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**:
      \[
      \begin{align*}
      H_2O & \rightarrow Na^+_{(aq)} + Cl^-_{(aq)} \\
      NaCl_{(s)} & \rightarrow Na^+_{(aq)} + Cl^-_{(aq)} \\
      Fe(NO_3)_{2(s)} & \rightarrow Fe^{2+}_{(aq)} + 2 NO_3^-_{(aq)} \\
      H_2O & \rightarrow Fe^{2+}_{(aq)} + 2 NO_3^-_{(aq)} \\
      Cr_2(SO_4)_{3(s)} & \rightarrow 2 Cr^{3+}_{(aq)} + 3 SO_4^{2-}_{(aq)} \\
      \end{align*}
      \]

   b) strong acids or bases:
   strong acids start with “H”: HCl, HNO_3, H_2SO_4, HClO_4
   strong bases end in “OH”: NaOH, Ca(OH)_2 (hydroxides)
   \[
   \begin{align*}
      H_2O & \rightarrow H^+_{(aq)} + NO_3^-_{(aq)} \\
      HNO_3_{(l)} & \rightarrow H^+_{(aq)} + NO_3^-_{(aq)} \\
      \end{align*}
   \]
\[
\begin{align*}
H_2O & \quad \text{NaOH} \rightarrow \text{Na}^{+} + \text{OH}^{-} \\
\text{bases produce} \text{OH}^{-} & \\
\end{align*}
\]

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

a) Weak or slightly/marginally soluble salts (*see Solubility Table)

\[
\begin{align*}
\text{H}_2\text{O} & \rightarrow 2 \text{Ag}^{+} + \text{S}^{2-} \\
\uparrow \text{equilibrium} & \text{ (There will be a balance, having some reactants and some products present.)}
\end{align*}
\]

b) Weak acids/bases:

\[
\begin{align*}
\text{HC}_2\text{H}_3\text{O}_2 & \rightarrow \text{H}^{+} + \text{C}_2\text{H}_3\text{O}_2^{-} \quad \text{starts with H} \\
\text{NH}_3 + \text{H}_2\text{O} & \rightarrow \text{NH}_4^{+} + \text{OH}^{-} \quad \text{has NH’s}
\end{align*}
\]

C. Nonelectrolytes: may dissolve in water, but produce no ions

\[
\begin{align*}
\text{H}_2\text{O} & \rightarrow \text{C}_6\text{H}_12\text{O}_6 \quad \text{sugar: C}_6\text{H}_12\text{O}_6 \rightarrow \text{C}_6\text{H}_12\text{O}_6\text{aq}
\end{align*}
\]

If time: Write the reactions for:

\[
\begin{align*}
a)\ \text{Ni(ClO}_2\text{)}_2 & \rightarrow \text{Ni}^{2+} + 2 \text{ClO}_2^{-} \\
b)\ \text{Cu}_3\text{(PO}_4\text{)}_2 & \rightarrow 3 \text{Cu}^{2+} + 2 \text{PO}_4^{3-}
\end{align*}
\]
#28 Notes  III. Concentration

Molarity = \( \frac{\text{mol solute}}{\text{L solution}} \)

1 ml = 1 cm\(^3\), 1 L = 1 dm\(^3\)

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

\[
\frac{5.23 \text{ g Fe(NO}_3\text{)}_2}{179.855 \text{ g}} = 0.029079 \text{ mol}
\]

\[
\frac{100.0 \text{ cm}^3}{1 \text{ cm}^3/\text{ml}} = 100.0 \text{ ml} = 0.100 \text{ L}
\]

\[
M = \frac{\text{mol}}{\text{L}} = \frac{0.029079 \text{ mol}}{0.100 \text{ L}} = 0.291 \text{ M Fe(NO}_3\text{)}_2
\]

Ex. 2) How do you make 500.0 cm\(^3\) of 4.00 M Fe(NO\(_3\))\(_2\) from solid Fe(NO\(_3\))\(_2\)?

**Find mols:**

\[
M = \frac{\text{mol}}{\text{L}} \quad \frac{500.0 \text{ cm}^3}{1 \text{ cm}^3} = \frac{1 \text{ ml}}{1 \text{ ml}} = 0.500 \text{ L}
\]

\[
4.00 \text{ M} = \frac{\text{mol}}{0.500 \text{ L}}
\]

2.00 = mol

**Find grams:**

\[
\frac{2.00 \text{ mol Fe(NO}_3\text{)}_2}{1 \text{ mol}} = \frac{179.855 \text{ g}}{1 \text{ mol}} = 360 \text{ g Fe(NO}_3\text{)}_2
\]

Add 360 g Fe(NO\(_3\))\(_2\) and dilute to 0.500 L with water.
Ex. 3) How would you make 1.5 L of 0.50 M NaOH from a 6.0 M stock solution of NaOH?

**This problem has 2 molarities!**

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

\[
M = \text{mol} / \text{L} \\
\text{mol undiluted (strong)} = \text{mol diluted (weak)}
\]

\[
M \cdot L = \text{mol}
\]

(substitute this in for mols undiluted and diluted)

\[
M_{\text{undiluted}} \cdot L_{\text{undiluted}} = M_{\text{diluted}} \cdot L_{\text{diluted}}
\]

\[
(6.0 \text{ M}) (L_{\text{undiluted}}) = (0.50 \text{ M}) (1.5 \text{ L})
\]

\[
L_{\text{undiluted}} = 0.13
\]

1.5 L solution

-0.13 L NaOH

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

Concentration of Ions

Ex. 4) Calculate the molarity of each ion for 76 g Fe(NO₃)₂ in 500.0 ml of solution.

\[
\frac{76 \text{ g Fe(NO}_3)_2}{179.855 \text{ g}} = 0.423 \text{ mol}
\]

\[
\frac{500.0 \text{ ml}}{1 \text{ ml}} = 0.500 \text{ L}
\]

\[
M = \text{mol} / \text{L} = \frac{0.423 \text{ mol}}{0.500 \text{ L}} = 0.85 \text{ M Fe(NO}_3)_2
\]

\[
\text{Fe(NO}_3)_2(s) \rightarrow \text{Fe}^{2+}(aq) + 2 \text{NO}_3^{-1}(aq)
\]

\[
0.85 \text{ M Fe(NO}_3)_2
\]

\[
\frac{X1}{X2} \rightarrow \frac{X2}{X1}
\]

\[
0.85 \text{ M Fe}^{2+} \rightarrow 1.7 \text{ M NO}_3^{-1}
\]
#29 Notes  IV. Precipitation Reactions:
-are reactions that form solids (precipitate).

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

**Solubility Table**
1. Nitrate (NO$_3^-$) salts are soluble.
2. Ammonium (NH$_4^+$) and Group I Alkali metal (Li$^+$, Na$^+$, K$^+$, Cs$^+$, Rb$^+$) salts are soluble.
3. Chloride, Bromide, Iodide (Cl$^-$, Br$^-$, I$^-$) salts are soluble, except for salts with Ag$^{+1}$, Pb$^{2+}$ and Hg$^{2+}$.
4. Sulfate (SO$_4^{2-}$) salts are soluble, except Ba$^{2+}$, Ca$^{2+}$, Pb$^{2+}$, and Hg$^{2+}$.
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 (CO$_3^{2-}$), Sulfides (S$^{2-}$), Chromates (CrO$_4^{2-}$) and Phosphates (PO$_4^{3-}$) are slightly soluble. (All with ammonium or Group 1 Alkali metals are soluble.)

Ex. 1) Write the **molecular** equation:

a) $\text{BaCl}_2(\text{aq}) + \text{Na}_2\text{SO}_4(\text{aq}) \rightarrow$

Step 1) Break compounds apart to their **Single ions**:

$\text{Ba}^{2+} \text{ Cl}^{-1} \text{ & Na}^{+1} \text{ SO}_4^{2-}$

2) Mix ions to make **new neutral compounds**:

$\text{Ba}^{2+} \text{ & SO}_4^{2-} \text{ Na}^{+1} \text{ & Cl}^{-1} = \text{BaSO}_4 + \text{NaCl}$

3) Use **Solubility Rules** from Table:

$\text{BaCl}_2(\text{aq}) + \text{Na}_2\text{SO}_4(\text{aq}) \rightarrow \text{BaSO}_4(\text{s}) + \text{NaCl}(\text{aq})$

↑ Rule #4  \ up ↑ Rule #2,3  **If no rule, then soluble (aq).**

4) Balance the equation:  $\text{BaCl}_2(\text{aq}) + \text{Na}_2\text{SO}_4(\text{aq}) \rightarrow \text{BaSO}_4(\text{s}) + 2 \text{NaCl}(\text{aq})$
Ex. 1b) \( \text{NaBr(aq)} + \text{Pb(NO}_3\text{)}_2(\text{aq}) \rightarrow \)

\[
\begin{align*}
\text{Na}^{+1} & \quad \text{Br}^{-1} & \quad \text{Pb}^{2+} & \quad \text{NO}_3^{-1} \\
\text{Na}^{+1} & \quad \text{NO}_3^{-1} & \quad \text{Pb}^{2+} & \quad \text{Br}^{-1} = \text{NaNO}_3 + \text{PbBr}_2
\end{align*}
\]

\[X2\]

\( \text{NaBr(aq)} + \text{Pb(NO}_3\text{)}_2(\text{aq}) \rightarrow \text{NaNO}_3(\text{aq}) + \text{PbBr}_2(\text{s}) \uparrow \uparrow \)

\[ \text{Rule } #1,2 \quad \text{Rule } #3 \]

\( 2 \text{NaBr(aq)} + \text{Pb(NO}_3\text{)}_2(\text{aq}) \rightarrow 2 \text{NaNO}_3(\text{aq}) + \text{PbBr}_2(\text{s}) \)

Ex. 1c) \( \text{Na}_2\text{CO}_3(\text{aq}) + \text{AlI}_3(\text{aq}) \rightarrow \)

\[
\begin{align*}
\text{Na}^{+1} & \quad \text{CO}_3^{2-} & \quad \text{Al}^{3+} & \quad \text{I}^{-1} \\
\text{Na}^{+1} & \quad \text{I}^{-1} & \quad \text{Al}^{3+} & \quad \text{CO}_3^{2-} = \text{NaI} + \text{Al}_2(\text{CO}_3)_3
\end{align*}
\]

\[X2 \quad X3\]

\( \text{Na}_2\text{CO}_3(\text{aq}) + \text{AlI}_3(\text{aq}) \rightarrow \text{NaI}(\text{aq}) + \text{Al}_2(\text{CO}_3)_3(\text{s}) \uparrow \uparrow \)

\[ \text{Rule } #2,3 \quad \text{Rule } #6 \]

\( 3 \text{Na}_2\text{CO}_3(\text{aq}) + 2 \text{AlI}_3(\text{aq}) \rightarrow 6 \text{NaI}(\text{aq}) + \text{Al}_2(\text{CO}_3)_3(\text{s}) \)

Ex. 1d) \( \text{KOH(aq)} + \text{Sr(NO}_3\text{)}_2(\text{aq}) \rightarrow \)

\[
\begin{align*}
\text{K}^{+1} & \quad \text{OH}^{-1} & \quad \text{Sr}^{2+} & \quad \text{NO}_3^{-1} \\
\text{K}^{+1} & \quad \text{NO}_3^{-1} & \quad \text{Sr}^{2+} & \quad \text{OH}^{-1} = \text{KNO}_3 + \text{Sr(OH)}_2
\end{align*}
\]

\[X2\]

\( \text{KOH(aq)} + \text{Sr(NO}_3\text{)}_2(\text{aq}) \rightarrow \text{KNO}_3(\text{aq}) + \text{Sr(OH)}_2(\text{s}) \uparrow \uparrow \)

\[ \text{Rule } #1,2 \quad \text{Rule } #5 \]

\( 2 \text{KOH(aq)} + \text{Sr(NO}_3\text{)}_2(\text{aq}) \rightarrow 2 \text{KNO}_3(\text{aq}) + \text{Sr(OH)}_2(\text{s}) \)

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Ex. 1) Write the molecular equation.

\[
\text{HClO}_4\text{(aq)} + \text{KOH}\text{(aq)} \rightarrow \text{H}^+ \text{ClO}_4^- \text{ & } \text{K}^+ \text{OH}^- \\
\text{H}^+ \text{ & OH}^- \text{ K}^+ \text{ & ClO}_4^- = \text{H}_2\text{O} + \text{KClO}_4 \\
\text{Water is a pure liquid.}
\]

\[
\text{HClO}_4\text{(aq)} + \text{KOH}\text{(aq)} \rightarrow \text{H}_2\text{O(l)} + \text{KClO}_4\text{(aq)} \\
\uparrow \\
\text{No rule, so soluble (aq)}
\]

Ex. 2) Write the molecular equation.

\[
\text{HNO}_3\text{(aq)} + \text{Sr(OH)}_2\text{(s)} \rightarrow \text{H}^+ \text{NO}_3^- \text{ & } \text{Sr}^{2+} \text{OH}^- \\
\text{H}^+ \text{ & OH}^- \text{ Sr}^{2+} \text{ and NO}_3^- = \text{H}_2\text{O} + \text{Sr(NO}_3)_2 \\
\text{X2}
\]

\[
2 \text{HNO}_3\text{(aq)} + \text{Sr(OH)}_2\text{(s)} \rightarrow 2 \text{H}_2\text{O(l)} + \text{Sr(NO}_3)_2\text{(aq)} \\
\uparrow \\
\text{Rule #1}
\]
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.

Remember:
\[
\text{HCl}_{(aq)} + \text{NaOH}_{(aq)} \rightarrow \text{H}_2\text{O}_{(l)} + \text{NaCl}_{(aq)}
\]

Acid    Base                   Water    Salt

**NEUTRAL**

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:
\[
m\text{ols } H^+ = m\text{ols } OH^- \\
M = \text{mol/L}
\]

\[
M_{H^+} \cdot L_{H^+} = M_{OH^-} \cdot L_{OH^-}
\]

Ex. 1) What is the concentration of a HCl solution, if 43.0 ml neutralizes 89.3 ml of 0.0400 M NaOH?

\[
M_{H^+} \cdot L_{H^+} = M_{OH^-} \cdot L_{OH^-} \\
(M_{H^+}) (43.0 \text{ ml}) = (0.0400 \text{ M NaOH}) (89.3 \text{ ml}) \\
(M_{H^+}) (43.0 \text{ ml}) = 3.572 \\
(M_{H^+}) = 0.0831 \text{ M HCl}
\]

Ex. 2) What volume of 0.250 M HNO₃ is needed to react completely with 29.4 ml of 0.311M Ca(OH)₂?

\[
M_{H^+} \cdot L_{H^+} = M_{OH^-} \cdot L_{OH^-} \\
(0.250 \text{ M HNO}_3) (L_{H^+}) = (2 \text{ OH}^-) (0.311 \text{ M Ca(OH)₂}) (29.4 \text{ ml}) \\
(L_{H^+}) = 73.1 \text{ ml HNO}_3
\]

H₂SO₄ ..... X₂ H⁺¹
Al(OH)₃ ..... X₃ OH⁻¹
VII. Oxidation Numbers (for individual atoms):

Oxidation Rules:

1) Any free element (without a charge) is zero.  Ex.)  Zn(s),  N₂(g)

2) Monoatomic ions have their charge.  Ex.)  Zn²⁺(aq),  F⁻(aq)

3) Hydrogen is always +1, unless it is alone with a positive metal, then the H is -1.
   Ex.)  NaH  =  Na⁺¹  H⁻¹

4) Oxygen is always -2, unless it is in a peroxide, then O is -1.
   Ex.)  H₂O₂  =  H⁺¹O⁻¹₂

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  NaCl = 0,  SO₄²⁻  so  SO₄ = -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:

a)  Na₂SO₄  \[ \text{Na}^{+1}_2 \text{S}^{2-} \text{O}^{-2}_4 \]
   ↑  ↑
   1ˢᵗ column  rule #4  **Next, add up charges and solve for S.
   +2 ?S -8 = 0 (since a neutral compound)
   S must equal +6

   \[ \text{Na}^{+1}_2 \text{S}^{+6} \text{O}^{-2}_4 \]

b)  MgCO₃  \[ \text{Mg}^{+2}_2 \text{C}^{+4} \text{O}^{+3}_3 \]
   ↑  ↑
   2ⁿᵈ column  rule #4
   +2 ?C -6 = 0  C must equal 4

   \[ \text{Mg}^{+2}_2\text{C}^{+4}\text{O}^{+3}_3 \]
c) \( \text{HSO}_4^- \quad (\text{H}^{+1} \quad \text{S}^{7} \quad \text{O}^{2-}_4)^{-1} \)

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

\[
\text{H}^{+1}\text{S}^{+6}\text{O}^{-2}_4^{-1}
\]

d) \( \text{Ca(NO}_2\text{)}_2 \quad \text{Ca}^{2+} \quad (\text{N}^{7} \quad \text{O}^{2-}_2)_2 \)

\[
+2 \quad ?\text{N(X2)} \quad -2(4) = 0
\]
\[
+2 \quad 2\text{N} \quad -8 = 0
\]
\[
2\text{N} = 6 \quad \text{N must equal +3}
\]

\[
\text{Ca}^{2+}(\text{N}^{+3}\text{O}^{-2}_2)_2
\]

VIII. Oxidation-Reduction Reactions (Redox)

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

Ex. 1) \( \text{Mg} + \text{Cu(NO}_3\text{)}_2 \rightarrow \text{Mg(NO}_3\text{)}_2 + \text{Cu} \)

\[
\downarrow \quad \text{Mg oxidized} \quad \downarrow \quad \text{Cu reduced}
\]
\[
\text{Mg}^{0} + \text{Cu}^{2+}(\text{NO}_3\text{)}_2 \rightarrow \text{Mg}^{2+}(\text{NO}_3\text{)}_2 + \text{Cu}^{0}
\]

\[
\uparrow \quad \text{Cu reduced, oxidizing agent} \quad \uparrow
\]

\[
\text{Mg oxidized, reducing agent}
\]

Ex. 2) \( \text{SnCl}_4 + \text{Fe} \rightarrow \text{SnCl}_2 + \text{FeCl}_2 \)

\[
\downarrow \quad \text{Sn reduced} \quad \downarrow
\]
\[
\text{Sn}^{4+}\text{Cl}_4 + \text{Fe}^{0} \rightarrow \text{Sn}^{2+}\text{Cl}_2 + \text{Fe}^{2+}\text{Cl}_2
\]

\[
\uparrow \quad \text{Fe oxidized} \quad \uparrow
\]

\[
\text{Sn reduced, oxidizing agent}
\]

\[
\text{Fe oxidized, reducing agent}
\]
Ex. 3) \(2 \text{Na} + \text{S} \rightarrow \text{Na}_2\text{S}\)

\[
\downarrow \text{Na oxid} \quad \downarrow \\
2 \text{Na}^0 + \text{S}^0 \rightarrow \text{Na}^{+1}_2\text{S}^{-2}
\]

\[
\uparrow \text{S red} \quad \uparrow
\]

Na oxidized, reducing agent
S reduced, oxidizing agent

Ex. 4) \(\text{HCl} + \text{NaOH} \rightarrow \text{NaCl} + \text{H}_2\text{O}\)

\[
\text{H}^{+1}\text{Cl}^{-1} + \text{Na}^{+1}\text{O}^{-2}\text{H}^{+1} \rightarrow \text{Na}^{+1}\text{Cl}^{-1} + \text{H}^{+1}_2\text{O}^{-2} \quad \text{not redox}, \text{nothing oxidizes or reduces}
\]
#33 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.

Ex. 1) MnO₄⁻¹ + H₂SO₃ → Mn²⁺ + HSO₄⁻¹ (acidic)

\[(\text{Mn}^{+7} \text{O}^{-2})^⁻ + \text{H}^{+1}_2 \text{S}^{+4}\text{O}^{-2}_3 \rightarrow \text{Mn}^{2+} + (\text{H}^{+1}\text{S}^{+6}\text{O}^{-2}_4)^⁻]\]

**Step 1:**

S Oxidizes
\[
\text{H}_2\text{S}^{+4}\text{O}_3 \rightarrow \text{HS}^{+6}\text{O}^{-1}_4
\]
Mn reduces
\[
\text{Mn}^{+7}\text{O}^{-1}_4 \rightarrow \text{Mn}^{2+}
\]

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

\[
\begin{align*}
+4 & \rightarrow +6 + 2e^- \\
\text{H}_2\text{S}^{+4}\text{O}_3 & \rightarrow \text{HS}^{+6}\text{O}^{-1}_4 + 2e^- \\
\text{Mn}^{+7}\text{O}^{-1}_4 & + 5 e^- \rightarrow \text{Mn}^{2+}
\end{align*}
\]

**Step 2b):** all balanced except H, O

**Step 2c): Add H₂O to balance O.**

\[
\text{H}_2\text{O} + \text{H}_2\text{SO}_3 \rightarrow \text{HSO}_4^{-1} + 2e^- \\
\text{MnO}_4^{-1} + 5 e^- \rightarrow \text{Mn}^{2+} + 4 \text{H}_2\text{O}
\]

**Step 2d): Add H⁺ to balance H.**

\[
\text{H}_2\text{O} + \text{H}_2\text{SO}_3 \rightarrow \text{HSO}_4^{-1} + 2e^- + 3 \text{H}^+ \\
8 \text{H}^+ + \text{MnO}_4^{-1} + 5 e^- \rightarrow \text{Mn}^{2+} + 4 \text{H}_2\text{O}
\]

**Step 3: Multiply to get common # of electrons.**

\[
5 (\text{H}_2\text{O} + \text{H}_2\text{SO}_3 \rightarrow \text{HSO}_4^{-1} + 2e^- + 3 \text{H}^+) \\
2 (8 \text{H}^+ + \text{MnO}_4^{-1} + 5 e^- \rightarrow \text{Mn}^{2+} + 4 \text{H}_2\text{O})
\]
Step 4: Add \( \frac{1}{2} \) reactions.

\[
5 \text{H}_2\text{O} + 5 \text{H}_2\text{SO}_3 \rightarrow 5 \text{HSO}_4^{-1} + 10\text{e}^- + 15 \text{H}^+
\]

\[
16 \text{H}^+ + 2 \text{MnO}_4^{-1} + 10 \text{e}^- \rightarrow 2 \text{Mn}^{2+} + 8 \text{H}_2\text{O}
\]

cancel \( \text{e}^- \), reduce \( \text{H}^+ \) & \( \text{H}_2\text{O} \)

\[
5 \text{H}_2\text{SO}_3 + \text{H}^+ + 2 \text{MnO}_4^{-1} \rightarrow 5 \text{HSO}_4^{-1} + 2 \text{Mn}^{2+} + 3 \text{H}_2\text{O}
\]

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

\[
\text{As}^{+3}_2\text{O}^{-2}_3 + (\text{N}^{+5}_\text{O}^{-2}_3)^{-1} \rightarrow \text{H}^{+1}_3\text{As}^{+5}\text{O}^{-2}_4 + \text{N}^{+2}\text{O}^{-2}
\]

As oxidizes

\[
\text{As}^{+3}_2\text{O}_3 \rightarrow \text{H}_3\text{As}^{+5}\text{O}_4
\]

N reduces

\[
\text{N}^{+5}\text{O}_3^{-1} \rightarrow \text{N}^{+2}\text{O}
\]

** Balance As before doing electrons, so then multiply charges by 2, since there are 2 Arsenics!!

\[
\begin{align*}
+6 & \rightarrow +10 +4\text{e}^- & 3\text{e}^- +5 & \rightarrow +2 \\
\text{As}^{+3}_2\text{O}_3 & \rightarrow 2 \text{H}_3\text{As}^{+5}\text{O}_4 + 4\text{e}^- & 3\text{e}^- + \text{N}^{+5}\text{O}_3^{-1} & \rightarrow \text{N}^{+2}\text{O}
\end{align*}
\]

\[
5 \text{H}_2\text{O} + \text{As}_2\text{O}_3 \rightarrow 2 \text{H}_3\text{AsO}_4 + 4\text{e}^- + \text{NO}_3^{-1} \rightarrow \text{NO} + 2 \text{H}_2\text{O}
\]

\[
5 \text{H}_2\text{O} + \text{As}_2\text{O}_3 \rightarrow 2 \text{H}_3\text{AsO}_4 + 4\text{e}^- + 4 \text{H}^+
\]

\[
4 \text{H}^+ + 3\text{e}^- + \text{NO}_3^{-1} \rightarrow \text{NO} + 2 \text{H}_2\text{O}
\]

3 ( \( 5 \text{H}_2\text{O} + \text{As}_2\text{O}_3 \rightarrow 2 \text{H}_3\text{AsO}_4 + 4\text{e}^- + 4 \text{H}^+ \))

4 ( \( 4 \text{H}^+ + 3\text{e}^- + \text{NO}_3^{-1} \rightarrow \text{NO} + 2 \text{H}_2\text{O} \))

\[
15 \text{H}_2\text{O} + 3 \text{As}_2\text{O}_3 \rightarrow 6 \text{H}_3\text{AsO}_4 + 12\text{e}^- + 12 \text{H}^+
\]

\[
16 \text{H}^+ + 12\text{e}^- + 4 \text{NO}_3^{-1} \rightarrow 4 \text{NO} + 8 \text{H}_2\text{O}
\]

cancel \( \text{e}^- \), reduce \( \text{H}_2\text{O} \) & \( \text{H}^+ \)

\[
7 \text{H}_2\text{O} + 3 \text{As}_2\text{O}_3 + 4 \text{H}^+ + 4 \text{NO}_3^{-1} \rightarrow 6 \text{H}_3\text{AsO}_4 + 4 \text{NO}
\]

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Ex.3 ) \( \text{MnO}_4^{2-} \rightarrow \text{MnO}_4^{-1} + \text{MnO}_2 \) (acidic)

\[
(Mn^{+6}O^{-2}_4)^{2-} \rightarrow (Mn^{+7}O^{-2}_4)^{-1} + Mn^{+4}O^{-2}_2
\]

Mn oxidizes
\[
\text{Mn}^{+6}O^{-2}_4 \rightarrow \text{Mn}^{+7}O^{-2}_4^{-1}
\]
\[
\begin{align*}
+6 & \rightarrow +7 +1e^- \\
\text{Mn}^{+6}O^{-2}_4 & \rightarrow \text{Mn}^{+7}O^{-2}_4^{-1} + 1e^- \\
2e^- +6 & \rightarrow +4 \\
\text{Mn}^{+6}O^{-2}_4 & \rightarrow \text{Mn}^{+4}O^{-2}_2
\end{align*}
\]

Mn reduces
\[
\text{Mn}^{+6}O^{-2}_4 \rightarrow \text{Mn}^{+4}O^{-2}_2
\]

\[
\begin{align*}
2e^- + \text{MnO}_4^{2-} & \rightarrow \text{MnO}_2 + 2 \text{H}_2\text{O} \\
4 \text{H}^+ + 2e^- + \text{MnO}_4^{2-} & \rightarrow \text{MnO}_2 + 2 \text{H}_2\text{O}
\end{align*}
\]

2( \( \text{MnO}_4^{2-} \rightarrow \text{MnO}_4^{-1} + 1e^- \))
1( \( 4 \text{H}^+ + 2e^- + \text{MnO}_4^{2-} \rightarrow \text{MnO}_2 + 2 \text{H}_2\text{O} \))

\[
\begin{align*}
2 \text{MnO}_4^{2-} & \rightarrow 2 \text{MnO}_4^{-1} + 2e^- \\
4 \text{H}^+ + 2e^- + \text{MnO}_4^{2-} & \rightarrow \text{MnO}_2 + 2 \text{H}_2\text{O} \\
3 \text{MnO}_4^{2-} + 4 \text{H}^+ & \rightarrow 2 \text{MnO}_4^{-31} + \text{MnO}_2 + 2 \text{H}_2\text{O}
\end{align*}
\]
Ex. 4) \( \text{NO}_2^- + \text{Al} \rightarrow \text{NH}_3 + \text{AlO}_2^- \) (basic)

\[
(N^{+3}O^{-2})^- + \text{Al}^0 \rightarrow N^{+3}H_3 + (\text{Al}^{+3}O^{-2})^-
\]

Al oxidizes
\[
\text{Al}^0 \rightarrow \text{Al}^{+3}O_2^- \\
0 \rightarrow +3 +3e^-
\]

N reduces
\[
N^{+3}O_2^- \rightarrow N^{+3}H_3 \\
6e^- +3 \rightarrow -3
\]

\[
2 \text{H}_2\text{O} + \text{Al} \rightarrow \text{AlO}_2^- + 3e^- \\
6e^- + \text{NO}_2^- \rightarrow \text{NH}_3 + 2 \text{H}_2\text{O}
\]

\[
2 \text{H}_2\text{O} + \text{Al} \rightarrow \text{AlO}_2^- + 3e^- + 4 \text{H}^+ \\
7 \text{H}^+ + 6e^- + \text{NO}_2^- \rightarrow \text{NH}_3 + 2 \text{H}_2\text{O}
\]

2 \(( 2 \text{H}_2\text{O} + \text{Al} \rightarrow \text{AlO}_2^- + 3e^- + 4 \text{H}^+)\)

1 \(( 7 \text{H}^+ + 6e^- + \text{NO}_2^- \rightarrow \text{NH}_3 + 2 \text{H}_2\text{O})\)

\[
4 \text{H}_2\text{O} + 2 \text{Al} \rightarrow 2 \text{AlO}_2^- + 6e^- + 8 \text{H}^+ \\
2 \text{H}_2\text{O} + 2 \text{Al} + \text{NO}_2^- \rightarrow 2 \text{AlO}_2^- + \text{H}^+ + \text{NH}_3
\]

\[
\uparrow 1 \text{H}^+
\]

Step 5: Note how many \( \text{H}^+ \).
Add this amount of \( \text{OH}^- \) to both sides, canceling the \( \text{H}^+ \).
\[
2 \text{H}_2\text{O} + 2 \text{Al} + \text{NO}_2^- \rightarrow 2 \text{AlO}_2^- + \text{H}^+ + \text{NH}_3
+ 1 \text{OH}^- + 1 \text{OH}^- \quad ** \text{H}^+ + \text{OH}^- \rightarrow \text{H}_2\text{O}
\]

\[
\text{OH}^- + 1 \text{H}_2\text{O} + 2 \text{Al} + \text{NO}_2^- \rightarrow 2 \text{AlO}_2^- + \text{H}_2\text{O} + \text{NH}_3
\]
Ex. 5) Do Ex. 3) as if it were BASIC.

\[ 3 \text{MnO}_4^{2-} + 4 \text{H}^+ \rightarrow 2 \text{MnO}_4^{-1} + \text{MnO}_2 + 2 \text{H}_2\text{O} \]

Note the number of \( \text{H}^+ \), and add this amount of \( \text{OH}^- \) to both sides.

\[ 3 \text{MnO}_4^{2-} + 4 \text{H}^+ \rightarrow 2 \text{MnO}_4^{-1} + \text{MnO}_2 + 2 \text{H}_2\text{O} \]

\[ + 4 \text{OH}^- \]

\[ 3 \text{MnO}_4^{2-} + 4 \text{H}_2\text{O} \rightarrow 2 \text{MnO}_4^{-1} + \text{MnO}_2 + 2 \text{H}_2\text{O} + 4 \text{OH}^- \]

\[ \text{so } 4 \text{H}^+ + 4 \text{OH}^- \rightarrow 4 \text{H}_2\text{O} \]

Cancel/Reduce \( \text{H}_2\text{O} \):

\[ 3 \text{MnO}_4^{2-} + 2 \text{H}_2\text{O} \rightarrow 2 \text{MnO}_4^{-1} + \text{MnO}_2 + 4 \text{OH}^- \]

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

*End of Notes*  (Assignments #35-37 are Review Assignments. There are no notes for these assignments.)