

Oxidation Numbers

The oxidation number of an element is the charge the element would have if it were an ion
It helps us keep track of electrons in an oxidation-reduction reaction; It may be real or make-believe

Oxidation: The process whereby the oxidation number of an element increases

- Becomes more positive
- Involves the **loss** of electrons
- Electrons are a **product**
- $M^{\circ} \rightarrow M^{n+} + ne^{-}$
- $X^{-} \rightarrow X^{\circ} + e^{-}$
- $M^{2+} \rightarrow M^{3+} + e^{-}$

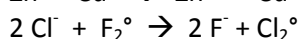
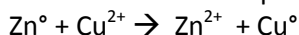
Reduction: The process whereby the oxidation number of an element decreases

- Becomes more negative
- Involves the **gain** of electrons
- Electrons are a **reactant**
- $M^{n+} + ne^{-} \rightarrow M^{\circ}$
- $X_2 + 2e^{-} \rightarrow 2X^{-}$
- $M^{4+} + 2e^{-} \rightarrow M^{2+}$

oxidation – reduction reactions

- Called “**redox**” reactions for short
- **Always** occur as a pair
- One element “loses” electrons
 - Oxidation
- One element “gains” electrons
 - Reduction

Redox reaction examples:



Determining oxidation numbers

1. The oxidation number of a free element is zero, regardless of it is “monatomic” or if it has a subscript

Examples: Mg, O₂, P₄, Zn

2. The oxidation number of a “monatomic” ion is the same as the charge of the ion

Na⁺ has an ox# of +1

S²⁻ has an ox# of -2

Fe³⁺ has an ox# of +3

I⁻ has an ox# of -1

3. The sum of all the oxidation numbers of all the elements in a substance is the same as the charge of the substance

The ox#'s in a neutral compound must all add up to zero

The ox#'s in a polyatomic ion must all add up to the charge of the polyatomic ion

In a compound...

4) The ox# of fluorine is -1

5) the ox# of hydrogen is +1

except in a hydride, where it is -1

ex: LiH, CaH₂

6) the ox# of oxygen is -2

except in peroxides and superoxides

ex: H₂O₂, KO₂, OF₂