Worksheet 6 – Displacement Reactions and Acid/Base reactions

Displacement reactions are those in which **ions recombine** in solution. If the products of this recombination are **insoluble** in water, a solid **precipitate** will form.

The **solubility rules** for common salts are summarized below:

1. Most nitrate salts are soluble.
2. Most salts of sodium, potassium and ammonium cations are soluble.
3. Most chloride salts are soluble.
    Exceptions: silver, lead (II), and mercury (I) chloride (Hg_2Cl_2)
4. Most sulfate salts are soluble.
Exceptions: calcium, barium and lead (II) sulfate
5. Most hydroxides are only slightly soluble
    Exceptions: sodium, potassium, calcium and barium hydroxides
6. Most sulfide, carbonate and phosphate salts are only slightly soluble.

Based on these rules, decide if each of the following compounds is soluble (s), or insoluble (i),and which rule this is based on:

compound	solubility	rule
silver chloride	i	3
manganese (II) hydroxide	i	5
calcium nitrate	s	1
sodium carbonate	s	2
nickel (II) sulfate	s	4
magnesium sulfide	i	6
calcium carbonate	i	6
potassium hydroxide	s	2
ammonium nitrate	s	1
potassium chloride	s	2
aluminum hydroxide	i	5

Each of the compounds in the grid below is water soluble. However, when they are mixed, insoluble compounds may be formed. First, write out the ions that form when the compounds dissolve in water. Then, in each cell of the grid, decide whether or not an insoluble precipitate will form. If a precipitate will form, write out its formula.

	K_3PO_4	$(NH_4)_2S$	$AgNO_3$	$CaCl_2$
Ag_2SO_4	Ag_3PO_4	Ag_2S	no ppt	$AgCl$, $CaSO_4$
$Pb(NO_3)_2$	$Pb_3(PO_4)_2$	PbS	no ppt	$PbCl_2$
KOH	no ppt	no ppt	$AgOH$	no ppt
$MgCl_2$	$Mg_3(PO_4)_2$	MgS	$AgCl$	no ppt

1.   100.0 mL of 1.0 M iron (III) nitrate is mixed with 100.0 mL of 1.0 M sodium hydroxide. A precipitate forms. Write the molecular equation, ionic equation and net ionic equation for this reaction.

   a.   Balanced molecular equation (leave ionic species as molecules)

   $Fe(NO_3)_3(aq) + 3NaOH(aq) \rightarrow Fe(OH)_3(s) + 3NaNO_3(aq)$
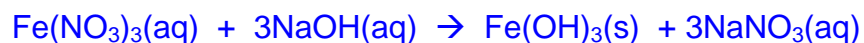
   b.   Ionic equation (allow soluble species to appear as ions)

   $Fe^{3+}(aq) + 3NO_3^-(aq) + 3Na^+(aq) + 3OH^-(aq) \rightarrow Fe(OH)_3(s) + 3Na^+(aq) + 3NO_3^-(aq)$

   c.   Net ionic equation (eliminate "spectator" ions)

   $Fe^{3+}(aq) + 3OH^-(aq) \rightarrow Fe(OH)_3(s)$

d. Determine the mass of the precipitate. Use the **net ionic** equation to make these calculations.

i. How many moles of the cation are present?

$$100.0 \text{ mL Fe(NO}_3)_3 \text{ soln} \times \frac{1\text{L}}{1000\text{mL}} \times \frac{1.0 \text{ mol Fe(NO}_3)_3}{1\text{L soln}} \times \frac{1\text{mol Fe}^{3+}}{1\text{mol Fe(NO}_3)_3} = 0.10 \text{ mol Fe}^{3+}$$

ii. How many moles of the anion are present?

$$100.0 \text{ mL NaOH soln} \times \frac{1\text{L}}{1000\text{mL}} \times \frac{1.0 \text{ mol NaOH}}{1\text{L soln}} \times \frac{1\text{mol OH}^-}{1\text{mol NaOH}} = 0.10 \text{ mol OH}^-$$

iii. Based on stoichiometry, is there a limiting reagent?

Yes – we need 3 moles of OH⁻ for every 1 mole of Fe³⁺, so the anion is limiting

iv. How many moles of precipitate will form?

$$0.10 \text{ mol OH}^- \times \frac{1\text{mol Fe (OH)}_3}{3 \text{ mol OH}^-} = 0.033 \text{ mol Fe (OH)}_3$$

v. How many grams of precipitate is this?

$$0.033 \text{ mol Fe (OH)}_3 \times \frac{106.9 \text{ Fe (OH)}_3}{\text{mol Fe (OH)}_3} = 3.5 \text{ g Fe (OH)}_3$$

vi. What is the concentration of the excess ion?

$$0.10 \text{ mol OH}^- \times \frac{1\text{mol Fe}^{3+}}{3 \text{ mol OH}^-} = 0.033 \text{ mol Fe}^{3+} \text{ used}$$

$$0.10 \text{ mol Fe}^{3+} - 0.033 \text{ mol Fe}^{3+} = 0.067 \text{ mol Fe}^{3+} \text{ remaining}$$

$$[\text{Fe}^{3+}] = \frac{0.067 \text{ mol Fe}^{3+}}{0.200 \text{ L soln}} = 0.33 \text{ M}$$

Acid/Base reactions

Strong acids (HA) dissociate completely in water to produce H^+ and A^-. They are called "proton donors" since they give off a proton, H^+. Strong bases (BOH) dissociate completely in water to form B^+ and OH^-. They are called "proton acceptors" since the OH^- will react with H^+. In a neutralization reaction, the H^+ and OH^- combine to form water (H_2O), leaving the components of salt, A^- and B^+ in solution.

1. What volume of 0.200 M HCl is needed to neutralize 20.0 mL of 0.350 M NaOH? HCl is a strong acid and NaOH is a strong base.

    a. How many moles of OH^- are present?

$$0.020 \text{ L NaOH soln} \times \frac{0.350 \text{ mol NaOH}}{1 \text{ L soln}} \times \frac{1 \text{ mol OH}^-}{1 \text{ mol NaOH}} = 0.00700 \text{ mol OH}^-$$

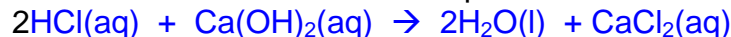   b. How many moles of H^+ are needed to neutralize the OH^-?
    mol H^+ = mol OH^- in a neutralization reaction = 0.00700 mol H^+

    c. What volume of 0.200 M HCl is needed?

$$0.00700 \text{ mol H}^+ \times \frac{1 \text{ mol HCl}}{1 \text{ mol H}^+} \times \frac{1 \text{ L soln}}{0.200 \text{ mol HCl}} = 0.0350 \text{ L soln} = 35.0 \text{ mL HCl}$$

2. When titrating 0.150 M HCl with a calcium hydroxide solution of unknown concentration, 35.0 mL of acid are required to neutralize 25.0 mL of the base. Calculate the molarity of the base.
    a. Write the balanced molecular equation.
    $2HCl(aq) + Ca(OH)_2(aq) \rightarrow 2H_2O(l) + CaCl_2(aq)$

    b. Write the balanced net ionic equation.
    $H^+(aq) + OH^-(aq) \rightarrow H_2O(l)$

    c. How many moles of H^+ are present?

$$0.035 \text{ L HCl soln} \times \frac{0.150 \text{ mol HCl}}{1 \text{ L soln}} \times \frac{1 \text{ mol H}^+}{1 \text{ mol HCl}} = 0.00525 \text{ mol H}^+$$

    d. How many moles of OH^- are required to neutralize the acid?
    mol OH^- = mol H^+ in a neutralization reaction = 0.00525 mol OH^-

    e. What is the molarity of the base?

$$0.00525 \text{ mol OH}^- \times \frac{1 \text{ mol Ca (OH)}_2}{2 \text{ mol OH}^-} = 0.00263 \text{ mol Ca (OH)}_2$$

$$[Ca(OH)_2] = \frac{0.00263 \text{ mol OH}^-}{0.025 \text{ L soln}} = 0.105 M$$

3.   A chemist dissolves 0.300 g of an unknown monoprotic (one acidic H) acid in water. She finds that 14.60 mL of 0.426 M NaOH are required to neutralize the acid. What is the molar mass of the acid?

a.   Units of molar mass are g/mol. How many grams of acid are present?
        0.300 g acid

b.   How many moles of base are required to reach the equivalence point?

$$0.01460 \text{ L NaOH soln} \times \frac{0.426 \text{ mol NaOH}}{1 \text{L soln}} \times \frac{1 \text{mol OH}^-}{1 \text{mol NaOH}} = 0.00622 \text{ mol OH}^-$$

c.   How many moles of acid are present?
        Monoprotic acid means mol acid = mol H$^+$ = mol OH$^-$
        0.00622 mol OH$^-$ → 0.00622 mol H$^+$ = 0.00622 mol HA

d.   What is the molar mass of the acid?

$$\frac{0.300 \text{ g HA}}{0.00622 \text{ mol HA}} = 48.2 \text{ g/mol}$$