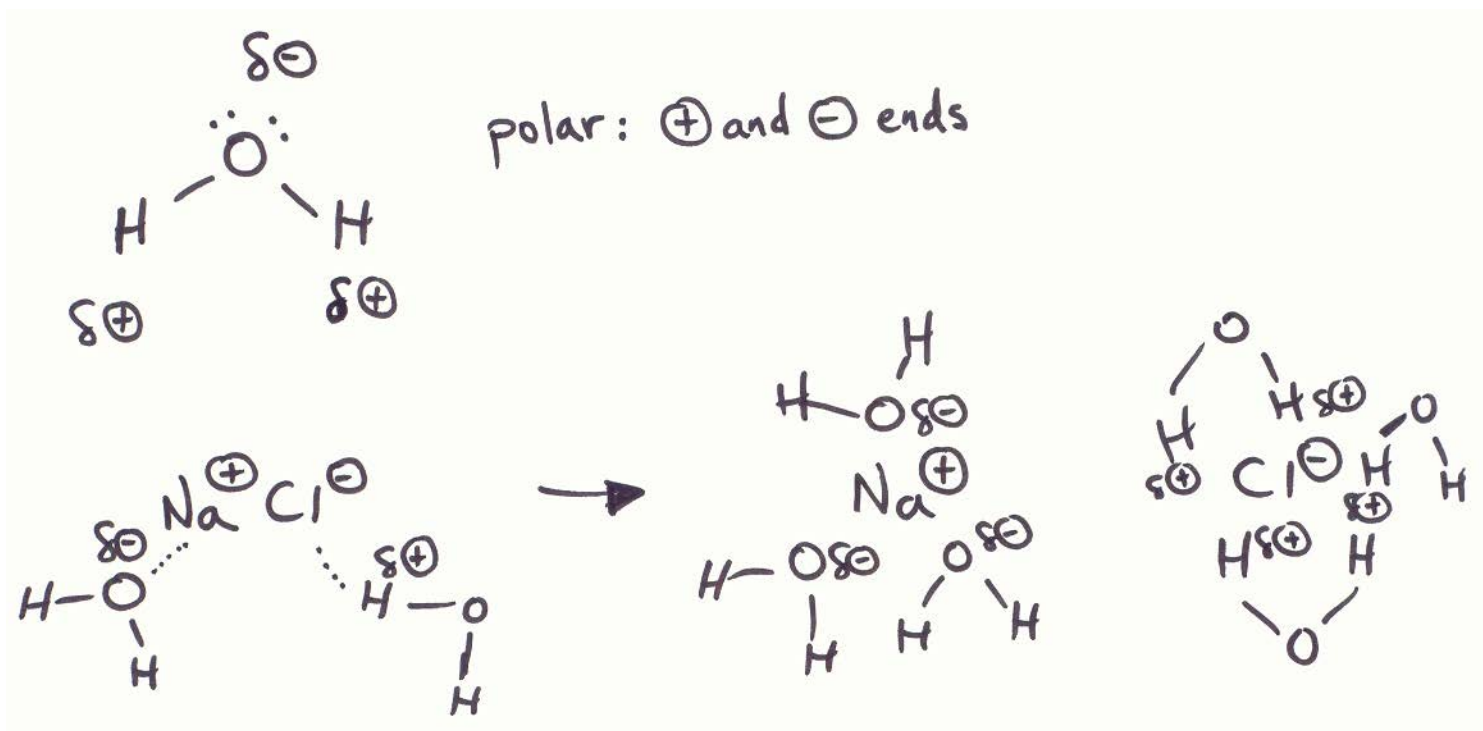


#19 Notes **Unit 3: Reactions in Solutions**
Ch. Reactions in Solutions

I. Solvation

-the act of dissolving (solute (salt) dissolves **in** the solvent (water))

Hydration: dissolving in water, the “universal” solvent.

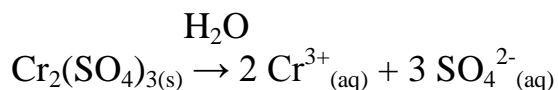


**The partially (-) oxygen pulls apart and surrounds the (+) cation.
The partially (+) hydrogen pulls apart and surrounds the (-) anion.

II. Electrolytes

A. Strong Electrolytes: completely ionize (fall apart) and will conduct electricity in a solution.

a) soluble salts**:



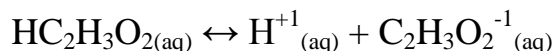
b) strong acids or bases:

strong acids start with “H” : HCl, HNO₃, H₂SO₄, HClO₄ {#O - #H ≥ 2 are strong acids}

strong bases end in “OH”: NaOH, Ca(OH)₂ (hydroxides)

B. Weak Electrolytes: only partially ionize (not all of the molecules fall apart) and conduct electricity poorly in solution.

a) slightly/marginally soluble salts, weak acids/bases:





Nonelectrolytes: may dissolve in water, but produce no ions H_2O
 (sugar: $\text{C}_6\text{H}_{12}\text{O}_{6(s)} \rightarrow \text{C}_6\text{H}_{12}\text{O}_{6(aq)}$)

III. Concentration

$$\text{Molarity} = \frac{\text{mol solute}}{\text{L solution}}$$

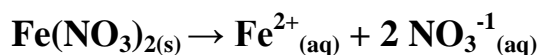
Ex. 1a) Calculate the molarity of 5.23 g $\text{Fe}(\text{NO}_3)_2$ in 100.0 cm^3 of solution.

$$\frac{5.23 \text{ g Fe}(\text{NO}_3)_2}{179.855 \text{ g}} \left| \frac{1 \text{ mol}}{179.855 \text{ g}} \right. = 0.029079 \text{ mol}$$

$$\frac{100.0 \text{ cm}^3}{1 \text{ cm}^3} \left| \frac{1 \text{ ml}}{1 \text{ cm}^3} \right| \left| \frac{1 \times 10^{-3} \text{ L}}{1 \text{ ml}} \right. = 0.100 \text{ L}$$

$$\text{M} = \text{mol} / \text{L} = \frac{0.029079 \text{ mol}}{0.100 \text{ L}} = \mathbf{0.291 \text{ M Fe}(\text{NO}_3)_2}$$

Ex. 1 b) Calculate the concentration of the ions in part “a” and write the solvation equation.



$$\begin{array}{ccc} 0.291 \text{ M Fe}(\text{NO}_3)_2 & & \\ \text{X1} \downarrow & \downarrow & \text{X2} \\ \mathbf{0.291 \text{ M Fe}^{2+}} & \downarrow & \\ & & \mathbf{0.582 \text{ M NO}_3^{-1}} \end{array}$$

Ex. 2) For 0.250 M $\text{Cr}_3(\text{PO}_4)_2$: Calculate the concentration of the ions and write the solvation equation.



$$\begin{array}{ccc} 0.250 \text{ M Cr}_3(\text{PO}_4)_2 & & \\ \text{X3} \downarrow & \downarrow & \text{X2} \\ \mathbf{0.750 \text{ M Cr}^{2+}} & \downarrow & \\ & & \mathbf{0.500 \text{ M PO}_4^{3-}} \end{array}$$

#20 Notes III. Concentration Continued:

Ex. 3) How do you make 500.0 cm³ of 4.00 M Fe(NO₃)₂ from solid Fe(NO₃)₂?

$$M = \text{mol} / \text{L} \quad \frac{500.0 \text{ cm}^3}{1 \text{ cm}^3} \left| \frac{1 \text{ ml}}{1 \text{ ml}} \right| \frac{1 \times 10^{-3} \text{ L}}{1 \text{ ml}} = 0.500 \text{ L}$$

$$4.00 \text{ M} = \frac{\text{mol}}{0.500 \text{ L}}$$

$$2.00 = \text{mol} \quad \frac{2.00 \text{ mol Fe(NO}_3)_2}{1 \text{ mol}} \left| \frac{179.855 \text{ g}}{1 \text{ mol}} \right| = 360 \text{ g Fe(NO}_3)_2$$

Add 360. g Fe(NO₃)₂ and dilute to 0.500 L with water.

Ex. 4) How would you make 1.5 L of 0.50 M NaOH from a 6.0 M stock solution of NaOH?

If we dilute orange juice, would we still have the same amount of actual orange stuff (just more water)? Yes, same # of orange particles, same # of mols

$$M = \text{mol} / \text{L} \quad \text{mol}_{\text{undiluted (strong)}} = \text{mol}_{\text{diluted (weak)}}$$

$$M \cdot L = \text{mol} \quad M_{\text{undiluted}} \cdot L_{\text{undiluted}} = M_{\text{diluted}} \cdot L_{\text{diluted}}$$
$$(6.0 \text{ M}) (L_{\text{undiluted}}) = (0.50 \text{ M}) (1.5 \text{ L})$$
$$L_{\text{undiluted}} = 0.13$$

$$1.5 \text{ L solution}$$
$$\underline{-0.13 \text{ L NaOH}}$$

Add 0.13 L of 6.0 M stock solution to 1.37 L H₂O.

IV. Precipitation Reactions:

-are reactions that form solids (precipitate).

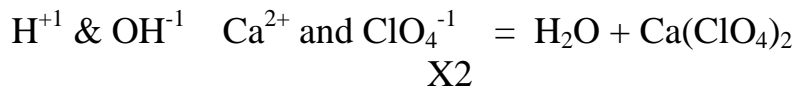
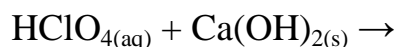
Solubility Rules ** soluble = aqueous
marginally/slightly soluble = solid

Solubility Table

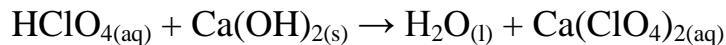
1. Nitrate (NO₃⁻¹) salts are soluble.
2. Ammonium (NH₄⁺¹) and Group I Alkali metal (Li⁺¹, Na⁺¹, K⁺¹, Cs⁺¹, Rb⁺¹) salts are soluble.
3. Chloride, Bromide, Iodide (Cl⁻¹, Br⁻¹, I⁻¹) salts are soluble, except for salts with Ag⁺¹, Pb²⁺ and Hg₂²⁺.
4. Sulfate (SO₄²⁻) salts are soluble, except Ba²⁺, Ca²⁺, Pb²⁺, and Hg₂²⁺.
5. Soluble Hydroxides are LiOH, NaOH and KOH. Marginally soluble: Ba(OH)₂, Sr(OH)₂ and Ca(OH)₂. The rest are slightly soluble.
6. Carbonates (CO₃²⁻), Sulfides (S²⁻), Chromates (CrO₄²⁻) and Phosphates (PO₄³⁻) are slightly soluble. (All with ammonium or Group 1 Alkali metals are soluble.)

#21 Notes V. Acid/Base Reactions

Ex. 1) Write the molecular, complete ionic and net ionic equations.

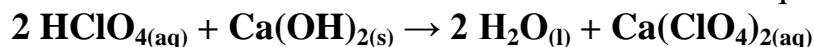


Water is a pure liquid.

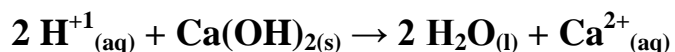
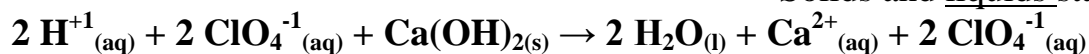


↑

No rule so aqueous.



Solids and liquids stay together.

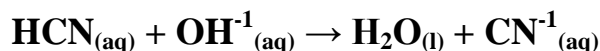
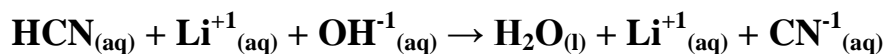
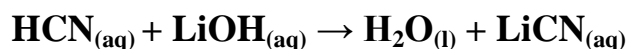


Ex. 2) Write the molecular, complete ionic and net ionic equations.

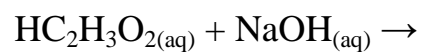


****weak acids stay together.**

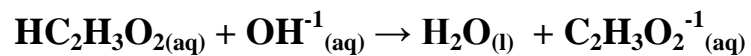
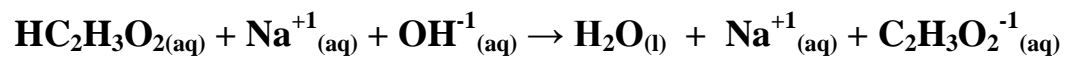
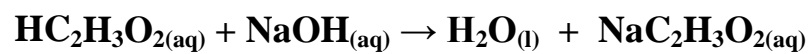
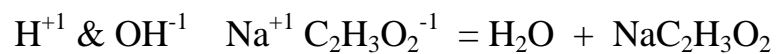
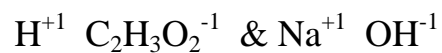
HCN is a weak acid.



Ex. 3) Write the molecular, complete ionic and net ionic equations.

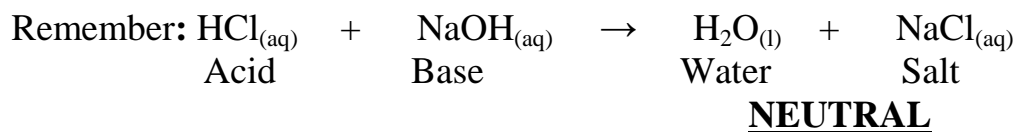


****HC₂H₃O₂ is a weak acid.**



#22 Notes VI. Titration Problems

Titration is a lab procedure in which acids and bases are mixed. If the concentration of one (acid or base) is known, the other (base or acid) can be calculated.



If there are equal amounts of acid and base, an equivalence point (neutral point) is reached. This can be determined by addition of an indicator, like the cabbage juice.

At the equivalence point: mols $\text{H}^{+1} = \text{mols OH}^{-1}$
 $M = \text{mol/L}$
 $M_{\text{H}^{+1}} \cdot L_{\text{H}^{+1}} = M_{\text{OH}^{-1}} \cdot L_{\text{OH}^{-1}}$

Ex. 1) What volume of 0.250 M HNO_3 is needed to react completely with 29.4 ml of 0.311M $\text{Ca}(\text{OH})_2$?

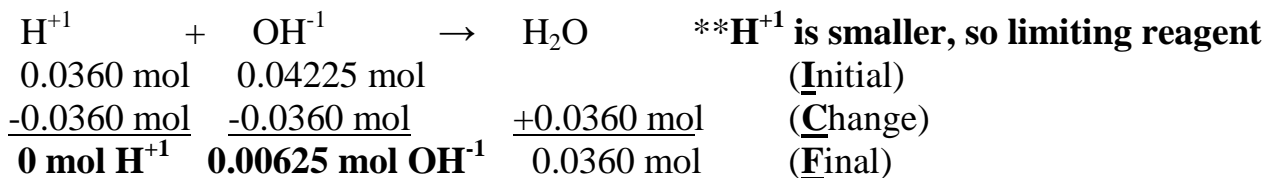
$$M_{\text{H}^{+1}} \cdot L_{\text{H}^{+1}} = M_{\text{OH}^{-1}} \cdot L_{\text{OH}^{-1}}$$

$$(0.250 \text{ M HNO}_3) (L_{\text{H}^{+1}}) = (2 \text{ OH}^{-1}) (0.311 \text{ M Ca}(\text{OH})_2) (29.4 \text{ ml})$$

$$(L_{\text{H}^{+1}}) = \mathbf{73.1 \text{ ml HNO}_3}$$

Ex. 2) 65 ml of 0.325 M $\text{Ca}(\text{OH})_2$ is added to 80 ml of 0.450 M HNO_3 . What is the concentration of excess H^{+1} or OH^{-1} ions?

M = mol/L	<u>OH^{-1}</u>	<u>H^{+1}</u>
	$(2 \text{ OH}^{-1}) (0.325 \text{ M}) = \frac{\text{mol}}{(0.065 \text{ L})}$	$(0.450 \text{ M}) = \frac{\text{mol}}{(0.080 \text{ L})}$
	$0.04225 = \text{mol OH}^{-1}$	$0.0360 = \text{mol H}^{+1}$



**0.00625 mol excess (left over) OH^{-1}
 Find concentration of excess OH^{-1} :

$$M = \text{mol/L} \quad M = \frac{0.00625 \text{ mol OH}^{-1} \text{ excess}}{0.065 \text{ L} + 0.080 \text{ L solution}}$$

$M = 0.043 \text{ M OH}^{-1}$, basic solution

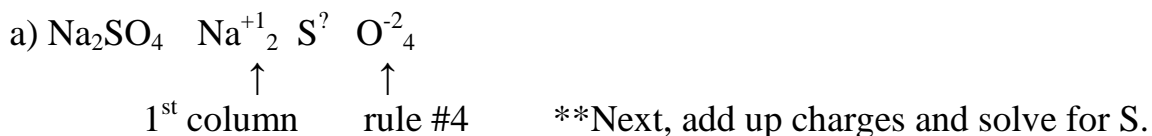
#23 Notes VII. Oxidation Numbers (for individual atoms):

Oxidation Rules:

- 1) Any free element (without a charge) is zero. Ex.) $\text{Zn}_{(s)}$, $\text{N}_{2(g)}$
- 2) Monoatomic ions have their charge. Ex.) $\text{Zn}^{2+}_{(aq)}$, $\text{F}^{-1}_{(aq)}$
- 3) Hydrogen is always +1, unless it is alone with a positive metal, then the H is -1.
Ex.) $\text{NaH} = \text{Na}^{+1} \text{H}^{-1}$
- 4) Oxygen is always -2, unless it is in a peroxide, then O is -1.
Ex.) $\text{H}_2\text{O}_2 = \text{H}^{+1} \text{O}^{-1}_2$
- 5) Fluorine is always -1.
- 6) The sum of oxidation numbers in a group of elements equals the charge of the group.
Ex.) NaCl so $\text{NaCl}^{=0}$, SO_4^{2-} so $\text{SO}_4^{=-2}$

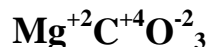
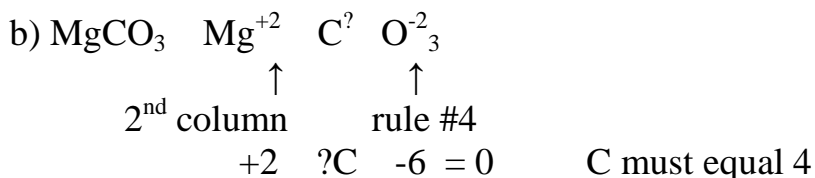
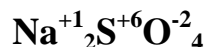
** Remember charges on Periodic Table Columns, for **single** element ions.
(1st column = +1, 2nd column = +2 etc.)

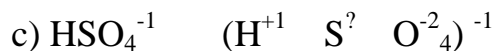
Ex.1) Find the charge on each element:



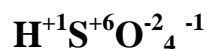
$$+2 \quad ?\text{S} \quad -8 = 0 \text{ (since a neutral compound)}$$

S must equal +6





$$+1 \quad ?\text{S} \quad -8 = -1 \quad \text{S must equal } +6$$

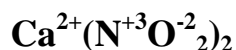


$$+2 \quad ?\text{N}(\times 2) \quad -2(4) = 0$$

$$+2 \quad 2\text{N} \quad -8 = 0$$

$$2\text{N} = 6$$

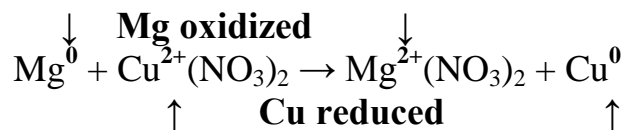
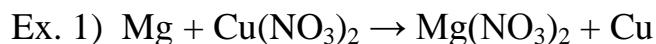
$$\text{N must equal } +3$$



VIII. Oxidation-Reduction Reactions (Redox)

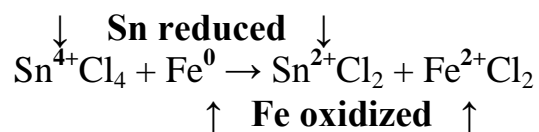
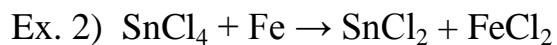
Oxidation- element gets more (+), it loses electrons

Reduction- element gets more (-), it gains electrons



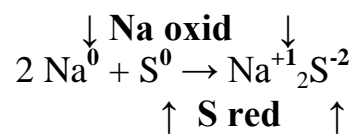
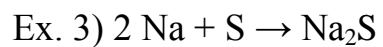
Mg oxidized, reducing agent

Cu reduced, oxidizing agent

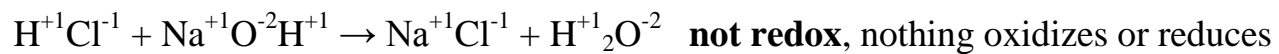


Sn reduced, oxidizing agent

Fe oxidized, reducing agent

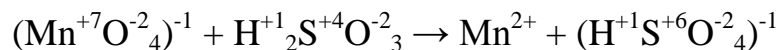
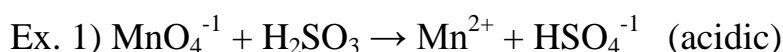


Na oxidized, reducing agent
S reduced, oxidizing agent



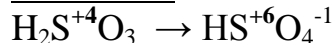
#24 Notes IX. Balancing Redox Reactions:

- 1) Write the half reactions. (one for oxidation, one for reduction)
- 2) For each half reaction:
 - a) Balance the electrons, by looking at the oxidation number of the elements that are oxidizing or reducing. (Make sure these elements are first balanced!)
 - b) Balance all other atoms except H and O. (If necessary, put in other compounds from the original reaction.)
 - c) Balance O by adding H₂O.
 - d) Balance H by adding H⁺.
 - e) Make sure all elements and charges balance.
- 3) Multiply the half reactions by numbers so that each has the same number of electrons.
- 4) Add half reactions and cancel what you can.
- 5) **If in basic solution:**
 - a) Notice how many H⁺ are left. Add the same amount of OH⁻ as H⁺, to both sides of the reaction.
 - b) The OH⁻ and H⁺ that are on the same side make H₂O.
 - c) Cancel any H₂O that will cancel.
- 6) Make sure all charges and atoms balance.

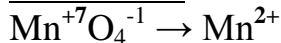


Step 1:

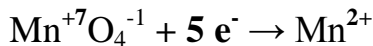
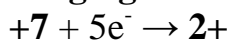
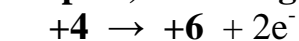
S Oxidizes



Mn reduces

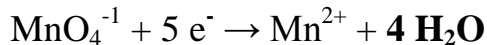
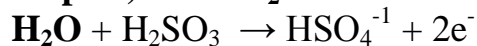


Step 2a): Add negative electrons to balance changing oxidation #'s.

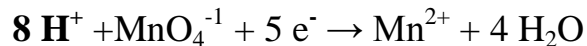


Step 2b): all balanced except H, O

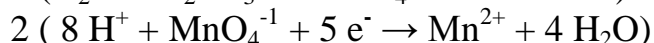
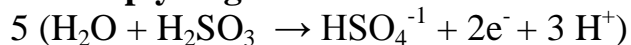
Step 2c): Add H₂O to balance O.

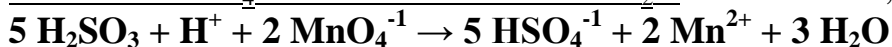
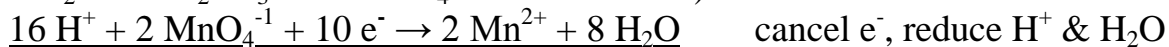
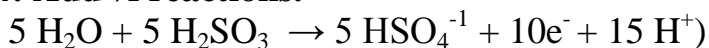


Step 2d): Add H⁺ to balance H.

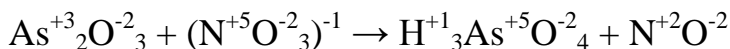


Step 3: Multiply to get common # of electrons.

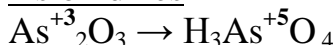


Step 4: Add 1/2 reactions.

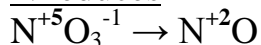
Ex. 2) $\text{As}_2\text{O}_3 + \text{NO}_3^{-1} \rightarrow \text{H}_3\text{AsO}_4 + \text{NO}$ (acidic)



As oxidizes

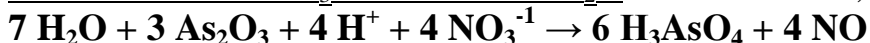
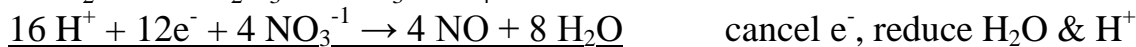
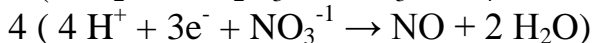
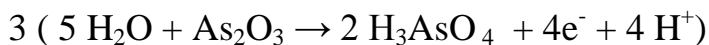
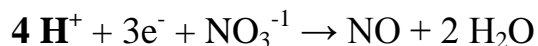
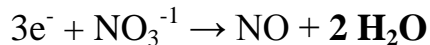
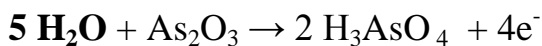
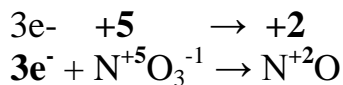
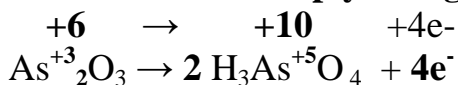


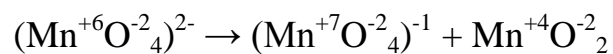
N reduces



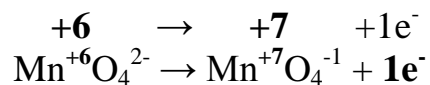
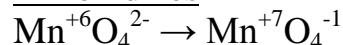
**** Balance As before doing electrons,**

so then multiply charges by 2, since there are 2 Arsenics!!

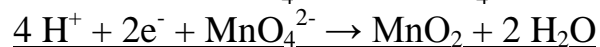
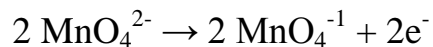
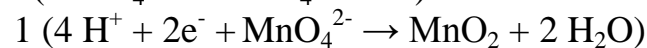
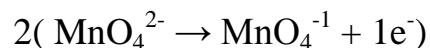
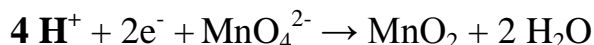
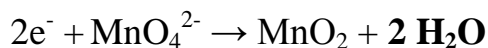
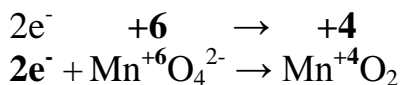
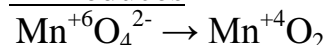




Mn oxidizes

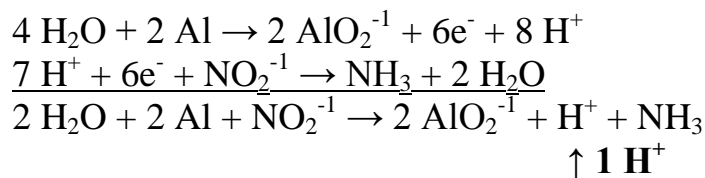
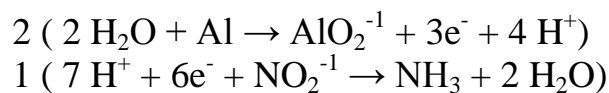
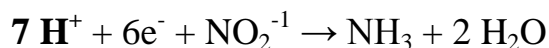
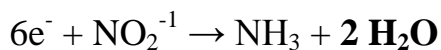
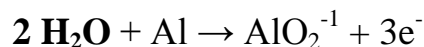
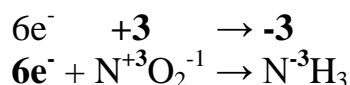
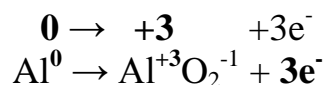
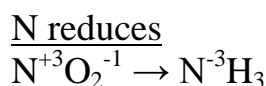
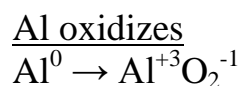
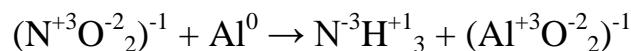
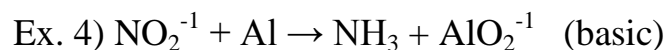


Mn reduces



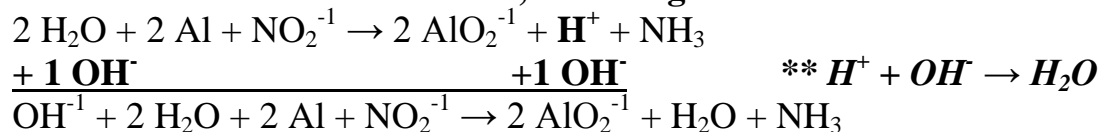
Pull down ions or compounds from the original reaction as needed to balance.

#25 Notes: Balancing continued.



Step 5: Note how many H^+ .

Add this amount of OH^- to both sides, canceling the H^+ .



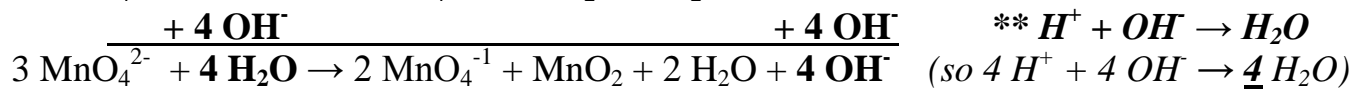
Cancel H_2O , if necessary.



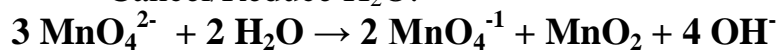
Ex. 5) Do Ex. 3) as if it were BASIC.



Note the number of H^+ , and add this amount of OH^- to both sides.



Cancel/Reduce H_2O :



End of Notes (Assignments #26-27 are Review Assignments. There are no notes for these assignments.)