Summing-up

All chemical reactions in which there is a transfer of electrons from one chemical to another are called **oxidation-reduction reactions** (or *redox* reactions).

Redox reactions are the result of two transformations that occur simultaneously:

- the oxidation half-reaction in which an atom of a chemical species donates electrons and increases its oxidation number;
- the reduction half-reaction in which an atom of a chemical species acquires electrons and decreases its oxidation number.

In other words, atoms can *oxidise* only if other atoms are simultaneously capable of being *reduced*, the species that oxidises is called the *reducing species* whilst that which is reduced is called the *oxidising species*.

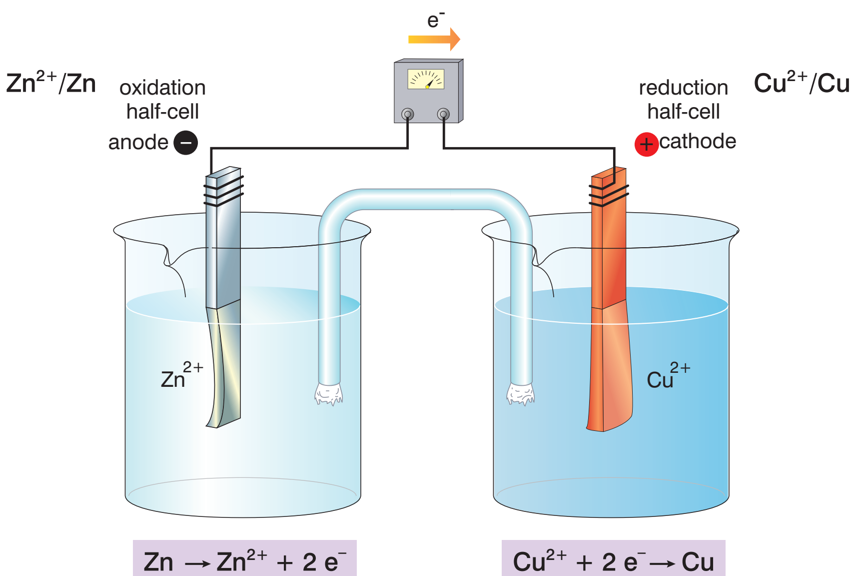
Therefore, all the reactions in which at least two atoms of the reacting chemical species change their oxidation number are redox reactions.

To balance a redox equation one must write the equations of the two half-reactions and insert coefficients such that the total number of electrons donated in the oxidation half-reaction are equal to those acquired in the reduction half-reaction.

A **battery** or **electrochemical cell** is any system capable of transforming chemical energy into electrical energy through a suitable oxidation-reduction reaction.

An electrochemical cell consists of:

- an oxidation half-cell, or anode which is the negative pole;
- a reduction half-cell or cathode which is the positive pole;
- a salt bridge that connects the two half-cells.



Every battery has an associated definite value of *potential difference (ddp)* or voltage.

Since it is not possible to directly measure the potential of a single half-cell it has been conventional practice to use a reference half-cell, the *standard hydrogen electrode*, which is assigned a zero electrical potential. In this way, it is possible to obtain the values of **standard reduction potentials (*E*°)** for all other chemical species. The higher the value of *E*°, the greater the tendency of a species to reduce. Comparing the values of standard potential for two species makes it possible to establish whether a redox reaction can take place or not.

These potential values enable the potential difference or *voltage* of any cell to be calculated by the algebraic difference between the potentials of the cathode and the anode.

An **electrolytic cell** is any electrochemical system in which electrical energy is converted into chemical energy.

For example, the *electrolysis of water* decomposes water into its constituent elements.

Quantitative studies conducted on electrolytic processes have led to the enunciation of both *Faraday's Laws of Electrolysis*:

- **Faraday's first law** states that the amount of substances that are formed at the electrodes during an electrolytic process is directly proportional to the amount of electricity that passes through the electrolytic cell;
- **Faraday's second law** states that if you provide the same amount of electricity to different electrolytic cells, varying quantities of substances are formed at the electrodes that are always in proportion to their electrochemical equivalents.

The *electrochemical equivalent* corresponds to the amount of substance that is produced at the electrodes when the quantity of electricity that passes through the circuit is equal to one mole of electrons: this quantity, corresponding to **96 485 C**, is defined as the **Faraday** (**F**).