Challenging questions on Ch. 3 (Stoichiometry)

Q: A mixture of KCl and KNO_3 is 44.20% potassium by mass. What is the percentage of KCl in the mixture?

A: X = mass of KCl in 100 g mixture
$100 - X$ = mass of KNO_3 in 100 g mixture

m.m. KCl = 74.55 g/mol
m.m. KNO_3 = 101.10 g/mol

mass of K in KCl = $X \text{ g KCl} \times \frac{1 \text{ mol KCl}}{74.55 \text{ g KCl}} \times \frac{1 \text{ mol K}}{1 \text{ mol KCl}} \times \frac{39.10 \text{ g K}}{1 \text{ mol K}} = 0.524\,X$

mass of K in KNO_3 = $(100-X)\text{ g } KNO_3 \times \frac{1 \text{ mol } KNO_3}{101.10 \text{ g } KNO_3} \times \frac{1 \text{ mol K}}{1 \text{ mol } KNO_3} \times \frac{39.10 \text{ g K}}{1 \text{ mol K}} =$
$= 38.7 - 0.387\,X$

mass of K in 100 g mixture = 44.20 g

$$0.524\,X + 38.7 - 0.387\,X = 44.20$$

$$X = 40.1 \text{ g} \implies \% \text{ KCl} = 40.1$$

Q: You heat 3.854 g of a mixture of Fe_3O_4 and FeO to form 4.148 g Fe_2O_3. What was the mass percent of FeO in the original mixture?

A: X = mass of FeO
$3.854 - X$ = mass of Fe_3O_4

m.m. FeO = 71.84 g/mol
m.m. Fe_3O_4 = 231.53 g/mol
m.m. Fe_2O_3 = 159.69 g/mol

Since the product mass is more than the starting mixture, clearly oxygen from the air reacted with the oxides to add mass. We don't need to figure out the reaction, but it does mean that we need to use Fe mass as the constant between the reactant oxides and the product.

mass of Fe in product = $4.148 \text{ g } Fe_2O_3 \times \frac{1 \text{ mol } Fe_2O_3}{159.69 \text{ g}} \times \frac{2 \text{ mol Fe}}{1 \text{ mol } Fe_2O_3} \times \frac{55.84 \text{ g Fe}}{1 \text{ mol Fe}}$
$= 2.901$ g Fe
= mass of Fe in reactants

mass of Fe in FeO = $X \text{ g FeO} \times \frac{1 \text{ mol FeO}}{71.84 \text{ g FeO}} \times \frac{1 \text{ mol Fe}}{1 \text{ mol FeO}} \times \frac{55.848 \text{ g Fe}}{1 \text{ mol Fe}} = 0.7773\,X$

mass of Fe in Fe_3O_4 = $(3.854 - X)\text{ g } Fe_3O_4 \times \frac{1 \text{ mol } Fe_3O_4}{231.53 \text{ g } Fe_3O_4} \times \frac{3 \text{ mol Fe}}{1 \text{ mol } Fe_3O_4} \times \frac{55.84 \text{ g Fe}}{1 \text{ mol Fe}} = 2.789 - 0.7235\,X$

$0.7773\,X + 2.789 - 0.7235\,X = 2.901 \implies X = 2.082$ g FeO

mass % of FeO = $\frac{2.082}{3.854} \times 100 = 54\%$

Q: A 2.80 g sample of an oxide of bromine is converted to 4.698 g of AgBr. Calculate the empirical formula of the oxide

A: There is no Ag in the reactant we know of (an oxide of bromine) and there is no oxygen in the product we know of (AgBr). We don't need to know the other reactants and products, or the exact reaction. We do know that all of the Br in the reactant ended up in AgBr.

We can find the number of moles of AgBr, from which we can find the number of moles of Br. We can then find the amount of oxygen in the bromine oxide because we know its mass. Then we can find the empirical formula.

m.m. AgBr = 187.78 g/mol m.m. Br = 79.90 g/mol m.m. O = 16.00 g/mol

$$4.698\text{ g AgBr} \times \frac{1\text{ mol AgBr}}{187.78\text{ g}} \times \frac{1\text{ mol Br}}{1\text{ mol AgBr}} \times \frac{79.90\text{ g Br}}{1\text{ mol Br}} = 2.00\text{ g Br}$$

mass of O in bromine oxide = 2.80 − 2.00 = 0.80 g O

Now that we have the masses of Br and O, we can use our usual method to find the empirical formula:

$$\left.\begin{array}{l} \text{Br: } 2.00\text{ g} \times \dfrac{1\text{ mol}}{79.90\text{ g}} = 0.0250\text{ mol} \rightarrow \dfrac{0.0250}{0.0250} = 1 \\ \text{O: } 0.80\text{ g} \times \dfrac{1\text{ mol}}{16.00\text{ g}} = 0.050\text{ mol} \rightarrow \dfrac{0.050}{0.0250} = 2 \end{array}\right\} \Rightarrow BrO_2$$

Q: A 0.4987 g sample of a compound containing only carbon, hydrogen, and oxygen was burned in oxygen to yield 0.9267 g of CO_2 and 0.1897 g of H_2O. What is the empirical formula of the compound?

A: Since oxygen is coming from both the compound and from O_2, it can't be used to tie back to the reactant. But we can use C and H amounts. Once we find the amounts of C and H (which are conserved and were present in the reactant in the same amounts), we can find the amount of O from the mass of sample. Then we can use our usual method to find the empirical formula.

m.m. CO_2 = 44.01 g/mol m.m. H_2O = 18.02 g/mol

$$\text{mass of C} = 0.9267\,g\,CO_2 \times \frac{1\,mol\,CO_2}{44.01\,g\,CO_2} \times \frac{1\,mol\,C}{1\,mol\,CO_2} \times \frac{12.01\,g\,C}{1\,mol\,C} = 0.2529\,g\,C$$

$$\text{mass of H} = 0.1897\,g\,H_2O \times \frac{1\,mol\,H_2O}{18.02\,g\,H_2O} \times \frac{2\,mol\,H}{1\,mol\,H_2O} \times \frac{1.008\,g\,H}{1\,mol} = 0.02122\,g\,H$$

$$\text{mass of O} = 0.4987 - (0.2529 + 0.02122) = 0.2246\,g\,O$$

To find the empirical formula:

$$C:\ 0.2529\,g\,C \times \frac{1\,mol\,C}{12.01\,g\,C} = 0.02106\,mol \quad \frac{0.02106}{0.01403} \simeq 1.5$$

$$H:\ 0.02122\,g\,H \times \frac{1\,mol\,H}{1.008\,g\,H} = 0.02105\,mol \quad \frac{0.02105}{0.01403} \simeq 1.5$$

$$O:\ 0.2246\,g\,O \times \frac{1\,mol\,O}{16.00\,g\,O} = 0.01403\,mol \quad \frac{0.01403}{0.01403} \simeq 1$$

$$\left.\begin{matrix}1.5\\1.5\\1\end{matrix}\right\} \times 2 = \left\{\begin{matrix}3\\3\\2\end{matrix}\right.$$

$\Rightarrow C_3H_3O_2$

Q: A 1.00 g sample of an alkaline earth metal chloride is treated with excess silver nitrate. All of the chloride is recovered as 1.38 g of silver chloride. Identify the metal.

A: If all of Cl^- is contained in 1.38 g of AgCl, we can find how many moles of Cl^- there was in the original sample. Then using the ratio of moles of metal and chloride in the alkaline earth metal chloride, we can find the molar mass of the metal.

The formula of the sample must be MCl_2 because alkaline earth metals form cations with +2 charge.

m.m. of AgCl = 143.35 g/mol   m.m. of Cl = 35.45

$$1.38\text{ g AgCl} \times \underbrace{\frac{1\text{ mol AgCl}}{143.35\text{ g AgCl}}}_{\text{grams to mols}} \times \underbrace{\frac{1\text{ mol Cl}^-}{1\text{ mol AgCl}}}_{\text{from AgCl formula to mols of Cl}} \times \frac{35.45\text{ g Cl}^-}{1\text{ mol Cl}^-} = 0.341\text{ g Cl}^- \text{ from } 1.00\text{ g sample}$$

$\Rightarrow$ mass of M = 1.00 - 0.341 = 0.659 g

$$\frac{0.659\text{ g M}^{2+}}{0.341\text{ g Cl}^-} \times \frac{35.45\text{ g Cl}^-}{1\text{ mol Cl}^-} \times \frac{2\text{ mol Cl}^-}{1\text{ mol M}^{2+}} = \frac{137.\text{ g M}^{2+}}{1\text{ mol M}^{2+}} \Rightarrow M = Ba$$

And a not-so-challenging example:

For the reaction shown, compute the theoretical yield of product (in grams) for each of the following initial amounts of reactants.

$$2\,Al\,(s) + 3\,Cl_2\,(g) \rightarrow 2\,AlCl_3\,(s)$$

a) 2.0 g Al, 2.0 g Cl_2

molar mass of Al = 26.98 g/mol
molar mass of Cl_2 = 70.90 g/mol

Al: $2.0\text{ g Al} \times \frac{1\text{ mol Al}}{26.98\text{ g Al}} \times \frac{2\text{ mol AlCl}_3}{2\text{ mol Al}} = 0.0741\text{ mol AlCl}_3$

Cl_2: $2.0\text{ g Cl}_2 \times \frac{1\text{ mol Cl}_2}{70.90\text{ g}} \times \frac{2\text{ mol AlCl}_3}{3\text{ mol Cl}_2} = 0.0188\text{ mol AlCl}_3$ ← smaller, so Cl_2 is limiting and the theoretical yield is 0.0188 mol

b) 7.5 g Al, 24.8 g Cl_2

Al: $7.5\text{ g Al} \times \frac{1\text{ mol Al}}{26.98\text{ g Al}} \times \frac{2\text{ mol AlCl}_3}{2\text{ mol Al}} = 0.278\text{ mol AlCl}_3$

Cl_2: $24.8\text{ g Cl}_2 \times \frac{1\text{ mol Cl}_2}{70.90\text{ g Cl}_2} \times \frac{2\text{ mol AlCl}_3}{3\text{ mol Cl}_2} = 0.233\text{ mol AlCl}_3$ ← smaller, so Cl_2 is limiting and the theoretical yield is 0.233 mol