

### 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  $HCl (H = +1); H_2S (H = +1); CaH_2 (H = -1)$   
O.N. = +1 (except metal hydrides = -1)

Oxygen **always** has  $Li_2O (O = -2); KNO_3 (O = -2); H_2O_2 (O = -1)$   
O.N. = -2 (except peroxides = -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<sup>1</sup>, **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  $CS_2 (S = -2), Li_3N (N = -3)$

**more electronegative atom** has O.N. = its usual charge

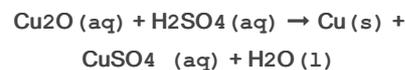
**Note**: The atoms **do not actually have these charges!**

<sup>[1]</sup> 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**

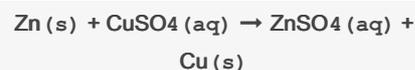


$Cu \rightarrow +1$  to  $0 \rightarrow$  gain  $1 e^- \rightarrow$  **reduced**

$Cu \rightarrow +1$  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* ( $\Delta E^\circ$ ): 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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