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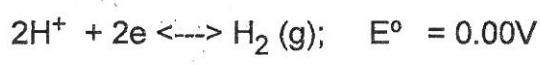
Chemistry 12
 Electrochem Lesson #9
 STANDARD REDUCTION POTENTIALS

The tendency of electrons to flow in an electrochemical cell is called the VOLTAGE or ELECTRICAL POTENTIAL to do work.

VOLTAGE is the **WORK DONE** per **ELECTRON TRANSFERRED!**

To measure the electrical potentials the DIFFERENCE between two electrical potentials of two different HALF-RXNS is measured.

note: * This is arbitrarily set * selected @ random w/o reason
 The ZERO POINT on the voltage scale the half-reaction of Hydrogen.

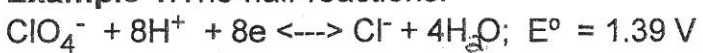


E° is the Standard REDUCTION potential (in Volts)
 "°" implies a **STANDARD STATE**

An electrochemical cell is in a **STANDARD STATE** if:

- i. IT IS AT 25°C
- ii. GASES ARE AT 101.3 kPa (1 atm)
- iii. ELEMENTS are in their STANDARD STATE (@ 25°C)
- iv. ALL SOLUTIONS ARE [1M]

Example 1. The half-reactions:

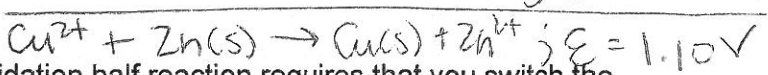
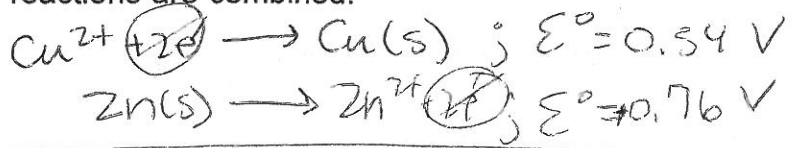
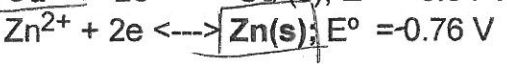
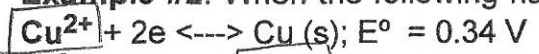


•• [ClO₄⁻] = 1M; [Cl⁻] = 1M; [Cu²⁺] = 1M,

NOTE:

★ The voltages for each half-reaction is relative to the **HYDROGEN HALF-CELL!!!**

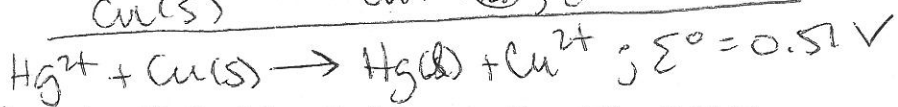
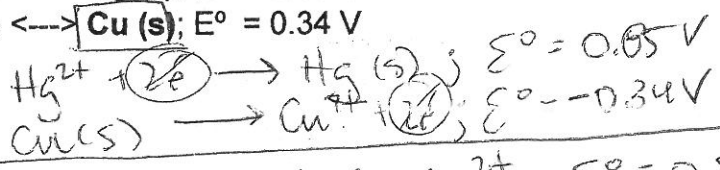
Example #2: When the following half reactions are combined:



Note: Writing the half-reaction as an oxidation half reaction requires that you switch the sign of E° .

Cu^{2+} reacts with $Zn(s)$ so that the resulting $E^{\circ} = 1.10V$

Example 3: Joining the following half-reactions:



Hg^{2+} reacts with Cu (s) so that the resulting $E^\circ = 0.51 \text{ V}$

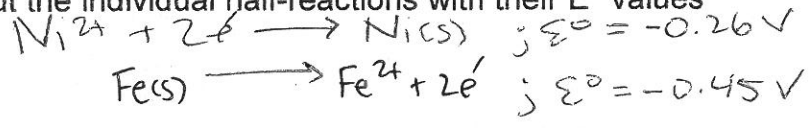
Another way to calculate E° is to use the values for each half-reaction and plugging them in (NOT CHANGING THEIR SIGN) to the following equation:

$$E^\circ_{\text{cell}} = (E^\circ_{\text{red}}) - (E^\circ_{\text{ox}})$$

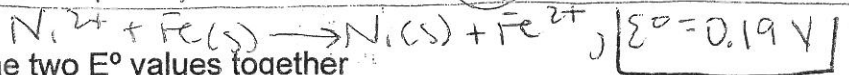
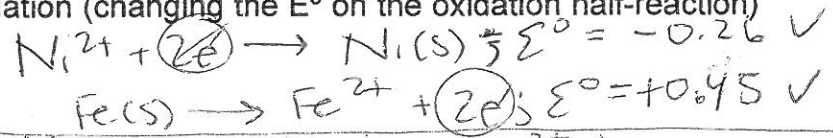
$$E^\circ_{\text{cell}} = (0.85) - (0.34 \text{ V}) = 0.51 \text{ V}$$

Example 4: Calculate the potential (E°_{cell}) of the cell $\text{Ni}^{2+} + \text{Fe} \rightarrow \text{Ni} + \text{Fe}^{2+}$

1. Write out the individual half-reactions with their E° values



2. re-Write the half-reactions according to how they are acting in the above redox equation (changing the E° on the oxidation half-reaction)

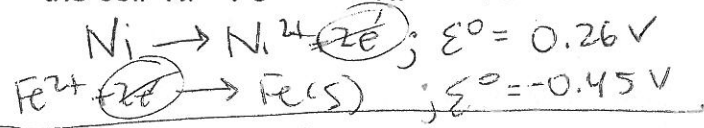


3. Add the two E° values together

4. OR use $E^\circ_{\text{cell}} = E^\circ_{\text{red}} - E^\circ_{\text{ox}}$

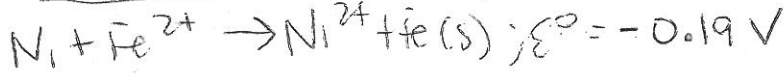
$$E^\circ_{\text{cell}} = (-0.26 \text{ V}) - (-0.45 \text{ V}) = 0.19 \text{ V}$$

Example 5: Use either of the two techniques above to calculate the potential for the cell $\text{Ni} + \text{Fe}^{2+} \rightarrow \text{Ni}^{2+} + \text{Fe}$



$$E^\circ_{\text{cell}} = (-0.45 \text{ V}) - (0.26 \text{ V})$$

$$E^\circ_{\text{cell}} = -0.19 \text{ V}$$



NOTICE: if the E° value is **POSITIVE**, the reaction is **SPONTANEOUS**
If the E° value is **NEGATIVE**, the reaction is **NON-SPONTANEOUS**

Seatwork/Homework: Read pgs 218- 224 before attempting Exercises 36- 41 (odd letters ie a,c, e etc) pg 224-225 and PLO's U6 - U10