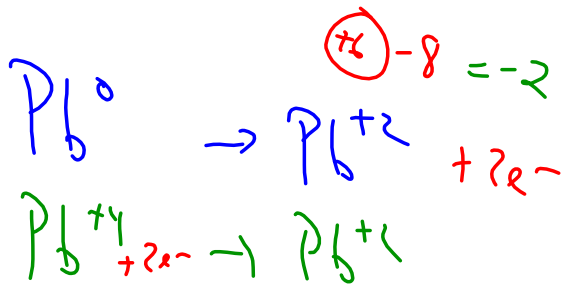
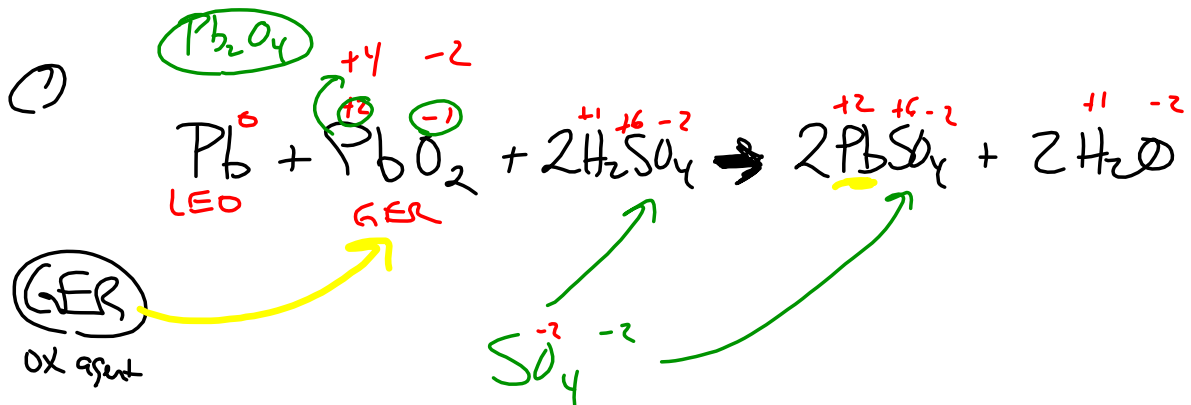
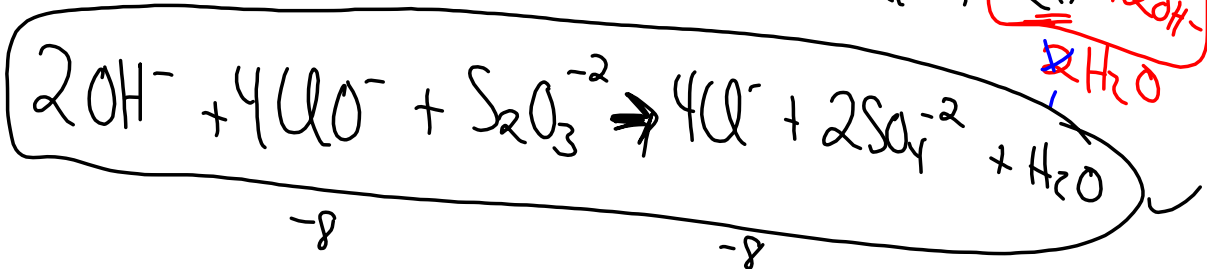
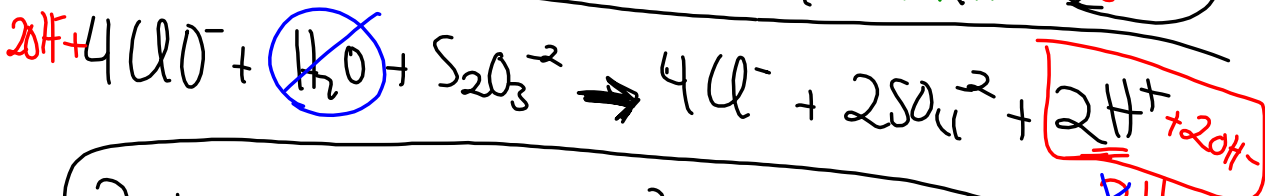
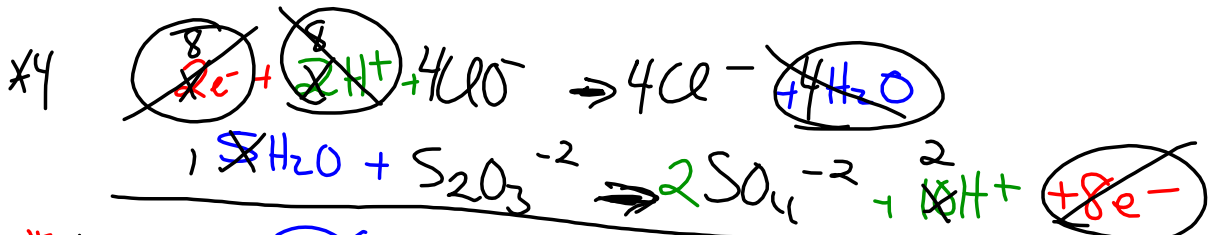
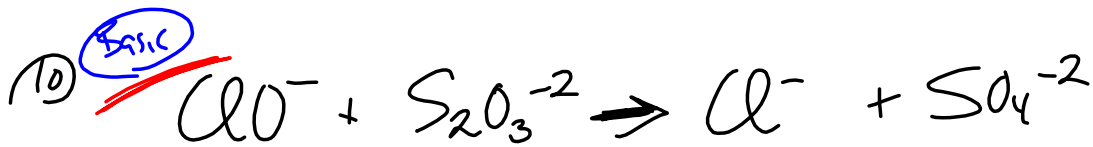
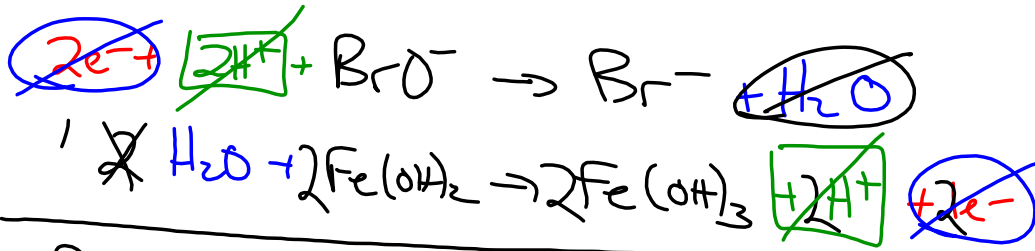
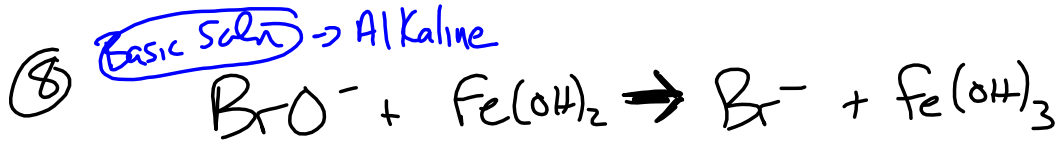


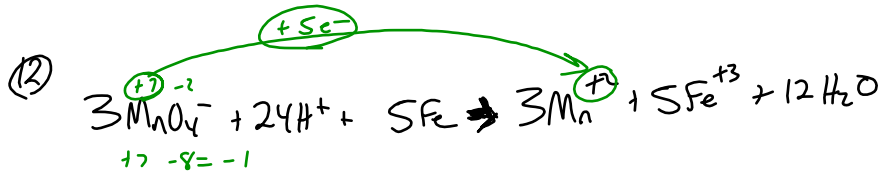
Reducing agent  $\Rightarrow$  LEO = OX  
 More  $e^- \rightarrow$  fewer  $e^-$   
 More  $\oplus$

$\text{S}^{-2}$  is undergoing OX

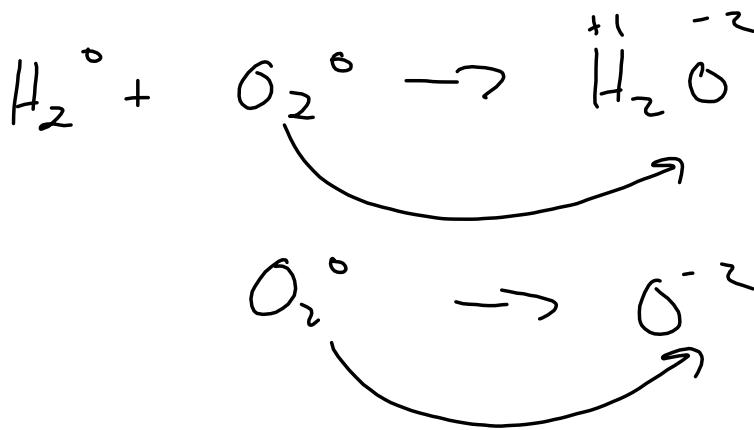
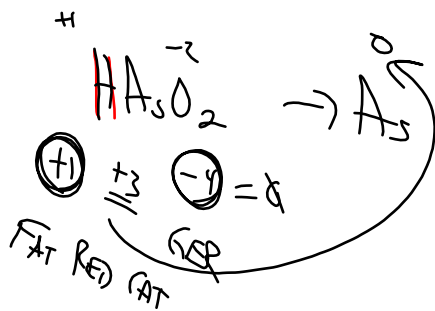
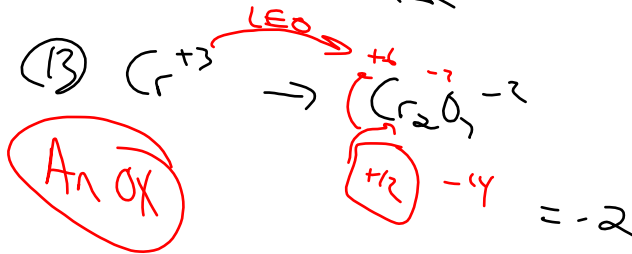
Always pick a reactant.







**Cathode** ⇒ FAT RED CAT  
 GER





$$\Delta G < 0$$

$$E^\circ > 0$$

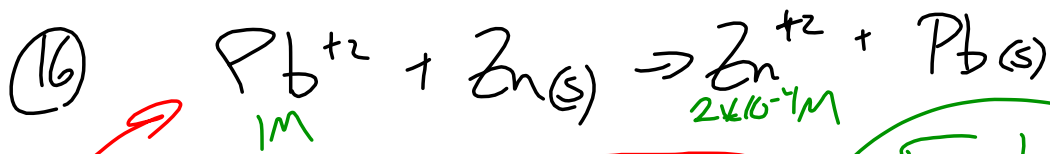
$$K > 1$$

(large)

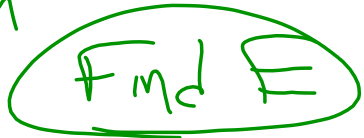
$$\ominus \Delta G$$

$$\oplus E^\circ$$

$$K = \frac{[P]}{[R]}$$



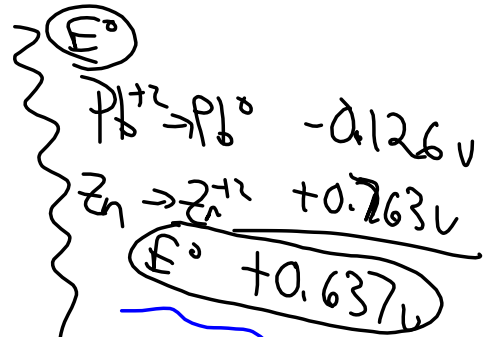
Short hand



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

$$= 0.637 - \frac{(8.314)(298)}{(2)(96500)} \ln \frac{2 \times 10^{-4}}{1}$$

$$E = 0.746V$$



$$K = \frac{[Zn^{+2}]}{[Pb^{+2}]}$$

Electrolysis - Need a Battery  
 Factor Label      Energy      NRG

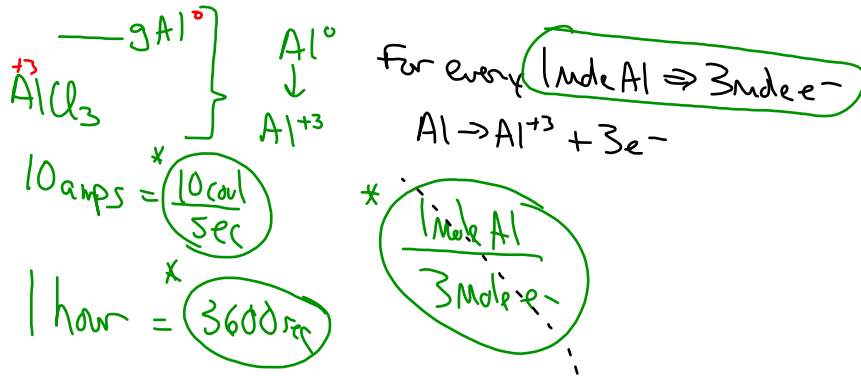
$$\frac{1 \text{ volt}}{1} = \frac{\text{Joule}}{\text{Coulomb}}$$

$$\frac{1 \text{ Faraday}}{1} = \frac{96500 \text{ coul}}{1 \text{ Mole of } e^-}$$

↑ value

$$\frac{\text{Ampere}}{1} = \frac{\text{Coulomb}}{\text{Sec}}$$

Ex) ? g Al is produced in 1 hour by electrolysis of molten  $\text{AlCl}_3$ . 10 amp current.



<del>1 mole Al</del>	27 g Al	<del>1 mole <math>e^-</math></del>	<del>10 coul</del>	<del>3600 sec</del>
3 mole $e^-$	<del>1 mole Al</del>	96500 coul	Sec	

3.36 g Al

Calc mass of Mg formed from <sup>molar</sup>  $MgCl_2$

60 amps over  $4 \times 10^3$  sec

$\frac{60 \text{ coul}}{\text{sec}}$ ,  $\frac{96500 \text{ coul}}{1 \text{ mole } e^-}$ ,  $\frac{1 \text{ mole } Mg}{24.3 \text{ Mg}}$ ,  $\frac{1 \text{ mole } Mg}{2 \text{ mole } e^-}$

$24.3 \text{ Mg}$	$1 \text{ mole } Mg$	$1 \text{ mole } e^-$	$60 \text{ coul}$	$4 \times 10^3 \text{ sec}$
$1 \text{ mole } Mg$	$2 \text{ mole } e^-$	$96500 \text{ coul}$	$\cancel{\text{sec}}$	$\cancel{\text{sec}}$

30.2 g Mg

HW PS 20-1

# 18, 19, 21, 23, 24, 26

Exam 3 soon