

## 21 • Electrochemistry

### BLUFFER'S GUIDE

1. Electrochemistry is all oxidation-reduction chemistry.  
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**Oxidation:** loss of  $e^-$ ; ox # increases  
**Reduction:** gain of  $e^-$ ; ox # decreases  
*example:*  $Fe^{2+} + 2e^- \rightarrow Fe(s)$  (reduction)

2. In a reaction, the **oxidizing agent** gets **reduced**; the **reducing agent** gets **oxidized**.

3. Balancing redox reactions:  
**oxidation number method**
- assign ox #'s to every atom
  - determine changes in ox #
  - balance changes
  - balance all atoms except H & O
  - balance O's (add  $H_2O$ 's)
  - balance H's (add  $H^+$ 's)
  - adjust for basic solution if needed

**half-reaction method.**

- determine oxidation & reduction
- write two separate half-reactions
- balance all atoms except H & O
- balance O's (add  $H_2O$ 's)
- balance H's (add  $H^+$ 's)
- add  $e^-$ 's to more positive side
- balance  $e^-$ 's between half-reactions
- combine half-reactions
- adjust for basic solution if needed

4. Electricity can either **cause** a reaction (electrolysis, electrolytic cell) or can be **produced by** the reaction (Galvanic cell, electrochemical cell, Voltaic cell).

5. **Electrolysis / Electroplating**  
 coulomb (C) = an amount of charge  
 amp = current = charge per second  
 $1 \text{ amp} \cdot 1 \text{ second} = 1 \text{ Coulomb}$   
 $1 \text{ C} / \text{amp} \cdot \text{s}$   
 Faraday constant, F:  
 $1 \text{ mole } e^- = 96,500 \text{ C}$

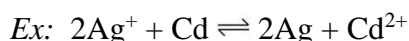
6. Electrolysis calculations begin with amp·s  
*Example:*

$$\begin{aligned} &\text{How many moles of copper metal can be plated using a 10 amp circuit for 30 s?} \\ &10 \text{ amp} \times 30 \text{ s} \times \frac{1 \text{ C}}{1 \text{ amp} \cdot \text{s}} \times \frac{1 \text{ mol } e^-}{96500 \text{ C}} \times \frac{1 \text{ mol Ag}}{1 \text{ mol } e^-} = \\ &= 3.1 \times 10^{-3} \text{ mole Ag} \end{aligned}$$

7. Spontaneous redox reactions (unlike electrolysis/electroplating) can simply occur (as in the ornament lab) or can be separated so the oxidation and reduction occur in different containers (half-cells). In this way, the electrons must move through an outside wire (this is an electrochemical cell—a battery).

8. Every atom has a different “potential” to accept electrons... “reduction potential”  
 $Ag^+(aq) + e^- \rightarrow Ag(s) \quad E^\circ = +0.80 \text{ v}$   
 $Cd^{2+}(aq) + 2e^- \rightarrow Cd(s) \quad E^\circ = -0.40 \text{ v}$   
 These are measured by comparing every chemical to the same “standard half-cell.”

The reduction with the more positive  $E^\circ$  value will occur as written; the other reaction will reverse (oxidation).



The **difference** in the  $E^\circ$  values is the voltage of a cell made using these two reactions.

$$\text{Ex: } +0.80 \text{ v} - (-0.40 \text{ v}) = 1.20 \text{ volts}$$

**NOTE that you do not multiply the Cd voltage by 2. Comparing every cell to the same standard cell accounts for this.**

9. Any change that drives the reaction forward will **increase** the cell's voltage.
10. In **all** electrochemical cells:  
**Oxidation** occurs at the **Anode**  
**Reduction** occurs at the **Cathode**

Neat online simulation/experiment for electrochem. I think it helps a lot!  
<https://tinyurl.com/mssmkps9>

