# N42 - Electrochemistry

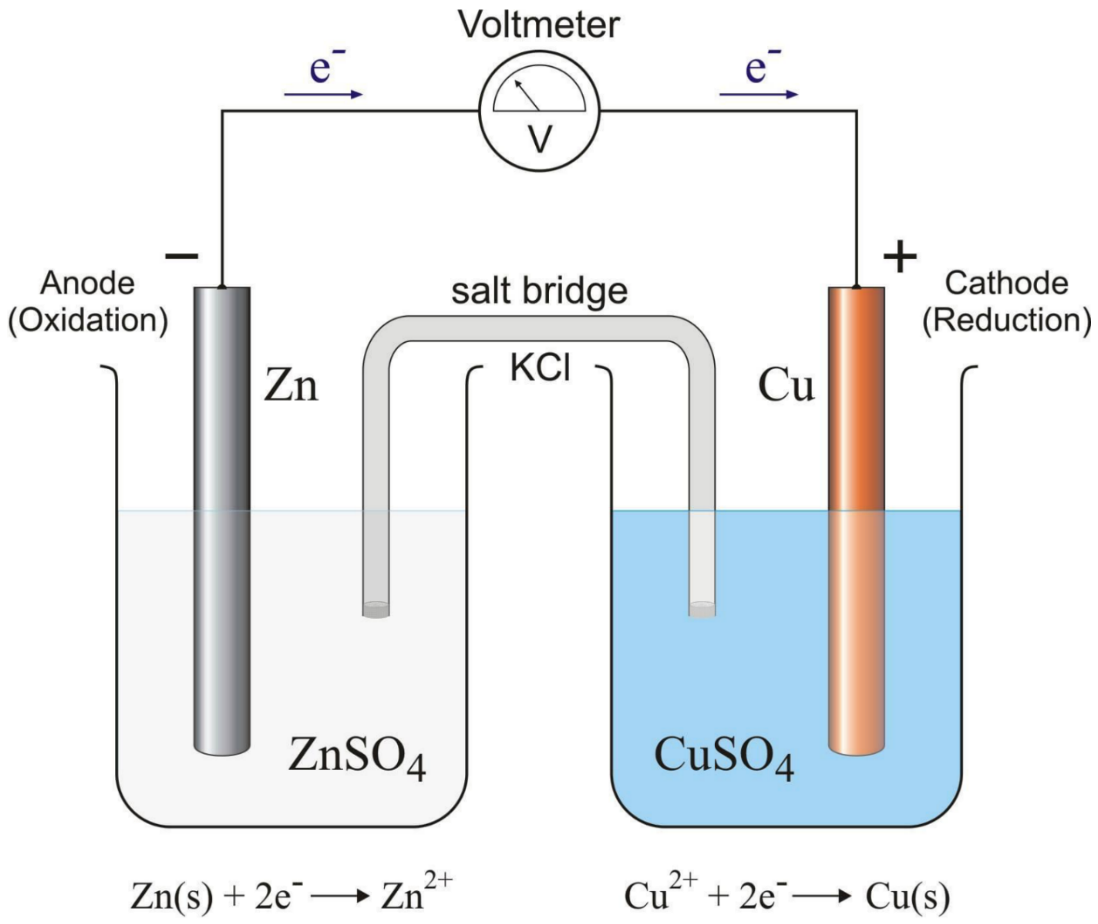
## Concepts

**Link to YouTube Presentation:** https://youtu.be/-y2xTX_BVsA

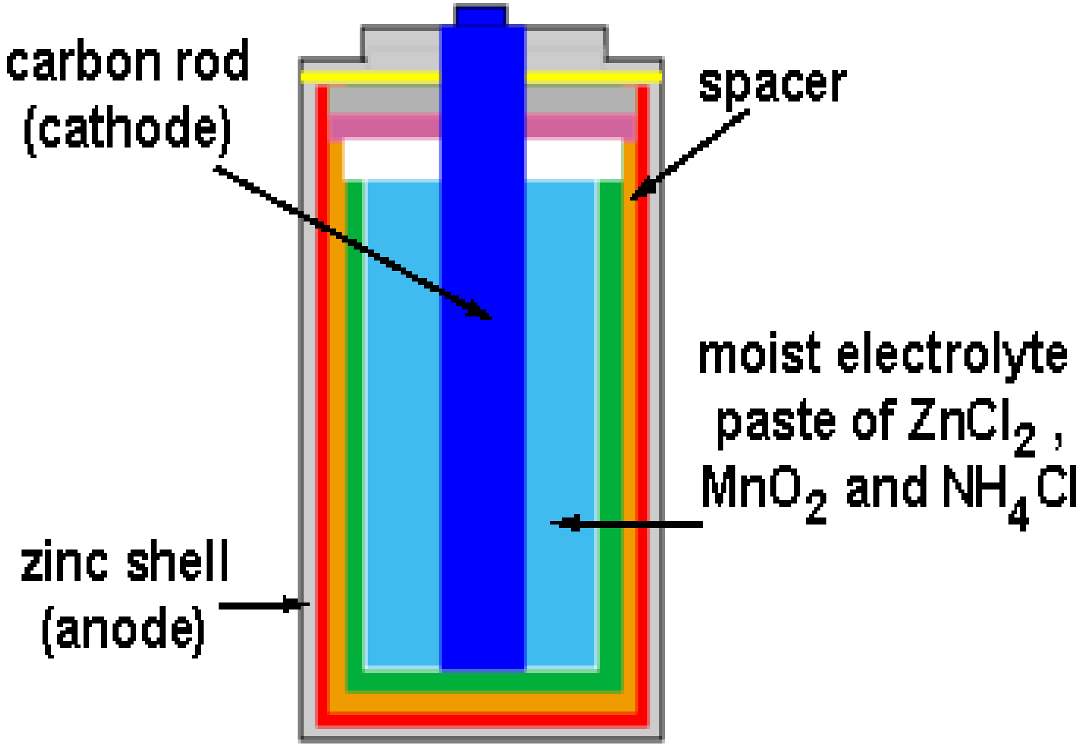
# N42 - Electrochemistry

## Concepts

**Target:** I can assign oxidation numbers and balance redox reactions.




$Zn(s) \mid ZnSO_4(aq) \parallel CuSO_4(aq) \mid Cu(s)$



# Electrochemistry

**The study of the interchange of chemical and electrical energy.**

**Concerned with:**

- **the generation of an electrical current from a spontaneous chemical reaction**
- **the use of an electrical current to produce a non-spontaneous chemical change**

# Mnemonics


## LEO goes GER

**Loss of Electrons is Oxidation**

**Gain of Electrons is Reduction**


## OIL RIG

**Oxidation is Loss of Electrons**

**Reduction is Gain of Electrons**

# Examples

## Oxidation

$Ag \rightarrow Ag^+ + e^-$

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

## Reduction

$Ag^+ + e^- \rightarrow Ag$

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

# Oxidation Numbers

**The charge an atom in a compound would have IF the compound was composed of ions.**

**Helps track how the electrons are moving around during a reaction.**

**Sometimes easy to determine, sometimes complex.**

$KF \rightarrow K^+ + F^-$ K ox # = +1, F ox # = -1

# Rules for Assigning Oxidation Numbers


1. Any uncombined element is 0.
2. Monatomic ion equals the charge on the ion.
3. The more-electronegative element in a binary compound is assigned the number equal to the charge it would have if it were an ion.
4. Fluorine in a compound is always -1
5. Oxygen is -2 unless it is combined with F, when it is +2, or it is in a peroxide, such as $H_2O_2$, when it is -1

6. Hydrogen in most of its compounds is +1 unless it is combined with a metal, in which case it is -1
7. In compounds, the elements of groups 1 and 2 as well as aluminum have oxidation numbers +1, +2 and +3 respectively.
8. The sum of the oxidation numbers of all atoms in a neutral compound is 0.
9. The sum of the oxidation numbers of all atoms in a polyatomic ion equals charge of the ion.

# Balancing Redox Equations


More complicated than balancing normal reactions.

You have to balance the electrons, not just the atoms!

## Steps

1. Assign oxidation numbers to determine which things are oxidized and which are reduced.
2. Split the rxn into two halves – oxidation half and reduction half. Include electrons.
3. Balance the atoms.
4. Balance the charge by balancing the number of electrons.
5. Add half reactions back together, simplify, and CHECK.

# Balancing Redox Reactions

Yes, I know this isn't balanced! That is what we are working on!

1

## Assign oxidation states

Determine the element oxidized and the element reduced.

oxygen -2, not in one of the exceptions.

$$Cl_2 + I^- \rightarrow 2Cl^- + IO_3^-$$

0 -1 -1 +5 -2

elements are 0

monoatomic ions match their charge

monoatomic ions match their charge

x + 3(-2) = -1
sum must equal the overall charge on ion
x = +5

# Balancing Redox Reactions

1

## Assign oxidation states

Determine the element oxidized
and the element reduced.

$$\underset{0}{Cl_2} + \underset{-1}{I^-} \rightarrow 2\underset{-1}{Cl^-} + \underset{+5}{I}\underset{-2}{O_3^-}$$

LEO goes GER

$I^-$ lost electrons = oxidized, -1 → +5

Cl gained electrons = reduced, 0 → -1

# Balancing Redox Reactions

2

**Write oxidation and reduction half-reactions, including electrons**

**Oxidation** electrons being lost, products

**Reduction** electrons being gained, reactants

I went from -1 → +5, that's a loss of…

Yes, I know these are not balanced! That is what we are still working on! It takes a while!

**Oxidation:** $I^- \rightarrow IO_3^- + \underline{\ 6e^-\ }$

**Reduction:** $Cl_2 + \underline{\ 2e^-\ } \rightarrow 2Cl^-$

Each Cl went from 0 → -1, that's a gain of…

# Balancing Redox Reactions

3

## Balance the atoms in the half reactions

- First balance elements other than H and O.
- Add $H_2O$ where O is needed.
- Add $H^+$ where H is needed

**Oxidation:** $I^- + \underline{\;3H_2O\;} \rightarrow IO_3^- + 6e^- + \underline{\;6H^+\;}$

**Reduction:** $Cl_2 + 2e^- \rightarrow 2Cl^-$

# Balancing Redox Reactions

4

## Balance the charge by balancing the # of $e^-$

- Balance electrons between half-reactions.
  *Least Common Multiple*

**Oxidation:** $1 \times ( I^- + 3H_2O \rightarrow IO_3^- + 6\,e^- + 6H^+ )$

**Reduction:** $3 \times ( Cl_2 + 2\,e^- \rightarrow 2\,Cl^- )$

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**Oxidation:** $I^- + 3H_2O \rightarrow IO_3^- + 6\,e^- + 6H^+$

**Reduction:** $3Cl_2 + 6\,e^- \rightarrow 6\,Cl^-$

# Balancing Redox Reactions

5

## Add half reactions together, simplify, check

- Make sure the atoms balance AND the charges

**Oxidation:** $I^- + 3H_2O \rightarrow IO_3^- + 6\ e^- + 6H^+$

**Reduction:** $3Cl_2 + 6\ e^- \rightarrow 6\ Cl^-$

$$I^- + 3H_2O + 3Cl_2 + \cancel{6\ e^-} \rightarrow IO_3^- + \cancel{6\ e^-} + 6H^+ + 6\ Cl^-$$

$$I^- + 3H_2O + 3Cl_2 \rightarrow IO_3^- + 6H^+ + 6\ Cl^-$$

**CHECK:** Atoms balanced – yes! Charges balanced – yes!

# Best Advice…

USE PENCIL!

DON'T CRAM YOUR WORK! USE LOTS OF SPACE!

DON'T PANIC!

STUCK??? ERASE AND START OVER.

# Oxidation and Reduction Recap


**Oxidation is the process that occurs when**

– the oxidation number of an element increases,
– an element loses electrons,
– a compound adds oxygen,
– a compound loses hydrogen, or
– a half-reaction has electrons as products.

**Reduction is the process that occurs when**

– the oxidation number of an element decreases,
– an element gains electrons,
– a compound loses oxygen,
– a compound gains hydrogen, or
– a half-reaction has electrons as reactants.

# A Few More Electrochemistry Terms

## Oxidizing agent

The substance that is **DOING the oxidizing** of the other substance. The substance that is **BEING REDUCED**!

## Reducing agent

The substance that is **DOING the reducing** of the other substance. The substance that is **BEING OXIDIZED!**

# A Few More Electrochemistry Terms


## Anode

The electrode where oxidation occurs


**Anode is Oxidation**

## Cathode

The electrode where reduction occurs


**Reduction at the Cathode**

# A Few More Electrochemistry Terms

**Current - the number of electrons that flow through the system per second.**

- Unit = ampere, amp, A

$$1\,A = \frac{1\,Coulomb}{1\,second} = \frac{6.242\,x\,10^{18} e^{-}}{1\,second}$$

**Electrode surface area dictates the # of e- that can flow.**

- Larger batteries produce larger currents.

# A Few More Electrochemistry Terms

**Cell Potential, potential difference, electromotive force**

**All terms for the difference in potential energy between the reactants and products**

- The voltage needed to drive electrons through the external circuit
- Amount of force pushing the electrons through the wire
- Unit = volts
- 1 V = 1 J of energy per coulomb of charge

# Useful Conversions


$$1\ A = \frac{1\ Coulomb}{1\ second} = \frac{6.242\ x\ 10^{18} e^{-}}{1\ second}$$

$$1\ Volt = \frac{1\ Joule}{1\ Coulomb}$$

$$1\ Faraday = \frac{96,500\ Coulombs}{1\ mol\ e^{-}}$$

**Example:** How many minutes does a **4.00 A** current need to be applied to a $Cu^{2+}$ solution to make 15.00 g of Cu?

| | | | | | |
|---|---|---|---|---|---|
| **15.00g** | **1 mol** | **2 mol e-** | **96,500 C** | **1 s** | **1 min** |
| | **63.55 g** | **1 mol Cu** | **1 mol e-** | **4 C** | **60 s** |

**= 189.8 min**

# YouTube Link to Presentation

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