

Worksheet 7 - Oxidation/Reduction Reactions

Oxidation number rules:

Elements have an oxidation number of **0**

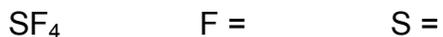
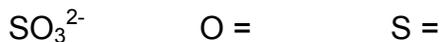
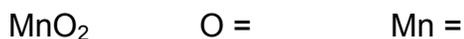
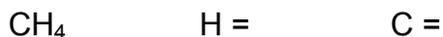
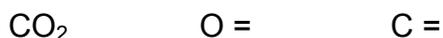
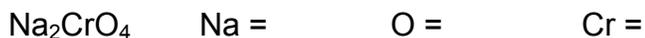
Group I and II – In addition to the elemental oxidation state of 0, Group I has an oxidation state of **+1** and Group II has an oxidation state of **+2**.

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:

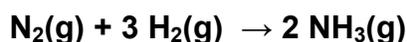


- What is the range of oxidation states for **carbon**?
- Which compound has C in a +4 state?
- Which compound has C in a -4 state?

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 ₂ OH	
nitrogen monoxide		
hydrazine	N ₂ H 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:



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

N (in N₂) = N (in NH₃) =

H (in H₂) = H (in NH₃) =

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?

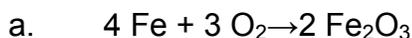
The species that is oxidized is called the **reducing agent** because it gives up an electron, so that another species can gain an electron (be reduced).

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**.



change in oxidation number

oxidizing agent

reducing agent



change in oxidation number

oxidizing agent

reducing agent



change in oxidation number

oxidizing agent

reducing agent



change in oxidation number

oxidizing agent

reducing agent



change in oxidation number

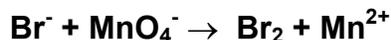
oxidizing agent

reducing agent

Balancing Redox Reactions

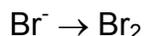
Oxidation/Reduction (Redox) reactions can be balanced using the oxidation state changes, as seen in the previous example. However, there is an easier method, which involves breaking a redox reaction into two **half-reactions**. This is best shown by working an example.

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

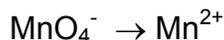


1. Break this into two **half-reactions**, one involving bromine and the other involving manganese.

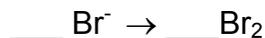
Bromine half-reaction



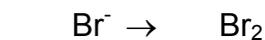
Manganese half-reaction



2. First balance the bromine half-reaction first.
 - a. Balance the **bromine** atoms of the reaction

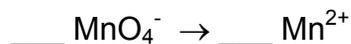


- b. Now balance **charge** by adding **electrons (e⁻)**

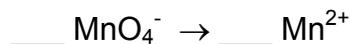


This half-reaction is **producing/consuming electrons**. This is an **oxidation/reduction** half-reaction. Confirm this by assigning oxidation numbers to the bromine species.

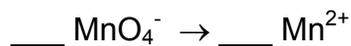
3. Next, balance the manganese half-reaction.
 - a. Balance the **manganese** atoms of the half-reaction



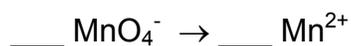
- b. Next, balance **oxygen** by adding water molecules (**H₂O**)



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



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



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.

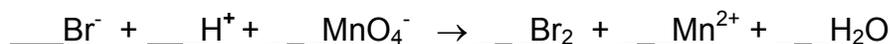
4. Now put the two **half-reactions** together. **The number of electrons produced must equal the number of electrons consumed.**



multiply this half-reaction by $\underline{\quad}$ multiply this half-reaction by $\underline{\quad}$



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

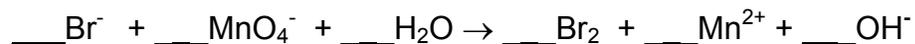


Which compound is the **oxidizing agent** ?

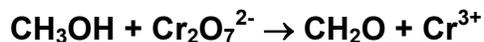
Which compound is the **reducing agent** ?

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:



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



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.

