Daniel C. Harris

Quantitative Chemical Analysis

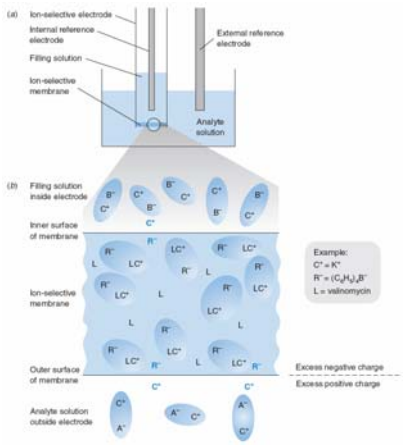
Seventh Edition

Chapter 15
Electrodes

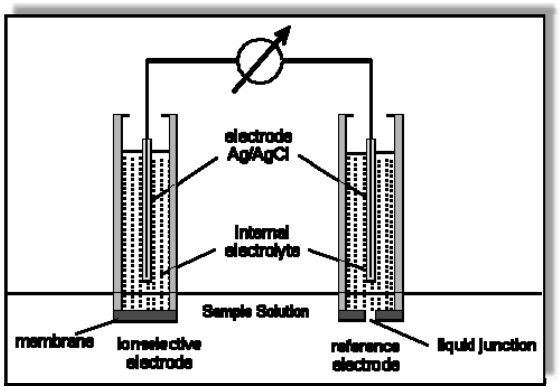
Ion selective electrode

$$E = k + \beta\left(\frac{RT\ln 10}{nF}\right)\log[\mathrm{X}]$$

Ion selective electrode



Ion selective electrode

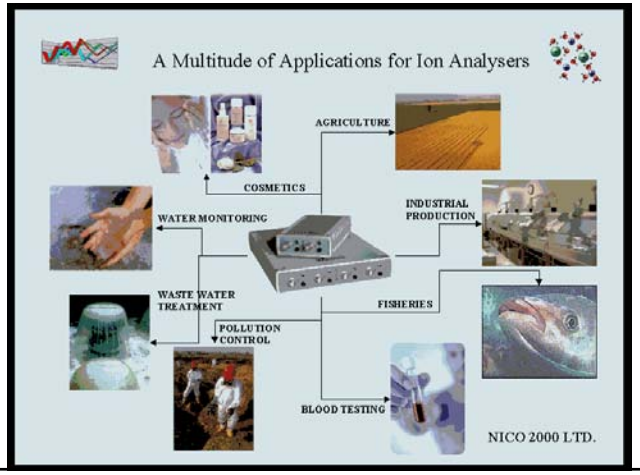


Ion selective electrode

$$E = k + \beta\left(\frac{RT\ln 10}{nF}\right)\log[X]$$

Ion selective electrode

Ion selective electrode



Ion selective electrode

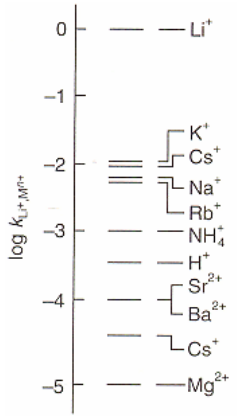
Ion selective electrode

CHEM-7 is a group of blood tests that provides information about your body's metabolism.
The test is commonly called a basic metabolic panel.

BUN: 7 to 20 mg/dL
CO2 (carbon dioxide): 20 to 29 mmol/L
Creatinine: 0.8 to 1.4 mg/dL
Glucose: 64 to 128 mg/dL
Serum chloride: 101 to 111 mmol/L
Serum potassium: 3.7 to 5.2 mEq/L
Serum sodium: 136 to 144 mEq/L

Ion selective electrode

Selectivity And interference



NERNST EQUATION

$$E = E_+ - E_-$$

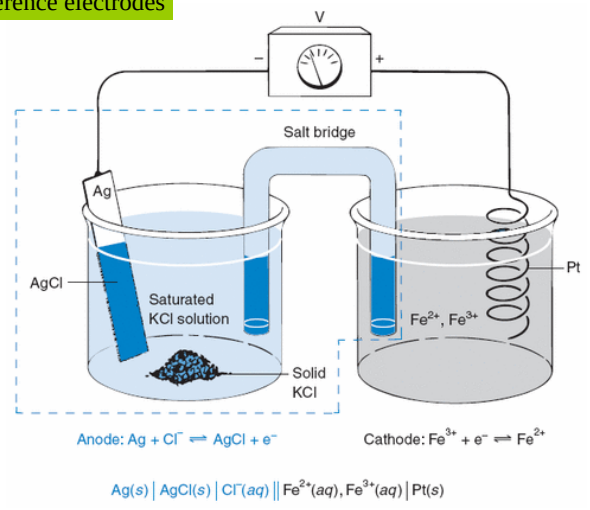
Right electrode: $aA + ne^- \rightleftharpoons cC$ E°_+

Left electrode: $dD + ne^- \rightleftharpoons bB$ E°_-

$$E = E^0 - \frac{RT}{nF} \ln Q$$

Measure E to determine 1 unknown concentration (…..so fix the other concentrations)

Reference electrodes



Reference electrodes

$$E_+ = 0.771 - 0.059\ 16 \log\left(\frac{[Fe^{2+}]}{[Fe^{3+}]}\right)$$

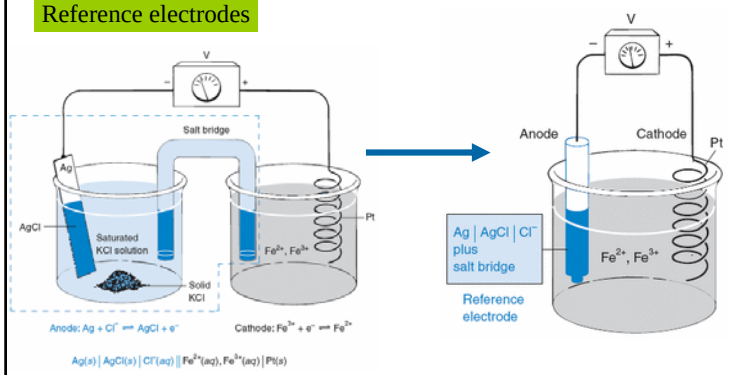
$$E_- = 0.222 - 0.059\ 16 \log[Cl^-]$$

$$E = \left\{0.771 - 0.059\ 16 \log\left(\frac{[Fe^{2+}]}{[Fe^{3+}]}\right)\right\} - \{0.222 - 0.059\ 16 \log[Cl^-]\}$$

My way:

$$E = E^\circ - 0.059\,16 \log \frac{[Fe^{2+}]}{[Fe^{3+}][Cl^-]}$$

Reference electrodes



Calomel electrode: $\frac{1}{2}Hg_2Cl_2(s) + e^- \rightleftharpoons Hg(l) + Cl^-$ $E^\circ = +0.268$ V

Mercury(I) chloride (calomel)

E(saturated KCl) $= +0.241$ V

Daniel C. Harris

Quantitative Chemical Analysis

Seventh Edition

Chapter 17

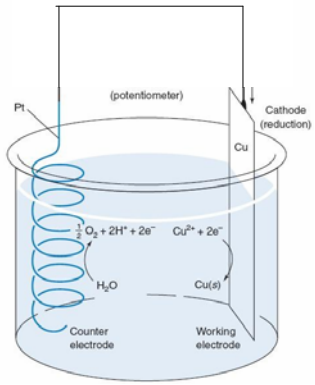
Electroanalytical techniques

Cathode:	$Cu^{2+} + 2e^- \rightleftharpoons Cu(s)$
Anode:	$H_2O \rightleftharpoons \frac{1}{2}O_2(g) + 2H^+ + 2e^-$
Net reaction:	$H_2O + Cu^{2+} \rightleftharpoons Cu(s) + \frac{1}{2}O_2(g) + 2H^+$

$E_+ = 0.339$ V

$-E_- = -1.229$ V

$E^o = -0.99$ V

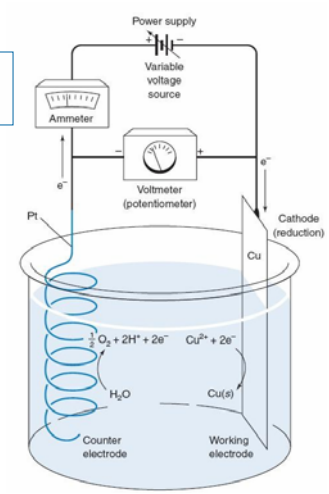


Electrolysis

Cathode: $Cu^{2+} + 2e^- \rightleftharpoons Cu(s)$

Anode: $H_2O \rightleftharpoons \frac{1}{2}O_2(g) + 2H^+ + 2e^-$

Net reaction: $H_2O + Cu^{2+} \rightleftharpoons Cu(s) + \frac{1}{2}O_2(g) + 2H^+$



Electrolysis

Cathode: $Cu^{2+} + 2e^- \rightleftharpoons Cu(s)$

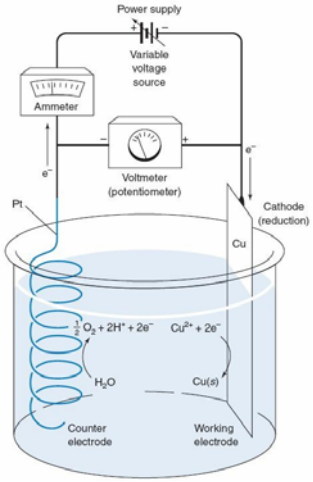
Anode: $H_2O \rightleftharpoons \frac{1}{2}O_2(g) + 2H^+ + 2e^-$

Net reaction: $H_2O + Cu^{2+} \rightleftharpoons Cu(s) + \frac{1}{2}O_2(g) + 2H^+$

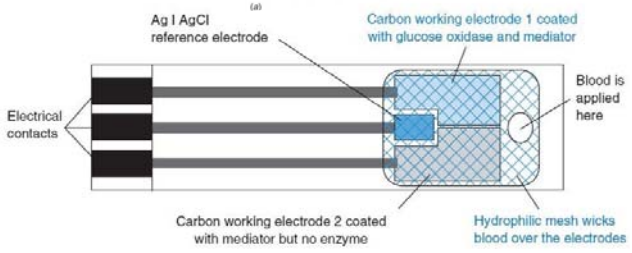
If a current of 0.17 A flows for 16 min through the cell how many grams of Cu(s) will be deposited?

$$\text{Moles of } e^- = \frac{I \cdot t}{F} = \frac{\left(0.17\frac{C}{s}\right)(16\text{ min})\left(60\frac{s}{\text{min}}\right)}{96\,485\left(\frac{C}{\text{mol}}\right)} = 1.6_9 \times 10^{-3}\text{ mol}$$

Half that number of moles
Converted to grams of Cu
→ 0.054 g

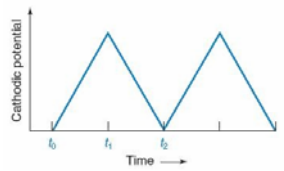


Electrolysis



Reaction at working electrode 1: $H_2O_2 \rightarrow O_2 + 2H^+ + 2e^-$

Voltametry



MAYBE LATER.....