Gr. 12 Electrochemistry Cheat Sheet

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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 **E**lectrons = **O**xidation
Gain **E**lectrons = **R**eduction

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.)

Oxidation States/Oxidation Numbers (cont)

Pure elements have O.N. = 0	$C(s) = 0$, $O_2(s) = 0$, $P_4(s) = 0$
Monoatomic ions have O.N. = their **charge**	$Al^{3+} = +3$, $Cl^- = -1$
Hydrogen **always** has O.N. = +1 (except metal hydrides = -1)	HCl (H = +1); H_2S (H = +1); CaH_2 (H = -1)
Oxygen **always** has O.N. = -2 (except peroxides = -1)	Li_2O (O = -2); KNO_3 (O = -2); H_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** compound[1], **ΣO.N. = 0**	CF_4 (C = +4, F = -1)
In a polyatomic ion, **ΣO.N. = ion's charge**	NO_3^- (N = 5, O = -2; Σ O.N. = -1)
In molecular compunds with **no O** or **H**, the **more electronegative atom** has O.N. = its usual charge	CS_2 (S = -2), Li_3N (N = -3)

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**

$$Cu_2O(aq) + H_2SO_4(aq) \rightarrow Cu(s) + CuSO_4(aq) + H_2O(l)$$

Cu → +1 to 0 → gain 1 e^- → **reduced**

Cu → +1 to +2 → lose 1 e^- → **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

$$Zn(s) + CuSO_4(aq) \rightarrow ZnSO_4(aq) + Cu(s)$$

$Zn(s) \rightarrow ZnSO_4(aq)$	$CuSO_4(aq) \rightarrow Cu(s)$

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 → 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 → high potential

Parts of an electrochemical cell

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

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:

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

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