Nernst Equasion

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Purpose:

- **Determine equalibrium constant for two equalibrium reactions $Ag^+ + 2NH_3 \leftrightarrows [Ag(NH_3)_2]$ and $Ag^+ + 2S_2O_3^{2-} \leftrightarrows [Ag(S_2O_3)_2]^{3-}$**
- **Evalutate accurracy of determine constants with percent error**
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Reference: Kateley, L. J., *Introduction to Chemistry in the Laboratory*, *20th Ed.*, Lake Forest College, **2021,** Experiment *xx*, Appendix *xx***.**

Equalibrium One

$Ag^+ + 2NH_3 \leftrightarrows [Ag(NH_3)_2]$

- $K_f = [Ag(NH_3)_2]/[Ag^+][NH_3]^2=$
- Assume:
    - The reaction shifts almost completely to the right and has a large K_f
    - The reaction reaches equalibrium quickly
    - The limiting reagent is Ag^+ and there is a large excess of NH_3 ligand

Equalibrium Two

$Ag^+ + 2S_2O_3^{2-} \leftrightarrows [Ag(S_2O_3)_2]^{3-}$

- $K_f = [Ag(S_2O_3)_2]^{3-}/[Ag^+][S_2O_3]^2$
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Standardization of pH Meter as voltmeter

- Probe calculated using a 1.61 +/-0.05V
- Battery registered as 1.56 V which is within the range noted on the battery

Set Up and Validation of a Voltic Cell

- Glassware cleaned to prevent source of error
- Seven strips of filter paper soaked in 1M KNO_3 to form salt bridge
- Black and red claws each attatched to silver wire

a)

- 10mL of 10.0mM Ag^+ placed in 20mL beaker for cathod
- 5mL of 10.0mM Ag^+ places in 30mL beaker for anode
- Wire attacked to black claw placed in anode and wire attached to red claw places in cathode
- Meter read 0.002V

b)

- Silver wires washed with deionized water
- Anode replaced with a solution of 1.00mL 10.0mM Ag^+ and 4.00mL of deionized water
- Salt bridge relapces with fresh 1M KNO_3 soaked strip of filter paper
- Meter read 0.036V

c)

- Silver wires washed with deionized water
- Anode replaced with a solution of 0.500mL of 10.0mM Ag^+ and 4.500mL of deionized water
- Salt bridge replaced with fresh 1.0M KNO_3 soaked strip of filter paper
- Meter reads .0.050V

TABLE 1: VALIDATION OF THE NERNST EQUATION

Cathode (red) vol 10.0 mM Ag^+	Anode (black) vol 10.0 mM Ag^+	Anode (black) H_2O	mM conc Ag^+ anode	mM conc Ag^+ cathode	Conc ratio: $[Ag^+_{dil}] \div [Ag^+_{conc}]$	Theoretical voltage (E_{cell})	Observed voltage (E_{cell})
10 mL	5 mL	0	10.0	10.0	1.00		
10 mL	1000 µL	4000 µL	2.00	10.0	0.200		
10 mL	500 µL	4500 µL	1.00	10.0	0.100		

Determine Formation Constants Using Nernst Equation

- Cathod of 10mL of 10.0mM Ag^+ solution

TABLE 2. ANODE CELL VOLUMES AND MILLIMOLES

ligand	volume 10.0 mM Ag^+	volume 100 mM ligand	total volume (mL)	mmol Ag^+ initially	mmol ligand initially	mmol complex at equilibrium	mmol ligand at equilibrium
NH_3	5000 µL	5000 µL	10.00	0.0500	0.500	0.0500	0.400
NH_3	7500 µL	2500 µL	10.00	0.0750	0.250	0.0750	0.100
$S_2O_3^{2-}$	5000 µL	5000 µL	10.00	0.0500	0.500	0.050	0.400
$S_2O_3^{2-}$	7500 µL	2500 µL	10.00	0.0750	0.250	0.0750	0.100

TABLE 3: DETERMINATION OF FORMATION CONSTANTS USING THE NERNST EQUATION

ligand	$M_{complex}$	M_{ligand}	$M_{Ag+dilute}$	E_{cell} volts	K_f
NH_3	0.00500	0.0400		0.099	
NH_3	0.00750	0.0100		0.220	
$S_2O_3^{2-}$	0.0050	0.040		0.606	
$S_2O_3^{2-}$	0.00500	0.0100		0.654	

Conclusion:

- **Silver ion is very low and hard to measure in moles per liter**
- **True values: 1.7×10^7 and 2.9×10^{13}**
- **Posible source of error from touching the filter paper**
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