

KEY

GENERAL & ANALYTICAL CHEMISTRY I

CHMG-141

With Dr. Bailey

Name _____

Recitation Week11 Chemical Reactions

SOLUBLE COMPOUNDS

Almost all salts of Na^+ , K^+ , NH_4^+

Salts of nitrate, NO_3^-
chlorate, ClO_3^-
perchlorate, ClO_4^-
acetate, CH_3CO_2^-

Almost all salts of Cl^- , Br^- , I^-

Compounds containing F^-

Salts of sulfate, SO_4^{2-}

EXCEPTIONS

Halides of Ag^+ , Hg_2^{2+} , Pb^{2+}

Fluorides of Mg^{2+} , Ca^{2+} , Sr^{2+} , Ba^{2+} , Pb^{2+}

Sulfates of Ca^{2+} , Sr^{2+} , Ba^{2+} , Pb^{2+}

INSOLUBLE COMPOUNDS

Most salts of carbonate, CO_3^{2-}
phosphate, PO_4^{3-}
oxalate, $\text{C}_2\text{O}_4^{2-}$
chromate, CrO_4^{2-}

Most metal sulfides, S^{2-}

Most metal hydroxides and oxides

EXCEPTIONS

Salts of NH_4^+ and the alkali metal cations

$\text{Ba}(\text{OH})_2$ is soluble

Rules for Assigning Oxidation States

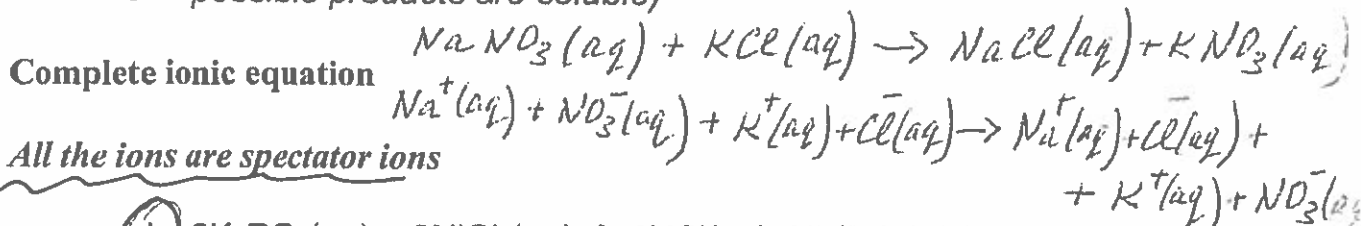
In their compounds, nonmetals have oxidation states according to the table below

✓ nonmetals higher on the table take priority

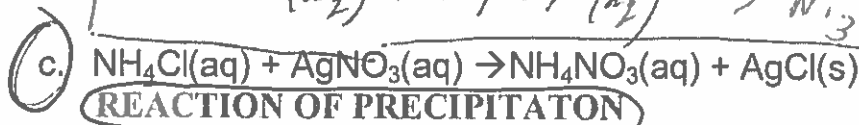
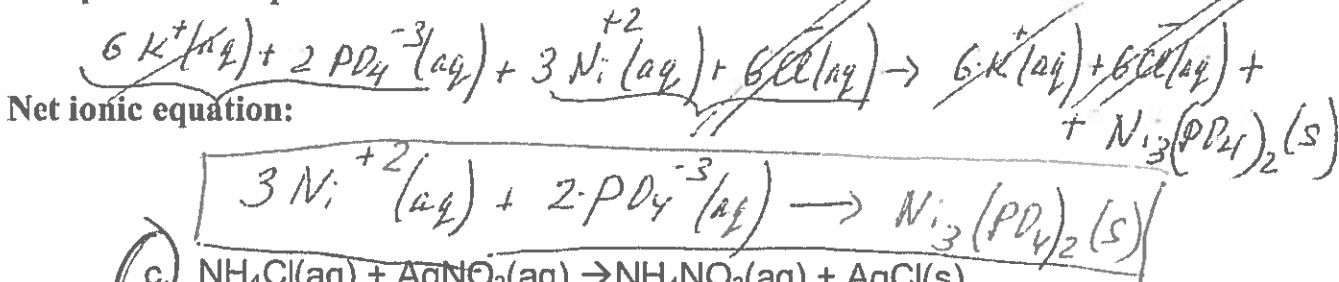
Nonmetal	Oxidation State	Example
F	-1	CF_4
H	+1	CH_4
O	-2	CO_2
Group 7A	-1	CCl_4
Group 6A	-2	CS_2
Group 5A	-3	NH_3

Problem 1:

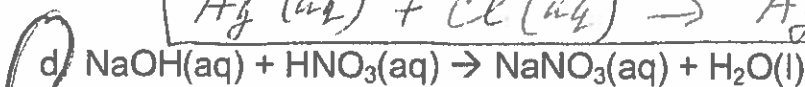
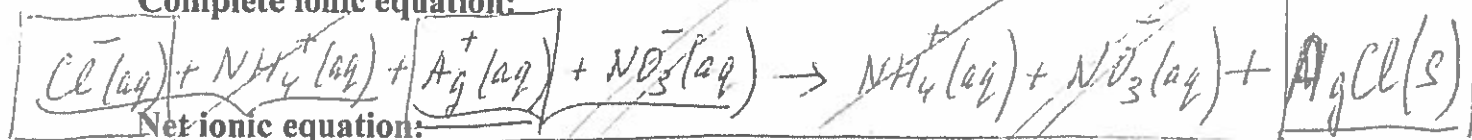
2) Complete and balance each of the following equations. If no reaction occurs, write NO REACTION:



Complete ionic equation:

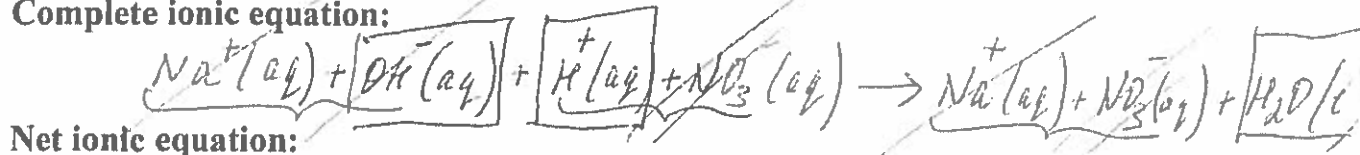


Complete ionic equation:

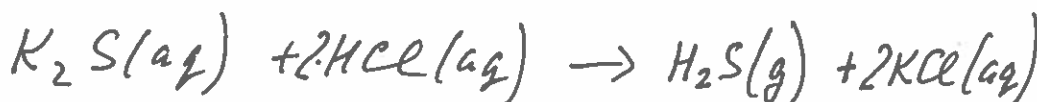


REACTION OF NEUTRALIZATION

Complete ionic equation:



Gas forming reaction



Problem 2 :

Write Balanced complete ionic and net ionic equation for each of the following reactions:

a. $\text{K}_2\text{SO}_4(\text{aq}) + \text{CaCl}_2(\text{aq}) \rightarrow \text{CaSO}_4(\text{s}) + 2\text{KCl}(\text{aq})$ – Molecular Equation; REACTION OF PRECIPITATION

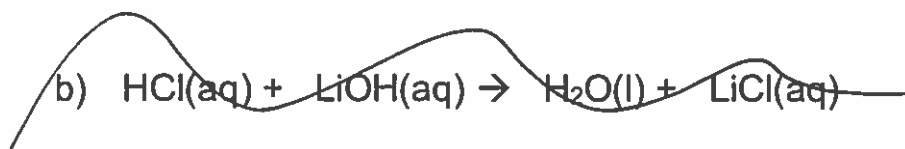
$2\text{K}^+(\text{aq}) + \text{SO}_4^{2-}(\text{aq}) + \text{Ca}^{2+}(\text{aq}) + 2\text{Cl}^-(\text{aq}) \rightarrow \text{CaSO}_4(\text{s}) + 2\text{K}^+(\text{aq}) + 2\text{Cl}^-(\text{aq})$ – Complete Ionic Equation

$\text{Ca}^{2+}(\text{aq}) + \text{SO}_4^{2-}(\text{aq}) \rightarrow \text{CaSO}_4(\text{s})$ – Net Ionic Equation

b. $\text{HCl}(\text{aq}) + \text{LiOH}(\text{aq}) \rightarrow \text{H}_2\text{O}(\text{l}) + \text{LiCl}(\text{aq})$ – Molecular Equation; REACTION of NEUTRALIZATION

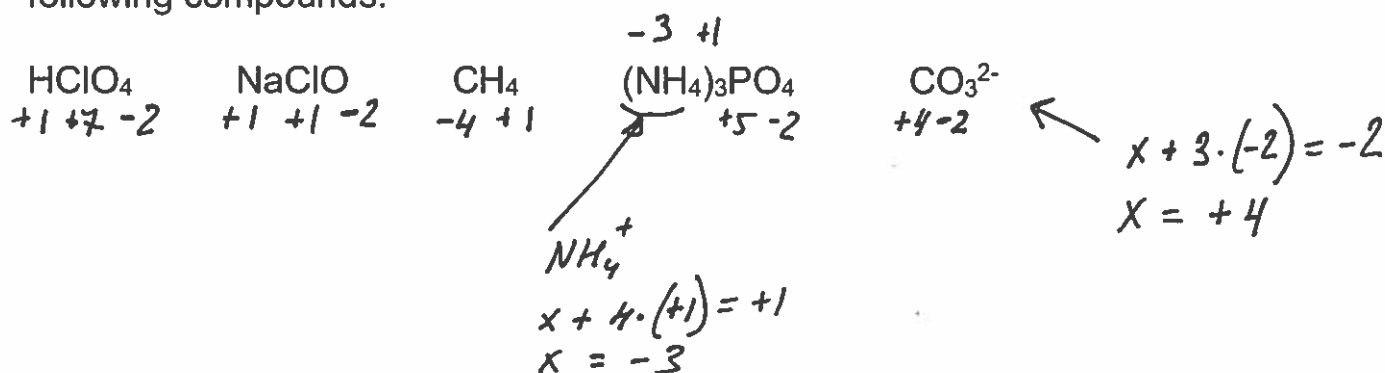
$\text{H}^+(\text{aq}) + \text{Cl}^-(\text{aq}) + \text{Li}^+(\text{aq}) + \text{OH}^-(\text{aq}) \rightarrow \text{Li}^+(\text{aq}) + \text{Cl}^-(\text{aq}) + \text{H}_2\text{O}(\text{l})$ – Complete Ionic Equation

$\text{H}^+(\text{aq}) + \text{OH}^-(\text{aq}) \rightarrow \text{H}_2\text{O}(\text{l})$ – Net Ionic Equation

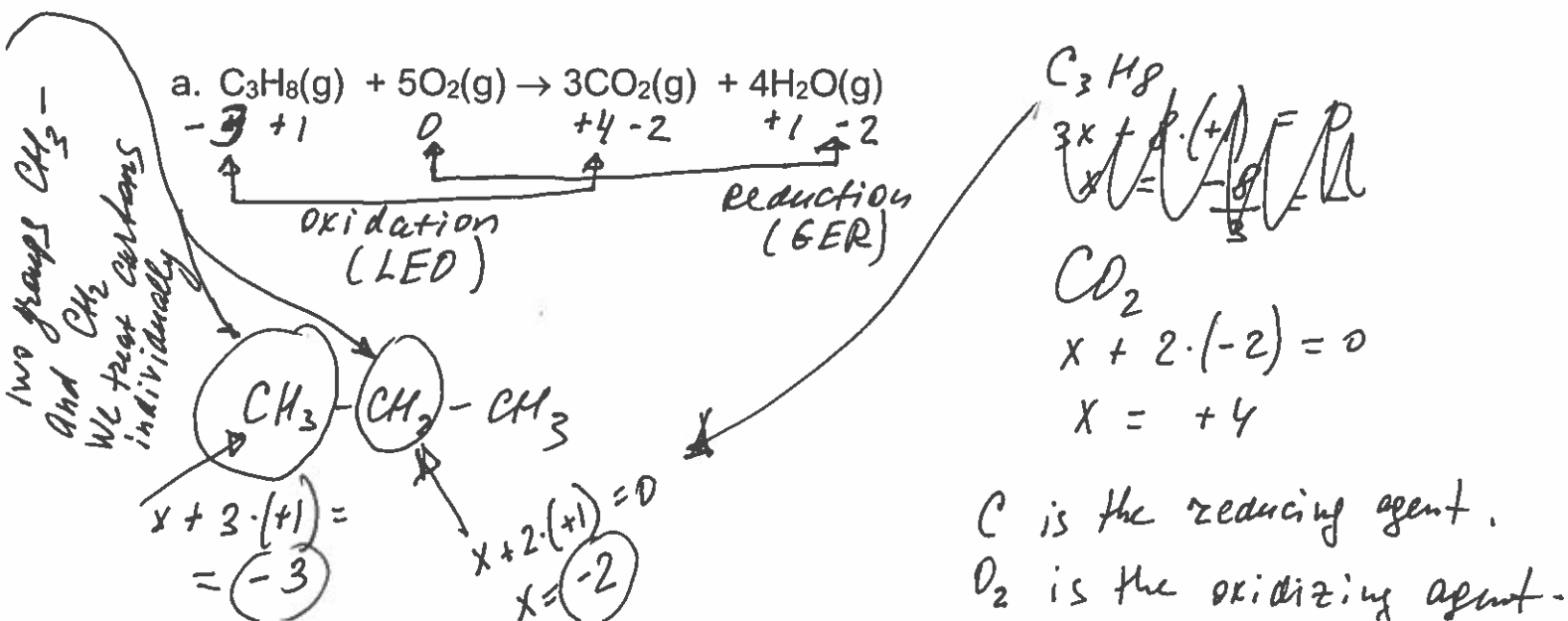


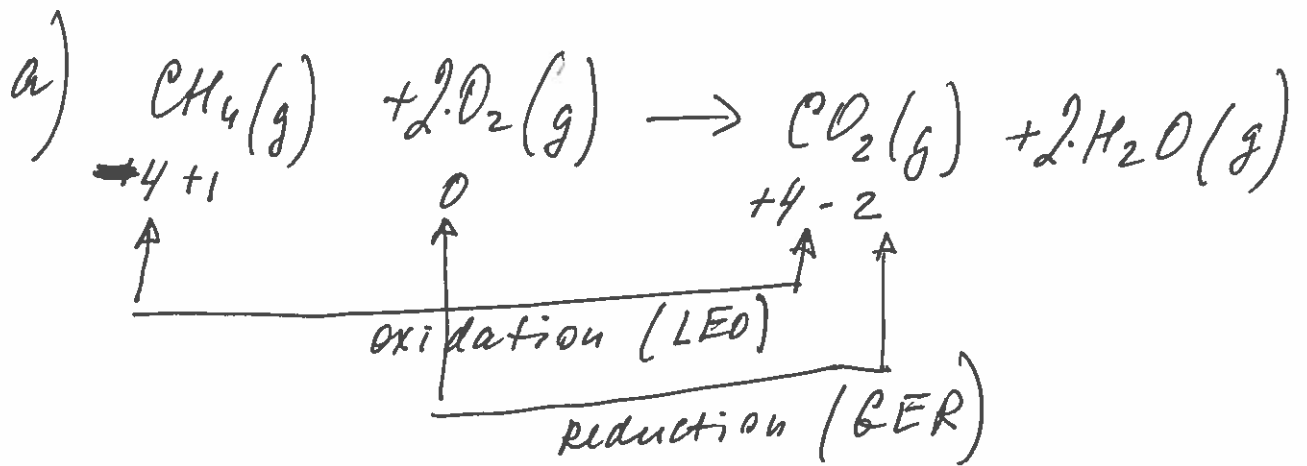
Part B

Problem 1: Determine the oxidation number for each element in the following compounds:



Problem 2: Determine if any of the following reactions are redox reactions. If so, identify the substance being oxidized and which is being reduced. Identify the oxidizing agent and the reducing agent.





Oxidation — ox. state increases
 Reduction — ox. state decreases.



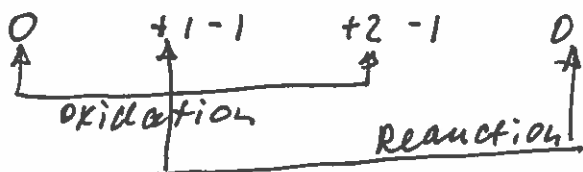
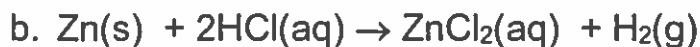
$$x + 4(+1) = 0$$

$$x = -4$$



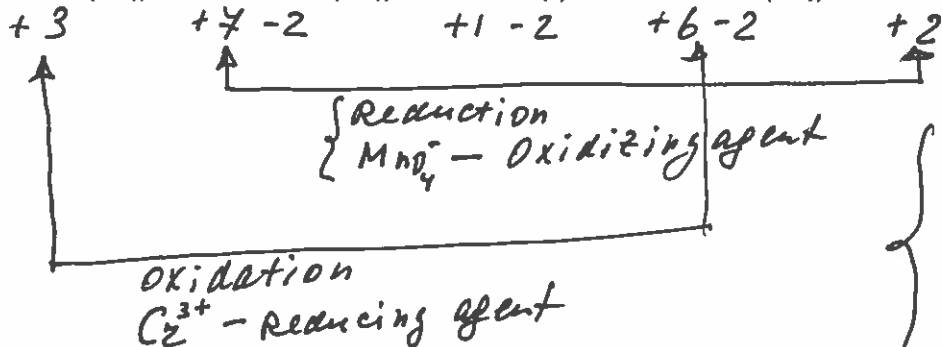
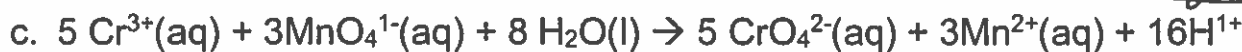
$$x + 2(-2) = 0$$

$$x = +4$$



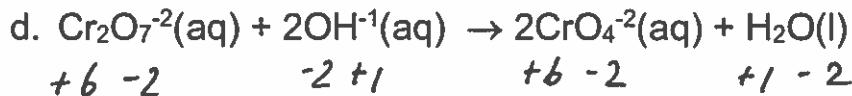
HCl is being reduced,
it is the oxidizing agent

Zn is being oxidized,
it is the reducing agent



MnO_4^-
 $x + 4(-2) = -1$
 $x = +7$

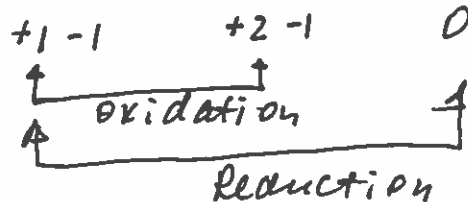
CrO_4^{2-}
 $x + 4(-2) = -2$
 $x = +6$



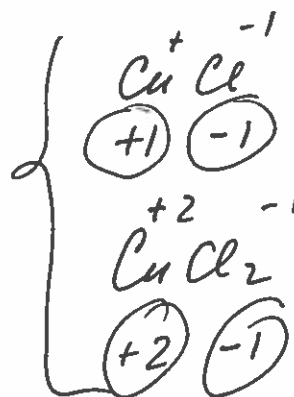
|| No change in the ox. states (numbers).
 || This is not a redox reaction.

$\text{Cr}_2\text{O}_7^{2-}$
 $2x + 7(-2) = -2$
 $x = \frac{+12}{2} = +6$

CrO_4^{2-}
 $x + 4(-2) = -2$
 $x = +6$



CuCl is being reduced,
it is the oxidizing agent
CuCl is being oxidized,
it is the reducing agent



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Part C

Balancing Ox/Red Reactions

Step 1. Write the unbalanced net ionic equation.

Step 2. Decide which atoms are oxidized and which are reduced, and write the two unbalanced half-reactions.

Step 3. Balance both half-reactions for all atoms except O and H.

Step 4. Balance each half-reaction for O by adding water to the side with less O, and balance for H by adding H^+ to the side with less H.

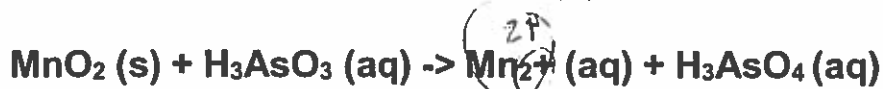
Step 5. Balance each half-reaction for charge by adding electrons to the side with greater positive charge, and then multiply by suitable factors to make the electron count the same in both half-reactions.

Step 6. Add the two balanced half-reactions together, and cancel electrons and other species that appear on both sides of the equation.

Check your answer by making sure the equation is balanced both for atoms and for charge.

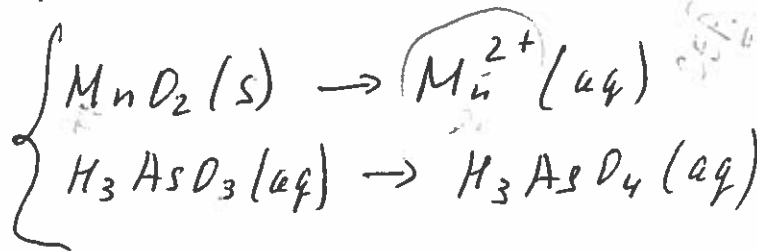
Balance the following reaction in (a) acidic solution and (b) basic solution

Balance the following reaction in (a) acidic solution and (b) basic solution



(a) acidic solution

- 1. Split the reaction into two half-reactions.**



2. Consider the following half reaction:

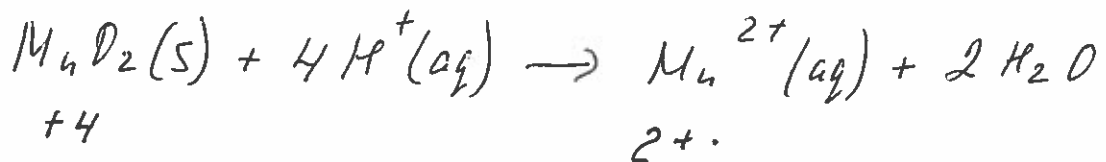


Balance everything but oxygen and hydrogen atoms.

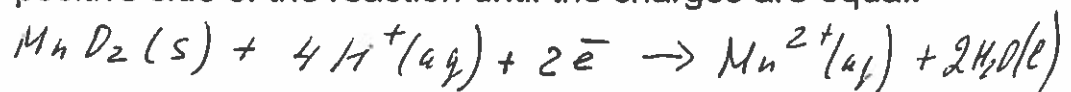
3. Take your answer to (2) and balance the oxygen by adding a water to add oxygen where needed.



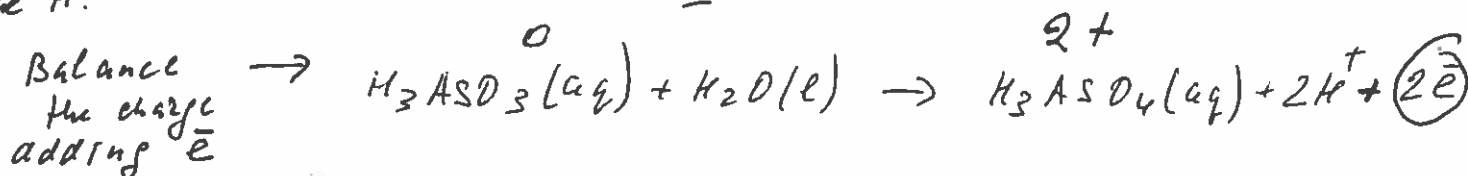
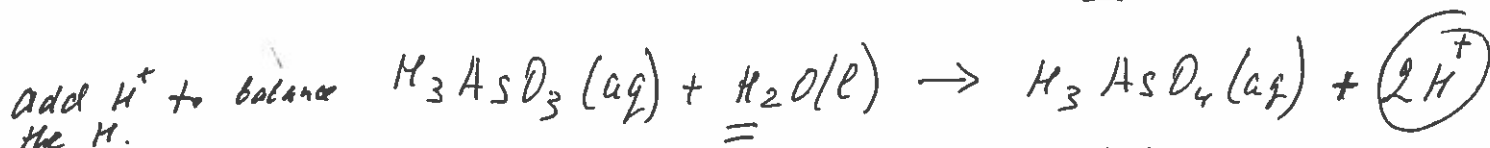
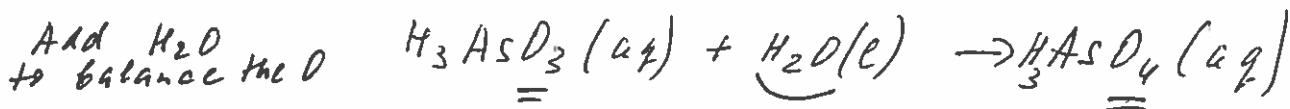
4. Take your answer to (3) and balance the hydrogen by adding in H^+ where needed (do not forget the charge on the H^+ !).



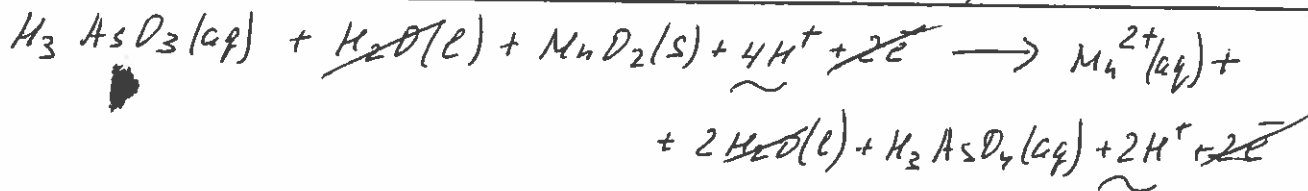
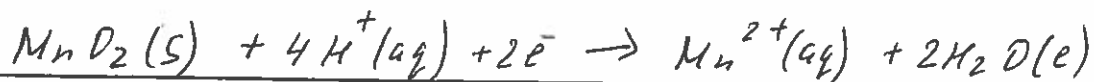
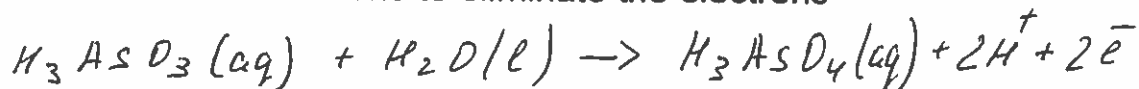
5. Take your answer to (4) and balance the charges by adding electrons to the more positive side of the reaction until the charges are equal.



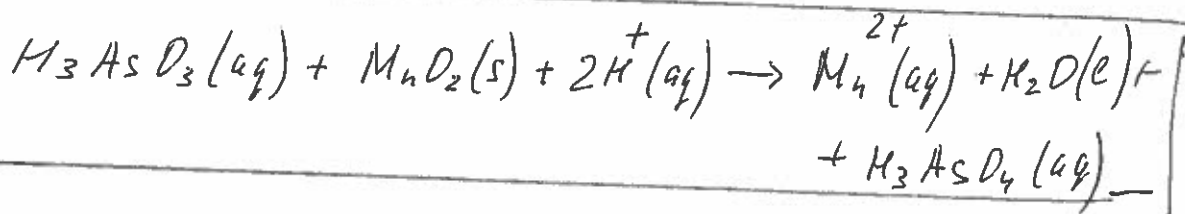
6. Balance the other half reaction in the same way



7. Combine the half reactions to eliminate the electrons

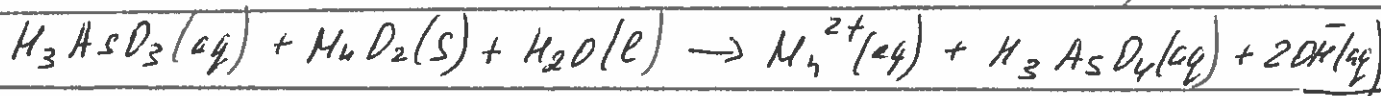
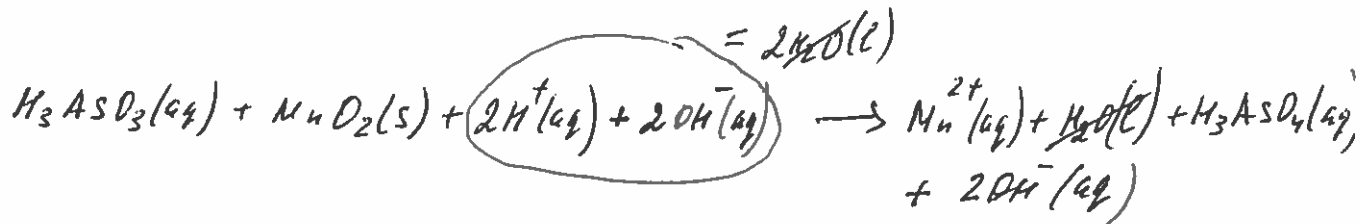


8. "Clean up": cancel other species that appear on both sides of the equation.



(b) Basic solution

Add OH⁻ to both sides of the equation to neutralize H⁺

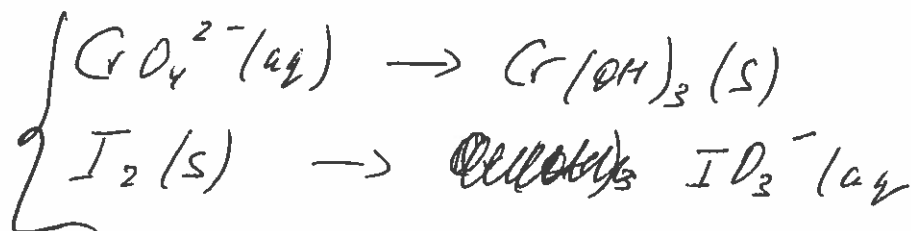


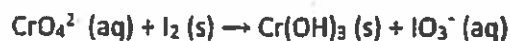
Problem B (additional, for your practice):

Balance the following reaction in basic solution

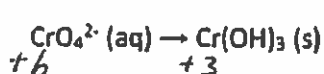


Half reactions:

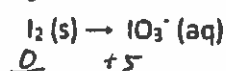




1. Separate into half reactions:

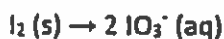


GER reduction



LEO oxidation

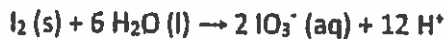
2. Balance everything but H and O



3. Balance O with H₂O



4. Balance H with H⁺

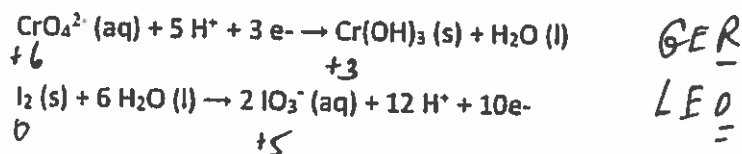


5. Balance charge with e⁻

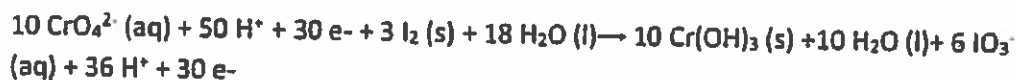
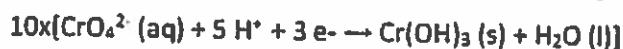
$$\begin{array}{l|l} \text{CrO}_4^{2-} & \text{Cr}(\text{OH})_3 \\ x + 4(-2) = -2 & x + 3(-2) + 3(+1) = \\ x = 8 - 2 = \boxed{+6} & x = \boxed{+3} = 0 \end{array}$$

$$\begin{array}{l} \text{IO}_3^- \\ x + 3(-2) = -1 \\ x = \boxed{+5} \end{array}$$

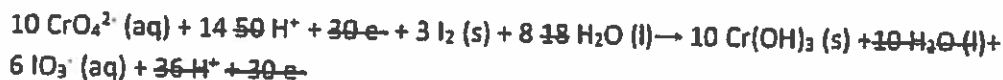
see also step 5



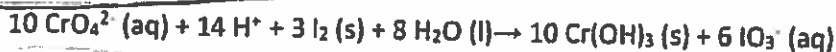
6. Combine to get rid of e-



Let's clean up a little bit:



In acidic solution



In Basic solution:

7. Add OH⁻ to neutralize H⁺ to both sides of the reaction:

