N42 - Electrochemistry

Concepts

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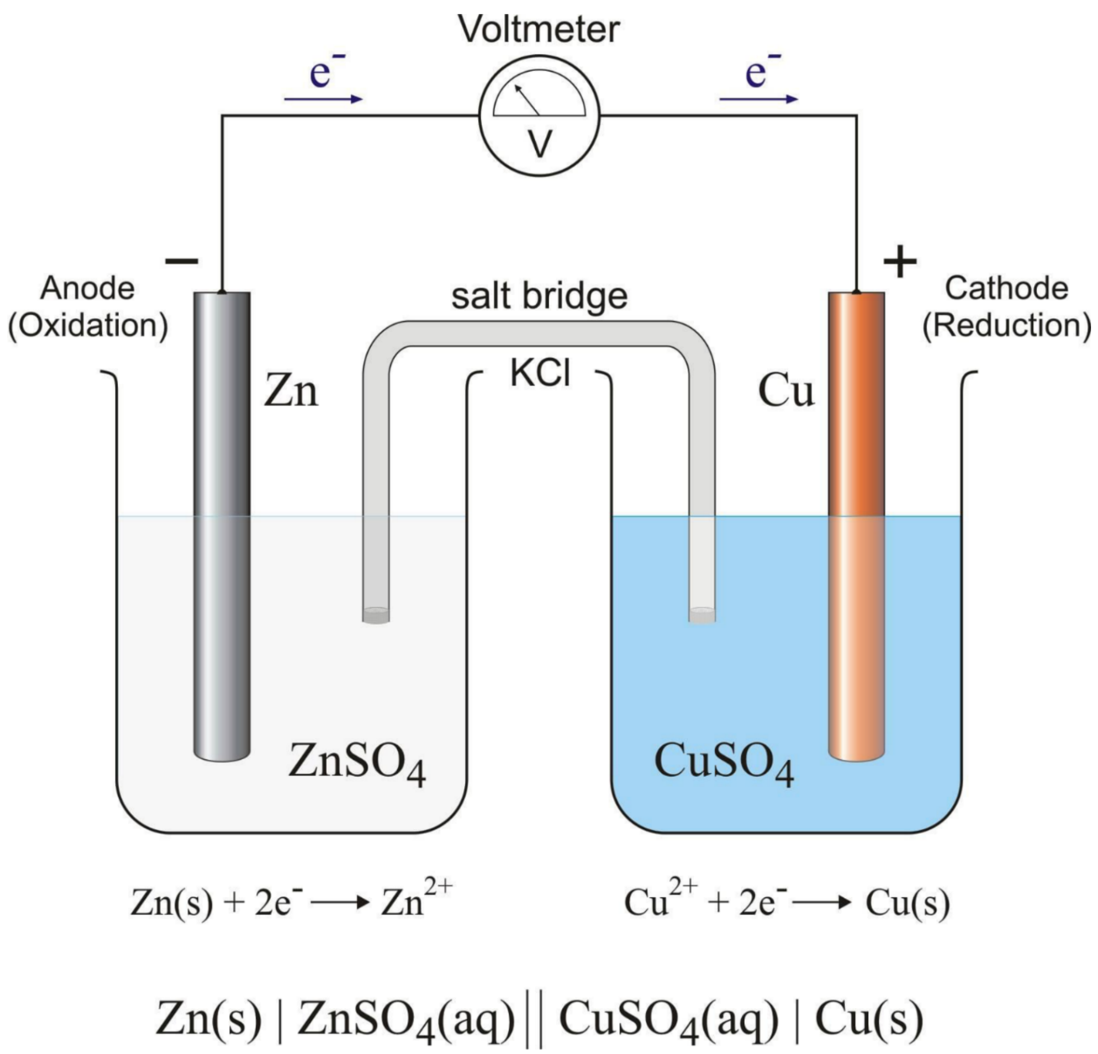
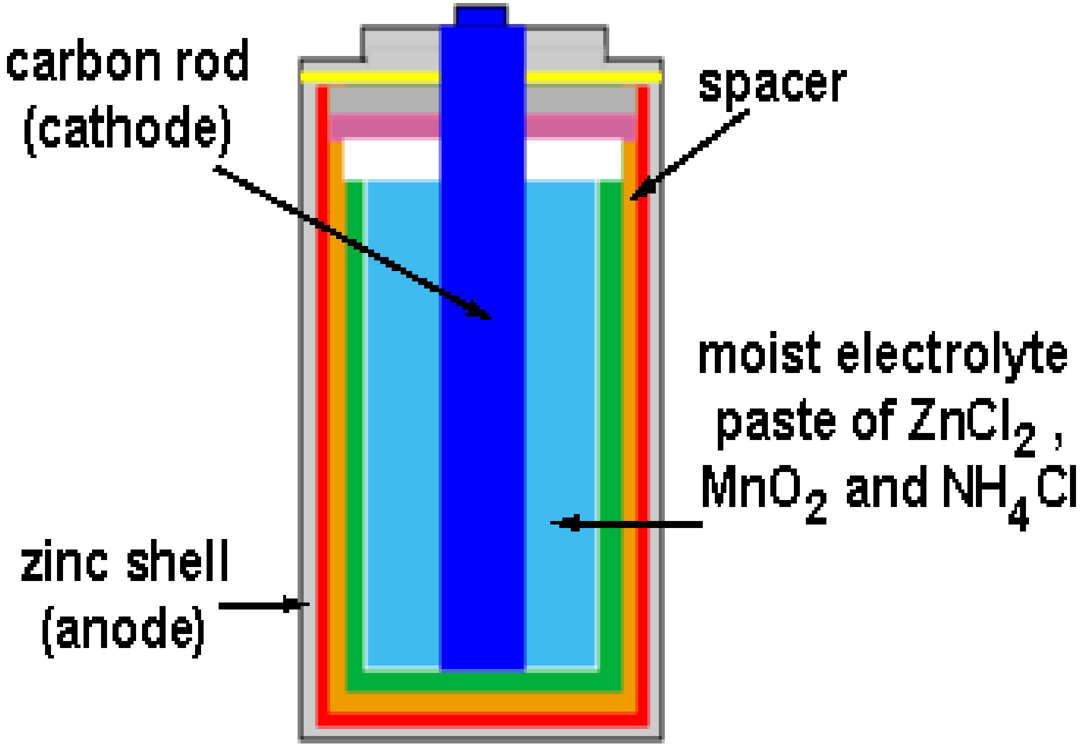
Concepts

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


Voltmeter
e-
e-
V
−
+
Anode
(Oxidation)
Cathode
(Reduction)
salt bridge
KCl
Zn
Cu
ZnSO4
CuSO4
Zn(s) + 2e- ⟶ Zn2+
Cu2+ + 2e- ⟶ Cu(s)
Zn(s) | ZnSO4(aq) || CuSO4(aq) | Cu(s)


carbon rod
(cathode)
spacer
moist electrolyte
paste of ZnCl2 ,
MnO2 and NH4Cl
zinc shell
(anode)

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 — elements are 0

-1 — monoatomic ions match their charge

-1 — monoatomic ions match their charge

+5 -2

$x + 3(-2) = -1$

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.

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

0   -1   -1   +5   -2

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 x ($I^- + 3H_2O \rightarrow IO_3^- + 6\ e^- + 6H^+$)

Reduction: 3 x ($Cl_2 + 2\ e^- \rightarrow 2\ Cl^-$)

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

https://youtu.be/-y2xTX_BVsA