CHEM 101A SOLUTIONS TO TOPIC B HOMEWORK

1) Based on the solubility rules, $\mathbf{K_2SO_4}$, $\mathbf{Fe_2(SO_4)_3}$, and $\mathbf{(NH_4)_2CO_3}$ are soluble in water. $\mathbf{H_2SO_4}$ is also soluble in water; the solubility rules do not cover acids, but you should know that the six common strong acids (HCl, HBr, HI, HNO_3, $HClO_4$, and H_2SO_4) and the weak acid $HC_2H_3O_2$ (acetic acid) are water-soluble.

2) Note that this problem asks you to identify ALL the species present in solution. Major species are those species that are present in high concentration, whereas minor species are present only in low concentration.

a) HNO_3 is a strong acid, so it dissociates completely in water. The solute species in this solution are **H^+ and NO_3^-**, and **both are present in high concentration**.

b) HNO_2 is a weak acid (all acids should be assumed to be weak unless they are one of the six strong acids listed in the handout). In a solution of a weak acid, only a small fraction of the molecules dissociate, but most of the acid molecules remain unchanged. Therefore, the solution will contain the following species:

HNO_2 (high concentration) **H^+ (low concentration)** **NO_2^- (low concentration)**

c) CH_3OH is a nonelectrolyte, so it does not dissociate at all. The only solute species in this solution is **CH_3OH**, which is (obviously) present **in high concentration**.

3) Sodium sulfate, like virtually all soluble ionic compounds, is a strong electrolyte and ionizes completely when it is dissolved in water:

$$Na_2SO_4(s) \rightarrow 2\ Na^+(aq) + SO_4^{2-}(aq)$$

Therefore, the solution will contain two species, **Na^+ and SO_4^{2-}**. Each mole of sodium sulfate forms <u>two</u> moles of Na^+, so the concentration of Na^+ in the solution is 0.25 M x 2 = **0.50 M**. Each mole of sodium sulfate forms <u>one</u> mole of SO_4^{2-}, so the concentration of SO_4^{2-} in the solution is 0.25 M x 1 = **0.25 M**.

4) a) Molarity is moles of solute divided by liters of solution. To calculate molarity, we must first determine the number of moles of solute (Na_2CO_3) and the number of liters of solution.

$$5.317\text{ g Na}_2\text{CO}_3 \times \frac{1\text{ mol Na}_2\text{CO}_3}{105.99\text{ g Na}_2\text{CO}_3} = 0.0\overline{5016}51\text{ mol Na}_2\text{CO}_3$$

$$170.0\text{ mL} \times \frac{1\text{ L}}{1000\text{ mL}} = 0.\overline{1700}\text{ L}$$

$$\frac{0.0\overline{5016}51\text{ mol Na}_2\text{CO}_3}{0.\overline{1700}\text{ L}} = \mathbf{0.2951\ M} \quad \text{(rounded from 0.295089 M)}$$

(Reminder: we use a bar to show which digits in an intermediate result are significant.)

b) We already know that the solution contains 0.0501651 moles of Na_2CO_3. What must be the total volume of the solution in order to get a concentration of 0.125 M (i.e. 0.125 moles of Na_2CO_3 per liter of solution)? We can calculate this volume by using the molarity as a conversion factor:

$$0.0501651 \text{ mol Na}_2\text{CO}_3 \times \frac{1 \text{ L solution}}{0.125 \text{ mol Na}_2\text{CO}_3} = 0.\overline{401}321 \text{ L solution} = \overline{401}.3 \text{ mL of solution}$$

We must end up with a total of 401 mL of solution (valid only to the nearest whole number). We started with 170.0 mL of solution, so we must add 401 mL – 170.0 mL = **231 mL of water**.

c) After we add Na_2SO_4, the solution will contain two sources of Na^+ ions, the Na_2SO_4 we just added and the Na_2CO_3 that was originally present. We need to start by figuring out how many moles of Na^+ are supplied by the Na_2CO_3.

$$0.0\overline{501}651 \text{ mol Na}_2\text{CO}_3 \times \frac{2 \text{ mol Na}^+}{1 \text{ mol Na}_2\text{CO}_3} = 0.\overline{1003}302 \text{ mol Na}^+$$

Next, we calculate the total number of moles of Na^+ that must be present if the solution is to contain 0.400 mol/L of this ion. Recall that the total volume of the solution in part b was 0.401321 L (ignoring rounding):

$$0.\overline{401}321 \text{ L} \times \frac{0.400 \text{ mol Na}^+}{1 \text{ L}} = 0.\overline{160}528 \text{ mol Na}^+$$

Now we calculate the number of moles of Na^+ that we must add:

$$0.\overline{160}528 \text{ mol Na}^+ - 0.\overline{1003}302 \text{ mol Na}^+ = 0.0\overline{60}198 \text{ mol Na}^+$$

Note that since the first number is known to only three decimal places, the answer is also only valid to three decimal places (i.e. if this were our final answer, we would round it to 0.060 mol). Recall that when you add or subtract, you keep the smallest number of <u>*decimal places*</u>*, not significant figures.*

The last step is to calculate the number of grams of Na_2SO_4 we need to add in order to get 0.060198 moles of Na^+.

$$0.0\overline{60}198 \text{ mol Na}^+ \times \frac{1 \text{ mol Na}_2\text{SO}_4}{2 \text{ mol Na}^+} \times \frac{142.04 \text{ g Na}_2\text{SO}_4}{1 \text{ mol Na}_2\text{SO}_4} = \mathbf{4.3 \text{ g Na}_2\text{SO}_4}$$

In this calculation, we are multiplying and dividing, so we go back to counting significant figures. The number of moles of Na^+ is known to just two sig figs, so our answer is only valied to two sig figs.

5) a) We need to make 250.0 mL (0.2500 L) of a solution that contains 0.210 moles of K_3PO_4 per liter of solution:

$$0.2500 \text{ L} \times \frac{0.210 \text{ mol K}_3\text{PO}_4}{1 \text{ L}} \times \frac{212.27 \text{ g K}_3\text{PO}_4}{1 \text{ mol K}_3\text{PO}_4} = \mathbf{11.1 \text{ g K}_3\text{PO}_4}$$

b) Each mole of K_3PO_4 dissociates into three moles of K^+ and one mole of PO_4^{3-}. Therefore, the concentration of K^+ is 0.210 M x 3 = **0.630 M** and the concentration of PO_4^{3-} is 0.210 M x 1 = **0.210 M**.

c) If we add 500.0 mL of water, the total volume of the solution will now be 250.0 mL + 500.0 mL = 750.0 mL. The number of moles of K^+ in the solution can be calculated from the original number of moles of K_3PO_4 (since adding water doesn't change the amount of solute):

$$0.2500\ \text{L} \times \frac{0.210\ \text{mol}\ K_3PO_4}{1\ \text{L}} \times \frac{3\ \text{mol}\ K^+}{1\ \text{mol}\ K_3PO_4} = 0.\overline{157}5\ \text{mol}\ K^+$$

Therefore, the new molarity of K^+ is:

$$\frac{0.\overline{157}5\ \text{mol}\ K^+}{0.\overline{7500}\ \text{L}} = \mathbf{0.210\ M}$$

6) For these net ionic equation problems, the solutions lists the major species present in solution, the reactivity (which substances will combine and why they will combine), and the net ionic equation for the reaction.

Note that the provided molarities are a way to clearly indicate that the reactants are aqueous solutions. The actually molarity values are relevant only when a reaction differs depending on the mole ratio of reactants present (such as for polyprotic acids.) For most cases, the molarity values are "extra information."

a) Species present: $Na^+(aq)$ $CO_3^{2-}(aq)$ $Ni^{2+}(aq)$ $Cl^-(aq)$
Ni^{2+} will combine with CO_3^{2-}, because $NiCO_3$ is insoluble in water.
Net ionic equation: $Ni^{2+}(aq) + CO_3^{2-}(aq) \rightarrow NiCO_3(s)$

b) Species present: $Ag^+(aq)$ $NO_3^-(aq)$ $K^+(aq)$ $PO_4^{3-}(aq)$
Ag^+ will combine with PO_4^{3-}, because Ag_3PO_4 is insoluble in water.
Net ionic equation: $3\ Ag^+(aq) + PO_4^{3-}(aq) \rightarrow Ag_3PO_4(s)$

c) Species present: $Mg^{2+}(aq)$ $Cl^-(aq)$ $NH_4^+(aq)$ $SO_4^{2-}(aq)$
None of these ions will combine, because all possible combinations are soluble in water and there is no acid present (so no acid/base reaction is possible, no weak acid formation is possible.)
No reaction.

d) Species present: $Na^+(aq)$ $PO_4^{3-}(aq)$ $K^+(aq)$ $S^{2-}(aq)$
None of these ions will combine, because all possible combinations are soluble in water and there is no acid present (so no acid/base reaction is possible, no weak acid formation is possible.)
. No reaction.

e) Species present: $Na^+(aq)$ $OH^-(aq)$ $H^+(aq)$ $Cl^-(aq)$
H^+ will combine with OH^- to form water, because water is extremely stable.
Net ionic equation: $H^+(aq) + OH^-(aq) \rightarrow H_2O(l)$

f) Species present: $HC_2H_3O_2(aq)$ $K^+(aq)$ $OH^-(aq)$

The H^+ ion will be transferred from acetic acid to OH^-, because H^+ bonds more tightly to OH^- than it does to other negative ions (that's why water is so stable!).

Net ionic equation: $HC_2H_3O_2(aq) + OH^-(aq) \rightarrow C_2H_3O_2^-(aq) + H_2O(l)$

Note that $HC_2H_3O_2$ is a weak acid, so most of the $HC_2H_3O_2$ molecules are not dissociated in aqueous solution. There is very little free H^+ present, so writing $H^+ + OH^- \rightarrow H_2O$ does not correctly describe the reaction! Weak acids must be written in their molecular form when writing net ionic equations.

g) Species present: $Ba^{2+}(aq)$ $OH^-(aq)$ $H^+(aq)$ $NO_3^-(aq)$

H^+ will combine with OH^- to form water, because water is extremely stable.

Net ionic equation: $H^+(aq) + OH^-(aq) \rightarrow H_2O(l)$

h) Species present: $H_3PO_4(aq)$ $Na^+(aq)$ $OH^-(aq)$

The H^+ ion will be transferred from phosphoric acid to OH^-, because H^+ bonds more tightly to OH^- than it does to other negative ions.

Net ionic equation: $H_3PO_4(aq) + OH^-(aq) \rightarrow H_2PO_4^-(aq) + H_2O(l)$

Note that H_3PO_4 is a polyprotic acid, so this reaction is complicated by the fact that once $H_2PO_4^-$ forms it could also react with OH^-. However, this second reaction will occur only once ALL the H_3PO_4 has first reacted and then only if there is excess OH^- available. In this case, we are mixing equal numbers of moles of H_3PO_4 and OH^-, so OH^- is not present in excess, so only the H_3PO_4 will react.

i) The reactants here are the same as they were in part h, but in this case OH^- is present in excess. Therefore, all three of the hydrogens will react and we will observe three reactions, one after the other.

Reaction 1: $H_3PO_4(aq) + OH^-(aq) \rightarrow H_2PO_4^-(aq) + H_2O(l)$

Reaction 2: $H_2PO_4^-(aq) + OH^-(aq) \rightarrow HPO_4^{2-}(aq) + H_2O(l)$

Reaction 3: $HPO_4^{2-}(aq) + OH^-(aq) \rightarrow PO_4^{3-}(aq) + H_2O(l)$

j) Species present: $Cu(OH)_2(s)$ $H^+(aq)$ $Cl^-(aq)$

The OH^- ions will be transferred from solid $Cu(OH)_2$ to H^+ and water will form. Water is a particularly stable compound; OH^- bonds more tightly to H^+ than it does to other positive ions (*"making water is a good thing"*).

Net ionic equation: $Cu(OH)_2(s) + 2\,H^+(aq) \rightarrow Cu^{2+}(aq) + 2\,H_2O(l)$

Note that initially, OH^- is tightly bonded to Cu^{2+}, so there is essentially no free OH^- in solution. Therefore, writing $OH^- + H^+ \rightarrow H_2O$ is not correct.

k) Species present: $Fe(OH)_3(s)$ $HC_2H_3O_2(aq)$

H^+ ions from acetic acid (a weak acid) will combine with OH^- ions from the solid iron (III) hydroxide to form water.

Net ionic equation:

$Fe(OH)_3(s) + 3\,HC_2H_3O_2(aq) \rightarrow Fe^{3+}(aq) + 3\,C_2H_3O_2^-(aq) + 3\,H_2O(l)$

7) For these net ionic equation problems, the solutions lists the major species present in solution, the reactivity (which substances will combine and why they will combine), and the net ionic equation for the reaction.

a) Species present: $Cr^{3+}(aq)$ $NO_3^-(aq)$ $NH_3(aq)$

This is a special reaction. Any metal ion that can form an insoluble hydroxide will react with aqueous ammonia. See the handout for more details on this type of reaction.

Net ionic equation: $Cr^{3+}(aq) + 3\ NH_3(aq) + 3\ H_2O(l) \rightarrow Cr(OH)_3(s) + 3\ NH_4^+(aq)$

b) Species present: $H^+(aq)$ $Cl^-(aq)$ $NH_3(aq)$

NH_3 is a weak base and reacts with H^+ to form NH_4^+ ion, which is quite stable.

Net ionic equation: $NH_3(aq) + H^+(aq) \rightarrow NH_4^+(aq)$

c) Species present: $HClO(aq)$ $NH_3(aq)$

The H^+ ion will be transferred from HClO to NH_3 (since NH_3 is a weak base.)

Net ionic equation: $HClO(aq) + NH_3(aq) \rightarrow ClO^-(aq) + NH_4^+(aq)$

d) Species present: $Na^+(aq)$ $HCO_3^-(aq)$ $H^+(aq)$ $HSO_4^-(aq)$

This is another special reaction. H^+ reacts with any source of HCO_3^- (or with any source of CO_3^{2-} if you use excess acid) to form CO_2 and water. See the handout for more details on this type of reaction.

Net ionic equation: $HCO_3^-(aq) + H^+(aq) \rightarrow CO_2(g) + H_2O(l)$

e) Species present: $Na^+(aq)$ $CO_3^{2-}(aq)$ $H^+(aq)$ $Cl^-(aq)$

H^+ combines with CO_3^{2-} to form HCO_3^-, which is a stable ion. Note that the problem uses equal numbers of moles of CO_3^{2-} and H^+, so there is no excess H^+ to react with the HCO_3^- ions we formed.

Net ionic equation: $CO_3^{2-}(aq) + H^+(aq) \rightarrow HCO_3^-(aq)$

f) The reactants here are the same as they were in part p, but in this case H^+ is present in excess. Therefore, we will observe two reactions, one after the other.

Reaction 1: $CO_3^{2-}(aq) + H^+(aq) \rightarrow HCO_3^-(aq)$

Reaction 2: $HCO_3^-(aq) + H^+(aq) \rightarrow CO_2(g) + H_2O(l)$

g) Species present $CaCO_3(s)$ $H^+(aq)$ $NO_3^-(aq)$

H^+ reacts with the CO_3^{2-} ion in $CaCO_3$. Since H^+ is present in excess, we will observe two reactions, one after the other.

Reaction 1: $CaCO_3(s) + H^+(aq) \rightarrow Ca^{2+}(aq) + HCO_3^-(aq)$

Reaction 2: $HCO_3^-(aq) + H^+(aq) \rightarrow CO_2(g) + H_2O(l)$

You can also write this as a single equation:

$$CaCO_3(s) + 2\ H^+(aq) \rightarrow Ca^{2+}(aq) + CO_2(g) + H_2O(l)$$

8) a) The two possible combinations we might form are $KC_2H_3O_2$ and Ag_2CrO_4. The solubility rules do not cover compounds containing CrO_4^{2-}, so we cannot use the rules to tell whether Ag_2CrO_4 is soluble. However, we know that $KC_2H_3O_2$ is water-soluble, because all compounds that contain K^+ dissolve in water. Therefore, the product cannot be $KC_2H_3O_2$, so it must be $\mathbf{Ag_2CrO_4}$.

b) $\mathbf{2\ Ag^+(aq) + CrO_4^{2-}(aq) \rightarrow Ag_2CrO_4(s)}$

9) a) Since the precipitate must contain a positive and a negative ion, the two possibilities are $NaNO_3$ and $PbBr_2$. However, $NaNO_3$ is water-soluble, so the precipitate must be $\mathbf{PbBr_2}$.

b) $\mathbf{Pb^{2+}(aq) + 2\ Br^-(aq) \rightarrow PbBr_2(s)}$

c) We start by calculating the number of moles of $PbBr_2$ that we formed.

$$3.006 \text{ g PbBr}_2 \times \frac{1 \text{ mol}}{367.0 \text{ g PbBr}_2} = 0.00\overline{8190}74 \text{ mol PbBr}_2$$

Each mole of $PbBr_2$ contains 2 moles of Br^- ions, which came from the original NaBr solution. Therefore, we can calculate the number of moles of Br^- in the original solution:

$$0.00\overline{8190}74 \text{ mol PbBr}_2 \times \frac{2 \text{ mol Br}^-}{1 \text{ mol PbBr}_2} = 0.0\overline{1638}14 \text{ mol Br}^-$$

The volume of the original NaBr solution was 250.0 mL (= 0.2500 L), so the concentration of Br^- in the original solution must have been:

$$\frac{0.0\overline{1638}14 \text{ mol Br}^-}{0.2500 \text{ L}} = \mathbf{0.06553\ M}$$

10) a) When a solid reacts with an acid to produce $CO_2(g)$, the solid must have contained either CO_3^{2-} ion or HCO_3^- ion, so the compound must be either $PbCO_3$ or $Pb(HCO_3)_2$. But which one is it? The key here is that we are told that the solid dissolves completely in 15 mL of 2.0 M HNO_3. If the solid is $PbCO_3$, the reaction is:

$$PbCO_3(s) + 2\ H^+(aq) \rightarrow Pb^{2+}(aq) + CO_2(g) + H_2O(l)$$

If the solid is $Pb(HCO_3)_2$, the reaction is:

$$Pb(HCO_3)_2(s) + 2\ H^+(aq) \rightarrow Pb^{2+}(aq) + 2\ CO_2(g) + 2\ H_2O(l)$$

Let's calculate how many moles of H^+ we have.

$$0.015 \text{ L} \times \frac{2.0 \text{ mol HNO}_3}{1 \text{ L}} \times \frac{1 \text{ mol H}^+}{1 \text{ mol HNO}_3} = 0.030 \text{ mol H}^+$$

Now let's see how many moles of CO_2 we can make if all of the H^+ is consumed (i.e. if H^+ is the limiting reactant). If the first reaction is correct, the mole ratio between H^+ and CO_2 is 2:1, so we would make…

$$0.030 \text{ mol H}^+ \times \frac{1 \text{ mol CO}_2}{2 \text{ mol H}^+} = 0.015 \text{ mol CO}_2$$

This is the maximum amount of CO_2 we could possibly make (if the first equation is correct), because once we have made 0.015 moles of CO_2, we run out of H^+ ions. However, according to the problem, we actually make 0.020 mol CO_2, which is more than the maximum amount. Therefore, the first equation cannot be right!

What about the second equation? Let's see how many moles of CO_2 we could make if the second reaction is correct. In the second reaction, the mole ration between H^+ and CO_2 is 2:2 (i.e. 1:1), so we can make…

$$0.030 \text{ mol H}^+ \times \frac{2 \text{ mol CO}_2}{2 \text{ mol H}^+} = 0.030 \text{ mol CO}_2$$

Again, this is the maximum amount of CO_2 we could make. In the case, the reaction produces less than this maximum amount, which is reasonable if the lead compound is the limiting reactant and H^+ is present in excess. Therefore, we can conclude that the second reaction is the right one, and **the solid must be $Pb(HCO_3)_2$.**

b) As we showed in part a, the chemical equation is

$$\mathbf{Pb(HCO_3)_2(s) + 2\,H^+(aq) \rightarrow Pb^{2+}(aq) + 2\,CO_2(g) + 2\,H_2O(l)}$$

c) In part a, we established that $Pb(HCO_3)_2$ is the limiting reactant. Therefore, we can use the number of moles of CO_2 to determine the number of moles of $Pb(HCO_3)_2$ that we started with, and then convert moles into grams.

$$0.020 \text{ mol CO}_2 \times \frac{1 \text{ mol Pb(HCO}_3)_2}{2 \text{ mol CO}_2} \times \frac{329.24 \text{ g Pb(HCO}_3)_2}{1 \text{ mol Pb(HCO}_3)_2} = \mathbf{3.3\ g\ Pb(HCO_3)_2}$$

11) In both cases, the H^+ ions from the sulfuric acid react with the CO_3^{2-} ions from the metal carbonate, forming CO_2 (the bubbles) and H_2O. However, when you use $BaCO_3$, the Ba^{2+} ions also combine with SO_4^{2-} ions, forming insoluble $BaSO_4$. In effect, you are exchanging one insoluble salt ($BaCO_3$) for another ($BaSO_4$), so you never see the solid go into solution. In the case of Mg^{2+}, though, $MgSO_4$ is water-soluble and does not form, so the product mixture does not contain any solids. The reactions are:

$$MgCO_3(s) + H^+(aq) + HSO_4^{2-}(aq) \rightarrow Mg^{2+}(aq) + SO_4^{2-}(aq) + CO_2(g) + H_2O(l)$$

$$BaCO_3(s) + H^+(aq) + HSO_4^{2-}(aq) \rightarrow BaSO_4(s) + CO_2(g) + H_2O(l)$$

12) The balanced equation for the reaction is $Mg^{2+}(aq) + 2\,OH^-(aq) \rightarrow Mg(OH)_2(s)$. The NO_3^- and Na^+ ions are spectators. Here is the material balance table:

Species	**Initial mmol**	**Final mmol**	**Final molarity or mass**
$Mg^{2+}(aq)$	4.607	0.7955	0.013 M
$NO_3^-(aq)$	9.214	9.214	0.145 M
$Na^+(aq)$	7.623	7.623	0.120 M
$OH^-(aq)$	7.623	0	0 M
$Mg(OH)_2(s)$	0	3.8115	0.223 g

Here are the details on how these numbers were determined. Note that molarity can be expressed as millimoles per milliliter (so 0.170 M = 0.170 mmol of solute per mL of solution).

Determining initial millimoles of Mg^{2+} and NO_3^-:

$$27.1\text{ mL} \times \frac{0.170\text{ mmol Mg(NO}_3)_2}{1\text{ mL}} = \overline{4.60}7\text{ mmol Mg(NO}_3)_2$$

$$\overline{4.60}7\text{ mmol Mg(NO}_3)_2 \times \frac{1\text{ mmol Mg}^{2+}}{1\text{ mmol Mg(NO}_3)_2} = \overline{4.60}7\text{ mmol Mg}^{2+}$$

$$\overline{4.60}7\text{ mmol Mg(NO}_3)_2 \times \frac{2\text{ mmol NO}_3^-}{1\text{ mmol Mg(NO}_3)_2} = \overline{9.21}4\text{ mmol NO}_3^-$$

Determining initial millimoles of Na^+ and OH^-:

$$36.3\text{ mL} \times \frac{0.210\text{ mmol NaOH}}{1\text{ mL}} = \overline{7.62}3\text{ mmol NaOH}$$

$$\overline{7.62}3\text{ mmol NaOH} \times \frac{1\text{ mmol Na}^+}{1\text{ mmol NaOH}} = \overline{7.62}3\text{ mmol Na}^+$$

$$\overline{7.62}3\text{ mmol NaOH} \times \frac{1\text{ mmol OH}^-}{1\text{ mmol NaOH}} = \overline{7.62}3\text{ mmol OH}^-$$

The limiting reactant here is OH^-. One way to work this out is to calculate the number of millimoles of $Mg(OH)_2$ that would be produced if all of the Mg^{2+} reacts and if all of the OH^- reacts; you find that you'll make 4.607 mmol $Mg(OH)_2$ if all of the Mg^{2+} reacts, but only 3.8115 mmol $Mg(OH)_2$ if all of the OH^- reacts. Since OH^- produces less product, it is the limiting reactant.

The limiting reactant calculation also gives us the final number of millimoles of $Mg(OH)_2$.

$$\overline{7.62}3\text{ mmol OH}^- \times \frac{1\text{ mmol Mg(OH)}_2}{2\text{ mmol OH}^-} = \overline{3.81}15\text{ mmol Mg(OH)}_2$$

The number of millimoles of Mg^{2+} that is consumed during the reaction is:

$$\overline{7.62}3\text{ mmol OH}^- \times \frac{1\text{ mmol Mg}^{2+}}{2\text{ mmol OH}^-} = \overline{3.81}15\text{ mmol Mg}^{2+}$$

So the number of millimoles of Mg^{2+} that remains after the reaction is:

$$\overline{4.60}7\text{ mmol} - \overline{3.81}15\text{ mmol} = \overline{0.79}55\text{ mmol Mg}^{2+}$$

To calculate the final molarities of the ions, we divide the number of millimoles of each ion by the total volume of solution (27.1 mL + 36.3 mL = 63.4 mL). For instance, for Mg^{2+} we have:

$$\frac{\overline{0.795}5\text{ mmol Mg}^{2+}}{63.4\text{ mL}} = 0.013\text{ M}$$

The product $Mg(OH)_2$ is a solid, so it has no molarity. Instead, we calculate its mass. The final mass of $Mg(OH)_2$ is:

$$3.8115\text{ mmol Mg(OH)}_2 \times \frac{1\text{ mol}}{1000\text{ mmol}} \times \frac{58.32\text{ g}}{1\text{ mol}} = 0.222\text{ g Mg(OH)}_2$$

13) The balanced equation for the reaction is $Zn(OH)_2(s) + 2\ H^+(aq) \rightarrow Zn^{2+}(aq) + 2\ H_2O(l)$. The Br^- ions are spectators. Here is the material balance table:

Species	**Initial mmol**	**Final mmol**	**Final molarity or mass**
$Zn(OH)_2(s)$	1.50896	0	0
$H^+(aq)$	4.599	1.58107	0.0866 M
$Br^-(aq)$	4.599	4.599	0.252 M
$Zn^{2+}(aq)$	0	1.50896	0.0827 M

You should be able to calculate the initial millimoles of the reactants on your own (note that the molar mass of $Zn(OH)_2$ is 99.406 g/mol), and you should also be able to work out that the limiting reactant is $Zn(OH)_2$. To calculate the final millimoles of H^+, you must first calculate the number of millimoles of H^+ that is consumed in the reaction:

$$\overline{1.50}896\text{ mmol Zn(OH)}_2 \times \frac{2\text{ mmol H}^+}{1\text{ mmol Zn(OH)}_2} = \overline{3.01}793\text{ mmol H}^+$$

The final number of mmol of H^+ is then

$$\overline{4.59}9\text{ mmol} - \overline{3.01}793\text{ mmol} = \overline{1.58}107\text{ mmol H}^+$$

To calculate the final molarities, divide the final mmol by 18.25 mL. The addition of the solid $Zn(OH)_2$ does not affect the volume of liquid.

13) We can use the first titration to determine the molarity of the NaOH solution. We start by calculating the number of moles of $HC_8H_4O_4^-$ that was used:

$$0.8280\text{ g KHC}_8\text{H}_4\text{O}_4 \times \frac{1\text{ mol KHC}_8\text{H}_4\text{O}_4}{204.22\text{ g KHC}_8\text{H}_4\text{O}_4} \times \frac{1\text{ mol HC}_8\text{H}_4\text{O}_4^-}{1\text{ mol KHC}_8\text{H}_4\text{O}_4} = 0.00\overline{4054}45\text{ mol HC}_8\text{H}_4\text{O}_4^-$$

The reaction between $HC_8H_4O_4^-$ and OH^- is 1:1, so the chemist must have added 0.00405445 moles of OH^- at the endpoint. It takes one mole of NaOH to supply one mole of OH^-, so the NaOH solution contained 0.00405445 moles of NaOH (before dissociation). The chemist used 30.96 mL (0.03096 L) of the NaOH solution, so the molarity of the NaOH solution was:

$$\frac{0.00\overline{4054}45\ \text{mol NaOH}}{0.03096\ \text{L}} = \mathbf{0.1310\ M} \qquad \text{(rounded from 0.130958 M)}$$

We can now use the NaOH molarity to calculate the molarity of the $H_2C_4H_4O_4$ solution. We start by calculating the number of moles of OH^- used in the second titration. Recall that this titration uses the same NaOH solution as the first one, so the NaOH molarity is still 0.1310 M. (We’ll use an unrounded value in the calculation.)

$$0.02523\ \text{L} \times \frac{0.\overline{1309}58\ \text{mol NaOH}}{1\ \text{L}} \times \frac{1\ \text{mol OH}^-}{1\ \text{mol NaOH}} = 0.00\overline{3304}06\ \text{mol OH}^-$$

From this value and the mole ratio in the balanced equation (2 moles of OH^- per 1 mole of $H_2C_4H_4O_4$), we can calculate the number of moles of $H_2C_4H_4O_4$.

$$0.00\overline{3304}06\ \text{mol OH}^- \times \frac{1\ \text{mol H}_2\text{C}_4\text{H}_4\text{O}_4}{2\ \text{mol OH}^-} = 0.00\overline{1652}03\ \text{mol H}_2\text{C}_4\text{H}_4\text{O}_4$$

The volume of the $H_2C_4H_4O_4$ solution was 15.45 mL (0.01545 L), so the molarity of the $H_2C_4H_4O_4$ solution was:

$$\frac{0.00\overline{1652}03\ \text{mol H}_2\text{C}_4\text{H}_4\text{O}_4}{0.01545\ \text{L}} = \mathbf{0.1069\ M}$$