

GETTING THE CHARGE OF ANYTHING

- Elements

- If the element has more than 4 electrons, then the charge is equal to “8 – valence number” and it is a **negative charge**
 - O \rightarrow O⁻²
 - F \rightarrow F⁻
 - N \rightarrow N⁻³
 - If the element has less than 4 electrons, then the charge is equal to “valence number” and it is a **positive charge**
 - Na \rightarrow Na⁺
 - Mg \rightarrow Mg⁺²
 - B \rightarrow B⁺³
 - Elements like sulfur, carry no charge

- Groups

- $q_f = V - N - B/2$
 - Where
 - q_f = formal charge
 - V = valence number
 - Number of bonding electrons
 - Number of lone electrons
 - Get q_f for each type of element in a group and then add them all together for example

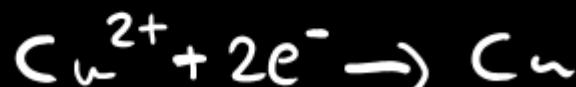
OXIDATION REACTION

It is the loss of one or more electrons.



REDUCTION REACTION

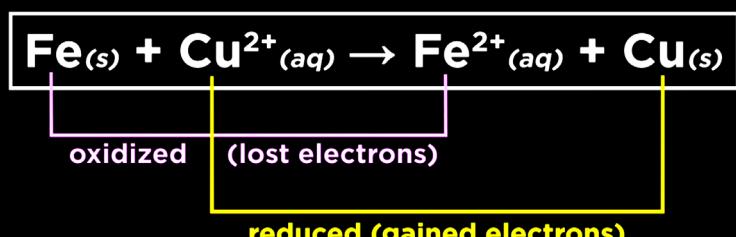
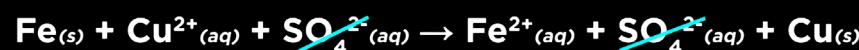
It is the gaining of one or more electrons



An oxidation equation turned backwards represents a reduction

REDOX REACTION

Is a reaction that works by transferring electrons between two metals, an element faces oxidation and loses the electrons which then reduce the other element
Here is a redox reaction, at first it just looks like a single replacement reaction, but once you look at the charges, and cancel out the SO_4^{2-} , you can see that the Fe got oxidized and the electrons reduced the Cu ion



The thing that faces oxidation is called a **reducing agent** because it **reduces** the other thing
The thing that faces reduction, is called an **oxidizing agent** because it **oxidizes** the other thing



We will use **oxidation numbers** to keep track of electron flow, these are hypothetical charge values, they are not actual electrical charges
Because in redox reactions, you need to assign the oxidation numbers for everything

Here are the rules

1. Oxidation number for **elements** is always **zero**, so O₂, H₂, He are 0
2. If it is a monoatomic "A molecule composed of just one atom" **ion** "not element", then the oxidation number is equal to the charge, so an ion like Fe⁺² will have an oxidation number of +2
3. Oxygen in compounds other than peroxides is always -2
4. For a neutral molecule, the oxidation numbers should add up to 0
 1. CO₂ tells us that the oxidation number for carbon is +4 and -2 for each oxygen
Because +4 -2 -2 = 0
5. for polyatomic "molecule consisting of more than 1 atom" ion, the oxidation numbers should add up to the charge of the ion
 1. In sulfate ion (SO₄)⁻², the sulfur atom has an oxidation number of +6+6 because each oxygen atom has an oxidation number of -2, and since there are four oxygen atoms, they contribute a total of (-2)×4=-8. Therefore, the sulfur atom must have a positive oxidation number of +6 to balance out the negative charge from the oxygen atoms and give the ion an overall charge of 2-.
 2. in the ammonium ion (NH₄)⁺. the nitrogen atom has an oxidation number of -3 because each hydrogen atom has an oxidation number of +1, and since there are four hydrogen atoms, they contribute a total of +1×4=+4. Therefore, the nitrogen atom must have a negative oxidation number of -3 to balance out the positive charge from the hydrogen atoms and give the ion an overall charge of +1.
6. Hydrogen is usually +1

To solve a redox reaction, you must balance everything, get the half equations, and then you can extract the reducer and the oxidizer

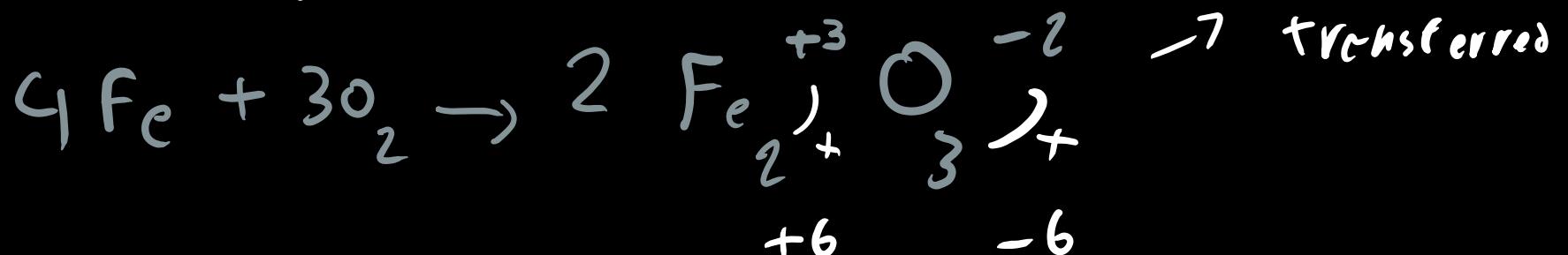
EXAMPLE



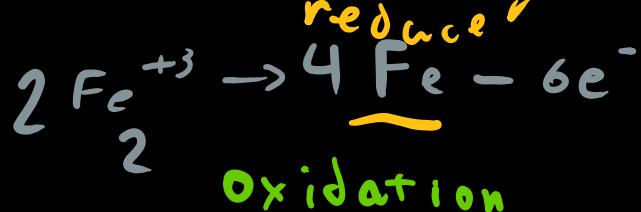
① Balance



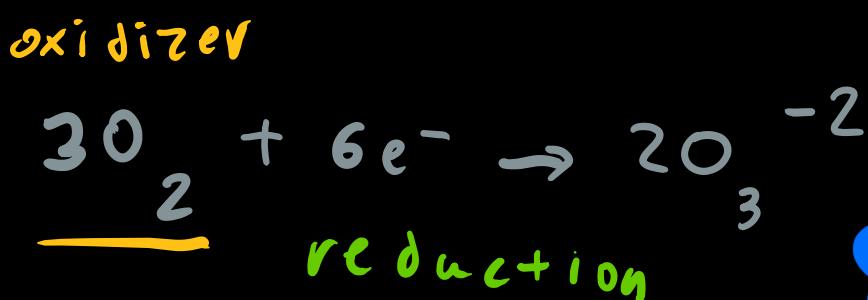
② Split (depending on charge)



③ half equations



oxidation

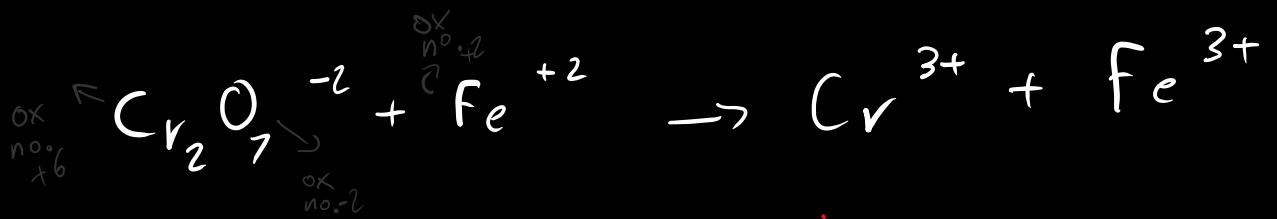


reduction

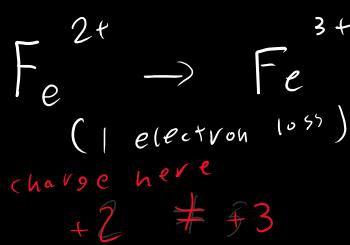


BALANCING COMPLEX REDOX REACTIONS

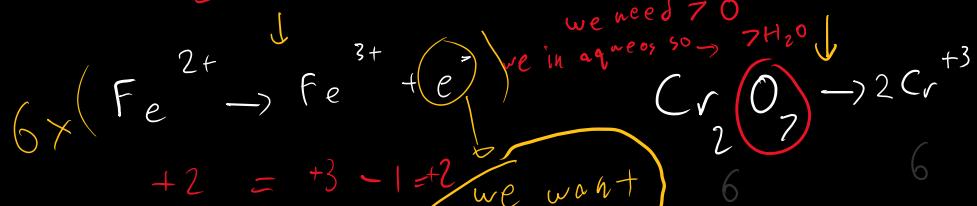
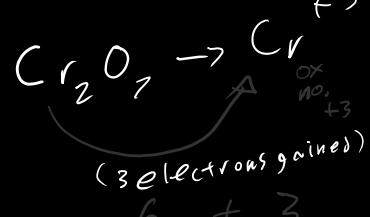
Redox reactions can become much more complicated when if the molecules of a solvent, are involved in the reaction for example : water molecules can involve themselves in redox reactions when it is done in an aqueous solution
in aqueous solution (acidic conditions)



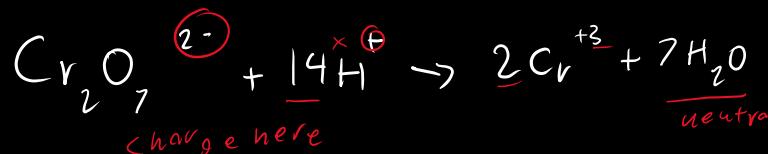
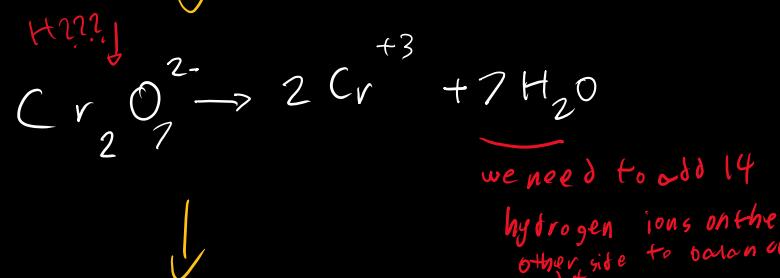
Oxidation



reduction

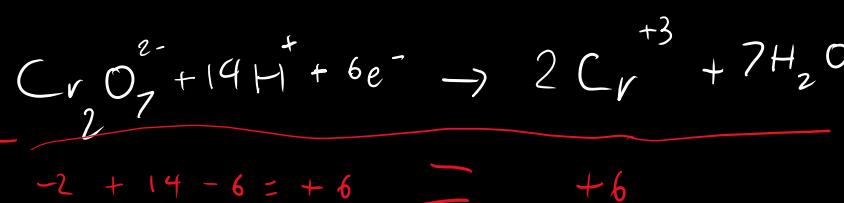


we want to balance those with the reduction equation so we can cancel the electrons



$$-2 + 14(+1) = 12 \neq 2(+3) = +6$$

now add them together for the full equation



EQUATION DONE under ACIDIC conditions

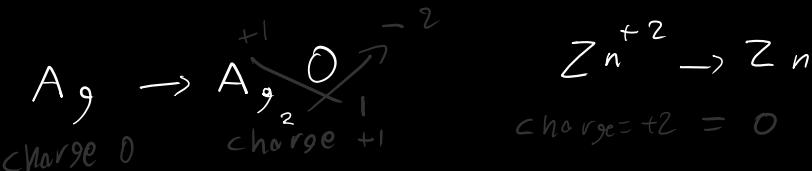
STEPS

1. Split into half reactions
2. On each half do
 1. Balance elements other than H and O
 2. Balance O with water molecules
 3. Balance H with hydrogen ions
 4. Balance Charge with electrons
 5. Make electron numbers equal
3. Combine the 2 equations
4. Cancel out the similarities (probably the electrons)

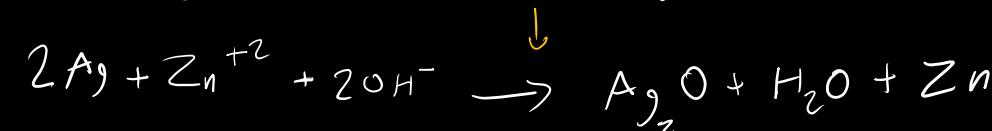
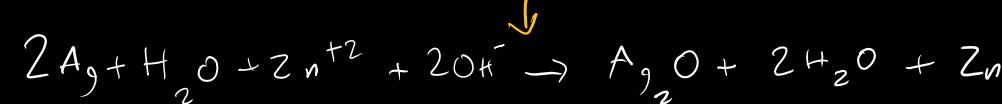
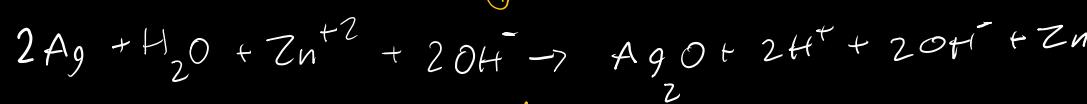
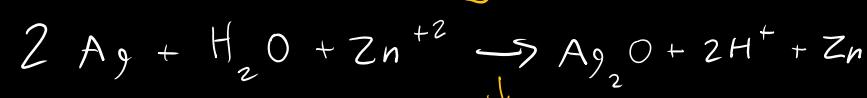
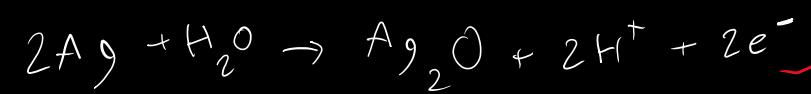
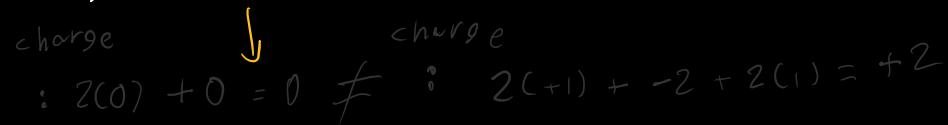
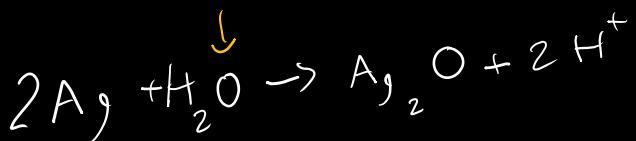
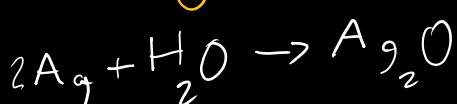
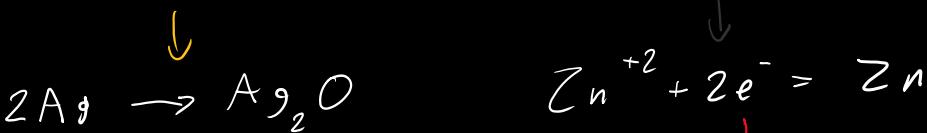
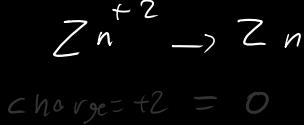
in aqueous solution (basic conditions)



oxidation



reduction



EQUATION DONE

under BASIC conditions

STEPS

1. Split into half reactions
2. On each half do
 1. Balance elements other than H and O
 2. Balance O with water molecules
 3. Balance H with hydrogen ions
 4. Balance Charge with electrons
 5. Make electron numbers equal
3. Combine the 2 equations
4. Cancel out the similarities (probably the electrons)

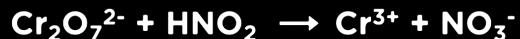
now this would be done if we were working with **ACIDIC** conditions, but basic needs a few more steps

5. Add hydroxides to both sides, with the same coefficient as the hydrogen
6. If a hydroxide and hydrogen exist on the same side, combine them with the same coefficient as one of them
7. Cancel out the water on both sides, but **leave the remainder**
so if there was 3 water on side A and 2 water on side B
side A will have 1 water
side B will have none

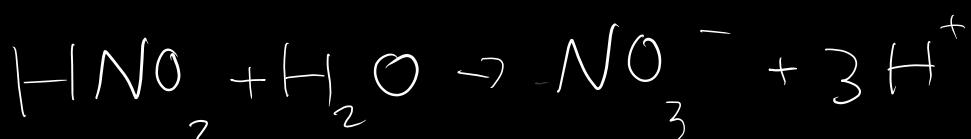
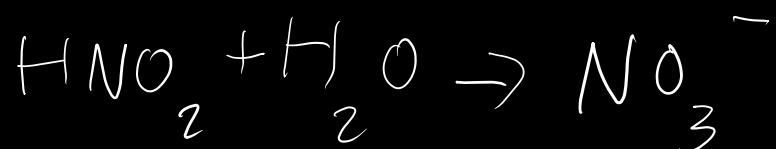


EXERCISE

Balance the following redox reaction under acidic conditions:

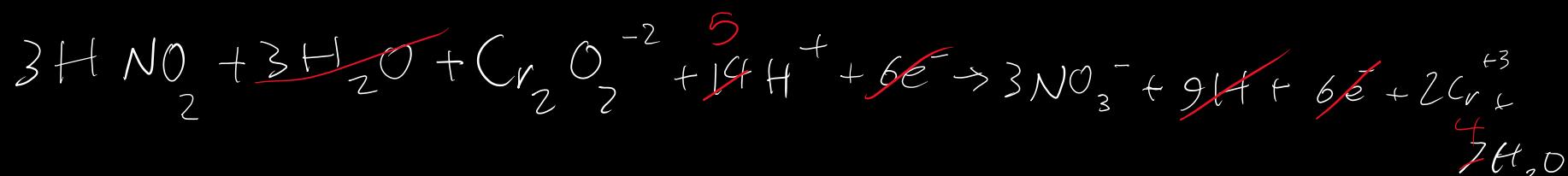
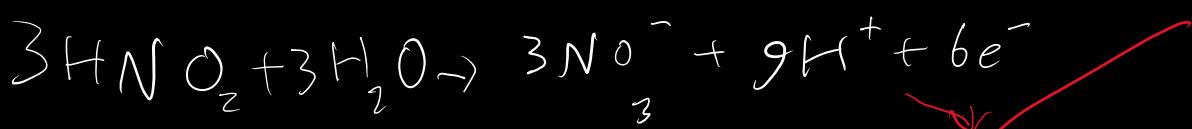


Oxidation

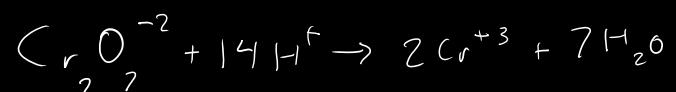


charge 0

charge = 2



Reduction

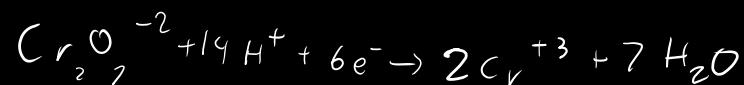


charge :

$$-2 + 14 = 12$$

charge :

$$2(3) = 6$$



Electrochemical cell

It is a device that does redox reactions to transfer electrons which generates an electric current, generating electrical energy from chemical reactions or vice versa

It consists of 4 main parts

1. **Electrolyte** : medium that provides ion transport mechanisms between the two electrodes

- There are two types of electrolytes

- **Strong electrolytes**
 - Those are completely dissociated in their aqueous solution into **cations** and **anions**, they have large conductivity
 - Strong acids & bases
 - Example : **HCl, HNO₃, NaOH, KOH, NaCl, KCl**



NaOH (aq)



NaCl (aq)

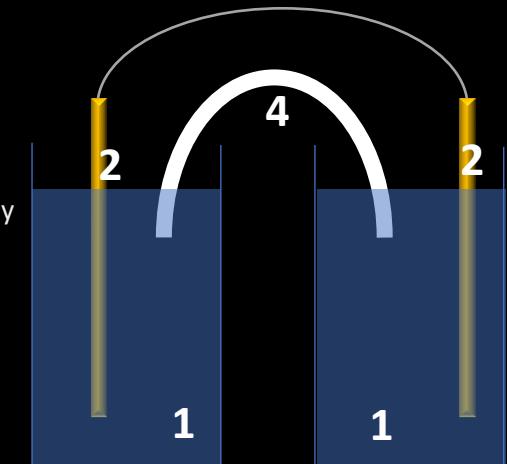
- **Weak electrolytes**

- They do not undergo complete dissociation in their aqueous solution, they are not that conductive, and they exist in an equilibrium of ions and molecules
- For example, the dissociation of acetic acid is shown like this
 - $\text{CH}_3\text{COOH} \rightleftharpoons \text{CH}_3\text{COO}^- + \text{H}^+$
 - \rightleftharpoons this symbol shows a cycle

2. **Electrode** : conductive materials with one being an anode and one being a cathode

3. **External pathway** : made from conductive materials to allow electron flow

4. **Salt bridge** : allows electrons to go in a cycle which won't buildup



FACTORS AFFECTING ELECTROLYTE CONDUCTANCE

Resistance:**r**

Length:**l**

Area:**a**

Where

$$r \propto l,$$

$$r \propto \frac{1}{a}$$

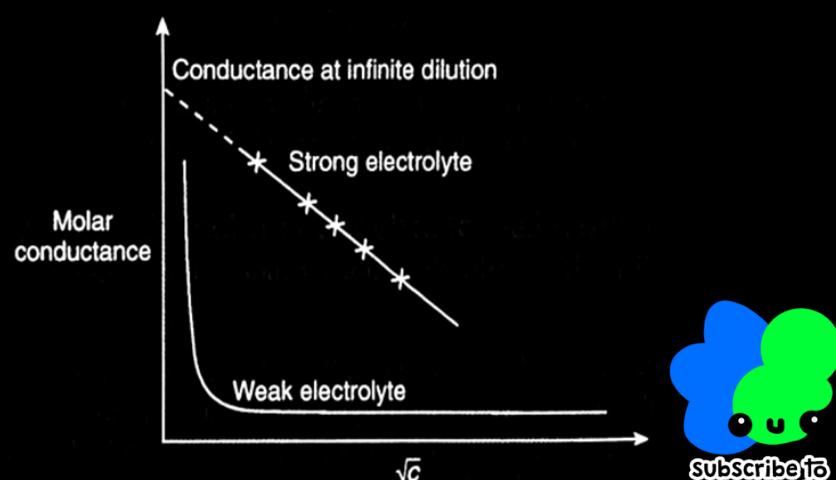
TEMPERATURE AND ELECTROLYTE CONDUCTANCE

In the case of strong electrolytes, increasing in temperature increases the speed of the ions. This results in increased conductivity. In the case of weak electrolytes, an increase in temperature increases the degree of dissociation. There is increase in both the number of ions as well as their speed, hence the conductivity increases.

CONCENTRATION AND ELECTROLYTE CONDUCTANCE

higher concentration -> higher attraction force -> lower speed of electrons -> **lower** conductivity

And vise versa



Galvanic (voltaic) cell

It converts chemical energy from a reaction to electric energy

The electrodes are dipped in a solution of its own ions, these makes current flow from the negative electrode to the positive electrode

where

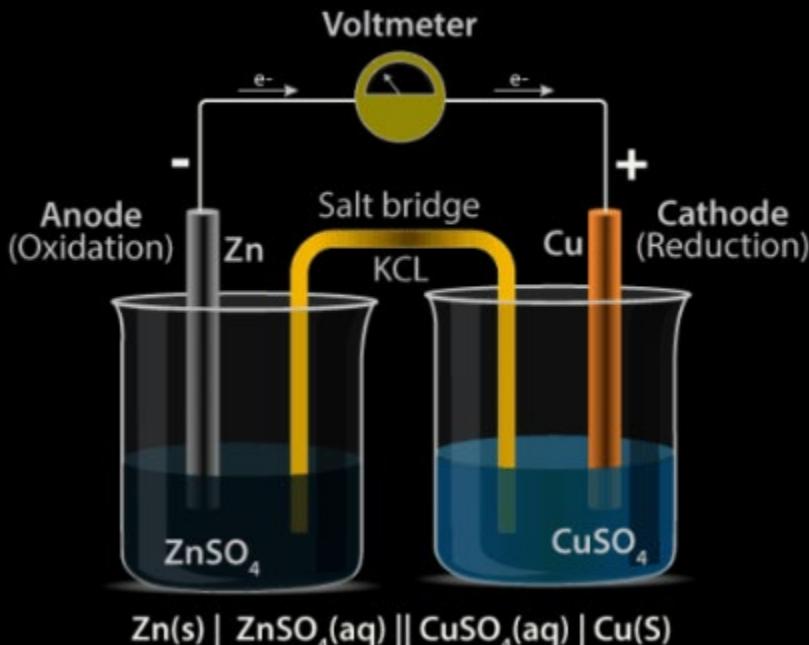
- The anode (has negative charge) electrode faces **oxidation**
- The cathode (has positive charge) electrode faces **reduction**

The most famous galvanic cell is the **Daniel cell** which is based on Zn^{2+} and Cu^{2+}

Potential of electrode

Let's say we have a Zn anode in a $ZnSO_4$ solution

oxidation happens to the Zn electrode, turning it into Zn^{2+} which then dissolves in the solution
the ions travel upwards the electrode, to then go to the cathode

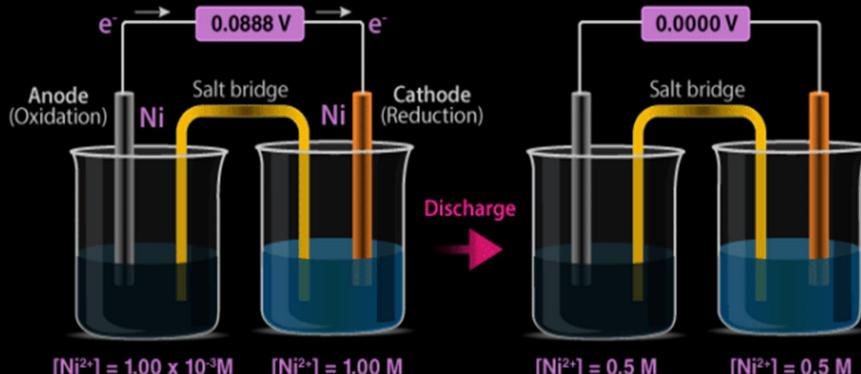


Concentration cell

it works by transferring electrons from the anode to a cathode made from the same material

In this nickel ion solution,

- oxidation will happen to the anode, leading to an increase in positive Nickle ions, which will increase the concentration in the anode compartment
- The electrons will travel to the cathode and into the solution
- The nickel would be reduced, leading to a decrease in the concentration of the solution
- Electricity stops flowing when the concentration is the same



Cell potential (ξ_{cell})

It is the ability of a cell to do electrical work

what does that mean

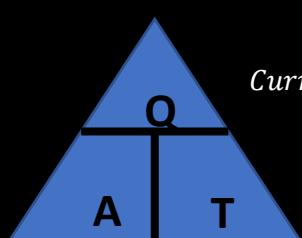
The galvanic cell consist of an oxidizing agent in a compartment and a reducing agent in the other compartment, the "pull" or driving force on the electrons, is called

- Cell potential
- Electromagnetic force (EMF) of the cell

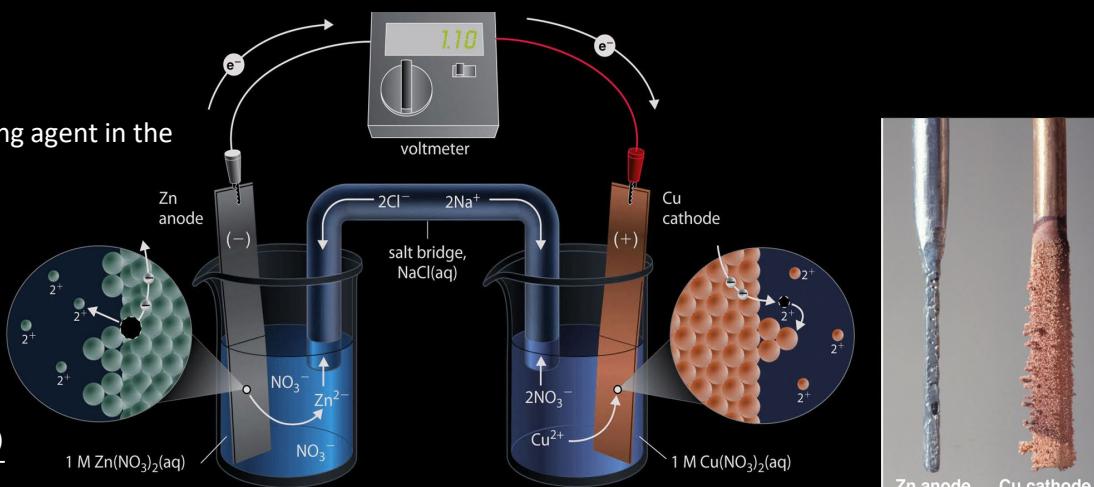
The unit of electrical potential is volts (V) which is defined as

1 joule of work / coulomb of charge transferred

LAWs

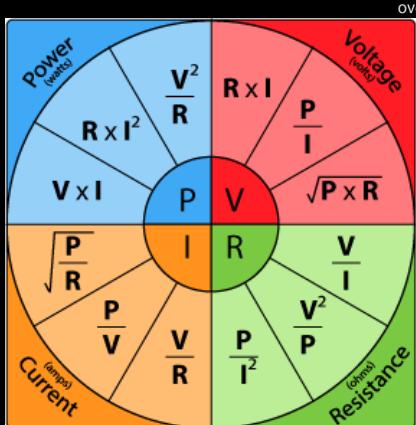
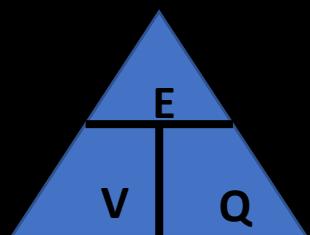


$$\text{Current intensity(ampere)} = \frac{\text{quentity(coulomb)}}{\text{time(second)}}$$



(b)

$$\text{potential difference or emf(volt)} = \frac{\text{work done(joule)}}{\text{quantity(coulomb)}}$$

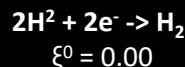


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Standard reduction potential (SRP)

It is the potential that exists between the **metal/gas** and its aqueous solution of unit concentration at standard conditions where the temperature is 298.15k (25C) and the sum of all pressures is 1 atm

All half reductions are compared to the hydrogen half-reactions that has the standard value of 0.00V under standard conditions, so that reaction is a reference



So ξ°_{red} (the STP notation of the standard potential) is the standard potential of 0.00 which is hydrogen, other stuff have their own potential that you can know by viewing this table

So when you get a reaction like this in a galvanic cell



You can get its potential by summing the individual potentials of each half equation

$$\xi^\circ_{\text{cell}} = \xi^\circ_{\text{H}^+ \rightarrow \text{H}_2} + \xi^\circ_{\text{Zn} \rightarrow \text{Zn}^{2+}}$$

$$0.76\text{V} \quad 0.00\text{V} \quad 0.76\text{V}$$

Rules for using the standard reduction potential

- Read the half reaction as written
- The more **POSITIVE** the potential, the greater the tendency is for a substance to be reduced, and therefore a better oxidizing agent
- The half-cell reactions are reversible (oxidation / reduction), to do that, change the **sign** of the ξ°_{cell}
- The ξ°_{cell} stays the same even if you change coefficients
- Under STP conditions, any species on the **left** of the half reaction, will react spontaneously with a species that is on the **RIGHT AND ABOVE IT**
- the biggest positive value for a ξ°_{cell} means that they are the strongest OXIDIZING AGENTS, and there fore are themselves reduced
- the biggest negative value for a ξ°_{cell} means that they are the strongest REDUCING AGENTS, and there fore are themselves oxidized

SALT BRIDGES

Without a salt bridge, electrons are gonna keep pouring into one side leading to inevitable polarization which will cause the cell to stop work

The negatively charged ions from salt bridges are needed to balance the solution on the anode and the positively charged ions from salt bridges are needed to balance the solution on the cathode

LINE NOTATION

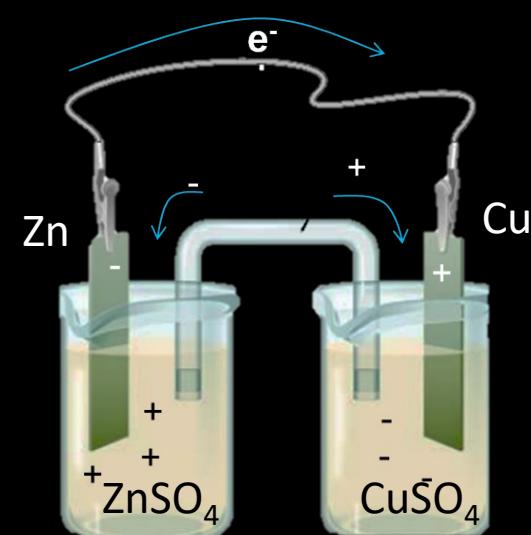
Here the

- Anode components are listed on the left
- cathode components are listed on the right
- The anode cathode components are separated by double vertical lines which indicate salt bridges
- A phase difference is indicated by a single vertical line,
 - Anode : it occurs between the Solid Zn metal and the Zn^{2+} in aqueous solution
 - Cathode : it occurs between the Solid Cu metal and the Cu^{2+} in aqueous solution



TABLE 18.1 | Standard Reduction Potentials at 25°C (298 K) for Many Common Half-Reactions

Half-Reaction	ξ° (V)	Half-Reaction	ξ° (V)
$\text{F}_2 + 2\text{e}^- \rightarrow 2\text{F}^-$	2.87	$\text{O}_2 + 2\text{H}_2\text{O} + 4\text{e}^- \rightarrow 4\text{OH}^-$	0.40
$\text{Ag}^{2+} + \text{e}^- \rightarrow \text{Ag}^+$	1.99	$\text{Cu}^{2+} + 2\text{e}^- \rightarrow \text{Cu}$	0.34
$\text{Co}^{3+} + \text{e}^- \rightarrow \text{Co}^{2+}$	1.82	$\text{Hg}_2\text{Cl}_2 + 2\text{e}^- \rightarrow 2\text{Hg} + 2\text{Cl}^-$	0.27
$\text{H}_2\text{O}_2 + 2\text{H}^+ + 2\text{e}^- \rightarrow 2\text{H}_2\text{O}$	1.78	$\text{AgCl} + \text{e}^- \rightarrow \text{Ag} + \text{Cl}^-$	0.22
$\text{Ce}^{4+} + \text{e}^- \rightarrow \text{Ce}^{3+}$	1.70	$\text{SO}_4^{2-} + 4\text{H}^+ + 2\text{e}^- \rightarrow \text{H}_2\text{SO}_3 + \text{H}_2\text{O}$	0.20
$\text{PbO}_2 + 4\text{H}^+ + \text{SO}_4^{2-} + 2\text{e}^- \rightarrow \text{PbSO}_4 + 2\text{H}_2\text{O}$	1.69	$\text{Cu}^{2+} + \text{e}^- \rightarrow \text{Cu}^+$	0.16
$\text{MnO}_4^- + 4\text{H}^+ + 3\text{e}^- \rightarrow \text{MnO}_2 + 2\text{H}_2\text{O}$	1.68	$2\text{H}^+ + 2\text{e}^- \rightarrow \text{H}_2$	0.00
$2\text{e}^- + 2\text{H}^+ + \text{IO}_4^- \rightarrow \text{IO}_3^- + \text{H}_2\text{O}$	1.60	$\text{Fe}^{3+} + 3\text{e}^- \rightarrow \text{Fe}$	-0.036
$\text{MnO}_4^- + 8\text{H}^+ + 5\text{e}^- \rightarrow \text{Mn}^{2+} + 4\text{H}_2\text{O}$	1.51	$\text{Pb}^{2+} + 2\text{e}^- \rightarrow \text{Pb}$	-0.13
$\text{Au}^{3+} + 3\text{e}^- \rightarrow \text{Au}$	1.50	$\text{Sn}^{2+} + 2\text{e}^- \rightarrow \text{Sn}$	-0.14
$\text{PbO}_2 + 4\text{H}^+ + 2\text{e}^- \rightarrow \text{Pb}^{2+} + 2\text{H}_2\text{O}$	1.46	$\text{Ni}^{2+} + 2\text{e}^- \rightarrow \text{Ni}$	-0.23
$\text{Cl}_2 + 2\text{e}^- \rightarrow 2\text{Cl}^-$	1.36	$\text{PbSO}_4 + 2\text{e}^- \rightarrow \text{Pb} + \text{SO}_4^{2-}$	-0.35
$\text{Cr}_2\text{O}_7^{2-} + 14\text{H}^+ + 6\text{e}^- \rightarrow 2\text{Cr}^{3+} + 7\text{H}_2\text{O}$	1.33	$\text{Cd}^{2+} + 2\text{e}^- \rightarrow \text{Cd}$	-0.40
$\text{O}_2 + 4\text{H}^+ + 4\text{e}^- \rightarrow 2\text{H}_2\text{O}$	1.23	$\text{Fe}^{2+} + 2\text{e}^- \rightarrow \text{Fe}$	-0.44
$\text{MnO}_2 + 4\text{H}^+ + 2\text{e}^- \rightarrow \text{Mn}^{2+} + 2\text{H}_2\text{O}$	1.21	$\text{Cr}^{3+} + \text{e}^- \rightarrow \text{Cr}^{2+}$	-0.50
$\text{IO}_3^- + 6\text{H}^+ + 5\text{e}^- \rightarrow \text{I}_2 + 3\text{H}_2\text{O}$	1.20	$\text{Cr}^{3+} + 3\text{e}^- \rightarrow \text{Cr}$	-0.73
$\text{Br}_2 + 2\text{e}^- \rightarrow 2\text{Br}^-$	1.09	$\text{Zn}^{2+} + 2\text{e}^- \rightarrow \text{Zn}$	-0.76
$\text{VO}_2^+ + 2\text{H}^+ + \text{e}^- \rightarrow \text{VO}^{2+} + \text{H}_2\text{O}$	1.00	$2\text{H}_2\text{O} + 2\text{e}^- \rightarrow \text{H}_2 + 2\text{OH}^-$	-0.83
$\text{AuCl}_4^- + 3\text{e}^- \rightarrow \text{Au} + 4\text{Cl}^-$	0.99	$\text{Mn}^{2+} + 2\text{e}^- \rightarrow \text{Mn}$	-1.18
$\text{NO}_3^- + 4\text{H}^+ + 3\text{e}^- \rightarrow \text{NO} + 2\text{H}_2\text{O}$	0.96	$\text{Al}^{3+} + 3\text{e}^- \rightarrow \text{Al}$	-1.66
$\text{ClO}_2 + \text{e}^- \rightarrow \text{ClO}_2^-$	0.954	$\text{H}_2 + 2\text{e}^- \rightarrow 2\text{H}^-$	-2.23
$2\text{Hg}^{2+} + 2\text{e}^- \rightarrow \text{Hg}_2^{2+}$	0.91	$\text{Mg}^{2+} + 2\text{e}^- \rightarrow \text{Mg}$	-2.37
$\text{Ag}^+ + \text{e}^- \rightarrow \text{Ag}$	0.80	$\text{La}^{3+} + 3\text{e}^- \rightarrow \text{La}$	-2.37
$\text{Hg}_2^{2+} + 2\text{e}^- \rightarrow 2\text{Hg}$	0.80	$\text{Na}^+ + \text{e}^- \rightarrow \text{Na}$	-2.71
$\text{Fe}^{3+} + \text{e}^- \rightarrow \text{Fe}^{2+}$	0.77	$\text{Ca}^{2+} + 2\text{e}^- \rightarrow \text{Ca}$	-2.76
$\text{O}_2 + 2\text{H}^+ + 2\text{e}^- \rightarrow \text{H}_2\text{O}_2$	0.68	$\text{Ba}^{2+} + 2\text{e}^- \rightarrow \text{Ba}$	-2.90
$\text{MnO}_4^- + \text{e}^- \rightarrow \text{MnO}_4^{2-}$	0.56	$\text{K}^+ + \text{e}^- \rightarrow \text{K}$	-2.92
$\text{I}_2 + 2\text{e}^- \rightarrow 2\text{I}^-$	0.54	$\text{Li}^+ + \text{e}^- \rightarrow \text{Li}$	-3.05



LET'S EXPLAIN THESE

1. Here it gives the equation for the E° cell's potential which is equal to the sum of the E° oxidation potential and E° reduction potential"
2. It represents the change in free energy (energy left in a cell before we can't get anymore)
3. It represents the change in free energy (energy left in a cell before we can't get anymore)
4. It represents the change in free energy of a reaction at a specific time (not necessarily an equilibrium)
5. Calculate the cell potential under non-standard conditions

Facts

- **Spontaneous Reactions:** A negative change in free energy ($\Delta G < 0$) indicates a spontaneous reaction, meaning it will naturally proceed without external input.



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- **Non-Spontaneous Reactions:** A positive change in free energy ($\Delta G > 0$) indicates a non-spontaneous reaction, meaning it won't occur on its own.



- **Equilibrium:** A ΔG of zero indicates the system is at equilibrium, where the forward and reverse reactions are happening at the same rate.



EMF

The difference in electrode potential value of the two electrodes of a cell in an open circuit. An open circuit means that no current flows through the circuit

Potential difference

If the difference in potential between the electrodes of the same cell is measured in a closed circuit, it is termed as potential difference. A closed circuit means an activity in which measurement is carried out under load

While measuring EMF, the cell reaction would be in equilibrium. Hence, electrical work done is the maximum

As the cell reaction is not in equilibrium when the potential difference is measured. The electric work done is less.

EMF is usually determined by potentiometric methods like a device named potentiometer

A simple device like a voltmeter is enough to measure the potential difference

EMF is the maximum potential available in a galvanic cell

The potential difference of the same cell is always less than the EMF

Laws of work, potential and free energy

- The following slides consist only of the laws without any discussion or proof.
- $E^\circ_{cell} = E^\circ_{reduction} + E^\circ_{oxidation}$
- $\Delta G^\circ = -RT \ln(K)$
- $\Delta G^\circ = -nFE^\circ$
- $\Delta G = \Delta G^\circ + RT \ln(Q)$
- $E = E^\circ - \frac{RT}{nF} \ln(Q)$
- $Q = \frac{[Products]}{[Reactants]}$

Faraday constant: "F"

96 485 . 332 9 s . A / mole

Universal gas constant:

"R"

8.3145 J / mole . K

General Symbols:

- Δ (Delta): This symbol denotes change in a quantity.
- $^\circ$ (Superscript degree sign): This indicates a standard state condition (usually temperature of 25°C and pressure of 1 atm).

Thermodynamic Properties:

- E (Electromotive force): Represents voltage or potential difference.
- G (Gibbs free energy): Represents the energy available to do non-expansion work in a system.
- R (Universal gas constant): A constant value relating pressure, volume, and temperature of a gas.
- T (Temperature): Typically measured in Kelvin (K).
- n (Number of electrons transferred): The number of electrons involved in a reaction.
- F (Faraday constant): A constant relating the amount of electric charge to the number of electrons.

Equilibrium and Concentrations:

- K (Equilibrium constant): The ratio of product concentrations to reactant concentrations at equilibrium.
- Q (Reaction quotient): The ratio of product concentrations to reactant concentrations at any given time (not necessarily equilibrium).

Specific to Cell Potential:

- E°_{cell} (Standard cell potential): The potential difference between anode and cathode under standard conditions.
- $E^\circ_{reduction}$ (Standard reduction potential): Potential difference between an oxidation half-reaction and a standard hydrogen electrode.
- $E^\circ_{oxidation}$ (Standard oxidation potential): Potential difference between a reduction half-reaction and a standard hydrogen electrode.



Formula equation:

All reactants and products are written as complete formulas

Complete ionic equation:

All reactants and products that are strong electrolytes are written as separated ions

Net ionic equation:

Only those compounds that undergo a change are written; spectator ions are not included

SPECTATOR IONS

They are ions that are present in a chemical reaction but don't actually participate, they are the ions that we cancel out on both sides in reactions like redox

- To understand the difference between the three types of equation study the following example:
 - $NaCl(aq) + AgNO_3(aq) \rightarrow AgCl(s) + NaNO_3(aq)$ "formula equation"
 - $Na^+ + Cl^- + Ag^+ + NO_3^- \rightarrow Na^+ + NO_3^- + AgCl(s)$ "complete ionic equation"
 - Spectator ions: ions that are present in the reactants and the products (does not change), then:
 - $\cancel{Na^+} + Cl^- + Ag^+ + \cancel{NO_3^-} \rightarrow \cancel{Na^+} + \cancel{NO_3^-} + AgCl(s)$
 - $Cl^- + Ag^+ \rightarrow AgCl(s)$ "net ionic equation"

Here, the NO_3 and Na ions are spectators

