

1. Vanillin, the dominant flavoring in vanilla, contains C, H, and O. When 1.05 g of this substance is completely combusted, 2.43 g of CO₂ and 0.50 g of H₂O are produced. What is the empirical formula of vanillin?

Yet another combustion analysis problem!

$$0.50 \text{ g H}_2\text{O} * \frac{1 \text{ mol H}_2\text{O}}{18.02 \text{ g H}_2\text{O}} * \frac{2 \text{ mol H}}{1 \text{ mol H}_2\text{O}} = 0.05549 \text{ mol H} * \frac{1.008 \text{ g H}}{1 \text{ mol H}} = 0.0559 \text{ g H}$$

$$2.43 \text{ g CO}_2 * \frac{1 \text{ mol CO}_2}{44.01 \text{ g CO}_2} * \frac{1 \text{ mol C}}{1 \text{ mol CO}_2} = 0.05521 \text{ mol C} * \frac{12.011 \text{ g C}}{1 \text{ mol C}} = 0.6632 \text{ g C}$$

$$\begin{aligned} \text{Mass O} &= \text{Mass sample} - \text{Mass C} - \text{Mass H} = 1.05 \text{ g} - 0.6632 \text{ g} - 0.0559 \text{ g} \\ &= 0.3309 \text{ g O} \end{aligned}$$

$$0.3309 \text{ g O} * \frac{1 \text{ mol O}}{15.9994 \text{ g O}} = 0.0207 \text{ mol O}$$

Take mole ratios to find empirical formula!

$$\text{C: } \frac{0.05521}{0.0207} = 2.67$$

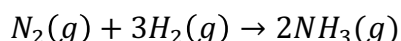
$$\text{H: } \frac{0.05549}{0.0207} = 2.67$$

$$\text{O: } \frac{0.0207}{0.0207} = 1$$

2.67 is equal to 8/3. Thus, multiply everything by 3 to get the empirical formula: C₈H₈O₃

2. A mixture of N₂(g) and H₂(g) reacts in a closed container to form ammonia, NH₃(g). The reaction ceases before either reactant has been totally consumed. At this stage, 3.0 mol N₂, 3.0 mol H₂, and 3.0 mol NH₃ are present. How many moles of N₂ and H₂ were present originally?

First write the balanced equation!



Since we assume that we started off with zero product (zero ammonia), let's use that information to find out how much reactant was needed to generate that much product.

$$3.0 \text{ mol } NH_3 * \frac{1 \text{ mol } N_2}{2 \text{ mol } NH_3} = 1.5 \text{ mol } N_2 \text{ needed}$$

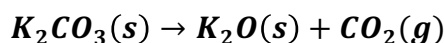
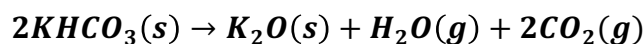
$$3.0 \text{ mol } NH_3 * \frac{3 \text{ mol } H_2}{2 \text{ mol } NH_3} = 4.5 \text{ mol } H_2 \text{ needed}$$

Thus, 1.5 mol N_2 and 4.5 mol H_2 were needed to bring the reaction to its current state, with 3 mol NH_3 . However, there are also current mole amounts of each of the reactants, as noted in the problem. Thus,

$$\text{Original } N_2 \text{ amount} = 1.5 \text{ mol} + 3.0 \text{ mol} = 4.5 \text{ mol } N_2$$

$$\text{Original } H_2 \text{ amount} = 4.5 \text{ mol} + 3.0 \text{ mol} = 7.5 \text{ mol } H_2$$

3. A mixture containing $KClO_3$, K_2CO_3 , $KHCO_3$, and KCl was heated, producing CO_2 , O_2 , and H_2O gases according to the following equations:



The KCl does not react under the conditions of the reaction. If 100.0 g of the mixture produces 1.80 g of H_2O , 13.20 g of CO_2 , and 4.00 g of O_2 , what was the composition of the original mixture? (Assume complete decomposition of the mixture.)

This is similar to the combustion analysis in some ways. Recognize that all of the O_2 produced is due to the first reaction, and all of the H_2O produced comes from the second reaction.

$$4.00 \text{ g } O_2 * \frac{1 \text{ mol } O_2}{31.9988 \text{ g } O_2} * \frac{2 \text{ mol } KClO_3}{3 \text{ mol } O_2} * \frac{122.55 \text{ g } KClO_3}{1 \text{ mol } KClO_3} = 10.21 \text{ g } KClO_3$$

$$1.80 \text{ g } H_2O * \frac{1 \text{ mol } H_2O}{18.02 \text{ g } H_2O} * \frac{2 \text{ mol } KHCO_3}{1 \text{ mol } H_2O} * \frac{100.12 \text{ g } KHCO_3}{1 \text{ mol } KHCO_3} = 20.00 \text{ g } KHCO_3$$

Next, we know how much CO_2 is produced, but that is produced in both the 2nd and 3rd reactions. So, since we just determined how much $KHCO_3$ was present, determine how much CO_2 that would have produced.

$$1.80 \text{ g } H_2O * \frac{1 \text{ mol } H_2O}{18.02 \text{ g } H_2O} * \frac{2 \text{ mol } CO_2}{1 \text{ mol } H_2O} * \frac{44.01 \text{ g } CO_2}{1 \text{ mol } CO_2} = 8.79 \text{ g } CO_2$$

Thus, the CO_2 produced from reaction 3 is:

$$13.20 \text{ g } CO_2 - 8.79 \text{ g } CO_2 = 4.41 \text{ g } CO_2$$

Then, this can be used to find how much K_2CO_3 was present in the original mixture.

$$4.41 \text{ g } CO_2 * \frac{1 \text{ mol } CO_2}{44.01 \text{ g } CO_2} * \frac{1 \text{ mol } K_2CO_3}{1 \text{ mol } CO_2} * \frac{138.21 \text{ g } K_2CO_3}{1 \text{ mol } K_2CO_3} = 13.84 \text{ g } K_2CO_3$$

$$\begin{aligned} \text{Mass } KCl &= \text{Mass sample} - \text{Mass } KClO_3 - \text{Mass } KHCO_3 - \text{Mass } K_2CO_3 \\ &= 100.0 - 10.21 - 20.00 - 13.84 = 56.0 \text{ g } KCl \end{aligned}$$

Composition:

$$\% KCl = \frac{56.0 \text{ g}}{100.0 \text{ g}} = 56.0\% KCl$$

$$\% KClO_3 = \frac{10.21 \text{ g}}{100.0 \text{ g}} = 10.2\% KClO_3$$

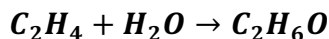
$$\% KHCO_3 = \frac{20.00 \text{ g}}{100.0 \text{ g}} = 20.0\% KClO_3$$

$$\% K_2CO_3 = \frac{13.84 \text{ g}}{100.0 \text{ g}} = 13.8\% K_2CO_3$$

4. Give the oxidation number of each element in the following compounds:

- a. BrO_3^- : O: -2; Br: +5
- b. H_2SO_4 : O: -2; S: +6; H: +1
- c. CrO_4^{2-} : O: -2; Cr: +6
- d. $LiAlH_4$: Li: +1; Al: +3; H: -1

5. Your friend has heard that she can make ethanol by reacting C_2H_4 with H_2O under acidic conditions, but she's not sure how much of each starting material she needs. So she randomly mixes 101.7 g of C_2H_4 with 55.19 g of H_2O .



- a. What is the theoretical yield of ethanol in mL (ethanol density = 0.789 g/mL)?

$$101.7 \text{ g } C_2H_4 * \frac{1 \text{ mol } C_2H_4}{28.05 \text{ g } C_2H_4} * \frac{1 \text{ mol } C_2H_6O}{1 \text{ mol } C_2H_4} = 3.626 \text{ mol } C_2H_6O$$

$$55.19 \text{ g } H_2O * \frac{1 \text{ mol } H_2O}{18.02 \text{ g } H_2O} * \frac{1 \text{ mol } C_2H_6O}{1 \text{ mol } H_2O} = 3.063 \text{ mol } C_2H_6O$$

Thus, H₂O is the limiting reactant

$$3.063 \text{ mol } C_2H_6O * \frac{46.07 \text{ g } C_2H_6O}{1 \text{ mol } C_2H_6O} * \frac{1 \text{ mL}}{0.789 \text{ g}} = 178.8 \text{ mL } C_2H_6O$$

b. How much (mass) excess reactant remains?

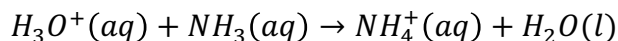
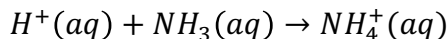
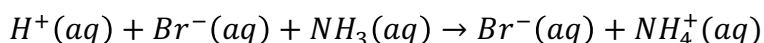
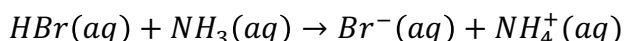
$$55.19 \text{ g } H_2O * \frac{1 \text{ mol } H_2O}{18.02 \text{ g } H_2O} * \frac{1 \text{ mol } C_2H_4}{1 \text{ mol } H_2O} = 3.063 \text{ mol } C_2H_4 \text{ used up}$$

$$101.7 \text{ g } C_2H_4 * \frac{1 \text{ mol } C_2H_4}{28.05 \text{ g } C_2H_4} = 3.626 \text{ mol } C_2H_4 \text{ initially}$$

$$3.626 \text{ mol} - 3.063 \text{ mol} = 0.563 \text{ mol } C_2H_4 \text{ remaining} * \frac{28.05 \text{ g}}{1 \text{ mol } C_2H_4} = 15.8 \text{ g } C_2H_4 \text{ remains}$$

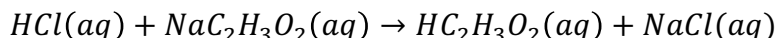
6. Complete and balance the following molecular equations, and then write the net ionic equation for each (note in past answer keys, we have written H⁺ in net ionic equations, but it is more correct to write H₃O⁺ instead. H⁺ doesn't actually exist itself in solution):

a. $HBr(aq) + NH_3(aq) \rightarrow$

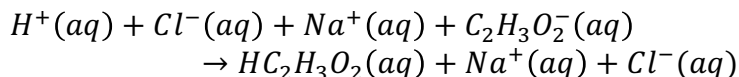


b. Aqueous hydrochloric acid and sodium acetate

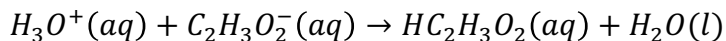
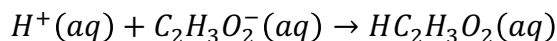
Overall balanced:



Total ionic:

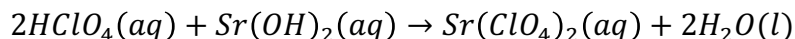


Net ionic:

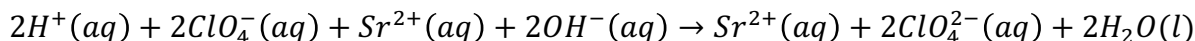


c. Aqueous perchloric acid and aqueous strontium hydroxide

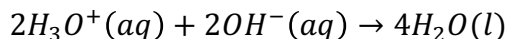
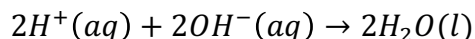
Overall balanced:



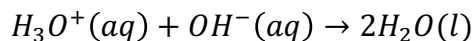
Total ionic:



Net ionic:



The above equation can be simplified to:



7. Starting with solid sucrose, $C_{12}H_{22}O_{11}$, describe how you would

a. Prepare 250 mL of a 0.250 M sucrose solution

$$0.250\text{ M} = \frac{0.250\text{ mol } C_{12}H_{22}O_{11}}{1\text{ L}} * 0.250\text{ L} = 0.0625\text{ mol } C_{12}H_{22}O_{11}$$

$$0.0625\text{ mol } C_{12}H_{22}O_{11} * \frac{342.3\text{ g } C_{12}H_{22}O_{11}}{1\text{ mol } C_{12}H_{22}O_{11}} = 21.39\text{ g } C_{12}H_{22}O_{11}$$

Measure out 21.39 g of sucrose, and add enough water to make 250 mL of solution.

b. Prepare 350.0 mL of 0.100 M $C_{12}H_{22}O_{11}$ starting with 3.00 L of 1.50 M $C_{12}H_{22}O_{11}$.

Determine how many moles of sucrose you need in the desired solution.

$$0.100\text{ M } C_{12}H_{22}O_{11} = \frac{0.100\text{ mol } C_{12}H_{22}O_{11}}{1\text{ L}} * 0.350\text{ L} = 0.035\text{ mol } C_{12}H_{22}O_{11}$$

You should check to make sure that you have at least this many moles of sucrose to begin with. Otherwise, you wouldn't have enough to prepare the desired solution.

$$1.50 \text{ M } C_{12}H_{22}O_{11} = \frac{1.50 \text{ mol } C_{12}H_{22}O_{11}}{1 \text{ L}} * 3.00 \text{ L} = 4.50 \text{ mol } C_{12}H_{22}O_{11}$$

This confirms that we have enough moles of sucrose to make the desired solution. We need 0.035 mol of $C_{12}H_{22}O_{11}$.

$$0.035 \text{ mol } C_{12}H_{22}O_{11} * \frac{1 \text{ L}}{1.50 \text{ mol } C_{12}H_{22}O_{11}} = 0.0233 \text{ L}$$

So, we need 23.3 mL of the 1.50 M solution, and then the rest of the 350.0 mL volume should be water. The final solution volume is 350.0 mL.