Electrolytic Cells

# Electrolysis of Aqueous Solutions

## Steps

1. Separate all of the “ingredients” within the solution into their respective ions
2. Create 2 headings: “Cathode / Reduction / -‘ve” and “Anode / Oxidation / +’ve”
3. Group all of the positive ions under “Cathode…” and all of the negative ions under “Anode…”
4. If there are oxy-ion(s) (i.e. ) present, calculate the oxidation charge for the non-oxygen element. If it possesses its highest oxidative charge, it cannot be further oxidized, and thus it becomes “useless” for the rest of the equation.

* **It is imperative to identify this reality, as it can mislead you for the remaining portion of the question**

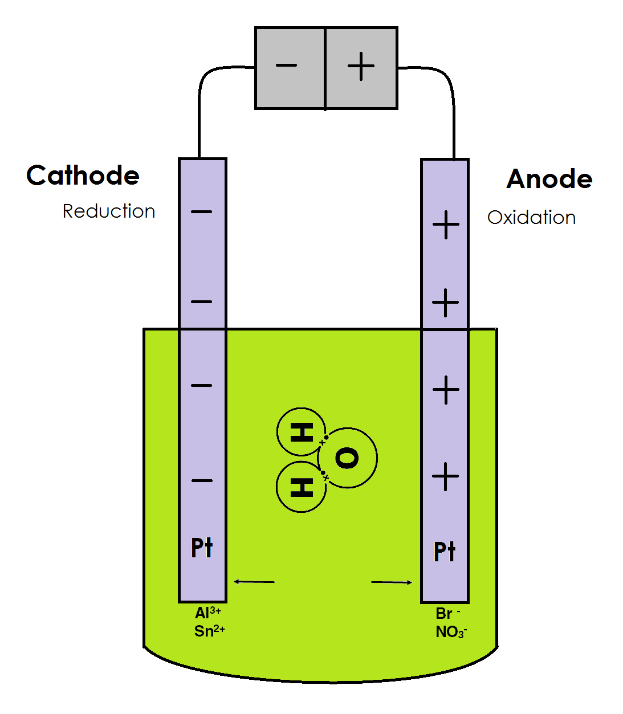
1. Identify reduction charges for each of the ions
2. “Cathode…” 🡪 Pick the ion with the highest reduction potential (i.e. -0.14 over -0.52, etc.)
3. “Anode…” 🡪 Pick the ion with the lowest reduction potential (i.e. +1.07 over +1.23 etc.)
4. Solve

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| --- | --- |
| 🡪 | x – 6 = -1  x= +5  Since, Nitrogen’s max oxidation number is +5, it cannot be further oxidized. |

# Example

In a solution with SnBr2 and Al(NO3)3

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| --- | --- | --- | --- |
| Cathode/Reduction/-‘ve |  | Anode/Oxidation/+’ve |  |
| Sn2+ 🡪  Al3+ 🡪  H+ 🡪 | -0.14  -1.66  -0.83 | Br- 🡪  🡪  O2- 🡪 | +1.07  X  +1.23 |



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| --- | --- |
| Cathode 🡪 | Sn2+ because it has the highest reduction potential |
| Anode 🡪 | Br - because it has the highest oxidation potential (lowest reduction potential) |

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| 2 Equations to be used: |