Stoichiometry - AP level
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Ten Examples

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Problem #1: A salt contains only barium and one of the halide ions. A 0.1480 g sample of the salt was dissolved in water and an excess of sulfuric acid was added to form barium sulfate, which was filtered, dried and weighed. Its mass was found to be 0.1660 g. What is the formula for the barium halide?

Solution:

1) Calculate mass of barium ion in precipitate:

0.166 g x (137.33 / 233.395) = 0.097674672 g

2) Determine mass of halide ion in dissolved sample:

0.1480 g - 0.097674672 g = 0.050325328 g

3) How many moles of barium ion were present in dissolved sample?

0.097674672 g / 137.33 g mol^{-1} = 0.000711241 mol

4) Barium halide compounds are known to take the formula BaX_2. How many moles of halide were present in the dissolved sample?

0.000711241 mol x 2 = 0.001422481 mol

5) Which halide has 0.001422 mole of it weigh 0.05032 grams?

F: 0.001422481 mol times 18.9984 g mol^{-1}

Cl: 0.001422481 mol times 35.453 g mol^{-1}

Br: 0.001422481 mol times 79.904 g mol^{-1}

I: 0.001422481 mol times 126.90447 g mol^{-1}

6) $BaCl_2$

Problem #2: A 4.000 g sample of M_2S_3 is converted to MO_2 and loses 0.277 g. What is the atomic weight of M?

Here is an alternate solution to the problem above which is a bit more compact. It might be clearer to you.

Solution:

1) Write the chemical equation:

$M_2S_3 + 5O_2$ ---> $2MO_2 + 3SO_2$

2) Some facts that I can't think of a good title for:

a) The grams of M in M_2S_3 equals the grams of M in $2MO_2$. (Notice the inclusion of the coefficient.)

b) Let x = the atomic weight of M.

3) Construct gravimetric factors for M:

$M_2S_3 \Rightarrow 2x / (2x + 96)$

$2MO_2 \Rightarrow 2x / (2x + 64)$

Comment: there are 2 M and 4 O in $2MO_2$.

4) Calculate grams of M_2O_3 and $2MO_2$:

$M_2S_3 \Rightarrow$ (4.000 g) times [2x / (2x + 96)]

$2MO_2 \Rightarrow$ (3.723 g) times [2x / (2x + 64)]

5) Set them equal to each other and solve:

(4.000 g) times [2x / (2x + 96)] = (3.723 g) times [2x / (2x + 64)]

You do the math.

x = 183

Here is an alternate solution to the problem above which is a bit more compact. It might be clearer to you.

In addition to the alternate solution, there are three additional forms of this problem, two which have solutions appended.

Problem #3a: A 5.000 gram sample of a dry mixture of potassium hydroxide, potassium carbonate and potassium chloride is reacted with 0.100 L of 2.00-molar HCl solution. A 249.0 mL sample of dry carbon dioxide gas, measured at 22.0 °C and 740.0 torr, is obtained from this reaction. What was the percentage of potassium carbonate in the mixture?

Solution to 3a:

1) Calculate moles of CO_2 using PV = nRT:

$$\boxed{n = \frac{(740.0\ \text{mmHg})\ (0.249\ \text{L})}{(760.0\ \frac{\text{mmHg}}{\text{atm}})\ (0.08206\ \frac{\text{L atm}}{\text{mol K}})\ (295.0\ \text{K})} = 0.0100153\ \text{mol}}$$

2) Recall the reaction between K_2CO_3 and HCl:

$K_2CO_3 + 2HCl \text{ ---> } 2KCl + CO_2 + H_2O$

3) Note the 1:1 molar ratio between K_2CO_3 and CO_2. From this we conclude:

0.0100153 mol K_2CO_3 present in the dry mixture

4) Calculate grams K_2CO_3, then its mass percentage:

0.0100153 mol x 138.2057 g/mol = 1.384 g

(1.384 g / 5.000 g) x 100 = 27.68%

Problem #3b: The excess HCl in problem 3a was found by titration to be chemically equivalent to 86.60 mL of 1.50-molar sodium hydroxide. What was the percentage of KOH and of KCl in the original mixture?

Solution to 3b:

1) Calculate excess HCl:

We know this: $HCl + NaOH \text{ ---> } NaCl + H_2O$

There is a 1:1 molar ratio between NaOH and HCl

Use NaOH data to determine moles of excess HCl:

$(0.08660\ L)\ (1.50\ mol\ L^{-1}) = 0.1299\ mol$

2) Calculate total moles of HCl in the 0.100 L of 2.00-molar HCl solution:

$(0.100\ L)\ (2.00\ mol\ L^{-1}) = 0.200\ mol$

3) Calculate amount of HCl used:

0.200 mol - 0.1299 mol = 0.0701 mol

However, this is the combined amount of HCl used to titrate K_2CO_3 **AND** KOH

4) Calculate HCl used to titrate KOH by subtracting the HCl used to titrate K_2CO_3:

0.0701 mol - 0.0200 mol = 0.0501 mol

The 0.02 comes from the fact that 2 HCl were required to neutralize one K_2CO_3. See step 2 in problem 3a.

5) Calculate grams of KOH

$KOH + HCl \text{ ---> } KCl + H_2O$

From the 1:1 ratio between the reactants, we know there are 0.0501 mol of KOH in the original sample.

$0.0501\ mol \times 56.1057\ g\ mol^{-1} = 2.811\ g$

6) Determine KCl by subtraction:

1.384 g + 2.811 g = 4.195 g

5.000 g - 4.756 g = 0.805 g

7) You may do the weight percentages for KOH and KCl on your own.

Problem #4a: For the reaction below, when 0.5000 g of XI_3 reacts completely, 0.2360 g of XCl_3 is obtained. Calculate the atomic weight of element X and identify it.

$2XI_3 + 3Cl_2 \text{--->} 2XCl_3 + 3I_2$

Solution:

1) From the balanced equation, we know that the moles XI_3 used equals moles XCl_3 produced. Therefore:

0.5000 g / (x + 381 g/mol) = 0.2360 g / (x + 106.5 g/mol)

The 381 is the weight of three iodines and the 106.5 is the weight of three chlorines.

2) Solving for x, we find it equal to 138.9. This is the atomic weight of lanthanum.

Problem #4b: If 0.520 grams of XCl_3 are treated with iodine, 0.979 g of XI_3 are produced. What is the chemical symbol for this element?

$2XCl_3 + 3I_2 \text{--->} 2XI_3 + 3Cl_2$

Solution:

1) From the balanced equation, we know that the moles XCl_3 used equals moles XI_3 produced. Therefore:

0.520 g / (x + 106.5 g/mol) = 0.979 g / (x + 381 g/mol)

The 106.5 is the weight of three chlorines and the 381 is the weight of three iodines.

2) Solving for x, we find it equal to 204.5. This is the atomic weight of thallium and its symbol is Tl.

Problem #4c: Consider the reaction involving unknown element X:

$F_2 + 2XBr \text{--->} Br_2 + XF$

When 5.500g of XBr reacts, 3.693g of Br_2 is produced. Identify element X.

Solution:

1) Moles of Br_2:

3.693 g / 159.808 g/mol = 0.023109 mol

2) Moles of XBr:

0.023109 mol times 2 = 0.046218 mol

3) Molar mass of XBr:

5.500 g / 0.046218 mol = 119.0 g/mol

4) Identity of X

119.0 - 79.904 = 39.0

X is potassium

Problem #5: A 2.077 g sample of an element, which has an atomic mass between 40 and 55, reacts with oxygen to form 3.708 g of an oxide. Determine the formula mass of the oxide (and identify the element).

Solution:

1) Determine moles of oxygen involved:

3.708 g minus 2.077 g = 1.631 g

1.631 g divided by 15.9994 g/mol = 0.10194 mol of oxygen in M_xO_y.

2) Determine the atomic weight of M, if the formula is MO:

2.077 g / 0.10194 mol = 20.37 g/mol

The formula of M_xO_y is not MO since the atomic weight of M is known to be between 40 and 55.

3) Determine the atomic weight of M, if the formula is MO_2:

2.077 g / 0.05097 mol = 40.75 g/mol

This is within the acceptable range.

Before going on, I would like to point out that a M_2O_3 formula leads to an atomic weight of approximately 34 and a M_2O formula leads to approxmately 10.2. You may do the math on those two possibilities.

4) The closest value to our desired range is potassium and yes, it does form the compound KO_2, known as potassium superoxide.

If we were to look for a +4 forming ion (in other words, something to satisfy the MO_2 requirement) in the 40-55 range, we find titanium. However, its atomic weight is about 48, which is too high a value predicted by the MO_2 formula.

Only KO_2 provides an atomic weight within the parameters specified by the problem.

Problem #6: A 12.5843 g sample of $ZrBr_4$ was dissolved and, after several steps, all of the combined bromine was precipitated as AgBr. The silver content of AgBr was found to be 13.2160 g. Assume the atomic masses of silver and bromine to be 107.868 and 79.904. What value was obtained for the atomic mass of Zr from this experiment?

Solution:

1) calculate the moles Ag in the AgBr:

13.2160 g / 107.868 g/mol = 0.12252 mole of Ag in the AgBr

2) Since AgBr has a 1:1 molar ratio of silver to bromine, Br in sample is 0.12252 mole. Calculate the grams Br in the sample:

0.12252 g times 79.904 g/mol = 9.78985 g of Br

3) calculate Zr in sample:

12.5843 - 9.78985 = 2.79445 g of Zr

4) determine moles of Zr present:

The 0.12252 mol of Br represents the 4 moles of Br in the formula $ZrBr_4$.

Therefore, 0.12252 / 4 equals the moles of Zr present.

moles Zr = 0.030635 mol

5) Determine molecular weight of Zr:

2.79445 g / 0.030635 mol = 91.2 g/mole

Problem #7: Two different chloride compounds of platinum are known, compound X and Y. When 3.45 g of compound X is heated, 2.72 g of compound Y is formed along with some chlorine gas. Upon further heating, the 2.72 g of compound Y is decomposed to 1.99 g of platinum metal and some more chlorine gas. Determine the formulas of compounds X and Y.

Solution:

1) Determine how much Cl_2 is produced in each reaction:

Heat 3.45 g X to get 2.72 g Y and Cl_2 gas
Mass of Cl_2 gas is 3.45 - 2.72 = 0.73 g

Heat 2.72 g Y to get 1.99 g Pt and Cl_2 gas
Mass of Cl_2 gas in compound Y is 2.72 - 1.99 = 0.73 g

2) Determine formula for compound Y:

moles of Pt atom = 1.99 g / 195.1 g/mol = 0.01 mol
moles of Cl atom (not chlorine molecules!) = 0.73 g / 35.45 g/mol = 0.02 mol

simplest mole ratio:
Pt = 0.01 / 0.01 = 1
Cl = 0.02 / 0.01 = 2

formula of Y = $PtCl_2$

3) Determine formula for compound X:

mass of Pt = 1.99 g
mass of Cl_2 = 0.73 g from 1st decomposition + 0.73 g from 2nd decomposition = 1.46 g total in X

mole of Pt atom = 1.99 g / 195.1 g/mol = 0.01 mol
mole of Cl atom (not chlorine molecules!) = 1.46 g / 35.45 g/mol = 0.04 mol

simplest mole ratio:
Pt = 0.01 / 0.01 = 1
Cl = 0.04 / 0.01 = 4

Formula of X = $PtCl_4$

Problem #8: The active ingredients of an antacid tablet contained only magnesium hydroxide and aluminum hydroxide. Complete neutralization of a sample of the active ingredients required 48.5 mL of 0.187 M hydrochloric acid. The chloride salts from this neutralization were obtained by evaporation of the filtrate from the titration; they weighed 0.4200 g. What was the percentage by mass of magnesium hydroxide in the active ingredients of the antacid tablet?

Solution:

1) Determine moles of chloride ion used:

$(0.187 \text{ mol L}^{-1})\ (0.0485 \text{ L}) = 9.0685 \times 10^{-3}$ mol

2) Detemine moles of chloride ion used to make the $MgCl_2$:

9.0685×10^{-3} mol x 0.40 = 3.6274×10^{-3} mol

3) Explanation of the 0.40:

Out of every 5 Cl^- used, three go to make one $AlCl_3$ and two go to make one $MgCl_2$

2 out of 5 is 40%.

4) Determine moles of $MgCl_2$ that are present:

3.6274×10^{-3} mol divided by 2 = 1.8137×10^{-3} mol

Remember, one $MgCl_2$ is present for every two Cl^-

5) Repeating (2), (3) and (4) with appropriate modifications yields 1.8137×10^{-3} mol for $AlCl_3$. (Use 0.60, not 0.40 and divide by 3, not 2.)

6) Determine moles of $Mg(OH)_2$ and $Al(OH)_3$ in original mixture:

First, write the equations which produce $MgCl_2$ and $AlCl_3$

$Mg(OH)_2 + 2HCl \text{ ---> } MgCl_2 + 2H_2O$

$Al(OH)_3 + 3HCl \text{ ---> } AlCl_3 + 3H_2O$

Second, note the following molar ratios:

$Mg(OH)_2 : MgCl_2$ is 1 : 1

$Al(OH)_3 : AlCl_3$ is 1 : 1

This means we have the following molar amounts in the original sample:

$Mg(OH)_2 = 1.8137 \times 10^{-3}$ mol

$Al(OH)_3 = 1.8137 \times 10^{-3}$ mol

7) Determine grams of $Mg(OH)_2$ and $Al(OH)_3$:

This is left to the reader.

8) Determine percent by mass of $Mg(OH)_2$ in the original sample.

This is left to the reader.

Problem #9: When the supply of oxygen is limited, iron metal reacts with oxygen to produce a mixture of FeO and Fe_2O_3. In a certain experiment, 20.00 g of iron metal was reacted with 11.20 g of oxygen gas. After the experiment the iron was totally consumed and 3.56 g oxygen gas remained. Calculate the amounts of FeO and Fe_2O_3 formed in this experiment.

Solution:

1) Determine the amount of oxygen gas (in grams, then moles) consumed:

11.20 g - 3.56 g = 7.64 g of O_2

7.64 g / 31.9988 g mol^{-1} = 0.238759 mol (I kept several guard digits.)

2) Determine how many moles of O_2 goes to form FeO and how many moles goes to form Fe_2O_3:

$2Fe + O_2$ ---> 2FeO
$4Fe + 3O_2$ ---> $2Fe_2O_3$

25% goes to form FeO
75% goes to form Fe_2O_3

0.238759 mol x 0.25 = 0.05968974 mol of O_2 used to form FeO
0.238759 x 0.75 = 0.17906922 mol of O_2 used to form Fe_2O_3

3) Determine moles of FeO and moles of Fe_2O_3 formed:

FeO ⇒ the O_2 to FeO ratio is 1:2, therefore double the amount of O_2 used to get FeO produced:

0.1193795 mol of FeO produced.

Fe_2O_3 ⇒ the O_2 to Fe_2O_3 ratio is 3:2, therefore double the amount of O_2 used to get Fe_2O_3 produced and then divide that value by three:

0.1193795 mol of Fe_2O_3 produced.

4) Determine grams of FeO and grams of Fe_2O_3 formed:

FeO ⇒ 0.1193795 mol x 71.844 g mol^{-1} = 8.5767 g
Fe_2O_3 ⇒ 0.1193795 mol x 159.687 g mol^{-1} = 19.06335 g

The above values (you may round them off on your own) are the answer to the problem, but I thought one more step would be fun.

5) Determine grams of Fe in FeO and grams of Fe in Fe_2O_3:

FeO ⇒ 8.5767 g x (55.845 / 71.844) = 6.667 g of Fe in FeO
Fe_2O_3 ⇒ 19.06335 g x (111.69 / 159.687) = 13.33

This, within rounding errors, totals to the 20.00 g of Fe mentioned in the problem.

Problem #10: 0.197 g of magnesium is burned in air:

2Mg + O_2 ---> 2MgO

However, some of the magnesium reacts with nitrogen in the air to form magnesium nitride instead:

3Mg + N_2 ---> Mg_3N_2

So you have a mixture of MgO and Mg_3N_2 weighing 0.315 g. Determine what percentage of the Mg formed the nitride in the initial reaction.

Solution:

1) This problem is solved with two simultaneous equations in two unknowns:

First equation: x + y = 0.315 g

Second equation: (24.305 / 40.304) x + (72.915 / 100.929) y = 0.197 g

Explanation:

x = the mass of MgO in the mixture of MgO and Mg_3N_2
y = the mass of Mg_3N_2 in the mixture of MgO and Mg_3N_2

(24.305 / 40.304) is the percentage of Mg in MgO
(72.915 / 100.929) is the percentage of Mg in Mg_3N_2

2) Substitute x = 0.315 - y into the second equation:

(24.305 / 40.304) (0.315 - y) + (72.915 / 100.929) y = 0.197

3) Solve:

0.189958 - 0.60304y + 0.72244y = 0.197

$0.1194y = 0.007042$

$y = 0.058978$ g

4) Percent of Mg_3N_2 in original sample:

(0.058978 g / 0.315 g) x 100 = 18.7% (to three sf)

Ten Examples

Problems #11-25

Problems #26-50

All Examples & Problems (no solutions)

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Problem #11: Hydroxylammonium chloride reacts with iron(III) chloride, $FeCl_3$, in solution to produce iron(II) chloride, HCl, H_2O and a compound of nitrogen. It was found that 2.00 g of iron(III) chloride reacted in this way with 31.0 mL of 0.200 M hydroxylammonium chloride. Suggest a possible formula for the compound of nitrogen so produced.

Solution:

1) Determine moles of hydroxylammonium chloride ($NH_3OH^+Cl^-$) and iron(III) chloride:

> (0.0310 L) (0.200 mol/L) = 0.0062 mol $NH_3OH^+Cl^-$
>
> 2.00 g / 162.204 g/mol = 0.01233 mol $FeCl_3$
>
> The key is to see that the moles of $FeCl_3$ are double that of the hydroxylammonium chloride.

2) Determine the oxidation number of N in hydroxylammonium chloride:

> N = -1
>
> The determination of this is left to the reader.
>
> By the way, we know that iron is reduced, so the nitrogen MUST be oxidized.

3) Allow the 0.0062 moles of nitrogen atoms to move from -1 oxidation state to zero:

> this liberates 0.0062 mol of electrons, which go to reduce 0.0062 mol of Fe^{3+} ions (which is only half of the ions available)

4) Allow the 0.0062 mole of N atoms (because of step 3 just above, now at an oxidation state of zero) to move from zero to an oxidation state of +1:

> this liberates another 0.0062 mol of electrons, which go to reduce 0.0062 mol of Fe^{3+} ions (which is the other half of the ions available)

5) We need nitrogen in the +1 oxidation state in our compound:

> N_2O

Problem #12: How many phosphate ions are in a sample of hydroxyapatite [$Ca_5(PO_4)_3OH$] that contains 5.50 x 10^{-3} grams of oxygen?

Solution:

1) Determine moles of oxygen:

5.50 x 10^{-3} g divided by 16.00 g/mol = 3.4375 x 10^{-4} mol

2) Determine moles of hydroxyapatite:

the molar ratio between hydroxyapatite and oxygen is 1:13

3.4375 x 10^{-4} mol divided by 13 = 2.64423 x 10^{-5} mol of hydroxyapatite

3) Determine moles of phosphate ions:

the molar ratio between hydroxyapatite and phosphate is 1:3

2.64423 x 10^{-5} mol times 3 = 7.9326923 x 10^{-5} mol of phosphate ions

4) Determine number of ions:

7.9326923 x 10^{-5} mol times 6.022 x 10^{23} mol^{-1} = 4.78 x 10^{19}

Problem #13: A mixture consisting of only sodium chloride (NaCl) and potassium chloride (KCl) weighs 1.0000 g. When the mixture is dissolved in water and an excess of silver nitrate is added, all the chloride ions associated with the original mixture are precipitated as insoluble silver chloride (AgCl). The mass of the silver chloride is found to be 2.1476 g. Calculate the mass percentages of sodium chloride and potassium chloride in the original mixture.

Solution #1:

1) Set up this equation:

(x) (grams Cl from NaCl) + (1 − x) (grams Cl from KCl) = total grams chloride

x = grams of NaCl in original mixture
1 − x = grams of KCl in original mixture

grams Cl from NaCl = 35.5/58.4
grams Cl from KCl = 35.5/74.6
total grams chloride = (2.1476 g) (35.5/143.3)

The numbers in the denominators are the molar masses of NaCl, KCl and AgCl. The three ratios are called "gravimetric factors."

(x) (35.5/58.4) + (1 − x) (35.5/74.6) = (2.1476 g) (35.5/143.3)

2) Solve:

(x) (35.5/58.4) + (1 − x) (35.5/74.6) = (2.1476 g) (35.5/143.3)

0.6079x + 0.4759 − 0.4759x = 0.532

$0.132x = 0.0561$

$x = 0.425$ g

This is the mass of NaCl in the original mixture. This computes to 43% of the original mixture.

Solution #2:

1) Set up this equation:

mass of NaCl + mass of KCl = 1.000 g

x = grams of NaCl in original mixture
1 − x = grams of KCl in original mixture

Therefore:

$(x) + (1 - x) = 1.000$ g

2) Transform x and 1 − x as follows:

(x divided by 58.442 g/mol) times (143.321 g/mol) = 2.452x

((1 − x) divided by 74.551 g/mol) times (143.321 g/mol) = 1.992(1 − x)

Comment: using NaCl as an example, the transformation does this:

a) First, we calculate the moles of NaCl.
b) Since there is a 1:1 molar ratio between NCl and AgCl, this is also the number of moles of AgCl produced.
c) Multiply by the molar mass of AgCl to get the grams of AgCl produced from x grams of NaCl.

3) Write (then solve) this equation:

$2.452x + 1.992(1 - x) = 2.1476$

$2.452x + 1.922 - 1.922x = 2.1476$

$0.53x = 0.2256$

$x = 0.426$ g

Solution #3:

Graph the theoretical AgCl yield from one gram of 100% KCl through one gram of 100% NaCl with a few mixtures in-between to demonstrate linearity (or not) and interpolate your answer.

Comment: This would be fun to do on a spreadsheet someday.

Problem #14: Ammonia is produce industrially by reacting:

$N_2 + 3H_2 \text{ ---> } 2NH_3$

Assuming 100% yield, what mass of ammonia will be produced from a 1:1 molar ratio mixture in a reactor that has a volume of 8.75×10^3 L under a total pressure of 2.75×10^7 Pa at 455 °C.

Solution:

1) A 1:1 molar ratio means hydrogen is the limiting reagent. This is because a 1:3 ratio of nitrogen to hydrogen is required to fully react all the nitrogen.

2) Determine the initial pressure of hydrogen:

2.75×10^7 Pa = 2.75×10^4 kPa

2.75×10^4 kPa / 101.325 kPa/atm = 271.404 atm

271.404 atm / 2 = 135.702 atm

The divide by two is done because hydrogen makes up 50% of the reacting mixture.

3) Use PV = nRT:

(135.702 atm) (8.75×10^3 L) = (x) (0.08206) (728 K)

x = 19876.111 mol of hydrogen

4) Convert to amount of ammonia:

3:2 molar ratio for H_2 : NH_3

moles of NH_3 = (19876.111 x 2) / 3 = 13250.74 mol

13250.74 mol x 17.0307 g/mol = 225669.4 g = 2.26×10^5 g

Problem #15: Upon heating, a 4.250 g sample loses 0.314 grams. Assuming the sample is $BaCl_2 \cdot 2H_2O$ and NaCl, calculate the mass percent of $BaCl_2 \cdot 2H_2O$.

Solution:

1) Upon heating, only water is lost. Determine the moles of water lost:

0.314 g / 18.015 g/mol = 0.01743 mol of water

2) From the formula $BaCl_2 \cdot 2H_2O$, we know:

2 moles of water per one mole of $BaCl_2$

therefore 0.008715 mole of $BaCl_2$

3) Determine grams, then percentage of barium chloride:

0.008715 mol x 244.2656 g/mol = 2.128775 g of $BaCl_2 \cdot 2H_2O$

2.128775 / 4.250 = 50.09%

Problem #16: A 0.6118 g sample containing only $MgCl_2$ and NaCl was analyzed by adding 145.0 mL of 0.1006 M $AgNO_3$. The precipitate of AgCl(s) formed had a mass of 1.7272 g. Calculate the mass of each component

($MgCl_2$ and NaCl) in the original sample.

Solution:

1) Using a gravimetric factor, determine the amount of chloride ion that preciptated:

1.7272 g times (35.453 / 143.321) = 0.42725366 g

2) Determine relative contribution of chloride by $MgCl_2$ and NaCl:

for every three Cl^- that react with Ag^+:

two come from $MgCl_2$
one comes from NaCl

Therefore:

magnesium chloride's contribution is 2/3
sodium chloride's contribution is 1/3

Please realize, this contribution is in terms of moles. So . . . .

3) Convert grams of chloride to moles:

0.42725366 g / 35.453 g/mol = 0.01205127 mol

4) Determine moles of NaCl in sample:

0.01205127 mol times one-third = 0.00401709 mol

5) Determine grams of NaCl in sample:

0.00401709 mol times 58.443 g/mol = 0.23477 g

to four sig figs: 0.2348 g

The mass of $MgCl_2$ may be obtained by subtraction.

Problem #17: Ammonium nitrate and potassium chlorate both produce oxygen gas when decomposed by heating. Without doing detailed calculations, determine which of the two yields the greater

(a) number of moles of O_2 per mole of solid and
(b) number of grams of O_2 per gram of solid.

The unbalanced equations are:

$NH_4NO_3(s) \text{ ---> } N_2(g) + O_2(g) + H_2O$
$KClO_3(s) \text{ ---> } KCl(s) + O_2(g)$

Solution:

1) Balance both equations:

$2NH_4NO_3(s)$ ---> $2N_2(g) + O_2(g) + 4H_2O$
$2KClO_3(s)$ ---> $2KCl(s) + 3O_2(g)$

2) Write molar ratios:

NH_4NO_3 to O_2 is 2:1
$KClO_3$ to O_2 is 2:3

3) Let the molar ratios be in terms of one mole of the solid:

NH_4NO_3 to O_2 is 1:0.5
$KClO_3$ to O_2 is 1:1.5

4) Answer to (a):

In terms of moles, $KClO_3$ produces more O_2 than NH_4NO_3. In fact, $KClO_3$ produces three times as much oxygen (compare 1.5 to 0.5).

5) Convert the moles of each molar ratio to grams:

NH_4NO_3 to O_2 is 80.04 to 16.00
$KClO_3$ to O_2 is 122.55 to 48.0

6) Let the gram ratios be in terms of one gram of the substance:

NH_4NO_3 to O_2 is 1 to 0.20
$KClO_3$ to O_2 is 1 to 0.39

4) Answer to (b):

In terms of grams, $KClO_3$ produces oxygen approximately twice as fast (0.30 to 0.20) as NH_4NO_3.

Problem #18: An element X forms both a dichloride (XCl_2) and a tetrachloride (XCl_4), Treatment of 10.00 g XCl_2 with excess chlorine forms 12.55 g XCl_4. Calculate the atomic mass of X, and identify X.

Solution:

1) Write a balanced equation for the reaction:

$XCl_2 + Cl_2$ ---> XCl_4

2) Determine grams, then moles of Cl_2 that react:

12.55 g minus 10.00 g = 2.55 g

2.55 g / 70.906 g/mol = 0.035963 mol

3) Determine moles of XCl_2 present:

Due to 1:1 molar ratio between XCl_2 and Cl_2, the moles of XCl_2 equals 0.035963 mol

4) Determine the molecular weight of XCl_2:

10.00 g / 0.035963 mol = 278.06 g/mol

5) Determine both atomic weight and identity of X:

278.06 g/mol minus 70.906 g/mol = 207.2 g/mol (rounded off to the 0.1 place)

X is lead.

Problem #19: Water is added to 4.267 g of UF_6. The only products of the reaction are 3.730 g of a solid containg only uranium, oxygen, and fluorine and 0.970 g of a gas. The gas is 95.0% fluorine and the remainder is hydrogen.

a) What fraction of the fluorine of the orginal is in the solid and what fraction in the gas after the reaction?
b) What is the formula of the solid product?

Solution to a:

1) Calculate moles UF_6 present:

4.267 g / 352.018 g/mol = 0.01212154 mol

2) Calculate grams of fluorine in UF_6:

(0.01212154 mol) (113.988 g/mol) = 1.38171 g

The 113.988 comes from the fact that 6 F are in UF_6

3) Calculate mass of fluorine in gas

(0.970 g) (0.950) = 0.9215 g

4) Calculate mass of fluorine in solid:

1.38171 g - 0.9215 g = 0.46021 g

5) Calculate percent fluorine in solid:

0.46021 g/ 1.38171 g = 33.307%

6) Calculate percent fluorine in gas:

100% - 33.307% = 66.693%

Solution to b:

1) Calculate mass of H_2O reacted:

(3.730 g + 0.970 g) - 4.267 g = 0.433 g

2) Calculate mass of oxygen in solid product:

(0.433 g / 18.015 g/mol)(15.999 g/mol) = 0.38454438 g

3) Calculate mass of uranium in solid product:

3.730 g - (0.38454438 g + 0.46021 g) = 2.88524562 g

4) Calculate moles of U, F and O in solid product:

U: 2.88524562 g / 238.029 g/mol = 0.01212 mol

F: 0.4604 g/ 18.998 g/mol = 0.02423 mol

O: 0.38454438 g/15.999 g/mol = 0.02403 mol

5) To more clearly see the 1:2:2 ratio, simply divide by the smallest number:

U: 0.01212/0.01212 = 1
F: 0.02423/0.01212 = 1.999
O: 0.02403/0.01212 = 1.98

The formula of the unknown is UF_2O_2 and the overall reaction is:

$UF_6 + 2H_2O \text{ --> } UF_2O_2 + 4HF$

Problem #20: A compound containing titanium and chlorine is analyzed by converting all the titanium into 1.20 g of titanium dioxide and all the chlorine into 6.45 g of AgCl. What is the simplest (empirical) formula for the original compound?

Solution:

By the way, note the use of millimoles rather than moles. Remember 1 mole equals 1000 millimoles.

1) This reaction happens:

$Ti_xCl_y \text{ ----> } x\ TiO_2 + y\ AgCl$

2) Determine moles TiO_2 formed:

1.20 g / 79.90 g mol^{-1} = 15.02 mmol

3) Determine moles of AgCl formed:

6.45 g / 143.32 g mol^{-1} = 45.00 mmol

4) Determine millimoles of Ti and Cl in original compound:

the Ti : TiO_2 molar ratio is 1:1, therefore 15.02 mmol of Ti
The Cl : AgCl molar ratio is 1:1, therefore 45.00 mmol of Cl

5) The mole ratio of Ti to Cl in the compound is 15:45 or 1:3. Therefore:

the compound's formula is $TiCl_3$

Problem #21: An unknown element X is found in two compounds, XCl_2 and XBr_2. In the following reaction:

$XBr_2 + Cl_2$ ---> $XCl_2 + Br_2$

when 1.5000 g XBr_2 is used, 0.8897 g XCl_2 is formed. Identify the element X.

Solution:

moles of XBr_2 = moles of XCl_2

1.500 / (x + 159.808) = 0.8897 / (x + 70.906)

(0.8897) (x + 159.808) = (1.500) (x + 70.906)

0.8897x + 142.1811776 = 1.5x + 106.359

0.6103x = 35.8221776

x = 58.70

Element X is Ni.

By the way, be careful. Take a look at Co and you'll see 58.93 and think that that is close enough. Nickel is 58.69. The Co/Ni pairing is one of three with the atomic weight goes down as you proceed from element to element. Ar/K and Te/I are the other two.

Problem #22: When the supply of oxygen is limited, iron metal reacts with oxygen to produce a mixture of FeO and Fe_2O_3. In a certain experiment, 20.00 g of iron metal was reacted with 11.20 g of oxygen gas. After the experiment the iron was totally consumed and 3.24 g oxygen gas remained. Calculate the amounts of FeO and Fe_2O_3 formed in this experiment.

Solution:

1) Determine moles of O_2 that reacted:

11.20 g - 3.24 g = 7.96 g

7.96 g / 31.99886 g /mol = 0.248759 mol

2) There are two independent reactions occurring simultaneously:

$2Fe + O_2$ ---> $2FeO$
$4Fe + 3O_2$ ---> $2Fe_2O_3$

By the way, you might be tempted to write one equation:

$3Fe + 2O_2$ --> $FeO + Fe_2O_3$

This would be incorrect as it masks the fact that the two oxygen to iron oxide molar ratios are different. Because the ratios are different, the calculation for FeO and Fe_2O_3 must be separate.

3) Let z be the mass fraction of Fe that produced FeO. Then 1-z is the mass fraction of Fe that produced Fe_2O_3 Determine the amount of oxygen consumed as (a) some Fe goes to FeO and (b) some Fe goes to Fe_2O_3:

(a) the FeO calculation:

$$(20.00 \text{ g Fe})(z) \times \frac{1 \text{ mol Fe}}{55.8450 \text{ g Fe}} \times \frac{1 \text{ mol } O_2}{2 \text{ mol Fe}} = 0.179067z \text{ (mol } O_2 \text{ consumed by producing FeO)}$$

(b) the Fe_2O_3 calculation

$$(20.00 \text{ g Fe})(1\text{-}z) \times \frac{1 \text{ mol Fe}}{55.8450 \text{ g Fe}} \times \frac{3 \text{ mol } O_2}{4 \text{ mol Fe}} = (0.268601 - 0.268601z) \text{ (mol } O_2 \text{ consumed by producing } Fe_2O_3)$$

Note that the amount of oxygen consumed is expressed in terms of the (unknown) mass fraction 'z.' The next step will determine the value of z.

4) Add the two amounts of O_2 consumed and set it equal to the total moles of O_2 reacted. Solve for 'z':

$(0.179067z) + (0.268601 - 0.268601z) = 0.248759$

$-0.089534z = -0.019842$

$z = 0.221614$

5) Determine (a) the amount of FeO produced and (b) the amount of Fe_2O_3 produced:

(a) the FeO calculation:

$$(20.00 \text{ g Fe})(0.221614) \times \frac{1 \text{ mol Fe}}{55.8450 \text{ g Fe}} \times \frac{2 \text{ mol FeO}}{2 \text{ mol Fe}} \times \frac{71.8444 \text{ g Fe}}{1 \text{ mol Fe}} = 5.702 \text{ g FeO produced}$$

(b) the Fe_2O_3 calculation:

$$(20.00 \text{ g Fe})(0.778386) \times \frac{1 \text{ mol Fe}}{55.8450 \text{ g Fe}} \times \frac{2 \text{ mol } Fe_2O_3}{4 \text{ mol Fe}} \times \frac{159.6882 \text{ g } Fe_2O_3}{1 \text{ mol Fe}} = 22.258 \text{ g } Fe_2O_3 \text{ produced}$$

6) Let's see if the Law of Conservation of Mass works:

20.00 g + 11.2 g - 3.24 g = 27.96 g
22.258 g + 5.702 g = 27.96 g

Yay!

7) Here's an incorrect solution to this problem.

Problem #23: A sheet of iron with a surface area of 525 cm^2 is covered with a coating of rust that has an average thickness of 0.0021 cm. What minimum volume of an HCl solution, in mL, having a density of 1.07 g/mL and consisting of 14% HCl by mass is required to clean the surface of the metal by reacting with the rust? Assume that the rust is Fe_2O_3(s), that it has a density of 5.2 g/cm^3, and that the reaction is:

$Fe_2O_3(s) + 6HCl(aq) \text{ ---> } 2FeCl_3(aq) + 3H_2O(\ell)$

Solution:

1) Volume of rust coating:

525 cm^2 x 0.0021 cm = 1.1025 cm^3

2) Mass of rust:

1.1025 cm^3 x 5.2 g/cm^3 = 5.733 g

3) Moles of rust:

5.733 g / 159.687 g/mol = 0.03590148 mol

4) Moles of HCl needed:

From the balanced equation, the Fe_2O_3 to HCl molar ratio is 1:6

1 is to 6 as 0.03590148 mol is to x

x = 0.21540888 mol

5) Mass of HCl needed:

0.21540888 mol x 36.4609 g/mol = 7.854 g

6) Mass of 14% solution required:

14 is to 100 as 7.854 is to x

x = 56.1 g

7) Volume of solution required:

56.1 g / 1.07 g/mL = 52.43 mL

Commentary: there are those teachers that absolutely insist on writing up a problem like the above in "dimensional analysis style." This means to string together all the calculations into one line. Here it is, done two different ways:

$(525\ cm^2 \times 0.0021\ cm)\ (5.2\ g/cm^3)\ (1\ mol / 159.687\ g)\ (6\ mol\ HCl / 1\ mol\ Fe_2O_3)\ (36.4609\ g / 1\ mol)\ (100 / 14)\ (1\ mL / 1.07\ g)$

$$(525\ cm^2 \times 0.0021\ cm) \frac{5.2\ g}{1\ cm^3} \times \frac{1\ mol}{159.687\ g} \times \frac{6}{1} \times \frac{36.4609\ g}{1\ mol} \times \frac{100}{14} \times \frac{1\ mL}{1.07\ g} = 52.43\ mL$$

Make sure that the 159.687 gets used in a division. Also, note the 100/14. You have to multiply by 100, then divide by 14 before moving on.

There are those who insist that the DA style is clearer, making it pedagogically sounder to teach. The ChemTeam is not among that group.

Problem #24: A 1.42 g sample of a pure compound, with formula M_2SO_4, was dissolved in water and treated with an excess of aqueous calcium chloride, resulting in the precipitation of all the sulfate ions as calcium sulfate. The precipitate was collected, dried, and found to weigh 1.36 g. Determine the atomic mass of M. What element is it?

Solution:

The -2 charge on the sulfate ion is ignored.

1) Moles of sulfate precipitated:

$(1.36\ g\ CaSO_4) / (136.1406\ g\ CaSO_4/mol) \times (1\ mol\ SO_4 / 1\ mol\ CaSO_4) = 0.0099897\ mol\ SO_4$

$(1.36\ g\ CaSO_4) / (136.1406\ g\ CaSO_4/mol)$ ---> converts grams to moles
$(1\ mol\ SO_4 / 1\ mol\ CaSO_4)$ ---> for every one mole of calcium sulfate produced, one mole of sulfate ion was used

2) Moles of M_2SO_4 holding that many moles of sulfate:

$(0.0099897\ mol\ SO_4) \times (1\ mol\ M_2SO_4 / 1\ mol\ SO_4) = 0.0099897\ mol\ M_2SO_4$

3) Molar mass of M_2SO_4:

$1.42\ g / 0.0099897\ mol = 142.1\ g/mol$

4) Molar mass of M:

$SO_4 = 96.0626\ g/mol$

$142.1\ g/mol\ M_2SO_4 - 96.0626\ g/mol\ SO_4 = 46.0\ g/mol\ M_2$

$(46.0\ g/mol\ M_2) / 2 = 23.0\ g/mol\ M$

sodium.

Problem #25: Calculate the volume change when iron is oxidized to Fe_2O_3 ($d = 5.24\ g/cm^3$). The density of Fe is $7.787\ g/cm^3$.

Solution:

1) Write the balanced chemical equation for the reaction:

$2Fe + \frac{3}{2}O_2 \text{ ---> } Fe_2O_3$

I decided to balance it with a fraction so as to use a Fe to Fe_2O_3 molar ratio of 2:1. If I had balanced the equation with whole numbers, the ratio I would have used would be 4:2. This would not affect the answer, my balancing choice was purely a stylistic one.

2) Let's start with 7.787 g (this is 1.00 cm^3 of iron). Convert it to moles:

7.787 g / 55.845 g/mol = 0.13944 mol

3) Use 2:1 molar ratio:

$$\frac{2}{1} = \frac{0.13944 \text{ mol}}{x}$$

x = 0.06972 mol (of Fe_2O_3 produced)

4) Convert moles of Fe_2O_3 to grams:

(0.06972 mol) (159.687 g/mol) = 11.1334 g

5) Convert to cm^3:

11.1334 g / 5.24 g/cm^3 = 2.12 cm^3

The volume changes from 1.00 cm^3 to 2.12 cm^3.

Ten Examples

Problems #1-10

Problems #26-50

All Examples & Problems (no solutions)

Return to Stoichiometry Menu

Stoichiometry - AP level
Problems #26-50

Ten Examples

Problems #1-10

Problems #11-25

All Examples & Problems (no solutions)

Return to Stoichiometry Menu

Problem #26: A 0.204 gram sample of a metal, M, reacts completely with sulfuric acid according to:

$M + H_2SO_4 \text{--->} MSO_4 + H_2$

A volume of 213 mL of hydrogen is collected over water; the water level in the collecting vessel is the same as the outside level. Atmospheric pressure is 756.0 torr and the temperature is 25.0 °C. Calculate the molar mass of the metal.

The vapor pressure of water at various temperatures can be found in this table.

Solution:

1) Use Dalton's Law of Partial Pressures to determine the pressure of the dry H_2:

756.0 - 23.8 = 732.2 torr

2) Converting that to atm gives:

732.2 torr / 760.0 torr/atm = 0.9634 atm

3) Use the ideal gas law:

PV = n RT

(0.9634 atm) (0.213 L) = (n) (0.08206 L atm / mol K) (298 K)

n = 0.0083915 mol H_2

4) From the balanced equation, 1 mole of H_2 is formed when 1 mol of metal M is reacted. Therefore, the moles of metal that reacted was:

0.0083915 mol M

5) The molar mass of M is:

0.204 g / 0.0083915 mol = 24.3 g/mol

The metal was magnesium.

Problem #27: A common way to obtain a pure metal from its impure metal oxide is to react the oxide with carbon, expressed generically as:

$2MO(s) + C(s) \text{ ---> } 2M(s) + CO_2(g)$

If 5.00 g of an unknown metal oxide (MO) reacted with excess carbon and formed 738 mL of CO_2 at 200.0 °C and 0.978 atm, what is the identity of the metal?

Solution:

1) Using PV = nRT, solve for moles of CO_2:

(0.978 atm) (0.738 L) = (n) (0.08206 L atm / mol K) (473 K)

n = 0.0186 mol

2) Determine moles of oxygen atoms in metal oxide:

For every one mole of CO_2 produced, two moles of MO were consumed.

1 is to 2 as 0.0186 is to x

x = 0.0372 mol

3) Determine molar mass of MO:

5.00 g / 0.0372 mol = 134.44 mol

4) Determine atomic weight of M:

x + 16.00 = 134.44

x = 118.44 g/mol

M is tin.

Problem #28: A compound of P and F was analyzed as follows: heating 0.2324 g of the compound in a 378 cm^3 flask turned all of it to gas, which had a pressure of 97.3 mmHg at 77 °C. Then, the gas was mixed with calcium chloride solution which turned all of the F to 0.2631 g of CaF_2. Determine the molecular formula of the compound.

Solution:

1) Molecular weight of the gas:

(97.3 / 760) (0.378) = (n) (0.08206) (350)

n = 0.001684967 mol

0.2324 g / 0.001684967 mol = 138 g/mol

We will use this to move from the empirical formula to the molecular formula.

2) Mass of F:

0.2631 g times (38.0 g / 78.074 g) = 0.128055 g

(38.0 g / 78.074 g) is called a gravimetric factor. It is the decimal percent of F in CaF_2.

3) Moles of F:

0.128055 g / 19.0 g/mol = 0.00674 mol

4) Moles of P:

0.2324 g minus 0.128055 g = 0.104345 g

0.104345 g / 31.0 g/mol = 0.003366 mol

5) Empirical formula:

P ---> 0.003366 / 0.003366 = 1
F ---> 0.00674 / 0.003366 = 2

PF_2

6) Molecular formula:

PF_2 weighs 31 + 38 = 69

138 / 69 = 2

PF_2 times 2 = P_2F_4

Another P_2F_4 problem.

Problem #29: A metal chloride reacts with silver nitrate solution to give a precipitate of silver chloride according to following equation:

$MCl_2 + 2AgNO_3$ ---> $M(NO_3)_2 + 2AgCl$

When a solution containing 0.4750 g of metal chloride is made to react with silver nitrate, 1.435 grams of silver chloride are formed. Identify the metal.

Solution:

1) Determine moles of AgCl produced:

1.435 g AgCl / 143.32 g/mol = 1.001256×10^{-2} mol AgCl

2) Determine moles of MCl_2 consumed:

1.001256×10^{-2} mol AgCl x (1 mol MCl_2 / 2 mol AgCl) = 5.00628×10^{-3} mol MCl_2

3) Determine molar mass of MCl_2:

0.4750 g / 5.00628×10^{-3} mol = 94.88 g/mol

4) Determine the atomic weight and identity of M:

The atomic of M will be:

94.88 minus the mass of 2 chlorine atoms:
94.88 minus (2 x 35.453) = 23.97 g/mol

M is magnesium

Problem #30: An unidentified metal M reacts with an unidentified halogen X to form a compound MX_2. When heated the compound decomposes by the reaction:

$2MX_2$ (s) ---> 2MX (s) + X_2 (g)

When 1.12 g of MX_2 is heated, 0.720 g of MX is obtained along with 56.0 mL of X_2 gas (at STP).

a) What is the atomic mass and the identity of the halogen X?
b) What is the atomic mass and identity of the metal M?

Solution:

1) Calculate moles of X_2 collected:

PV = nRT

(1.00 atm) (0.0560 L) = (n) (0.08206 L atm / mol K) (273 K)

n = 0.00250 mol

2) Calculate mass of X_2 collected:

1.12 g minus 0.720 g = 0.40 g

3) Calculate molecular weight of X_2 and determine identity of X

0.40 g/0.00250 moles = 160 g/mol

half of 160 is 80, the atomic weight of X

bromine

4) Determine moles of MX:

MX : X_2 molar ratio is 2 : 1

therefore, 0.0050 moles of MX

5) Calculate molar mass of MX:

0.720 g/0.0050 moles = 144 g/mole

6) Calculate atomic weight of M:

144 g/mol minus 80 g/mol = 64 g/mole

copper

Problem #31: A metal sulfate has the formula M_2SO_4. 10.99 g of the compound was dissolved in water to make 500.0 cm^3 of solution. A 25.0 cm^3 sample was removed and reacted with an excess of $BaCl_2(aq)$ to produce a precipitate of $BaSO_4$, which when dried had a mass of 1.167 g.

a) Determine the number of moles of $BaSO_4$ precipitated.
b) Determine the concentration of M_2SO_4
c) Identify M

Solution:

1) Moles of $BaSO_4$ precipitated from 25 cm^3:

1.167 g / 233.4 g/mol = 0.005 mol

2) Moles of $BaSO_4$ that would have precipitated from 500 cm^3:

0.005 mol is to 25 as x is to 500

x = 0.100 mol of $BaSO_4$

3) The molarity of the M_2SO_4 is this:

0.100 mol / 0.500 L = 0.200 M

The 0.100 mole comes from the fact that there is a 1:1 molar ratio between M_2SO_4 and $BaSO_4$

$M_2SO_4 + CaCl_2 \text{--->} BaSO_4 + 2MCl$

4) What is M?

0.1 mol of sulfate weighs 96.061 x 0.1 = 9.6061 g

10.99 minus 9.6061 = 1.3839 g of M^+ in solution

There were 0.2 mol of M^+ in the 500 cm^3 [from this: $M_2SO_4(s) \text{--->} 2M^+(aq) + SO_4^{2-}(aq)$]

1.3839 g / 0.2 mol = 6.9 g/mol

Li_2SO_4

Problem #32: An element, X, forms two compounds with bromine: XBr_2 and XBr_4. When 10.00 grams of the XBr_2 is reacted with excess bromine, 14.35 g of XBr_4 is formed. Identify X.

Solution:

14.35 - 10.00 = 4.35 g of Br_2 reacted

4.35 g / 159.808 g/mol = 0.02722 mol of Br_2

$XBr_2 + Br_2$ ---> XBr_4

XBr_2 and Br_2 react in a 1:1 molar ratio

Therefore, 10.00 g represents 0.02722 mol of XBr_2

10.00 g / 0.02722 mol = 367 g/mol

In 367.377 g of XBr_2, there is 159.808 g of bromine

367.377 - 159.808 = 207.569 <--- this is the atomic weight of X

X is lead.

Problem #33: Exactly 4.32 g of oxygen gas was required to completely combust a 2.16 g sample of a mixture of methanol and ethanol:

(1) How many moles of ethanol are contained within the sample?
(2) What is the percentage by weight of methanol in the sample?

Solution:

1) Balanced chemical equations:

$CH_3OH + 1.5O_2$ ---> $CO_2 + 2H_2O$
$C_2H_5OH + 3O_2$ ---> $2CO_2 + 3H_2O$

2) Let x = mass of methanol; let y = mass of ethanol

$x + y = 2.16$

$[(1.5)(x/32)] + [(3)(y/46)] = 4.32/32$

3) Using x = 2.16 - y:

$[(1.5)((2.16-y)/32)] + [(3)(y/46)] = 0.135$

$[(3.24 - 1.5y) / 32] + [3y / 46] = 0.135$

4) Multiply each side by 32, then by 46:

$(46)(3.24 - 1.5y) + (32)(3y) = 198.72$

$149.04 - 69y + 96y = 198.72$

$27y = 49.68$

y = 1.84 g

x = 0.32 g

5) Answers to (1) and (2):

moles ethanol ---> 1.84 g / 46 g/mol = 0.040 mol percent methanol ---> (0.32 g / 2.16 g) * 100 = 14.8%

Problem #34: A 3.41 g sample of a metallic element, M, reacts completely with 0.0158 mol of a gas, X_2, to form 4.53 g MX. What are the identities of M and X?

Solution:

$2M + X_2$ ---> 2MX

We know that M and X_2 react in a 2:1 molar ratio.

Therefore, 0.0158 mol of X_2 reacts with twice that many moles of M, 0.0316 mol

We can now determine the atomic weight of M:

3.41 g / 0.0316 mol = 107.9 g/mol

From the periodic table, we see that M is silver.

4.53 - 3.41 = 1.12 g <--- this is the mass of X in MX

We know that there is a 1:1 molar ratio between M and MX, therefore 0.0316 of MX was produced. Since we know the formula is MX, we know that 0.0316 mol of X is involved.

1.12 g / 0.0316 mol = 35.44 g/mol

This atomic weight is within experimental error for X to be chlorine.

Problem #35: When 2.3 moles of X reacts with 1.6 moles of Y, 71 grams of Z are produced. What is the molar mass of Z?

3X + 4Y ---> 5Z

This reaction has a 50% yield.

Solution:

1) Determine the limiting reagent:

X: 2.3 / 3 = 0.77
Y: 1.6 / 4 = 0.4

Y is the limiting reagent.

2) How much Z is produced at 100% yield:

4 is to 5 as 1.6 is to x

x = 2.0 mol

3) The reaction has only 50% yield, so 1.0 mole of Z was produced. The molar mass of Z is:

71 g / 1.0 mol = 71 g/mol

Problem #36: 20.0 mL of solution containing NaCl and KCl gave, on evaporation to dryness, 0.180 g of the mixed chlorides. 20.0 mL of the same solution gave 0.370 g of AgCl on treatment with a slight excess of the $AgNO_3$ solution. Calculate, for the original solution, the mass per liter of both chlorides.

Solution:

1) Let x = mass NaCl and let y = mass KCl. from that, we get our first equation:

$x + y = 0.180$

2) Using AgCl, determine the moles of chloride ion in the solution:

moles AgCl ---> 0.370 g / 143.321 g/mol = 0.00258 mol

0.00258 mol gives the moles of chloride based on this:

$Ag^+(aq) + Cl^-(aq)$ ---> AgCl(s)

For every 1 mole of AgCl that precipitated, there was 1 mole of chloride in solution.

3) We can now write our second equation:

x / 58.4428 g/mol + y / 74.553 = 0.00258

x / 58.4428 ---> moles of NaCl
y / 74.553 ---> moles of KCl

Based on the 1:1 molar ratios between NaCl and Cl^- (as well as KCl and Cl^-), we know that the two above divisions give the moles of chloride contributed from the NaCl and the KCl.

4) We will now substitute the first equation into the second and solve:

$x = 0.180 - y$

$[(0.180 - y) / 58.4428] + [y / 74.553] = 0.00258$

$13.4 - 74.553y + 58.4428y = 11.2$

$16.1y = 2.2$

y = 0.137 g <--- mass KCl in 20.0 mL
x = 0.180 − 0.137 = 0.0430 g <--- mass NaCl in 20.0 mL

mass KCl in 1 L ---> 0.137 g / 0.020 L = 6.85 g
mass NaCl in 1 L ---> 0.043 / 0.020 L = 2.15 g

Problem #37: You are given a mixture of three hydrated salts: $Na_2CO_3 \cdot 10H_2O$, $MgSO_4 \cdot 7H_2O$, and $CuSO_4 \cdot 5H_2O$. The total mass of the mixture is 12.123 grams. When the mixture is heated gently, the following two reactions occur:

$Na_2CO_3 \cdot 10H_2O(s)$ ---> $Na_2CO_3 \cdot 7H_2O(s) + 3H_2O(g)$
$MgSO_4 \cdot 7H_2O(s)$ ---> $MgSO_4 \cdot H_2O(s) + 6H_2O(g)$

After these reactions are complete, the mass of the mixture has decreased to 9.049 grams. This mixture is then heated more strongly, and the following additional reactions occur:

$Na_2CO_3 \cdot 7H_2O(s) \text{--->} Na_2CO_3(s) + 7H_2O(g)$
$MgSO_4 \cdot H_2O(s) \text{--->} MgSO_4(s) + H_2O(g)$
$CuSO_4 \cdot 5H_2O(s) \text{--->} CuSO_4(s) + 5H_2O(g)$

After this final heating, the mass of the mixture has decreased to 6.412 grams. From this information, calculate the masses of each of the three compounds in the original mixture.

Solution:

Comment: three simultaneous equations in three unknowns are required.

1) The first equation:

$X + Y + Z = 12.123$ g

where:

X = the mass of $Na_2CO_3 \cdot 10H_2O$
Y = the mass of $MgSO_4 \cdot 7H_2O$
Z = the mass of $CuSO_4 \cdot 5H_2O$

2) The second equation is developed based on the information gained from the first heating. Here is the second equation:

$(232.0916 / 286.136) (X) + (138.3808 / 246.4696) (Y) + Z = 9.049$ g

(232.0916 / 286.136) ---> this gravimetric factor is the decimal percent decrease in the original mass (the X) of the sodium carbonate decahydrate as it is changed to the heptahydrate. 232 is the molar mass of the heptahydrate and 286 is the molar mass of the decahydrate.

(138.3808 / 246.4696) ---> this gravimetric factor is the decimal percent decrease in the original mass (the Y) of magnesium heptahydrate as it is changed to the monohydrate. 246 is the molar mass of the heptahydrate and 138 is the molar mass of the monohydrate

There is no gravimetric factor for copper(II) sulfate pentahydrate since it did not lose any mass. It started with Z grams present and ended with Z grams present.

By the way, after the heating, the sample is now composed of $Na_2CO_3 \cdot 7H_2O$, $MgSO_4 \cdot H_2O$, and $CuSO_4 \cdot 5H_2O$.

3) The third equation is developed based on the information gained from the second heating. Here is the third equation:

$(105.988 / 286.136) (X) + (120.366 / 246.4696) (Y) + (159.607 / 249.681) (Z) = 6.412$ g

Notice that, in each gravimetric factor, the denominator is the molar mass of the original hydrate. This is because the 6.412 value references the entire loss of mass from the starting value of 12.123 g.

4) I'm going to rewrite the three equations, but I will use decimals rather than fractions:

$X + Y + Z = 12.123$

$0.8111234X + 0.5614518Y + Z = 9.049$

$0.3704113X + 0.4883604Y + 0.6392437Z = 6.412$

5) To calculate the answers, I used a 3 Equation System Solver. When I did that, I got the following answers (which I rounded off):

1.373 grams of $Na_2CO_3 \cdot 10H_2O$
6.418 grams of $MgSO_4 \cdot 7H_2O$
4.332 grams of $CuSO_4 \cdot 5H_2O$

6) I will start a step-by-step solution:

Rewrite the first equation:

$Z = 12.123 - (X + Y)$

Substitute into the second and third equations:

$0.8111234X + 0.5614518Y + [12.123 - (X + Y)] = 9.049$

$0.3704113X + 0.4883604Y + 0.6392437[12.123 - (X + Y)] = 6.412$

We now have two simultaneous equations in 2 unknowns.

And that is where I will leave it.

7) I must admit that I did not solve this problem on my own. When I found the problem in my notes, I did an Internet search which yielded some discussion about this problem. Look for links to 'www.chemicalforums.com'

Problem #38: A mixture of $CuSO_4 \cdot 5H_2O$ and $MgSO_4 \cdot 7H_2O$ is heated until all the water is lost. If 5.020 g of the mixture gives 2.988 g of the anhydrous salts, what is the percent by mass of $CuSO_4 \cdot 5H_2O$ in the mixture?

Solution:

1) For every mole of $CuSO_4$ there are 5 moles of H_2O, and for every mole of $MgSO_4$ there are 7 moles of H_2O. Therefore:

(5 * moles of $CuSO_4$) + (7 * moles of $MgSO_4$) = moles of H_2O

2) Determine moles of water lost:

$5.020 - 2.988 = 2.032$ g

2.032 g / 18.015 g/mol = 0.112795 mol

3) Let X be the grams of anhydrous $CuSO_4$. Therefore:

$2.988 - X$ ---> the grams of anhydrous $MgSO_4$

3) We can now substitute into the equation in step 1:

$5 * (X / 159.6096) + 7 * [(2.988 - X) / 120.3686] = 0.112795$

$X / 159.6096$ ---> moles of anhydrous $CuSO_4$
$(2.988 - X) / 120.3686$ ---> moles of anhydrous $MgSO_4$

4) Solve it:

$(159.6096) (120.3686) * 5 * (X / 159.6096) + (159.6096) (120.3686) * 7 * [(2.988 - X) / 120.3686] = (0.112795) (159.6096) (120.3686)$

$(120.3686) * 5 * X + (159.6096) * 7 * (2.988 - X) = 2167.016$

$601.843X + 3338.3944 - 1117.2672X = 2167.016$

$515.4242X = 1171.3784$

$x = 2.27265$ g of anhydrous $CuSO_4$

5) Determine moles of anhydrous $CuSO_4$:

2.27265 g / 159.6096 g/mol = 0.0142388 mol

6) Determine mass of water associated with 0.0142388 mol of anhydrous $CuSO_4$:

(0.0142388 mol) (5) = 0.071194 mol of H_2O

(0.071194 mol) (18.015 g/mol) = 1.28256 g

7) Determine mass of $CuSO_4 \cdot 5H_2O$ in the original sample and its percentage:

2.27265 g + 1.28256 g = 3.55521 g

(3.55521 / 5.020) * 100 = 70.82%

Ten Examples

Problems #1-10

Problems #11-25

All Examples & Problems (no solutions)

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