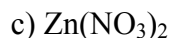
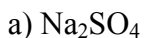


CHEM 101A

SOLUTIONS TO TOPIC A HOMEWORK

WARNING! You should be able to solve all (or at least most) assigned problems on your own. If you are struggling with a homework problem, **do not** immediately look at the solutions key. Instead, take a short break, or talk with a classmate, or “sleep on it”. Then come back to the problem later. Problem solving is a skill; it gets better with practice. Struggling with a problem may not be fun, but it builds problem solving skills!

1) Note that this should be review!



2) Note that this should be review!

a) calcium fluoride

b) potassium phosphate

c) copper(II) carbonate

d) sulfuric acid (“hydrogen sulfate” is not used)

e) silver sulfide

3) Note that this should be review! Note that the molar mass of $(\text{NH}_4)_2\text{Cr}_2\text{O}_7$ is 252.084 g/mol.

$$\text{a) } 12.50 \text{ g } (\text{NH}_4)_2\text{Cr}_2\text{O}_7 \times \frac{1 \text{ mol } (\text{NH}_4)_2\text{Cr}_2\text{O}_7}{252.084 \text{ g } (\text{NH}_4)_2\text{Cr}_2\text{O}_7} = \mathbf{0.04959 \text{ mol } (\text{NH}_4)_2\text{Cr}_2\text{O}_7}$$

(rounded from 0.049586646 mol)

$$\text{b) } 0.049586646 \text{ mol } (\text{NH}_4)_2\text{Cr}_2\text{O}_7 \times \frac{7 \text{ mol O}}{1 \text{ mol } (\text{NH}_4)_2\text{Cr}_2\text{O}_7} = \mathbf{0.3471 \text{ mol O}}$$

$$\text{c) } 0.049586646 \text{ mol } (\text{NH}_4)_2\text{Cr}_2\text{O}_7 \times \frac{8 \text{ mol H}}{1 \text{ mol } (\text{NH}_4)_2\text{Cr}_2\text{O}_7} \times \frac{6.022 \times 10^{23} \text{ H atoms}}{1 \text{ mol H}}$$

$$= \mathbf{2.389 \times 10^{23} \text{ H atoms}}$$

$$\text{d) } 0.049586646 \text{ mol } (\text{NH}_4)_2\text{Cr}_2\text{O}_7 \times \frac{2 \text{ mol Cr}}{1 \text{ mol } (\text{NH}_4)_2\text{Cr}_2\text{O}_7} \times \frac{52.00 \text{ g Cr}}{1 \text{ mol Cr}}$$

$$= \mathbf{5.157 \text{ g Cr}}$$

$$\text{e) } 0.049586646 \text{ mol } (\text{NH}_4)_2\text{Cr}_2\text{O}_7 \times \frac{2 \text{ mol NH}_4^+}{1 \text{ mol } (\text{NH}_4)_2\text{Cr}_2\text{O}_7} \times \frac{6.022 \times 10^{23} \text{ NH}_4^+ \text{ ions}}{1 \text{ mol NH}_4^+}$$

$$= \mathbf{5.972 \times 10^{22} \text{ NH}_4^+ \text{ ions}}$$

Notes on rounding and significant figures in stoichiometric calculations: In our solution keys, we use four sig fig atomic weights (as given in the periodic table in the inside front cover of the textbook), and we do not round the resulting molar masses unless they are the final answer. In multiple-step calculations, we keep extra digits, rounding only at the final answer. Unless we specify otherwise, all final answers are rounded in accordance with the significant figure rules.

4) Note that this should be review! One mole of $(\text{NH}_4)_3\text{PO}_4$ weighs 149.096 g. A mole of this compound contains 12 moles of hydrogen atoms, which weigh 12.096 g. The mass percentage of hydrogen is therefore

$$\frac{12.096 \text{ g}}{149.096 \text{ g}} \times 100\% = \mathbf{8.1130\%}.$$

5) Note that this should be review!

$$\text{a) } 3.516 \text{ g SCl}_5 \times \frac{1 \text{ mol SCl}_5}{209.31 \text{ g SCl}_5} \times \frac{5 \text{ mol Al}}{3 \text{ mol SCl}_5} \times \frac{26.98 \text{ g Al}}{1 \text{ mol Al}} = \mathbf{0.7554 \text{ g Al}}$$

$$\text{b) } 65.0 \text{ g AlCl}_3 \times \frac{1 \text{ mol AlCl}_3}{133.33 \text{ g AlCl}_3} \times \frac{5 \text{ mol Al}}{5 \text{ mol AlCl}_3} \times \frac{26.98 \text{ g Al}}{1 \text{ mol Al}} = \mathbf{13.2 \text{ g Al}}$$

c) You must determine which reactant is limiting. There are several ways to do this; it is best to pick a method, learn it well, and use it consistently. One technique that is not very elegant but tends to be easily remembered is to calculate the mass of product that will be formed based on each reactant; whichever reactant produces the least product is the limiting reactant. You can select any product you choose. In this case, let's use the mass of sulfur that is produced.

If all of the aluminum is used up, how much sulfur will we make?

$$12.5 \text{ g Al} \times \frac{1 \text{ mol Al}}{26.98 \text{ g Al}} \times \frac{3 \text{ mol S}}{5 \text{ mol Al}} \times \frac{32.06 \text{ g S}}{1 \text{ mol S}} = 8.912 \text{ g S}$$

If all of the SCl_5 is used up, how much sulfur will we make?

$$25.0 \text{ g SCl}_5 \times \frac{1 \text{ mol SCl}_5}{209.31 \text{ g SCl}_5} \times \frac{3 \text{ mol S}}{3 \text{ mol SCl}_5} \times \frac{32.06 \text{ g S}}{1 \text{ mol S}} = 3.829 \text{ g S}$$

Since we run out of SCl_5 when we have only made 3.829 g of sulfur, SCl_5 is the limiting reactant and **Al is in excess**. Now we can work out how much Al we actually use, based on the knowledge that we use 25.0 g of SCl_5 :

$$25.0 \text{ g SCl}_5 \times \frac{1 \text{ mol SCl}_5}{209.31 \text{ g SCl}_5} \times \frac{5 \text{ mol Al}}{3 \text{ mol SCl}_5} \times \frac{26.98 \text{ g Al}}{1 \text{ mol Al}} = \overline{5.371} \text{ g Al}$$

The bar over the first three digits indicates that this number is valid to 3 sig figs.

This is the mass of aluminum that is consumed in the reaction, so we will be left with:

$$12.5 \text{ g Al} - 5.371 \text{ g Al} = \mathbf{7.1 \text{ g Al}}$$

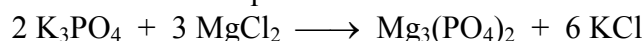
Note that this answer must be rounded to one decimal place, because the original mass of Al was only given to one decimal place. When adding and subtracting, you keep the fewest number of decimal places, not significant figures.

d) The theoretical yield of AlCl_3 is the mass that is formed when all of the limiting reactant is consumed:

$$25.0 \text{ g SCl}_5 \times \frac{1 \text{ mol SCl}_5}{209.31 \text{ g SCl}_5} \times \frac{5 \text{ mol AlCl}_3}{3 \text{ mol SCl}_5} \times \frac{133.33 \text{ g AlCl}_3}{1 \text{ mol AlCl}_3} = \mathbf{26.5 \text{ g AlCl}_3}$$

6) Note that this should be review!

First, we must write a balanced equation:



Now we can do stoichiometry.

$$715 \text{ kg K}_3\text{PO}_4 \times \frac{10^3 \text{ g}}{1 \text{ kg}} \times \frac{1 \text{ mol K}_3\text{PO}_4}{212.27 \text{ g K}_3\text{PO}_4} \times \frac{1 \text{ mol Mg}_3(\text{PO}_4)_2}{2 \text{ mol K}_3\text{PO}_4} \times \frac{262.87 \text{ g Mg}_3(\text{PO}_4)_2}{1 \text{ mol Mg}_3(\text{PO}_4)_2} \times \frac{1 \text{ kg}}{10^3 \text{ g}} \\ = \mathbf{443 \text{ kg Mg}_3(\text{PO}_4)_2}$$

7) Note that this should be review!

a) We can use a slight variation of the technique in problem 5c to answer this question: The reactant (KMnO_4 or $\text{C}_2\text{H}_5\text{OH}$) that uses less H_2SO_4 must be the limiting reactant.

If we use up all of the KMnO_4 , how much H_2SO_4 will we need?

$$4.175 \text{ g KMnO}_4 \times \frac{1 \text{ mol KMnO}_4}{158.04 \text{ g KMnO}_4} \times \frac{6 \text{ mol H}_2\text{SO}_4}{4 \text{ mol KMnO}_4} \times \frac{98.076 \text{ g H}_2\text{SO}_4}{1 \text{ mol H}_2\text{SO}_4} = 3.886 \text{ g H}_2\text{SO}_4$$

If we use up all of the $\text{C}_2\text{H}_5\text{OH}$, how much H_2SO_4 will we need?

$$2.318 \text{ g C}_2\text{H}_5\text{OH} \times \frac{1 \text{ mol C}_2\text{H}_5\text{OH}}{46.068 \text{ g C}_2\text{H}_5\text{OH}} \times \frac{6 \text{ mol H}_2\text{SO}_4}{5 \text{ mol C}_2\text{H}_5\text{OH}} \times \frac{98.076 \text{ g H}_2\text{SO}_4}{1 \text{ mol H}_2\text{SO}_4} = 5.922 \text{ g H}_2\text{SO}_4$$

Since it takes less H_2SO_4 to use up all of the KMnO_4 , KMnO_4 is the limiting reactant and **3.886 g of H_2SO_4 will be consumed.**

b) To calculate the percent yield, first we need to know the theoretical yield of acetic acid. Using the mass of the limiting reactant, we get:

$$4.175 \text{ g KMnO}_4 \times \frac{1 \text{ mol KMnO}_4}{158.04 \text{ g KMnO}_4} \times \frac{5 \text{ mol HC}_2\text{H}_3\text{O}_2}{4 \text{ mol KMnO}_4} \times \frac{60.052 \text{ g HC}_2\text{H}_3\text{O}_2}{1 \text{ mol HC}_2\text{H}_3\text{O}_2} = 1.983 \text{ g HC}_2\text{H}_3\text{O}_2$$

We expect to make 1.983 g of $\text{HC}_2\text{H}_3\text{O}_2$, but we actually make only 1.720 g of $\text{HC}_2\text{H}_3\text{O}_2$. The percent yield is

$$\frac{1.720 \text{ g}}{1.983 \text{ g}} \times 100\% = \mathbf{86.74\%}$$

8) The percentage of urea in the fertilizer is the mass of urea divided by the mass of fertilizer, then multiplied by 100%. (In general, the percentage of X in Y is $X/Y \times 100\%$.) We are given the mass of fertilizer, so we need to find the mass of urea. We can get this from the mass of ammonia that is formed when the urea breaks down.

$$0.978 \text{ g NH}_3 \times \frac{1 \text{ mol NH}_3}{17.034 \text{ g NH}_3} \times \frac{1 \text{ mol N}_2\text{H}_4\text{CO}}{2 \text{ mol NH}_3} \times \frac{60.062 \text{ g N}_2\text{H}_4\text{CO}}{1 \text{ mol N}_2\text{H}_4\text{CO}} = \underline{\quad} \text{ g N}_2\text{H}_4\text{CO}$$

The percentage of urea in the fertilizer is:

$$\frac{\underline{\quad} \text{ g N}_2\text{H}_4\text{CO}}{7.225 \text{ g fertilizer}} \times 100\% = \mathbf{23.86\%}$$

9) The standard way to convert a percent composition into an empirical formula is to work with a 100 g sample of the compound. The percentages tell us that 100 g of compound contains 87.73 g of carbon and 12.27 g of hydrogen. We begin by converting these masses into moles.

$$87.73 \text{ g C} \times \frac{1 \text{ mol C}}{12.01 \text{ g C}} = 7.30475 \text{ mol C}$$

$$12.27 \text{ g H} \times \frac{1 \text{ mol H}}{1.008 \text{ g H}} = 12.17262 \text{ mol H}$$

Next, we divide both of these by the smaller one, to obtain a whole number ratio.

$$\frac{12.17262 \text{ mol H}}{7.30475 \text{ mol C}} = 1.66640 \text{ moles of H per mole of C}$$

$$\frac{7.30475 \text{ mol C}}{7.30475 \text{ mol C}} = 1 \text{ mole of C per mole of C}$$

Elemental analyses are ordinarily very precise, so we are not justified in rounding 1.66640 up to 2. However, 1.66640 is very close to 1 and $\frac{2}{3}$, or $\frac{5}{3}$ (1.666666...), so we can state the mole ratio as $\frac{5}{3}$ of a mole of H per mole of C. To convert this into a whole number ratio, we can multiply everything by 3.

$$1.66640 \times 3 = 4.99920 \approx 5 \text{ mol H} \qquad 1 \times 3 = 3 \text{ mol C}$$

The empirical formula of our compound is therefore **C₃H₅**. (You can write the elements in either order, but it is traditional to list nonmetals in alphabetical order.)

(Note: since you are looking for whole number coefficients, it is not necessary to keep careful track of significant figures in elemental analysis problems, but you should keep a lot of digits (at least five significant figures) throughout the problem to avoid rounding errors.)

10) a) This is similar to problem 9. From the percentages, we know that 100 g of arginine contains 41.36 g of carbon, 8.10 g of hydrogen, and 32.17 g of nitrogen. We can get the oxygen by difference:

$$100 \text{ g total} - (41.36 \text{ g C} + 8.10 \text{ g H} + 32.17 \text{ g N}) = 18.37 \text{ g of oxygen}$$

Next, we convert all of these element masses into moles.

$$41.36 \text{ g C} \times \frac{1 \text{ mol C}}{12.01 \text{ g C}} = 3.44380 \text{ mol C}$$

$$8.10 \text{ g H} \times \frac{1 \text{ mol H}}{1.008 \text{ g H}} = 8.03571 \text{ mol H}$$

$$32.17 \text{ g N} \times \frac{1 \text{ mol N}}{14.01 \text{ g N}} = 2.29622 \text{ mol N}$$

$$18.37 \text{ g O} \times \frac{1 \text{ mol O}}{16.00 \text{ g O}} = 1.14813 \text{ mol O}$$

To convert these into whole number ratios, we divide each of them by the smallest:

$$\frac{3.44380 \text{ mol C}}{1.14813 \text{ mol O}} = 2.99950 \text{ moles of C per mole of O} \approx 3 \text{ mol C/mol O}$$

$$\frac{8.03571 \text{ mol H}}{1.14813 \text{ mol O}} = 6.99899 \text{ moles of H per mole of O} \approx 7 \text{ mol H/mol O}$$

$$\frac{2.29622 \text{ mol N}}{1.14813 \text{ mol O}} = 1.99997 \text{ moles of N per mole of O} \approx 2 \text{ mol N/mol O}$$

$$\frac{1.14813 \text{ mol O}}{1.14813 \text{ mol O}} = 1 \text{ mole of O per mole of O (obviously)}$$

The empirical formula must match these ratios, so the empirical formula is **C₃H₇N₂O**.

b) The molecular formula must be a whole-number multiple of the empirical formula. The possibilities are:

C₃H₇N₂O (molar mass = 87 g/mol)

C₆H₁₄N₄O₂ (molar mass = 174 g/mol)

C₉H₂₁N₆O₃ (molar mass = 261 g/mol)

etc....

Only the second formula has a molar mass between 100 and 200 g/mol, so the molecular formula of arginine is **C₆H₁₄N₄O₂**.

11) Working a combustion analysis problem involves a specific technique.

- Calculate the number of moles and the number of grams of each element *other than oxygen*, using the product masses.
- Calculate the number of grams of oxygen by subtracting the other elements from the original mass of compound, then convert grams of oxygen to moles of oxygen.
- Once you have the moles of all of the elements, proceed as with a percent composition problem.

Here are the details. We start by calculating the numbers of moles and the numbers of grams of S, C, H and N in our penicillin sample.

$$0.4503 \text{ g SO}_2 \times \frac{1 \text{ mol SO}_2}{64.06 \text{ g SO}_2} \times \frac{1 \text{ mol S}}{1 \text{ mol SO}_2} = 0.00702935 \text{ mol S}$$

$$0.007029347 \text{ mol S} \times \frac{32.06 \text{ g S}}{1 \text{ mol S}} = 0.22536 \text{ g S in the original sample of penicillin}$$

$$2.1654 \text{ g CO}_2 \times \frac{1 \text{ mol CO}_2}{44.01 \text{ g CO}_2} \times \frac{1 \text{ mol C}}{1 \text{ mol CO}_2} = 0.0492025 \text{ mol C}$$

$$0.0492025 \text{ mol C} \times \frac{12.01 \text{ g C}}{1 \text{ mol C}} = 0.59092 \text{ g C in the original sample of penicillin}$$

$$0.6965 \text{ g H}_2\text{O} \times \frac{1 \text{ mol H}_2\text{O}}{18.016 \text{ g H}_2\text{O}} \times \frac{2 \text{ mol H}}{1 \text{ mol H}_2\text{O}} = 0.0773202 \text{ mol H}$$

$$0.0773202 \text{ mol H} \times \frac{1.008 \text{ g H}}{1 \text{ mol H}} = 0.07794 \text{ g H in the original sample of penicillin}$$

$$0.2109 \text{ g NO} \times \frac{1 \text{ mol NO}}{30.01 \text{ g NO}} \times \frac{1 \text{ mol N}}{1 \text{ mol NO}} = 0.00702766 \text{ mol N}$$

$$0.00702766 \text{ mol N} \times \frac{14.01 \text{ g N}}{1 \text{ mol N}} = 0.09846 \text{ g N in the original sample of penicillin}$$

Next, we calculate the mass of oxygen in the penicillin, and from that, the number of moles of oxygen.

$$1.2177 \text{ g penicillin} - (0.22536 \text{ g S} + 0.59092 \text{ g C} + 0.07794 \text{ g H} + 0.09846 \text{ g N}) \\ = 0.22502 \text{ g O}$$

$$0.22502 \text{ g O} \times \frac{1 \text{ mol O}}{16.00 \text{ g O}} = 0.0140638 \text{ mol O}$$

Now we divide all of the numbers of moles by the smallest number (0.00702766 mol N):

$$\frac{0.00702935 \text{ mol S}}{0.00702766 \text{ mol N}} = 1.0002 \text{ moles of S per mole of N} \approx 1 \text{ mol S/mol N}$$

$$\frac{0.0492026 \text{ mol C}}{0.00702766 \text{ mol N}} = 7.0013 \text{ moles of C per mole of N} \approx 7 \text{ mol C/mol N}$$

$$\frac{0.0773202 \text{ mol S}}{0.00702766 \text{ mol N}} = 11.0023 \text{ moles of H per mole of N} \approx 11 \text{ mol H/mol N}$$

$$\frac{0.00702766 \text{ mol N}}{0.00702766 \text{ mol N}} = 1 \text{ mol N/mol N}$$

$$\frac{0.0140638 \text{ mol O}}{0.00702766 \text{ mol N}} = 2.0012 \text{ moles of O per mole of N} \approx 2 \text{ mol O/mol N}$$

The empirical formula of this compound is **C₇H₁₁NO₂S**.

12) There are several strategies you can use to solve this problem. The simplest is trial and error; X must be one of the five halogens, so try each halogen until you find the one that agrees with the masses in the problem. For instance, let's see if X is fluorine. If it is, then AlX₃ is actually AlF₃ (molar mass 83.98 g/mol) and AgX is actually AgF (molar mass 126.9 g/mol). If we start with 275 mg (0.275 g) of AlF₃, what mass of AgF do we make?

$$0.275 \text{ g AlF}_3 \times \frac{1 \text{ mol AlF}_3}{83.98 \text{ g AlF}_3} \times \frac{3 \text{ mol AgF}}{1 \text{ mol AlF}_3} \times \frac{126.9 \text{ g AgF}}{1 \text{ mol AgF}} = 1.247 \text{ g AgF}$$

We make 1.247 grams of AgF (1247 mg), which does not match the final mass given in the problem, so we conclude that element X is not fluorine. Continuing in this way, you will eventually find that the mystery element is **bromine**:

$$0.275 \text{ g AlBr}_3 \times \frac{1 \text{ mol AlBr}_3}{266.68 \text{ g AlBr}_3} \times \frac{3 \text{ mol AgBr}}{1 \text{ mol AlBr}_3} \times \frac{187.8 \text{ g AgBr}}{1 \text{ mol AgBr}} = 0.581 \text{ g AgBr}$$

You can also solve this problem algebraically. Let x be the molar mass of element X. Therefore, the molar mass of AlX₃ is $26.98 + 3x$, and the molar mass of AgX is $107.9 + x$. Use these molar masses in the stoichiometry setup:

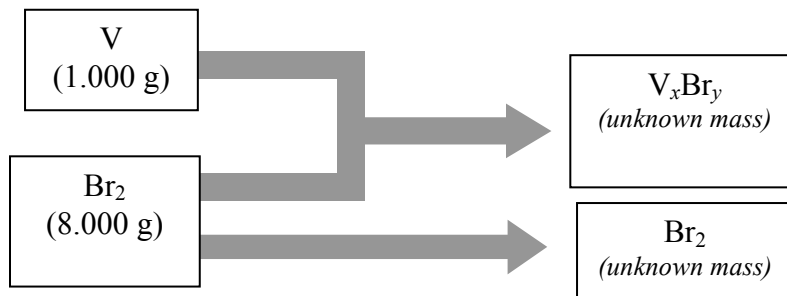
$$\begin{aligned} 0.275 \text{ g AlX}_3 \times \frac{1 \text{ mol AlX}_3}{(26.98 + 3x) \text{ g AlX}_3} \times \frac{3 \text{ mol AgX}}{1 \text{ mol AlX}_3} \times \frac{(107.9 + x) \text{ g AgX}}{1 \text{ mol AgX}} \\ = \frac{0.825(107.9 + x)}{(26.98 + 3x)} \text{ g AgX} \end{aligned}$$

We are given that the final mass of AgX is 581 mg (0.581 g), so...

$$\frac{0.825(107.9 + x)}{(26.98 + 3x)} = 0.581$$

Solving this equation gives $x = 79.89$ g/mol, which matches the atomic weight of bromine. Therefore, element X must be bromine, just as we found by trial and error.

12) For a problem like this one, it is very important to be sure you understand and organize the information you are given. There are two reactions going on here. In the first reaction, the vanadium reacts with part of the bromine, producing a compound that we can call V_xBr_y (where x and y are unknown coefficients).



In the second reaction, all of the remaining bromine reacts with Na_2SO_3 and NaOH , producing (among other things) 1.534 g of Na_2SO_4 . Since we know the mass of Na_2SO_4 that is produced, we can work out the mass of Br_2 that was consumed in this second reaction.

$$1.534 \text{ g Na}_2\text{SO}_4 \times \frac{1 \text{ mol Na}_2\text{SO}_4}{142.04 \text{ g Na}_2\text{SO}_4} \times \frac{1 \text{ mol Br}_2}{1 \text{ mol Na}_2\text{SO}_4} \times \frac{159.8 \text{ g Br}_2}{1 \text{ mol Br}_2} = \overline{1.7258 \text{ g Br}_2}$$

We now know the second reaction consumed 1.7258 g of Br_2 , so we can conclude that there was 1.7258 g of Br_2 left over after the first reaction. Therefore, the first reaction must have consumed $8.000 \text{ g} - 1.7258 \text{ g} = 6.274 \text{ g}$ of Br_2 . This in turn tells us that the product in the first reaction (V_xBr_y) contained 1.000 g V and 6.274 g Br. Using this information, we can work out the empirical formula, using the same technique we use when we are given a percent composition.

$$1.000 \text{ g V} \times \frac{1 \text{ mol V}}{50.94 \text{ g V}} = 0.0196309 \text{ mol V}$$

$$6.274 \text{ g Br} \times \frac{1 \text{ mol Br}}{79.90 \text{ g Br}} = 0.0785232 \text{ mol Br}$$

$$\frac{0.0785232 \text{ mol Br}}{0.0196309 \text{ mol V}} = 4.0000 \text{ mol Br/mol V}$$

The empirical formula of this compound is **VBr_4** .

(Note: you may be wondering why we switched from Br_2 to Br halfway through this problem. Bromine is diatomic only when it is a pure element, not combined with anything else. In chemical compounds, bromine atoms are not bonded to one another, so bromine must be treated as monatomic. The same reasoning holds true for the other diatomic elements: H, N, O, F, Cl, I and At. A good rule is to treat all elements as monatomic when you are finding the formula of a compound from mass information.)

14) To identify the element, we must find its atomic weight (i.e. its molar mass). We know that the compound contains 0.250 g of the lanthanide element. Since the compound weighs 0.427 g and contains only the lanthanide and chlorine, the compound must contain $0.427 \text{ g} - 0.250 \text{ g} = 0.177 \text{ g}$ of chlorine. We can convert this to moles:

$$0.177 \text{ g Cl} \times \frac{1 \text{ mol Cl}}{35.45 \text{ g Cl}} = 0.0049929 \text{ mol Cl}$$

If we knew the atom ratio in the formula, we could use stoichiometry to determine the molar mass of the element directly. However, we have three possible chemical formulas. Since three is not a large number, it is easiest to solve this problem by trial and error.

Let's start by guessing that the chemical formula of the compound is MCl_2 . We know that the compound contains 0.0049929 moles of Cl, so it must contain half this amount of element "M":

$$0.0049929 \text{ mol Cl} \times \frac{1 \text{ mol M}}{2 \text{ mol Cl}} = 0.0024965 \text{ mol M}$$

The mass of element "M" is 0.250 g, so we know that 0.0024965 moles of M weighs 0.250 g (if the chemical formula is really MCl_2). Therefore, the molar mass of M is:

$$\frac{0.250 \text{ g}}{0.0024965 \text{ mol}} = 100 \text{ g/mol}$$

We have figured out that if the chemical formula of the compound is MCl_2 , then the atomic weight of M must be 100. Unfortunately, none of the lanthanides has this atomic weight, so we conclude that the chemical formula is not MCl_2 .

Let's try MCl_3 next. Again, we start with the knowledge that the compound contains 0.0049929 moles of Cl. If the formula is MCl_3 , the compound must contain one third as many moles of M.

$$0.0049929 \text{ mol Cl} \times \frac{1 \text{ mol M}}{3 \text{ mol Cl}} = 0.0016643 \text{ mol M}$$

The molar mass of M would then be:

$$\frac{0.250 \text{ g}}{0.0016079 \text{ mol}} = 150 \text{ g/mol}$$

This is the atomic weight of samarium (Sm, element #62), so the formula could be MCl_3 . However, we should check the possibility that the formula might be MCl_4 also.

If the formula of the compound is MCl_4 , the compound contains one fourth as many moles of M as it contains moles of Cl.

$$0.0049929 \text{ mol Cl} \times \frac{1 \text{ mol M}}{4 \text{ mol Cl}} = 0.0012482 \text{ mol M}$$

The molar mass of M would be:

$$\frac{0.250 \text{ g}}{0.0012482 \text{ mol}} = 200 \text{ g/mol}$$

No lanthanide has this atomic weight (this is roughly the molar mass of mercury, but mercury isn't a lanthanide), so the compound cannot have the formula MCl_4 . Therefore, we can conclude that **the metal is Sm and the compound is SmCl_3** .

Optional challenge problem:

Please prioritize your study time!! If you still have lingering questions about any of the previous problems, spend your time and energy on those required problems before tackling this optional challenge problem.

From the information in the problem, we know the following:

- In the original metal sample, the mass of Mg and the mass of Al add up to 0.317 g.
- Both Mg and Al react with HCl to form H_2 .
- The total mass of H_2 formed is:

$$337 \text{ mL} \times \frac{1 \text{ L}}{1000 \text{ mL}} \times \frac{0.08158 \text{ g}}{1 \text{ L}} = 0.02749246 \text{ g}$$

However, we do not know the individual masses of Mg and Al, and we do not know how much of the H_2 was produced by each reaction. A good method in such cases is to use a pair of variables to represent the quantities we want to determine (the masses of the two metals).

Let x = mass of Mg (in grams)

Let y = mass of Al (in grams)

Now we use stoichiometry to relate the masses of Mg and Al to the mass of hydrogen.

$$x \text{ g Mg} \times \frac{1 \text{ mol Mg}}{24.31 \text{ g Mg}} \times \frac{1 \text{ mol H}_2}{1 \text{ mol Mg}} \times \frac{2.016 \text{ g H}_2}{1 \text{ mol H}_2} = 0.0829288x \text{ g H}_2$$

(This tells us that when we start with x grams of Mg, we must make $0.0829288x$ grams of H_2 .)

$$y \text{ g Al} \times \frac{1 \text{ mol Al}}{26.98 \text{ g Al}} \times \frac{3 \text{ mol H}_2}{2 \text{ mol Al}} \times \frac{2.016 \text{ g H}_2}{1 \text{ mol H}_2} = 0.112083y \text{ g H}_2$$

(This tells us that when we start with y grams of Al, we must make $0.112083y$ grams of H_2 .)

We can use the information we've gathered to write two equations. We know that the total mass of metal was 0.317 g, so our first equation is:

$$x + y = 0.317$$

We also know that we make a total of 0.02749246 g of H₂, so our second equation is:

$$0.0829288x + 0.112083y = 0.02749246$$

The rest of the solution is algebra. To solve a system of two equations with two unknowns, one good technique is to solve one equation for one of the variables (it doesn't matter which).

Solving the first equation for x gives us:

$$x = 0.317 - y$$

Then we replace x in the second equation with $0.317 - y$, giving us an equation that only has one variable.

$$0.0829288(0.317 - y) + 0.112083y = 0.02749246$$

Solving this equation gives us $y = 0.041298$, so the original metal sample contained 0.041298 grams of Al (ignoring significant figures). To find x , we substitute this value of y into either of the original equations. Doing so gives $x = 0.275702$, so the metal sample contained 0.275702 grams of Mg. The percentage of each metal in the original sample is therefore:

$$\frac{0.275702 \text{ g Mg}}{0.317 \text{ g metal}} = \mathbf{87.0\% \text{ Mg}}$$

$$\frac{0.041298 \text{ g Al}}{0.317 \text{ g metal}} = \mathbf{13.0\% \text{ Al}}$$