

Redox Reactions

Redox = "reduction and oxidation"

Oxidation: lose electrons (e^-)

(Classical: any reaction involving oxygen gas (O_2) - rusting, combustion, etc.)

Reduction: gain electrons (e^-)

(Classical: reducing a metal ore into pure metal)

Both reactions **always occur together**

Reducing agent: substance that causes **another substance** to become **reduced**

Oxidizing agent: substance that causes **another substance** to become **oxidized**

Reducing agents **NEVER** reduce themselves; is always **oxidized** to promote reduction

Oxidizing agents **NEVER** oxidize themselves; is always **reduced** to promote oxidation

Reducing/oxidizing agents are **ALWAYS** reactants

Remember: **LEO** the lion says **GER**

Lose Electrons = Oxidation

Gain Electrons = Reduction

Not new reactions: many synthesis, decomposition, combustion and single displacement reactions are often redox reactions

Oxidation States/Oxidation Numbers

Net charge that an atom **would have** if the e^- pairs in covalent bonds belonged **entirely** to the **more electronegative ion**

All redox reactions require a **change in O.N.**

↑ O.N. = oxidation (loss of e^-) ↓ O.N. = reduction (gain e^-)

Rules to determine oxidation numbers (O.N.)

Pure elements $C(s) = 0, O_2(s) = 0, P_4(s) = 0$
have O.N. = 0

Oxidation States/Oxidation Numbers (cont)

Monoatomic ions have $Al^{3+} = +3, Cl^- = -1$
O.N. = their **charge**

Hydrogen **always** has $HCl (H = +1);$
O.N. = +1 (except $H_2S (H = +1);$
metal hydrides = -1) $CaH_2 (H = -1)$

Oxygen **always** has $Li_2O (O = -2);$
O.N. = -2 (except $KNO_3 (O = -2); H$
peroxides = -1) $2O_2 (O = -1)$

In a compound, **groups I, II, and IV** usually have O.N. = **ionic charge** $NaCl (Na = +1, Cl = -1); MgO (Mg = +2)$

In a **neutral** compound¹, **ΣO.N. = 0** $CF_4 (C = +4, F = -1)$

In a polyatomic ion, **ΣO.N. = ion's charge** $NO_3^- (N = 5, O = -2; \Sigma O.N. = -1)$

In molecular compounds with **no O or H**, the $CS_2 (S = -2),$
more electronegative atom has O.N. = its $Li_3N (N = -3)$
usual charge

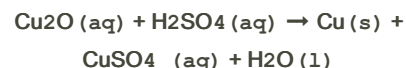
Note: The atoms **do not actually have these charges!**

^[1] If a compound contains a polyatomic ion, the **charge on the other ion** is the **opposite** to the **polyatomic ion's charge** (ex. $KIO_3 - K = +1$ because IO_3 is 1-)

Format for O.N.: "**±#**" (not "**#±**" - ionic charges)

Half-Reactions and Disproportionate Reactions

Most often, one atom is reduced and another is oxidized, but sometimes the same atom can be **oxidized and reduced in the same redox reaction**

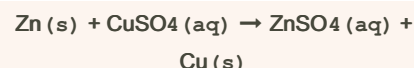


$Cu \rightarrow +1$ to $0 \rightarrow$ gain $1 e^- \rightarrow$ **reduced**

$Cu \rightarrow +1$ to $+2 \rightarrow$ lose $1 e^- \rightarrow$ **oxidized**

Since both of these happen in the **same reaction**, it is **disproportionate**

Half-reaction: reactions made from overall net ionic equations that focus on 1 specific atom



Electrochemical Cells

Two types of electrochemical cells:

Galvanic cells arrangement of 2 connected *half-cells* that **spontaneously** produce an electric current; e^- always flow from high potential \rightarrow low potential

Electrolytic cells arrangement of 2 connected *half-cells* that uses electrical energy to produce a **non-spontaneous** electric current; e^- always flow from low potential \rightarrow high potential

Parts of an electrochemical cell

Electrolyte: solution that contains aqueous ions (**cations** (+) and **anions** (-))



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Electrochemical Cells (cont)

Electrode: solid metal conductor where redox reactions occur (**cathode** (oxidation) and **anode** (reduction))

Salt bridge: tube that **contains an electrolyte solution** and **connects the 2 half-cells**; used to **maintain electrical neutrality**

During the lifespan of the cell, the anode **decreases in mass**, while the cathode **increases in mass**

Cell Potential

The measure of the **electric potential difference** (voltage) between 2 half-cells

Standard cell: galvanic cell in which all entities are at **SATP** and all solution concentrations are **1.0 mol/L**

Standard cell potential (ΔE°): the ability of each half cell to **gain e^-** (reduction)

Cell potential formula:

$$\Delta E^\circ (\text{cell}) = E^\circ \text{cathode} - E^\circ \text{anode}$$

Batteries

Cell: **2 connected electrodes** in contact with an **electrolyte**

Battery: **2 or more** cells connected **in series**; voltage of battery is the **sum of the voltage of all the cells**

Different kinds of batteries are made for **different sized devices**; the **bigger** the battery, the **more electrolytic solution**, and the **longer it lasts**

Alkaline battery: a battery that uses an **alkaline (basic)** electrode rather than an acid

Primary vs secondary cells

Batteries (cont)

Primary cells: **non-rechargeable** cells that run until reactants are **used up** (galvanic cells)

Secondary cells: cells that can be **recharged** by adding an **electric current** (galvanic when being used, electrolytic when being recharged)

Corrosion

The **breakdown/deterioration of metal** by a **redox reaction**

Conditions required for corrosion	Conditions that accelerate corrosion
Oxygen (O_2)	High temperature
Water (H_2O (l))	Salt and/or other electrolytes
	Decrease in pH (more acidity)

Rusting: the corrosion of iron (Fe) **specifically**

Corrosion Prevention

Method 1: Galvanize the metal

Galvanizing: process where a metal (usually steel) is **coated with a thick layer of zinc (Zn)** to prevent corrosion

Method 2: Cathodic protection

Cathodic protection: form of corrosion prevention in which **e^- are continuously supplied** to the metal being protected, making it a **cathode**

Two forms:

Corrosion Prevention (cont)

Sacrificial anode the oxidation of a more active metal attached to the metal being protected prevents the protected metal from corrosion

Impressed current e^- from a direct current (DC) power source are continuously supplied to the protected metal



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