

Name:

Atomic Structure	Objectives
5. Oxidation and Reduction	<ul style="list-style-type: none"> -define oxidation and reduction in terms of electron transfer -use simple examples , e.g. Na with Cl_2, Mg with O_2, Zn with Cu^{2+} to describe oxidation and reduction in terms of electron transfer -apply knowledge of oxidation and reduction to explain the rusting of iron -define oxidising agent and reducing agent -carry out an experiment to show that halogens act as oxidising agents(reactions with bromides, iodides, Fe^{2+} and sulfites; half equations only required) -carry out an experiment to demonstrate the displacement reactions of metals (Zn with Cu^{2+}, Mg with Cu^{2+})
Chemical Bonding	
6. Oxidation Numbers	<ul style="list-style-type: none"> -define oxidation number, oxidation state -define oxidation and reduction in terms of change of oxidation numbers -state the rules for oxidation numbers (exclude peroxides, except for hydrogen peroxide) -calculate oxidation numbers of transition metals in their compounds and of other elements -use oxidation numbers in nomenclature of transition metal compounds -give an example of an oxidising and a reducing bleach

Oxidation and reduction can be described in four ways:

In terms of:

1. Addition/removal of oxygen.
2. Addition/removal of hydrogen.
3. Electron transfer.
4. Change in oxidation number.

1. Addition/Removal of Oxygen:

Defⁿ: Oxidation is the addition of oxygen.

E.g. $\text{C} + \text{O}_2 \rightarrow \text{CO}_2$ The carbon gains oxygen, therefore the carbon is oxidised.

Defⁿ: Reduction is the removal of oxygen.

E.g. $\text{CuO} + \text{H}_2 \rightarrow \text{Cu} + \text{H}_2\text{O}$ The copper loses oxygen, therefore the copper is reduced.

2. Addition/Removal of Hydrogen:

Defⁿ: Oxidation is the removal of hydrogen.

E.g. $\text{H}_2\text{S} + \text{Cl}_2 \rightarrow \text{S} + 2\text{HCl}$ The sulphur loses hydrogen, therefore the S is oxidised.

Defⁿ: Reduction is the addition of hydrogen.

E.g. $\text{CO} + 2\text{H}_2 \rightarrow \text{CH}_3\text{OH}$ The carbon monoxide gains hydrogen, therefore the CO is reduced.

3. Electron Transfer:

Defⁿ: Oxidation is the loss of electrons.

E.g. $\text{Zn} + \text{Cu}^{2+} \rightarrow \text{Zn}^{2+} + \text{Cu}$ The zinc loses 2e^- , therefore the zinc is oxidised.

Defⁿ: Reduction is the gain of electrons.

E.g. $\text{Zn} + \text{Cu}^{2+} \rightarrow \text{Zn}^{2+} + \text{Cu}$ The copper gains 2e^- , therefore the copper is reduced.

Remember, for electron transfer: **O**_{xidation} **I**_s **L**oss **R**eduction **I**_s **G**ain.

4. Change in Oxidation Number: (Finding oxidation numbers mentioned later in these notes)

Defⁿ: Oxidation is an increase in oxidation number.

E.g. $\text{H}_2 + \frac{1}{2}\text{O}_2 \rightarrow \text{H}_2\text{O}$ The O.N. of H increased from 0 to +1, therefore H is oxidised.

$\begin{array}{ccc} 2(0) & 2(0) & 2(+1) \quad (-2) \\ | & & \uparrow \\ \hline & & \end{array}$

Defⁿ: Reduction is a decrease in oxidation number.

E.g. $\text{H}_2 + \frac{1}{2}\text{O}_2 \rightarrow \text{H}_2\text{O}$ The O.N. of O decreased from 0 to -2, therefore O is reduced.

$\begin{array}{ccc} 2(0) & 2(0) & 2(+1) \quad (-2) \\ | & & \uparrow \\ \hline & & \end{array}$

Oxidising/Reducing Agents:

Defⁿ: An **Oxidising Agent** is a substance that brings about oxidation in other substances by being reduced.

Defⁿ: A **Reducing Agent** is a substance that brings about reduction in other substances by being oxidised.

E.g. $\text{Zn} + \text{Cu}^{2+} \rightarrow \text{Zn}^{2+} + \text{Cu}$

$\begin{array}{c} +2\text{e}^- \rightarrow \text{reduced} \rightarrow \text{oxidising agent} \\ | \qquad \qquad \qquad \downarrow \\ \hline \end{array}$

Assigning Oxidation Numbers:

Defⁿ: **Oxidation Number** is the apparent charge an atom has when electrons are distributed according to certain rules.

Rules:

- The oxidation number of any element that is not bonded to another different element is 0.
E.g. O.N. of H in H_2 is 0. O.N. of Na is 0. O.N. of Cl in Cl_2 is 0.
- The oxidation number of an ion of an element is the same as its charge.
E.g. O.N. of H in H^+ is +1. O.N. of Ca in Ca^{2+} is +2. O.N. of Cl in Na^+Cl^- is -1.
- The total of the oxidation numbers in a neutral compound must add to give 0.
E.g. For H_2O , each H is +1 and O is -2. This gives $2(+1) + (-2) = 0$.

4. Oxygen has an oxidation number of -2.

Exceptions:

- (a) Peroxides like H_2O_2 , as O.N of H = +1, each O must have an O.N. of -1 to give a total sum of 0.
 (b) In OF_2 , as O.N. of F = -1, O must have an O.N. of +2.

5. Hydrogen has an oxidation number of +1.

Exception:

- (a) In metal hydrides like NaH , CaH_2 (H after a metal). Here H has an O.N. of -1.

6. Halogens (Group VII elements) have an oxidation number of -1.

Exception:

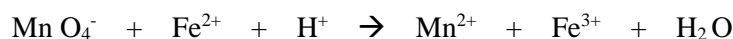
- (a) When bonded to a more electronegative element (O or F), halogens can have other oxidation numbers which are calculated using rule 3.

7. The total of the oxidation numbers in a complex ion must add to give the charge on the ion.

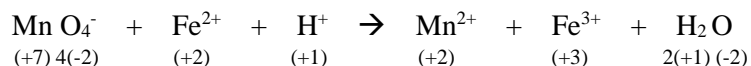
E.g. For NO_3^- , each O has an O.N. of -2, giving $3(-2) = -6$. The total must give -1 as this is the charge on the ion. Therefore N must have an O.N. of +5.

Balancing Redox Equations:

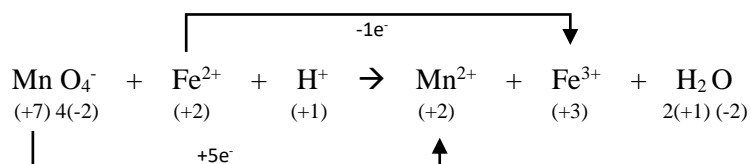
Using oxidation numbers, balance the following equation:



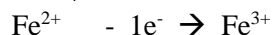
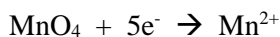
1. Assign oxidation numbers:



2. Show the number of electrons gained/lost:



3. Write the half-equations for the oxidation/reduction of the two substances above:

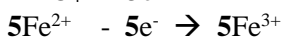
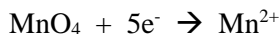


Balance so that:

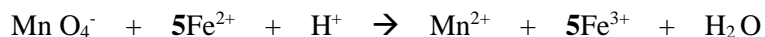
- (a) No. of Mn atoms is equal on both sides

- (b) No. of Fe atoms is equal on both sides

- (c) electrons gained = electrons lost:



4. Write the coefficients from step 3 into the original equation:



5. Balance the rest of the equation by inspection (leave H until last):

