TOPIC 10.1

Defining Oxidation and Reduction

How can we keep track of electron loss/gain by elements?

To fully understand the reactions we'll be discussing this unit, it is essential to understand atomic structure and bonding behavior. Let's start things off by looking at atoms and ions:

To get from **Al to Al^{3+}**, what has to happen? lose 3 electrons

- Draw a diagram:

$$\dot{A}l\cdot \xrightarrow{3e^-} [Al]^{3+}$$

- This is called: OXIDATION ("charge" goes up)

To get from **Cu^{2+} to Cu**, what has to happen? gain 2 electrons

- Draw a diagram:

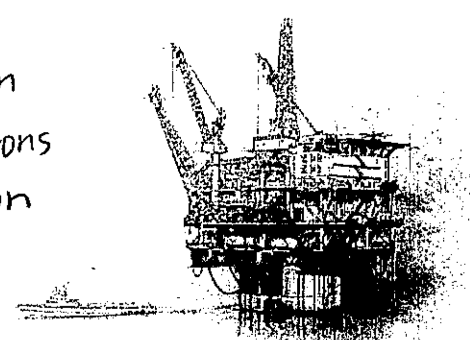
$$[Cu]^{2+} + 2e^- \longrightarrow \dot{C}u\cdot$$

- This is called: REDUCTION ("charge" goes down)

<u>SUMMARIZE</u>:

Process	Change in Electrons	Change in "Charge" (**<u>Oxidation State</u>**)
Oxidation (oxidized)	loss	up
Reduction (reduced)	gain	down

Losing Electrons is Oxidation
Gaining Electrons is Reduction



OIL RIG

Oxidation Is Losing
Reduction Is Gaining

TOPIC 10.2

Half Reactions

How can we represent oxidation and reduction separately?

A **redox reaction** is defined as a chemical reaction that involves the transfer of electrons. Let's be a bit more specific:

Redox reaction: chemical change where one reactant species is oxidized and another is reduced.

EXAMPLE: $Al + Cu^{2+} \rightarrow Al^{3+} + Cu$

Let's represent the oxidation and reduction that are occurring separately by writing half reactions.

1. Which species is being reduced in the example redox reaction? Cu^{2+}

   How do you know, in terms of charge? Charge goes down

   How do you know, in terms of electrons? gains electrons

2. Which species is being oxidized in the example redox reaction? Al

   How do you know, in terms of charge? Charge goes up

   How do you know, in terms of electrons? loses electrons

Reduction half – reaction:

Picture	*Equation*
$[Cu]^{2+} + 2e^- \longrightarrow \dot{C}u\cdot$	$\underbrace{Cu^{2+} + 2e^-}_{\text{total charge} = 0} \longrightarrow \underbrace{Cu}_{\substack{\text{neutral atom} \\ \text{total charge} = 0}}$

Oxidation half – reaction:

Picture	*Equation*
$\dot{A}l\cdot \longrightarrow [Al]^{3+} + 3e^-$	$\underbrace{Al}_{\substack{\text{neutral atom} \\ \text{total charge} = 0}} \longrightarrow \underbrace{Al^{3+} + 3e^-}_{\text{total charge} = 0}$

Is the example redox reaction balanced? Consider what has to be conserved in all chemical changes...

$$\underbrace{Al + Cu^{2+}}_{\text{total charge} = +2} \rightarrow \underbrace{Al^{3+} + Cu}_{\text{total charge} = +3}$$

No! Charge is not conserved...

TOPIC 10.3

Balancing Half Reactions

How can we add up half reactions to show a complete redox reaction?

Recall that the redox reaction from the previous lesson was not balanced….let's try writing a balanced redox reaction from the two half-reactions we wrote in the last mini-lesson:

Reduction half: $[Cu^{2+} + 2e^- \rightarrow Cu] \times 3$

$= 3\,Cu^{2+} + 6\,e^- \rightarrow \boxed{3\,Cu}$

Oxidation half: $[Al \rightarrow Al^{3+} + 3e^-] \times 2$

$= 2\,Al \rightarrow \boxed{2Al^{3+} + 6e^-}$

Redox reaction: $3\,Cu^{2+} + \cancel{6}e^- + 2Al \rightarrow 2\,Al^{3+} + \cancel{6e^-} + 3Cu$

add up the 2 ½ reactions!

cancel out electrons

$$\boxed{3\,Cu^{2+} + 2Al \rightarrow 2\,Al^{3+} + 3\,Cu}$$

total charge = +6

total charge = +6 ✓

In a **redox** reaction, the **number of (moles of) e⁻** gained is **ALWAYS** EQUAL the **number of (moles of) e⁻** lost.

TOPIC 10.4

Identifying Redox Reactions

How can we tell if a redox reaction is occurring?

redox reaction involves the transfer of electrons. Somebody loses them, and somebody else gains them. And just to recap – the moles of electrons lost should equal the moles of electrons gained. So with all this change in electrons, what happens to oxidation state (or charge)? It changes as well! Remember this nifty table?

Process	Change in Electrons	Change in "Charge" (**Oxidation State**)
Oxidation (oxidized)	loss	up
Reduction (reduced)	gain	down

Earlier in this unit, you worked on a POGIL to determine the rules for assigning oxidation numbers. Let's see what you can do...

ZnF_2

Zn: +2 F: −1

F_2

F: 0

CuO

Cu: +2 O: −2

+4 −2 −2

CO_2

C: +4 O: −2

+1 +1 +4 −2 −2 −2

H_2CO_3

H: +1 C: +4 O: −2

x −2

SO_4^{2-}

x + 4(−2) = −2

x − 8 = −2
+8 +8
x = 6

S: +6 O: −2

So, if there's been a change in oxidation number for a specific element, that indicates a redox reaction has occurred!

Reaction	Redox?			Reaction Type
$2H_2 + O_2 \rightarrow 2H_2O$ (0, 0 → +1 +1 −2)	Y	/	N	synthesis
$H_2CO_3 \rightarrow H_2O + CO_2$ (+1 +1 +4 −2 −2 −2 → +1 +1 −2 +4 −2 −2)	Y	/	N	decomposition
$Mg + 2HBr \rightarrow H_2 + MgBr_2$ (0 +1 −1 → 0 +2 −1 −1)	Y	/	N	single replacement (ALWAYS redox)
$2KBr + CuF_2 \rightarrow 2KF + CuBr_2$ (+1 −1 +2 −1 −1 → +1 −1 +2 −1 −1)	Y	/	N	double replacement (NEVER redox)

Synthesis and decomposition: sometimes redox

TOPIC 10.5

Table J and Spontaneous Reactions

How can we predict whether a redox reaction will take place or not?

Before we can start predicting if a redox reaction will occur spontaneously (happen naturally), we have to talk about **element activity**.

Element activity: how likely an element is to react (gain or lose e^-)

Summarize:

- More **active metals** are likely to easily lose (*gain / lose ?*) electrons, so they have low (*low / high ?*) electronegativity values. This means more active metals get oxidized (*oxidized / reduced ?*).

- More **active nonmetals** are likely to easily gain (*gain / lose ?*) electrons, so they have high (*low / high ?*) electronegativity values. This means more active nonmetals get reduced (*oxidized / reduced ?*).

Table J
Activity Series**

Most Active ↑ easy ox. (Metals) — Most Active ↑ easy red. (Nonmetals)

Metals	Nonmetals
Li	F_2
Rb	Cl_2
K	Br_2
Cs	I_2
Ba	
Sr	
Ca	
Na	
Mg	
Al	
Ti	
Mn	
Zn	
Cr	
Fe	
Co	
Ni	
Sn	
Pb	
H_2	
Cu	
Ag	
Au	

Least Active — Least Active

**Activity Series is based on the hydrogen standard. H_2 is *not* a metal.

A single-replacement redox reaction will take place **spontaneously** if a MORE active element is able to "replace" a LESS active ION (aq).

1. $Mg + AgNO_3 \rightarrow Mg(NO_3)_2 + Ag$ — *SPONTANEOUS* / *NONSPONTANEOUS*
2. $Mg + LiNO_3 \rightarrow Mg(NO_3)_2 + Li$ — *SPONTANEOUS* / *NONSPONTANEOUS*
3. $Cl_2 + NaBr \rightarrow Br_2 + NaCl$ — *SPONTANEOUS* / *NONSPONTANEOUS*
4. $Na + H_2O \rightarrow Na_2O + H_2$ — *SPONTANEOUS* / *NONSPONTANEOUS*

$ZnSO_4 + H_2 \rightarrow H_2SO_4 + Zn$ — *SPONTANEOUS* / *NONSPONTANEOUS*

TOPIC 10.6

Voltaic Cells

How is electricity generated in a battery (voltaic cell)?

Voltaic cell: simple battery that spontaneously converts chemical energy into electrical energy

3 Primary Components:

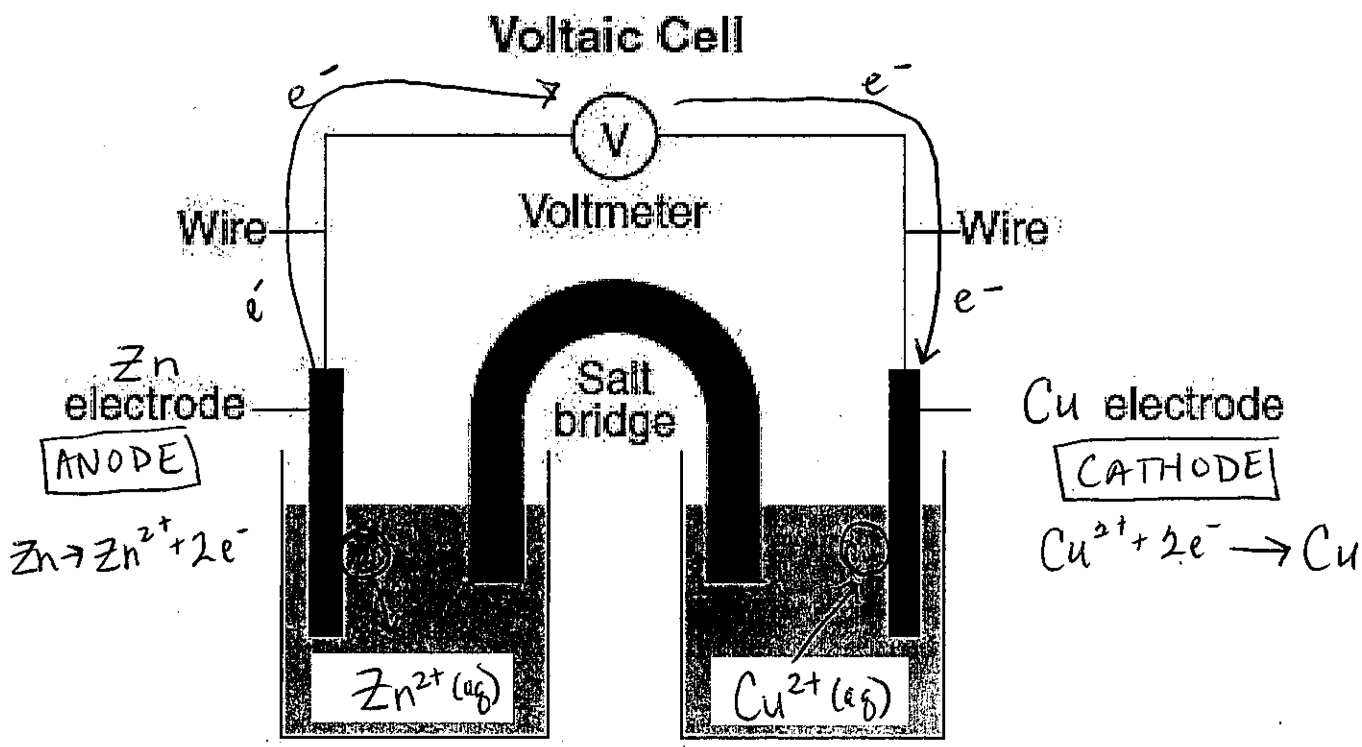
1) Positive terminal/electrode:
metal # 1 (CATHODE)

2) Negative terminal/electrode:
metal # 2 (ANODE)

3) Electrolyte exchange:
salt bridge (allows ions to move and therefore prevents charge build-up)

How does it work? Recall that an electric current requires the flow of charged particles. (ions or e^-)

AN OX and a **BIG RED CAT**



What is the "AN OX"?

Anode gets OXidized (metal that is higher on Table J)

What is the "BIG RED CAT"?

- Reduction occurs at the cathode
- "big" b/c mass increases as ions (aq) turn into solid atoms

TOPIC 10.7

Electrolytic Cells

How can we force an electrochemical cell to run in "reverse"?

Electrolytic cell: uses electrical energy to bring about a NONspontaneous chemical reaction

★ converts electrical energy to chemical energy ★

Important Distinctions from Voltaic Cells:

- Usually no salt bridge
- 1 aqueous solution (instead of 2)
- ALWAYS need a power source in order to power a **nonspontaneous** chemical change
- The cathode and anode are OPPOSITE

↓ oxidation still happens here, but the metal is lower on Table J

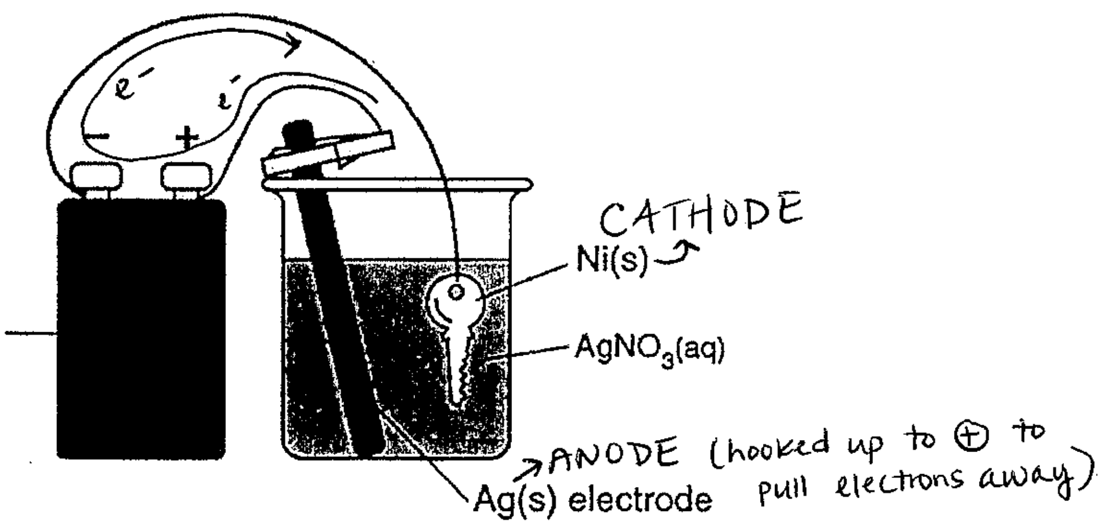
Uses of Electrolytic Cells

- Electroplating (coating a thin layer of metal over another metal)

The diagram below represents an operating electrolytic cell used to plate silver onto a nickel key. As the cell operates, oxidation occurs at the silver electrode and the mass of the silver electrode decreases.

Ni vs Ag

↑ lower on Table J, so it's the anode



Explain, in terms of Ag atoms and $Ag^+(aq)$ ions, why the mass of the silver electrode decreases as the cell operates.

Solid silver atoms turn into aqueous ions and get dispersed in the solution.

$Ag \rightarrow Ag^{2+} + 2e^-$