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. $NO_3^-, ClO_4^-, etc.$) 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**
5. Identify reduction charges for each of the ions
6. "Cathode…" → Pick the ion with the highest reduction potential (i.e. -0.14 over -0.52, etc.)
7. "Anode…" → Pick the ion with the lowest reduction potential (i.e. +1.07 over +1.23 etc.)
8. Solve

EXAMPLE

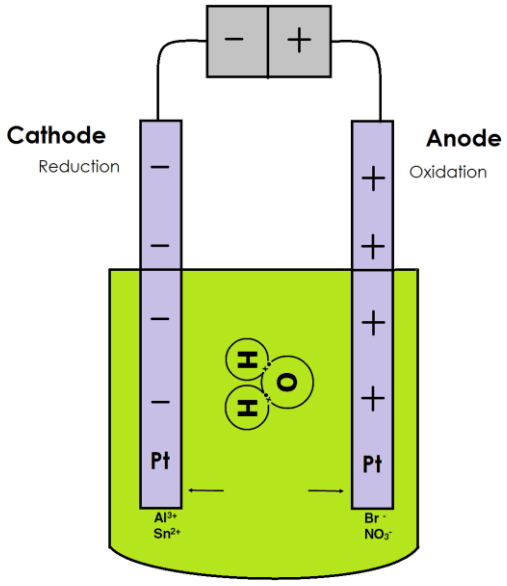
In a solution with $SnBr_2$ and $Al(NO_3)_3$

Cathode/Reduction/-'ve	E^o	Anode/Oxidation/+'ve	E^o
Sn^{2+} →	-0.14	Br →	+1.07
Al^{3+} →	-1.66	NO_3^- →	X
H^+ →	-0.83	O^{2-} →	+1.23

NO_3^- →	x – 6 = -1 x= +5 Since, Nitrogen's max oxidation number is +5, it cannot be further oxidized.

Cathode → Sn^{2+} because it has the highest reduction potential

Anode → Br^- because it has the highest oxidation potential (lowest reduction potential)



2 Equations to be used:

$$Sn^{2+} + 2\ e^- \rightarrow Sn \quad E^o = -0.14$$
$$Br_2 + 2\ e^- \rightarrow 2\ Br^- \quad E^o = +1.07$$

$$Sn^{2+} + 2\ e^- \rightarrow Sn \quad E^o = -0.14$$
$$2\ Br^- \rightarrow Br_2 + 2\ e^- \quad E^o = -1.07$$

$$Sn^{2+} + 2\ Br^- \rightarrow Sn + Br_2 \quad E^o = -0.61$$