

1. Some fundamental calculations:

- a. Calculate the number of oxygen atoms in a 288 amu sample.

$$288 \text{ amu} \times \frac{1 \text{ atom O}}{16.00 \text{ amu}} = 18 \text{ atoms O}$$

- b. Calculate the mass (in amu) of 51 Na atoms.

$$51 \text{ atoms Na} \times \frac{22.99 \text{ amu}}{1 \text{ atom Na}} = 1173 \text{ amu}$$

- c. Calculate the number of moles in a sample with
- 5.00×10^{20}
- atoms of Cr.

$$5.00 \times 10^{20} \text{ atoms Cr} \times \frac{1 \text{ mol}}{6.022 \times 10^{23} \text{ atom}} = 8.30 \times 10^{-4} \text{ mol Cr}$$

- d. Calculate the mass of the sample in part (c).

$$5.00 \times 10^{20} \text{ atoms Cr} \times \frac{1 \text{ mol}}{6.022 \times 10^{23} \text{ atom}} \times \frac{52.00 \text{ g Cr}}{1 \text{ mol Cr}} = 4.32 \times 10^{-2} \text{ g Cr}$$

- e. How many atoms of Si are in a 5.68 mg sample?

$$5.68 \text{ mg Si} \times \frac{1 \text{ g}}{1000 \text{ mg}} \times \frac{1 \text{ mol Si}}{28.09 \text{ g Si}} \times \frac{6.022 \times 10^{23} \text{ atom}}{1 \text{ mol}} = 1.22 \times 10^{20} \text{ atoms Si}$$

- f. How many grams of aluminum sulfate are in a 0.630 mol sample?

$$0.630 \text{ mol Al}_2(\text{SO}_4)_3 \times \frac{342.14 \text{ g Al}_2(\text{SO}_4)_3}{1 \text{ mol Al}_2(\text{SO}_4)_3} = 216 \text{ g Al}_2(\text{SO}_4)_3$$

- g. How many moles are in 50.0 g sample of ammonium carbonate?

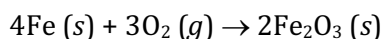
$$50.0 \text{ g (NH}_4)_2\text{CO}_3 \times \frac{1 \text{ mol (NH}_4)_2\text{CO}_3}{96.094 \text{ g (NH}_4)_2\text{CO}_3} = 0.520 \text{ g (NH}_4)_2\text{CO}_3$$

- h. What is the mass of one molecule of dinitrogen tetroxide?

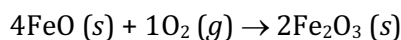
$$1 \text{ molecule N}_2\text{O}_4 \times \frac{1 \text{ mol}}{6.022 \times 10^{23} \text{ molecules}} \times \frac{92.02 \text{ g N}_2\text{O}_4}{1 \text{ mol N}_2\text{O}_4} = 1.528 \times 10^{-22} \text{ g N}_2\text{O}_4$$

2. Translate the following descriptions into a balanced chemical equation.

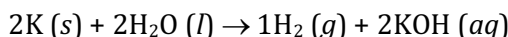
- a. Solid iron metal reacts with oxygen gas to produce solid iron (III) oxide.



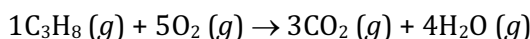
- b. The combustion of solid iron (II) oxide produces solid iron (III) oxide.



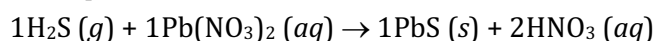
- c. Solid potassium metal reacts with water to make hydrogen gas and aqueous potassium hydroxide.



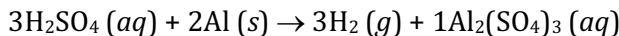
- d. Propane gas combusts to give off carbon dioxide and water vapor.



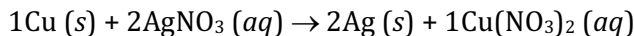
- e. Dihydrogen sulfide gas is bubbled through an aqueous solution of lead (II) nitrate and solid lead (II) sulfide forms alongside aqueous nitric acid.



- f. Sulfuric acid is poured onto solid aluminum to give off hydrogen gas and a solution of aluminum sulfate.



- g. A copper wire dipped in a solution of silver (I) nitrate produces silver metal and copper (II) nitrate solution.



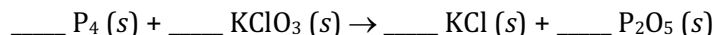
3. If you have equal masses of each compound (H_2SO_4 , $\text{C}_{12}\text{H}_{22}\text{O}_{11}$, KClO_3), which sample has the greatest number of oxygen atoms?

(Note: There are two ways you can do this problem!)

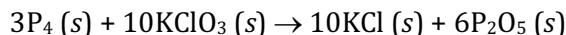
Method 1: Determine mass percent of O in each. $\rightarrow \text{H}_2\text{SO}_4$: 65 % O | $\text{C}_{12}\text{H}_{22}\text{O}_{11}$: 51% O | KClO_3 : 39% O

Method 2 (much longer): Determine number of O atoms in each compound.

4. Striking a match results in the following unbalanced chemical reaction:



- a. Balance the equation above.



- b. If 15.0 mg of P_2O_5 was produced in this reaction, what masses of P_4 and KClO_3 were required?

Start by converting 15.0 mg of P_2O_5 to mol of P_2O_5 :

$$15.0 \text{ mg } \text{P}_2\text{O}_5 \times \frac{1 \text{ mg}}{1000 \text{ mg}} \times \frac{1 \text{ mol } \text{P}_2\text{O}_5}{141.94 \text{ g } \text{P}_2\text{O}_5} = 1.057 \times 10^{-4} \text{ mol } \text{P}_2\text{O}_5$$

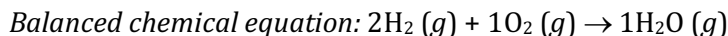
Now determine how much P_4 is needed to produce this much P_2O_5 :

$$1.057 \times 10^{-4} \text{ mol } \text{P}_2\text{O}_5 \times \frac{3 \text{ mol } \text{P}_4}{6 \text{ mol } \text{P}_2\text{O}_5} \times \frac{123.88 \text{ g } \text{P}_4}{1 \text{ mol } \text{P}_4} = 6.55 \times 10^{-3} \text{ g } \text{P}_4$$

Now determine how much KClO_3 is needed to produce this much P_2O_5 :

$$1.057 \times 10^{-4} \text{ mol } \text{P}_2\text{O}_5 \times \frac{10 \text{ mol } \text{KClO}_3}{6 \text{ mol } \text{P}_2\text{O}_5} \times \frac{122.55 \text{ g } \text{KClO}_3}{1 \text{ mol } \text{KClO}_3} = 2.16 \times 10^{-2} \text{ g } \text{KClO}_3$$

5. You react 10.0 g of hydrogen gas with 60.0 g of oxygen gas to form water. Determine the amount of water formed and the amount of excess reactant (both in grams) after the reaction is complete.



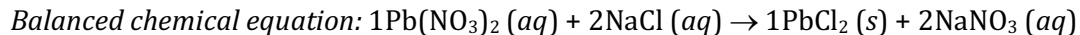
Now determine whether the limiting reactant is H_2 or O_2 gas (two ways):

<p><u>Method 1:</u> <i>How much H_2O can we make from all H_2?</i> $10.0 \text{ g } \text{H}_2 \times \frac{1 \text{ mol } \text{H}_2}{2.016 \text{ g } \text{H}_2} \times \frac{2 \text{ mol } \text{H}_2\text{O}}{2 \text{ mol } \text{H}_2} = 4.96 \text{ mol } \text{H}_2\text{O}$ <i>How much H_2O can we make from all O_2?</i> $60.0 \text{ g } \text{O}_2 \times \frac{1 \text{ mol } \text{O}_2}{32.00 \text{ g } \text{O}_2} \times \frac{2 \text{ mol } \text{H}_2\text{O}}{1 \text{ mol } \text{O}_2} = 3.75 \text{ mol } \text{H}_2\text{O}$ <i>Therefore, O_2 is limiting (makes less H_2O)!</i></p>	<p><u>Method 2:</u> <i>We have 4.96 mol H_2 and 1.875 mol O_2. How much H_2 do we need to react with all of the O_2?</i> $60.0 \text{ g } \text{O}_2 \times \frac{1 \text{ mol } \text{O}_2}{32.00 \text{ g } \text{O}_2} \times \frac{2 \text{ mol } \text{H}_2}{1 \text{ mol } \text{O}_2} = 3.75 \text{ mol } \text{H}_2$ <i>We <u>need</u> 3.75 mol H_2, but <u>have</u> 4.96 mol H_2, so H_2 is in excess.</i> <i>Therefore, O_2 is limiting!</i></p>
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H_2 is in excess and we need to use 3.75 mol. Leftover = $4.96 \text{ mol} - 3.75 \text{ mol} = 1.21 \text{ mol } \text{H}_2 = 2.44 \text{ g } \text{H}_2$.

From method 1, we know we make 3.75 mol H_2O or 67.6 g H_2O .

6. Consider mixing an excess of lead (II) nitrate (aq) with 0.0800 mol sodium chloride. Determine the mass of solid lead (II) chloride formed assuming a complete reaction.



Since we are told $\text{Pb}(\text{NO}_3)_2$ is in excess, the NaCl is the limiting reactant.

$$0.0800 \text{ mol NaCl} \times \frac{1 \text{ mol PbCl}_2}{2 \text{ mol NaCl}} \times \frac{278.10 \text{ g PbCl}_2}{1 \text{ mol PbCl}_2} = 11.1 \text{ g PbCl}_2$$

7. The percent by mass of nitrogen is 46.7% for a species containing only nitrogen and oxygen. Which of the following could this species be?



Assuming a 100 g sample of N_xO_y compound: 46.7 g N and 53.3 g O

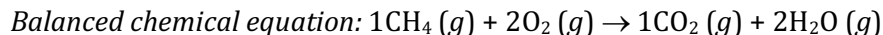
Determine the empirical formula:

$$46.7 \text{ g N} \times \frac{1 \text{ mol N}}{14.01 \text{ g N}} = 3.333 \text{ mol N} \Rightarrow \frac{3.333 \text{ mol N}}{3.331} = 1 \text{ mol N}$$

$$53.3 \text{ g O} \times \frac{1 \text{ mol O}}{16.00 \text{ g O}} = 3.331 \text{ mol O} \Rightarrow \frac{3.331 \text{ mol O}}{3.331} = 1 \text{ mol O}$$

Empirical formula: NO

8. If 5.00 g of CH_4 is burned, what mass of water can be produced?



$$5.00 \text{ g CH}_4 \times \frac{1 \text{ mol CH}_4}{16.04 \text{ g CH}_4} \times \frac{2 \text{ mol H}_2\text{O}}{1 \text{ mol CH}_4} \times \frac{18.02 \text{ g H}_2\text{O}}{1 \text{ mol H}_2\text{O}} = 11.2 \text{ g H}_2\text{O}$$