Name: Key
Blk:_____Date:_____

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.

$$2H^+ + 2e \longleftrightarrow H_2(g); \quad E^o = 0.00V$$

E^o is the Standard REDUCTION potential (in Volts)
" o " 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 (1atm)
iii. ELEMENTS are in their STANDARD STATE (@ 25°C)
iv. ALL SOLUTIONS ARE [1M]

Example 1.The half-reactions:
$ClO_4^- + 8H^+ + 8e \longleftrightarrow Cl^- + 4H_2O$; $E^o = 1.39$ V
$Cu^{2+} + 2e \longleftrightarrow Cu(s)$; $E^o = 0.34$ V
∴ $[ClO_4^-] = 1M$; $[Cl^-] = 1M$; $[Cu^{2+}] = 1M$,

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

Example #2: When the following half reactions are combined:
Cu^{2+} + 2e <---> Cu (s); $E^o = 0.34$ V
Zn^{2+} + 2e <---> **Zn(s)**; $E^o = -0.76$ V

$Cu^{2+} + 2e \rightarrow Cu(s)$; $E^o = 0.34$ V
$Zn(s) \rightarrow Zn^{2+} + 2e$; $E^o = +0.76$ V
$Cu^{2+} + Zn(s) \rightarrow Cu(s) + Zn^{2+}$; $E = 1.10V$

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

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

Example 3: Joining the following half-reactions:

Hg^{2+} + 2e <---> Hg (l); E° = 0.85 V

Cu^{2+} + 2e <---> **Cu (s)**; E° = 0.34 V

$Hg^{2+} + 2e \rightarrow Hg(l)$; $E° = 0.85$ V

$Cu(s) \rightarrow Cu^{2+} + 2e$; $E° = -0.34$ V

$Hg^{2+} + Cu(s) \rightarrow Hg(l) + Cu^{2+}$; $E° = 0.51$ V

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

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

E° cell = (E° red) - (E° ox)

$E°cell = (0.85) - (0.34V) = 0.51V$

Example 4: Calculate the potential (E° cell) of the cell Ni^{2+} + Fe ---> Ni + Fe^{2+}

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

$Ni^{2+} + 2e \rightarrow Ni(s)$; $E° = -0.26$ V

$Fe(s) \rightarrow Fe^{2+} + 2e$; $E° = -0.45$ V

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)

$Ni^{2+} + 2e \rightarrow Ni(s)$; $E° = -0.26$ V

$Fe(s) \rightarrow Fe^{2+} + 2e$; $E° = +0.45$ V

$Ni^{2+} + Fe(s) \rightarrow Ni(s) + Fe^{2+}$; $E° = 0.19$ V

3. Add the two E° values together

4. **OR** use **E° cell = E° red - E° ox**

$E°cell = (-0.26V) - (-0.45V) = 0.19V$

Example 5: Use either of the two techniques above to calculate the potential for the cell Ni + Fe^{2+} --> Ni^{2+} + Fe

$Ni \rightarrow Ni^{2+} + 2e$; $E° = 0.26$ V

$Fe^{2+} + 2e \rightarrow Fe(s)$; $E° = -0.45$ V

$Ni + Fe^{2+} \rightarrow Ni^{2+} + Fe(s)$; $E° = -0.19$ V

$E°cell = (-0.45V) - (-0.26V)$

$E°cell = -0.19V$

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**