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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 amp · 1 second = 1 Coulomb
   1 C / amp·s
   Faraday constant, F:
   1 mole e^- = 96,500 C

6. Electrolysis calculations begin with amp·s
   Example:
   How many moles of copper metal can be plated using a 10 amp circuit for 30 s?
   $$10\text{amp} \times 30\text{s} \times \frac{1\text{C}}{1\text{ amp}\cdot\text{s}} \times \frac{1\text{ mol e-}}{96500\text{C}} \times \frac{1\text{ mol Ag}}{1\text{ mol e-}} =$$
   $= 3.1 \times 10^{-3}$ mole Ag

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"
   $Ag^+(aq) + e^- \rightarrow Ag(s)$ E° = +0.80 v
   $Cd^{2+}(aq) + 2e^- \rightarrow Cd(s)$ E° = –0.40 v
   These are measured by comparing every chemical to the same "standard half-cell."

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

   Ex: $2Ag^+ + Cd \rightleftharpoons 2Ag + Cd^{2+}$

   The **difference** in the E° values is the voltage of a cell made using these two reactions.

   Ex: +0.80 v – (-0.40 v) = 1.20 volts
   NOTE that you do <u>not</u> 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**