO_2 \rightarrow N_2 + 2\ H_2O$

change in oxidation number

oxidizing agent

reducing agent

d. $P_4 + 10\ Cl_2 \rightarrow 4\ PCl_5$

change in oxidation number

oxidizing agent

reducing agent

e. $2\ Cr^{3+} + H_2O + 6\ ClO_3^- \rightarrow Cr_2O_7^{2-} + 6ClO_2 + 2\ H^+$

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,

$$\mathbf{Br^- + MnO_4^- \rightarrow Br_2 + Mn^{2+}}$$

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

   Bromine half-reaction | **Manganese half-reaction**

   $Br^- \rightarrow Br_2$ | $MnO_4^- \rightarrow Mn^{2+}$

2. First balance the bromine half-reaction first.

   a. Balance the **bromine** atoms of the reaction

   $$\_\_\_ Br^- \rightarrow \_\_\_ Br_2$$

   b. Now balance **charge** by adding **electrons** (**e**$^-$)

   $$\_\_\_ Br^- \rightarrow \_\_\_ Br_2$$

   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

   $$\_\_\_ MnO_4^- \rightarrow \_\_\_ Mn^{2+}$$

   b. Next, balance **oxygen** by adding water molecules (**H_2O**)

   $$\_\_\_ MnO_4^- \rightarrow \_\_\_ Mn^{2+}$$

c. Next, balance **hydrogen** by adding protons (**H^+**)

$$\_\_\_ \ MnO_4^- \rightarrow \_\_\_ \ Mn^{2+}$$

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

$$\_\_\_ \ MnO_4^- \rightarrow \_\_\_ \ Mn^{2+}$$

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**.

$2\ Br^- \rightarrow Br_2 + \mathbf{2e^-}$ $\qquad$ $\mathbf{5e^-} + 8H^+ + MnO_4^- \rightarrow Mn^{2+} + 4H_2O$

multiply this half-reaction by ____ $\qquad$ multiply this half-reaction by ___

$\_\_\_Br^- \rightarrow \_\_\_Br_2 + \_\_\_e^-$ $\qquad$ $\_\_\_e^- + \_\_\_H^+ + \_\_\_MnO_4^- \rightarrow \_\_\_Mn^{2+} + \_\_\_H_2O$

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

$$\_\_\_Br^- + \_\_\_H^+ + \_\_\_MnO_4^- \rightarrow \_\_\_Br_2 + \_\_\_Mn^{2+} + \_\_\_H_2O$$

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:

$$\_\_\_Br^- + \_\_\_MnO_4^- + \_\_\_H_2O \rightarrow \_\_\_Br_2 + \_\_\_Mn^{2+} + \_\_\_OH^-$$

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

$$\mathbf{CH_3OH + Cr_2O_7^{2-} \rightarrow CH_2O + Cr^{3+}}$$

First, separate this into two half-reactions
Then, balance the redox active species.
Then, balance oxygens with H_2O
Balance hydrogen with 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 **OH^-** to neutralize the acid, **H^+**

Balance the following redox-reaction which takes place in **basic** solutions.

$$Zn\ (s) + NO_2^- \rightarrow NH_3 + Zn(OH)_4^{2-}$$