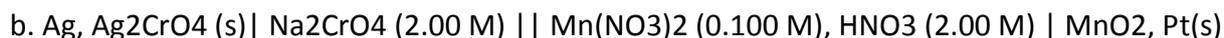
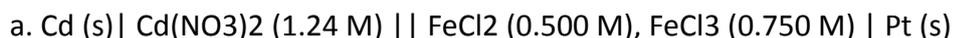


## Answer on Question #56849 - Chemistry - General Chemistry

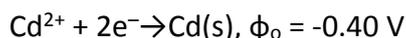
### Question:

For each of the following electrochemical cells, calculate the  $E_o$ , write the balanced cell reaction equation and identify what is being oxidized and reduced, and calculate the  $Q$  and  $E_{cell}$ . Note that the substances with  $M$  given are in solution. All other substances are solids.

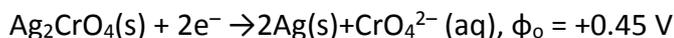
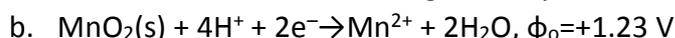


### Answer:

Standard potential of the cell is the difference between right half-cell potential, where the reduction occurs, and left half-cell potential, where the oxidation occurs. Let's find the potentials of the half-cells for each case and then, subtracting the one for the left half-cell from the one for the right half-cell, find the  $E_o$ :

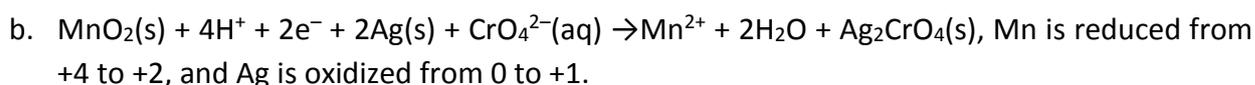


$$E_o = E_{right} - E_{left} = +0.77 - (-0.40) = +1.17 \text{ V}$$



$$E_o = E_{right} - E_{left} = +1.23 - 0.45 = 0.78 \text{ V}$$

Balanced cell reaction equations are:



Now, let's calculate the reaction quotient:

a.  $Q = \frac{c(\text{Fe}^{2+})^2 \cdot c(\text{Cd}^{2+})}{c(\text{Fe}^{3+})^2} = \frac{0.5^2 \cdot 1.24}{0.75^2} = 0.55$

b.  $Q = \frac{c(\text{Mn}^{2+})}{c(\text{CrO}_4^{2-}) \cdot c(\text{H}^+)^4} = \frac{0.1}{2 \cdot 2^4} = 3.13 \cdot 10^{-3}$

And finally, we can calculate  $E_{cell}$ :

a.  $E_{cell} = E_o - \frac{RT}{nF} \ln Q = 1.17 - \frac{0.05916}{2} \cdot \lg(0.55) = 1.18 \text{ V}$

b.  $E_{cell} = E_o - \frac{RT}{nF} \ln Q = 0.78 - \frac{0.05916}{2} \cdot \lg(3.13 \cdot 10^{-3}) = 0.85 \text{ V}$