### 2.1.5 Redox

**oxidation** is the process of electron loss: \[ \text{Zn} \rightarrow \text{Zn}^{2+} + 2e^- \]

It involves an *increase in oxidation* number

**reduction** is the process of electron gain: \[ \text{Cl}_2 + 2e^- \rightarrow 2\text{Cl}^- \]

It involves a *decrease in oxidation* number

### Rules for assigning oxidation numbers

1. All uncombined elements have an oxidation number of zero eg. Zn, Cl, O, Ar all have oxidation numbers of zero

   In NaCl  \( \text{Na} = +1 \)  \( \text{Cl} = -1 \)
   
   Sum = +1 -1 = 0

2. The oxidation numbers of the elements in a compound add up to zero

   e.g. \( \text{Zn}^{2+} = +2 \) \( \text{Cl}^- = -1 \)

   e.g. in \( \text{CO}_3^{2-} \) \( C = +4 \) and \( O = -2 \)
   
   \[ \text{sum} = +4 + (3 \times -2) = -2 \]

3. The oxidation number of a monoatomic ion is equal to the ionic charge

4. In a polyatomic ion (\( \text{CO}_3^{2-} \)) the sum of the individual oxidation numbers of the elements adds up to the charge on the ion

5. Several elements have invariable oxidation numbers in their common compounds.

| Group 1 metals | +1 |
| Group 2 metals | +2 |
| Al | +3 |
| H = +1 (except in metal hydrides where it is –1 eg NaH) | |
| F = -1 | |
| Cl, Br, I = –1 except in compounds with oxygen and fluorine | |
| O = -2 except in peroxides (\( \text{H}_2\text{O}_2 \)) where it is –1 and in compounds with fluorine. | |

#### Naming using Roman Numerals

Use a Roman numeral to indicate the magnitude of the oxidation state of an element, when a name may be ambiguous.

- \( \text{FeCl}_2 \): Iron(II) Chloride
- \( \text{FeCl}_3 \): Iron(III) Chloride

In IUPAC convention the various forms of sulphur, nitrogen and chlorine compounds where oxygen is combined are all called sulfates, nitrates and chlorates with relevant oxidation number given in roman numerals. If asked to name these compounds remember to add the oxidation number.

- \( \text{NaClO} \): sodium chlorate(I)
- \( \text{NaClO}_3 \): sodium chlorate(V)
- \( \text{K}_2\text{SO}_4 \): potassium sulfate(VI)
- \( \text{K}_2\text{SO}_3 \): potassium sulfate(IV)
- \( \text{NaNO}_2 \): Sodium nitrate(III)
- \( \text{NaNO}_3 \): Sodium nitrate(V)

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Redox Reactions

Metals generally form ions by losing electrons with an increase in oxidation number to form positive ions:
\[
\text{Zn} \rightarrow \text{Zn}^{2+} + 2\text{e}^{-}
\]

Non-metals generally react by gaining electrons with a decrease in oxidation number to form negative ions:
\[
\text{Cl}_2 + 2\text{e}^{-} \rightarrow 2\text{Cl}^{-}
\]

Oxygen is reducing because its oxidation number is decreasing from 0 to -2
\[
4\text{Li} + \text{O}_2 \rightarrow 2\text{Li}_2\text{O}
\]
Lithium is oxidising because its oxidation number is increasing from 0 to +1

Nitrogen is reducing because its oxidation number is decreasing from +5 to +4
\[
2\text{Sr(NO}_3)_2 \rightarrow 2\text{SrO} + 4\text{NO}_2 + \text{O}_2
\]
Oxygen is oxidising because its oxidation number is increasing from -2 to 0

Nitrogen is oxidising because its oxidation number is increasing from -3 to -2
\[
2\text{NH}_3 + \text{NaClO} \rightarrow \text{N}_2\text{H}_4 + \text{NaCl} + \text{H}_2\text{O}
\]
Chlorine is reducing because its oxidation number is decreasing from +1 to -1

Tungsten is reducing because its oxidation number is decreasing from +6 to 0
\[
\text{WO}_3 + 3\text{H}_2 \rightarrow \text{W} + 3\text{H}_2\text{O}
\]
Hydrogen is oxidising because its oxidation number is increasing from 0 to +1

Redox Reactions of Metals and acid

**ACID + METAL → SALT + HYDROGEN**

Hydrogen is reducing because its oxidation number is decreasing from +1 to 0
\[
2\text{HCl} + \text{Mg} \rightarrow \text{MgCl}_2 + \text{H}_2
\]
Magnesium is oxidising because its oxidation number is increasing from 0 to +2

Be able to write equations for reactions of metals with hydrochloric acid and sulphuric acid
\[
\text{Fe} + \text{H}_2\text{SO}_4 \rightarrow \text{FeSO}_4 + \text{H}_2
\]

Observations: These reactions will effervesce because \(\text{H}_2\) gas is evolved and the metal will dissolve