## UNIT 4 – CHAPTER 4 STUDENT NOTES: TYPES OF REACTIONS

AP chemistry recognizes three types of chemical reactions:

1) **Precipitate reactions**: $AX_{(aq)} + BY_{(aq)} \rightarrow AY_{(aq)} + BX_{(s)}$

2) **Redox reactions**: $Ca + FeCl_2 \rightarrow CaCl_2 + Fe$

3) **Acid-Base reactions**: $HA_{(aq)} + BOH_{(aq)} \rightarrow HOH_{(l)} + BA_{(aq)}$

### 1) Precipitate reactions

Molarity ($M$) – is an expression of concentration

$M$ = moles of solute / liters of solution

EX 1: Calculate the molarity of the following solutions

a) 11.5 grams of NaOH in 1.50 liters of water

b) 6.2 grams of $ZnCl_2$ in 450 milliliters of water

A) 11.5g NaOH | 1 mol NaOH / 40g NaOH = 0.2875 mol NaOH / 1.50 L Sol'n = 0.192M NaOH

B) 6.2g $ZnCl_2$ | 1 mol $ZnCl_2$ / 136.4g $ZnCl_2$ = 0.04545 mol $ZnCl_2$ / 0.45 L Sol'n = 0.101M $ZnCl_2$

EX 2: Calculate the moles of zinc ions and chlorine ions in solution b.

$ZnCl_2 \rightarrow Zn^{2+} + 2Cl^-$

$Zn^{2+}$ → 0.101M $ZnCl_2$ = 0.101 mol $ZnCl_2$ / 1 L Sol'n | 0.45 L Sol'n = 0.04545 mol $ZnCl_2$ | 1 mol $Zn^{2+}$ / 1 mol $ZnCl_2$ = 0.04545 mol $Zn^{2+}$ ions

$Cl^{-1}$ → 0.101M $ZnCl_2$ = 0.101 mol $ZnCl_2$ / 1 L Sol'n | 0.45 L Sol'n = 0.04545 mol $ZnCl_2$ | 2 mol $Cl^{-1}$ / 1 mol $ZnCl_2$ = 0.091 mol $Cl^-$ ions

EX 3: Blood is about 0.14 $M$ NaCl. What volume of blood contains 1.0 mg of NaCl?

1.0mg NaCl | 1g NaCl / 1000mg NaCl | 1 mol NaCl / 58.5g NaCl = $1.7\times10^{-9}$ mol NaCl

0.14M NaCl = $1.7\times10^{-9}$ mol NaCl / X L Sol'n

X = $1.22\times10^{-4}$ L or 0.122mL

**Dilutions**: amounts needed to make a solution

Solid

Liquid

Solubility Rules – pg. 144 of your book

1. Most nitrate ($NO_3^-$) salts are soluble.
2. Most salts containing the alkali metal ions ($Li^+$, $Na^+$, $K^+$, $Cs^+$, $Rb^+$) and the ammonium ion ($NH_4^+$) are soluble.
3. Most chloride, bromide, and iodide salts are soluble. Notable exceptions are salts containing the ions $Ag^+$, $Pb^{2+}$, and $Hg_2^{2+}$.
4. Most sulfate salts are soluble. Notable exceptions are $BaSO_4$, $PbSO_4$, $Hg_2SO_4$, and $CaSO_4$.
5. Most hydroxide salts are only slightly soluble. The important soluble hydroxides are NaOH and KOH. The compounds $Ba(OH)_2$, $Sr(OH)_2$, and $Ca(OH)_2$ are marginally soluble.
6. Most sulfide ($S^{2-}$), carbonate ($CO_3^{2-}$), chromate ($CrO_4^{2-}$), and phosphate ($PO_4^{3-}$) salts are only slightly soluble.

EX 4: Complete and balance the equation: lead (II) nitrate + potassium iodide →

$Pb(NO_3)_{2(aq)} + KI_{(aq)} \rightarrow PbI_{2(s)} + 2KNO_{3(aq)}$

Complete ion equation:

$Pb^{2+}_{(aq)} + 2NO_3^{-1}{}_{(aq)} + 2K^+_{(aq)} + 2I^{-1}_{(aq)} \rightarrow PbI_{2(s)} + 2K^+_{(aq)} + 2NO_3^{-1}{}_{(aq)}$

Net ionic equation:

$Pb^{2+}_{(aq)} + 2I^-_{(aq)} \rightarrow PbI_{2(s)}$ *

EX 5: copper (II) sulfate + sodium hydroxide

EX 6: potassium nitrate + barium chloride

$2KNO_{3(aq)} + BaCl_{2(aq)} \rightarrow 2KCl_{(aq)} + Ba(NO_3)_{2(aq)}$

NO REACTION - ALL SPECTATORS

**Precipitate stoichiometry** (Non-Redox)

Calculate the mass of solid NaCl that must be added to 1.50 liters of 0.100 *M* $AgNO_3$ solution to precipitate all of the $Ag^+$ ions in the form of AgCl.

$NaCl_{(aq)} + AgNO_{3(aq)} \rightarrow AgCl_{(s)} + NaNO_{3(aq)}$

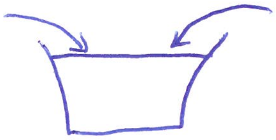
? 25.5g

$0.100M\ AgNO_3 = \frac{X \text{ mol } AgNO_3}{1.5 \text{ L soln}}$

$X = 0.15 \text{ mol } AgNO_3 \times \frac{1 \text{ mol NaCl}}{1 \text{ mol } AgNO_3} \times \frac{58.5 \text{ g NaCl}}{1 \text{ mol NaCl}} = \boxed{8.77 \text{ g NaCl}}$

What if $BaCl_2$ was used to precipitate the silver ions?

$Pb(NO_3)_2$ 0.500M 1.25L



$Na_2SO_4$ 0.0250M 2.0L

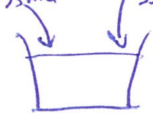
When aqueous solutions of $Na_2SO_4$ and $Pb(NO_3)_2$ are mixed, a precipitate is formed. Calculate the mass of precipitate formed when 1.25 liters of a 0.500 *M* $Pb(NO_3)_2$ solution is mixed with 2.00 liters of a 0.0250 *M* $Na_2SO_4$ solution.

$Na_2SO_{4(aq)} + Pb(NO_3)_{2(aq)} \rightarrow Pb(SO_4)_{(s)} + 2NaNO_{3(aq)}$

$SO_{4(aq)}^{-2} \rightarrow 0.0250M = \frac{0.0250\ mol\ Na_2SO_4}{1L} \times 2.0L = 0.05\ mol\ Na_2SO_4 \times \frac{1\ mol\ Pb(SO_4)}{1\ mol\ Na_2SO_4} \times \frac{303.2g\ PbSO_4}{1\ mol\ PbSO_4} = $ **15.2g $PbSO_4$** ← LIMITING

$Pb_{(aq)}^{+2} \rightarrow 0.500M = \frac{0.500\ mol\ Pb(NO_3)_2}{1L} \times 1.25L = 0.625\ mol\ Pb(NO_3)_2 \times \frac{1\ mol\ Pb(SO_4)}{1\ mol\ Pb(NO_3)_2} \times \frac{303.2g\ PbSO_4}{1\ mol\ PbSO_4} = 189.5g\ PbSO_4$ ← EXCESS

KOH .180M 55mL — $Fe(NO_3)_3$ .250M 35mL



What mass of precipitate is produced when 35 mL of 0.250 *M* $Fe(NO_3)_3$ solution is mixed with 55 mL of 0.180 *M* KOH solution?

$Fe(NO_3)_{3(aq)} + 3KOH_{(aq)} \rightarrow Fe(OH)_{3(s)} + 3KNO_{3(aq)}$

$Fe^{3+} \rightarrow 0.25M = \frac{.25\ mol\ Fe(NO_3)_3}{1L} \times .035L = 0.00875\ mol\ Fe(NO_3)_3 \times \frac{1\ mol\ Fe(OH)_3}{1\ mol\ Fe(NO_3)_3} \times \frac{106.9g\ Fe(OH)_3}{1\ mol\ Fe(OH)_3} = 0.935g\ Fe(OH)_3$ ← EXCESS

$OH^- \rightarrow 0.180M = \frac{.180\ mol\ KOH}{1L} \times .055L = .0099\ mol\ KOH \times \frac{1\ mol\ Fe(OH)_3}{3\ mol\ KOH} \times \frac{106.9g\ Fe(OH)_3}{1\ mol\ Fe(OH)_3} = $ **0.352g $Fe(OH)_3$** ← LIMITING

**2) Oxidation – Reduction Reactions (Redox)** **(OIL RIG!)**

$Fe + S \rightarrow FeS$

**Oxidation:**

**Reduction:**

**Rules for Assigning Oxidation States**

The oxidation state (OS) of an element corresponds to the number of electrons, $e^-$, that an atom loses, gains, or appears to use when joining with other atoms in compounds. In determining the oxidation state of an atom, there are seven guidelines to follow:

1. The oxidation state of an individual atom is 0.
2. The total oxidation state of all atoms in: a *neutral species* is 0 and in an *ion* is equal to the ion charge.
3. Group 1 metals have an oxidation state of +1 and Group 2 an oxidation state of +2
4. The oxidation state of fluorine is -1 in compounds
5. Hydrogen generally has an oxidation state of +1 in compounds
6. Oxygen generally has an oxidation state of -2 in compounds
7. In binary metal compounds, Group 17 elements have an oxidation state of -1, Group 16 elements of -2, and Group 15 elements of -3.

Assign oxidation states for the following:

$CO_2$

$NO_3^-$

$AgNO_3$

Assign oxidations states:

$AgNO_3 + NaCl \rightarrow NaNO_3 + AgCl$

$CH_4 + 2\ O_2 \rightarrow CO_2 + 2\ H_2O$

$2\ PbS + 3\ O_2 \rightarrow 2\ PbO + 2\ SO_2$

**Balancing redox reactions**

**1. Neutral solutions**

$Cu^{+}_{(aq)} + Fe_{(s)} \rightarrow Fe^{3+}_{(aq)} + Cu_{(s)}$

Step 1: Separate the half-reactions according to oxidation and reduction reactions.

$Cu^{+}_{(aq)} \rightarrow Cu_{(s)}$

$Fe_{(s)} \rightarrow Fe^{3+}_{(aq)}$

Step 2: Balance the charges by adding $e^-$ to the other side.

$Cu^{+}_{(aq)} + e^{-} \rightarrow Cu_{(s)}$

$Fe_{(s)} \rightarrow Fe^{3+}_{(aq)} + 3\ e^{-}$

Step 3: Balance the electrons by multiplying the entire equation by a multiplier.

$3\ Cu^{+}_{(aq)} + 3\ e^{-} \rightarrow 3\ Cu_{(s)}$

$Fe_{(s)} \rightarrow Fe^{3+}_{(aq)} + 3\ e^{-}$

Step 4: Add the equations.

$3\ Cu^{+}_{(aq)} + 3\ e^{-} + Fe_{(s)} \rightarrow 3\ Cu_{(s)} + Fe^{3+}_{(aq)} + 3\ e^{-}$

Step 5: Cancel out the electrons

$3\ Cu^{+}_{(aq)} + Fe_{(s)} \rightarrow 3\ Cu_{(s)} + Fe^{3+}_{(aq)}$

## 2. Acidic solutions

Step 1: Separate the half-reactions according to oxidation and reduction reactions.

Step 2: Balance elements OTHER than H and O.

Step 3: Balance O by adding $H_2O$ to the other side.

Step 4: Balance H by adding $H^+$ to the other side.

Step 5: Balance the charges by adding $e^-$ to the other side.

Step 6: Balance the electrons by multiplying the entire equation by a multiplier.

Step 7: Add the equations.

Step 8: Cancel out the electrons.

## 3. Basic solutions

Step 1: Separate the half-reactions according to oxidation and reduction reactions.

Step 2: Balance elements OTHER than H and O.

Step 3: Balance O by adding $H_2O$ to the other side.

Step 4: Balance H by adding $H^+$ to the other side.

Step 5: Balance the charges by adding $e^-$ to the other side.

Step 6: Balance the electrons by multiplying the entire equation by a multiplier.

Step 7: Add the equations.

Step 8: Cancel out the electrons.

Step 9: Add $OH^-$ to eliminate $H^+$ ions.

****Remember: If it is a basic solution, do all of the steps above, then add $OH^-$ to eliminate the $H^+$ ions****

EX 7: $HCl_{(aq)} + KMnO_{4\,(aq)} \rightarrow H_2O_{(l)} + KCl_{(aq)} + MnCl_{2(aq)} + Cl_{2\,(g)}$

EX 8: $Zn_{(s)} + HNO_{3\,(aq)} \rightarrow Zn(NO_3)_{2\,(aq)} + NO_{(l)} + H_2O_{(l)}$

0 +1 +5 −2 +4 −2 +2 −2 +1 −2

EX 9: $S_{(s)} + HNO_{3\,(aq)} \rightarrow SO_{2\,(g)} + NO_{(g)} + H_2O_{(l)}$

RED: $4(3e^- + 4H^+ + NO_3^- \rightarrow NO + 2H_2O)$

OXID: $3(2H_2O + S \longrightarrow SO_2 + 4H^+ + 4e^-)$

$12e^- + 16H^+ + 4NO_3^- + 6H_2O + 3S \rightarrow 4NO + 8H_2O + 3SO_2 + 12H^+ + 12e^-$

4 H⁺ rem. 2 H₂O rem.

$$3S + 4HNO_3 \rightarrow 3SO_2 + 4NO + 2H_2O$$

+1 +7 −2 +2 +6 −2 +1 +6 −2 +1 +6 −2 +2 +6 −2 +3 +6 −2 +1 −2

EX 10: $KMnO_{4(aq)} + FeSO_{4(aq)} + H_2SO_{4(aq)} \rightarrow K_2SO_{4(aq)} + MnSO_{4(aq)} + Fe_2(SO_4)_{3(aq)} + H_2O_{(l)}$

RED: $2(5e^- + 8H^+ + MnO_4^- \longrightarrow Mn^{2+} + 4H_2O)$

OXID: $10(Fe^{2+} \longrightarrow Fe^{3+} + 1e^-)$

$10e^- + 16H^+ + 2MnO_4^- + 10Fe^{2+} \rightarrow 2Mn^{2+} + 8H_2O + 10Fe^{3+} + 10e^-$

$$2KMnO_4 + 10FeSO_4 + 8H_2SO_4 \rightarrow K_2SO_4 + 2MnSO_4 + 5Fe_2(SO_4)_3 + 8H_2O$$

SINCE YOU CAN'T HAVE IONS FREE-FLOATING, NEED K⁺ + SO₄²⁻ BACK IN. TOUGH BALANCING!

**Net ionic equations**:

$Cr_2O_7{}^{2-}{}_{(aq)} + C_2H_5OH_{(l)} \rightarrow Cr^{3+} + CO_{2(g)}$

**Reminder on balancing redox reactions in <u>basic solutions</u>**

1) Use all the steps for an acid solution to obtain the final balanced equations

   - Identify the red-ox
   - Balance the elements
   - Balance the O with $H_2O$
   - Balance the H with $H^+$
   - Balance the charge with $e^-$
   - Equalize the $e^-$

2) Add $OH^-$ to eliminate the $H^+$ ions

EX 11: The following reaction occurs in a basic solution. Balance this equation using the half-reaction method.

$$Zn_{(s)} + NO_3^-{}_{(aq)} \rightarrow NO_{2(g)} + Zn^{2+}{}_{(aq)}$$

## 3) Acid-Base Reactions

Arrhenius definition:

Acid –

Base –

Brønsted-Lowry definition:

Acid –

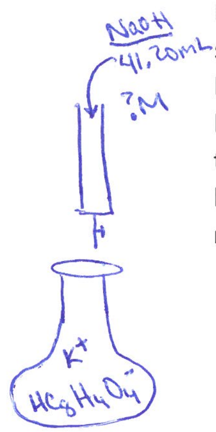
Base –

**Strong Acid-Base reaction:**

**Weak Acid-Base reaction:**

EX 12: What volume of 0.100 *M* HCl solution is needed to neutralize 25.0 mL of 0.350 *M* NaOH? Then replace HCl with $H_2SO_4$.

EX 13: In an experiment, 28 mL of 0.250 *M* $HNO_3$ and 53 mL of 0.320 *M* KOH are mixed. Calculate the amount of water formed in the reaction and calculate the concentration of $H^+$ or $OH^-$ in excess after the reaction goes to completion.




EX 14: A student carries out an experiment to standardize (determine the exact concentration of) a sodium hydroxide solution. To do this, the student masses out a 1.3009-gram sample of potassium hydrogen phthalate ($KHC_8H_4O_4$, often abbreviated KHP). KHP, molar mass 204.22 g/mol, has one acidic hydrogen. The student dissolves the KHP in distilled water, adds phenolphthalein as an indicator, and titrates the resulting solution with the sodium hydroxide solution to the end-point. The difference between the final and initial burette readings indicates that 41.20 mL of the sodium hydroxide solution is required to react exactly with the KHP. Calculate the concentration of the sodium hydroxide solution.

$$NaOH + KHC_8H_4O_4 \rightarrow Na^+ + OH^- \rightarrow K^+ + HC_8H_4O_4^-$$

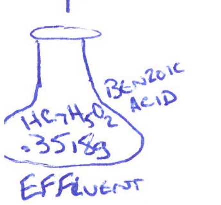
NET IONIC EQUATION

$$HC_8H_4O_4^- + OH^- \rightarrow H(OH) + C_8H_4O_4^{2-}$$

$$M = \frac{\text{MOL SOLUTE}}{\text{L SOL'N}}$$

$$1.3009\,g\ HC_8H_4O_4^- \times \frac{1\ \text{MOL KHP}}{204.22\,g\ \text{KHP}} \times \frac{1\ \text{MOL}\ OH^-}{1\ \text{MOL KHP}} = 0.006371\ \text{MOL}\ OH^-$$

$$M = \frac{0.006371\ \text{MOL}\ OH^-}{0.04120\,L} = \boxed{0.155\,M\ NaOH}$$




EX 15: An environmental chemist analyzed the effluent (the released waste material) from an industrial process known to produce the compound carbon tetrachloride ($CCl_4$) and benzoic acid ($HC_7H_5O_2$), a weak acid that has an acidic hydrogen per molecule. A sample of effluent massing 0.3518 grams was shaken with water, and the resulting aqueous solution required 10.59 mL of 0.1546 *M* NaOH for neutralization. Calculate the mass percent of benzoic acid in the original sample.

$$NaOH + HC_7H_5O_2 \rightarrow Na^+ + OH^- + HC_7H_5O_2$$

NET IONIC

$$OH^- + HC_7H_5O_2 \rightarrow H(OH) + C_7H_5O_2^-$$

.1546M    Xg

10.59mL

$$.1546M = \frac{.1546\ \text{MOL}\ OH^-}{1\ \text{L SOL'N}} \times 0.01059\,L \times \frac{1\ \text{MOL}\ HC_7H_5O_2}{1\ \text{MOL}\ OH^-} \times \frac{122.12\,g\ HC_7H_5O_2}{1\ \text{MOL}\ HC_7H_5O_2} = 0.1999\,g\ HC_7H_5O_2$$

$$\text{MASS\%} = \frac{0.1999\,g\ HC_7H_5O_2}{0.3159\ \text{EFFLUENT}} \times 100\% = \boxed{56.82\%}$$