

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

$$M = \text{moles of solute} / \text{liters of solution}$$

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

$$\begin{aligned} \text{a) } 11.5g \text{ NaOH} & \left| \frac{1 \text{ mol NaOH}}{40g \text{ NaOH}} \right| = \frac{0.2875 \text{ mol NaOH}}{1.50 \text{ L Sol'n}} \\ & = 0.192 \text{ M NaOH} \end{aligned} \quad \left| \quad \begin{aligned} \text{b) } 6.2g \text{ ZnCl}_2 & \left| \frac{1 \text{ mol ZnCl}_2}{136.4g \text{ ZnCl}_2} \right| = \frac{0.04545 \text{ mol ZnCl}_2}{0.45 \text{ L Sol'n}} \\ & = 0.10 \text{ M ZnCl}_2 \end{aligned}$$

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

$$\begin{aligned} \text{ZnCl}_2(aq) & \rightarrow \text{Zn}^{2+} + 2 \text{Cl}^- \\ \text{Zn}^{2+} \rightarrow 0.10 \text{ M ZnCl}_2 & = \frac{0.10 \text{ mol ZnCl}_2}{1 \text{ L Sol'n}} \left| \frac{0.45 \text{ L Sol'n}}{1 \text{ mol ZnCl}_2} \right| = 0.04545 \text{ mol ZnCl}_2 \left| \frac{1 \text{ mol Zn}^{2+}}{1 \text{ mol ZnCl}_2} \right| = 0.045 \text{ mol Zn}^{2+} \\ \text{Cl}^- \rightarrow 0.10 \text{ M ZnCl}_2 & = \frac{0.10 \text{ mol ZnCl}_2}{1 \text{ L Sol'n}} \left| \frac{0.45 \text{ L Sol'n}}{1 \text{ mol ZnCl}_2} \right| = 0.04545 \text{ mol ZnCl}_2 \left| \frac{2 \text{ mol Cl}^-}{1 \text{ mol ZnCl}_2} \right| = 0.091 \text{ mol Cl}^- \end{aligned}$$

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

$$1.0 \text{ mg NaCl} \left| \frac{1g \text{ NaCl}}{1000 \text{ mg NaCl}} \right| \left| \frac{1 \text{ mol NaCl}}{58.45g \text{ NaCl}} \right| = 1.7 \times 10^{-5} \text{ mol NaCl}$$

$$0.14 \text{ M} = \frac{1.7 \times 10^{-5} \text{ mol NaCl}}{x \text{ L Sol'n}} \rightarrow x = 1.2 \times 10^{-4} \text{ L} \text{ or } 0.12 \text{ mL}$$

Dilutions: amounts needed to make a solution

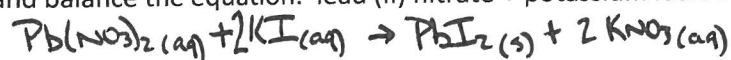
Solid

Liquid

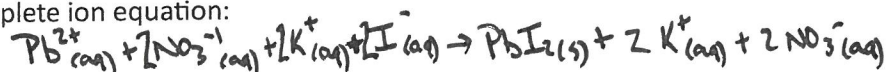
Solubility Rules – pg. 144 of your book

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 $\text{Ba}(\text{OH})_2$, $\text{Sr}(\text{OH})_2$, and $\text{Ca}(\text{OH})_2$ 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.

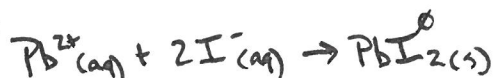
EX 4: Complete and balance the equation: lead (II) nitrate + potassium iodide $\rightarrow$



Complete ion equation:

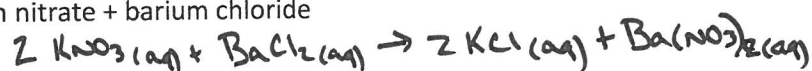


Net ionic equation:



EX 5: copper (II) sulfate + sodium hydroxide

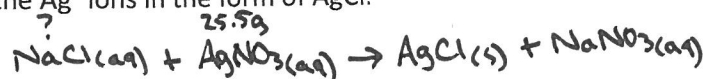
EX 6: potassium nitrate + barium chloride



NO REACTION - ALL SPECTATORS

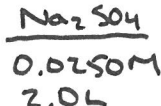
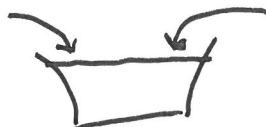
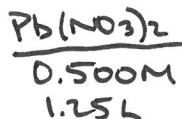
Precipitate stoichiometry (Non-Redox)

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 .

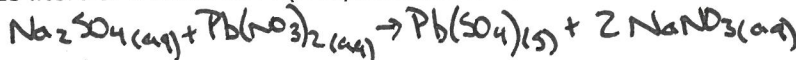


$$0.100\text{M AgNO}_3 = \frac{x \text{ mol AgNO}_3}{1.5 \text{ L SOLN}} \rightarrow x = 0.15 \text{ mol AgNO}_3 \quad \frac{1 \text{ mol NaCl}}{1 \text{ mol AgNO}_3} \quad \frac{58.5 \text{ g NaCl}}{1 \text{ mol NaCl}} = \boxed{8.77 \text{ g NaCl}}$$

What if BaCl_2 was used to precipitate the silver ions?



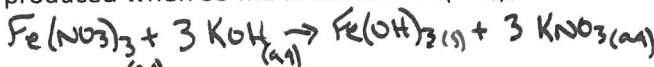
When aqueous solutions of Na_2SO_4 and $\text{Pb}(\text{NO}_3)_2$ are mixed, a precipitate is formed. Calculate the mass of precipitate formed when 1.25 liters of a 0.500 M $\text{Pb}(\text{NO}_3)_2$ solution is mixed with 2.00 liters of a 0.0250 M Na_2SO_4 solution.



$$\text{SO}_4^{2-}(\text{aq}) \rightarrow 0.0250\text{M} = \frac{0.0250\text{mol Na}_2\text{SO}_4}{1\text{L SOLN}} \times \frac{2.0\text{L}}{1\text{mol Na}_2\text{SO}_4} = 0.05\text{mol Na}_2\text{SO}_4 \left| \frac{1\text{mol PbSO}_4}{1\text{mol Na}_2\text{SO}_4} \right| \frac{303.2\text{g PbSO}_4}{1\text{mol PbSO}_4} = 15.2\text{g PbSO}_4 \quad \text{LR}$$

$$\text{Pb}^{2+}(\text{aq}) \rightarrow 0.500\text{M} = \frac{0.500\text{mol Pb}(\text{NO}_3)_2}{1\text{L SOLN}} \times \frac{1.25\text{L}}{1\text{mol Pb}(\text{NO}_3)_2} = 0.625\text{mol Pb}(\text{NO}_3)_2 \left| \frac{1\text{mol PbSO}_4}{1\text{mol Pb}(\text{NO}_3)_2} \right| \frac{303.2\text{g PbSO}_4}{1\text{mol PbSO}_4} = 189.5\text{g PbSO}_4 \quad \text{ER}$$

What mass of precipitate is produced when 35 mL of 0.250 M $\text{Fe}(\text{NO}_3)_3$ solution is mixed with 55 mL of 0.180 M KOH solution?



$$\text{Fe}^{3+} \rightarrow 0.25\text{M} = \frac{0.25\text{mol Fe}(\text{NO}_3)_3}{1\text{L SOLN}} \times \frac{35\text{mL}}{1000\text{mL/L}} = 0.00875\text{mol Fe}(\text{NO}_3)_3 \left| \frac{1\text{mol Fe}(\text{OH})_3}{1\text{mol Fe}(\text{NO}_3)_3} \right| \frac{106.9\text{g Fe}(\text{OH})_3}{1\text{mol Fe}(\text{OH})_3} = 0.935\text{g Fe}(\text{OH})_3 \quad \text{ER}$$

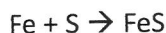
$$\text{OH}^- \rightarrow 0.180\text{M} = \frac{0.180\text{mol KOH}}{1\text{L SOLN}} \times \frac{55\text{mL}}{1000\text{mL/L}} = 0.0099\text{mol KOH} \left| \frac{1\text{mol Fe}(\text{OH})_3}{3\text{mol KOH}} \right| \frac{106.9\text{g Fe}(\text{OH})_3}{1\text{mol Fe}(\text{OH})_3} = 0.352\text{g Fe}(\text{OH})_3 \quad \text{LR}$$

END 4.2

PROG. 4.3

2) Oxidation – Reduction Reactions (Redox)

(OIL RIG!)



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:



Assign oxidations states:

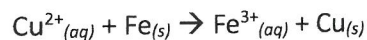


Balancing redox reactions

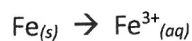
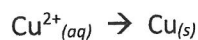
END 4.3

BEGIN 4.4

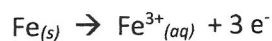
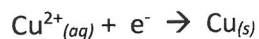
1. Neutral solutions



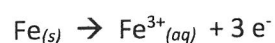
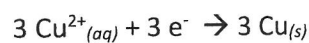
Step 1: Separate the half-reactions according to oxidation and reduction reactions.



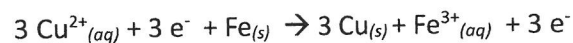
Step 2: Balance the charges by adding e^- to the other side.



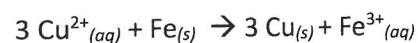
Step 3: Balance the electrons by multiplying the entire equation by a multiplier.



Step 4: Add the equations.



Step 5: Cancel out the electrons



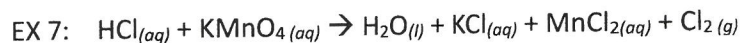
2. Acidic solutions

- 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 H₂O 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 H₂O 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.

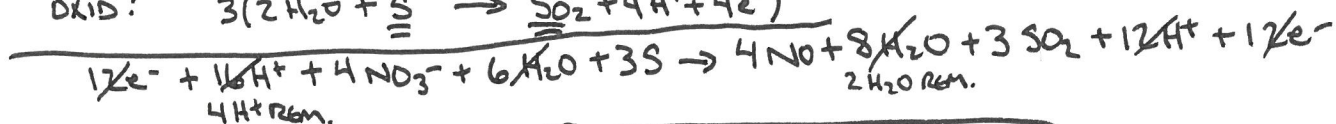
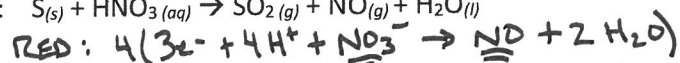
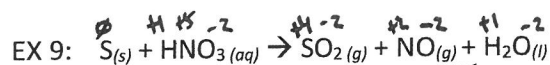
****Remember: If it is a basic solution, do all of the steps above, then add OH⁻ to eliminate the H⁺ ions****



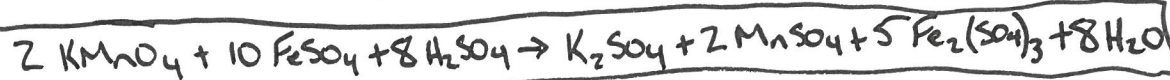
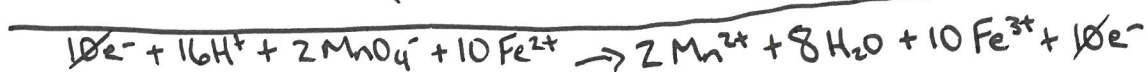
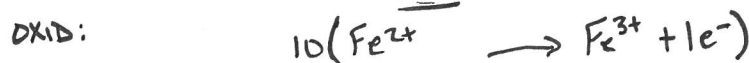
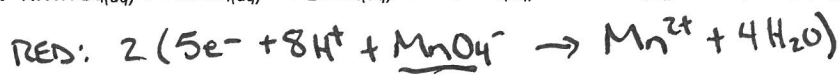
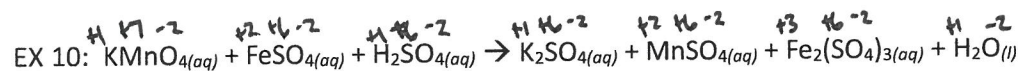
(Acid)



(Acid)

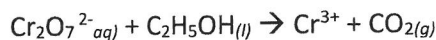


(Acid)



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

Net ionic equations:



****Reminder on balancing redox reactions in basic solutions****

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.



3) Acid-Base Reactions

Arrhenius definition:

Acid –

Base –

Brønsted-Lowry definition:

Acid –

Base –

Strong Acid-Base reaction:

Weak Acid-Base reaction:

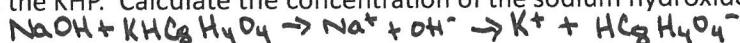
End Section
4.4
BEGIN 4.5

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 H₂SO₄.

EX 13: In an experiment, 28 mL of 0.250 M HNO₃ 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.

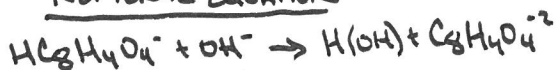
NaOH
41.20 mL
? M
KHC₈H₄O₄

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 (KHC₈H₄O₄, 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.



$$M = \frac{\text{mol Solute}}{\text{L Sol'n}}$$

NET IONIC EQUATION



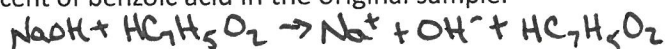
$$\frac{1.3009 \text{ g HC}_8\text{H}_4\text{O}_4^-}{204.22 \text{ g/KHP}} \times \frac{1 \text{ mol KHP}}{1 \text{ mol OH}^-} = 0.006371 \text{ mol OH}^-$$

$$M = \frac{0.006371 \text{ mol OH}^-}{0.04120 \text{ L}}$$

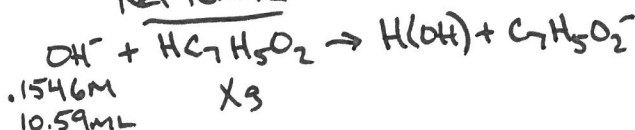
$$= 0.155 \text{ M NaOH}$$

NaOH
0.1546 M
10.59 mL
H

EX 15: An environmental chemist analyzed the effluent (the released waste material) from an industrial process known to produce the compound carbon tetrachloride (CCl₄) and benzoic acid (HC₇H₅O₂), 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.



NET IONIC



$$0.1546 \text{ M} = \frac{0.1546 \text{ mol NaOH}}{1 \text{ L Sol'n}} \times \frac{0.01059 \text{ L}}{1 \text{ mol OH}^-} \times \frac{1 \text{ mol HC}_7\text{H}_5\text{O}_2}{1 \text{ mol HC}_7\text{H}_5\text{O}_2} = 0.1999 \text{ g HC}_7\text{H}_5\text{O}_2$$

$$\text{MASS \%} = \frac{0.1999 \text{ g HC}_7\text{H}_5\text{O}_2}{0.3518 \text{ g}} \times 100\% = 56.82\%$$

END 4.5