

Quantitative Stoichiometry

<u>Reading:</u> Ch 3 sections 7 – 9 Ch 4 sections 1 – 4	<u>Homework:</u> Chapter 3: 55*, 57*, 59*, 63, 65*, 67, 75, 77, 79*, 81, 83 Chapter 4: 25*, 27, 29, 31, 33*, 37*, 41*, 43, 47
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* = 'important' homework question

Formula Weight

Question: What is the *average* mass (in amu) of 1 molecule of water (H₂O)?
Where would you find the required information?

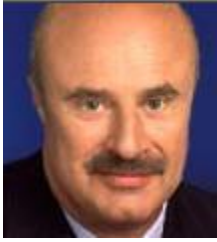


**A molecule can be considered the 'sum of its parts' (atoms).
Thus, simply add the masses of all the atoms in a single molecule
to find the molecule's mass (formula weight) in amu**

Determine formula weights for the following molecular materials:

<u>Name</u>	<u>Molecular Formula</u>	<u>Formula Weight (amu per molecule)</u>
Carbon dioxide		
	H ₂ SO ₄	
Diphosphorous pentoxide		

Percent Composition from Formulas



Analogy: If Dr. Phil weighs 250 pounds and his head alone weighs 25 pounds, then what % by mass (% composition or mass percent) is Dr. Phil's head of his entire body?

Dr. Phil – a man with a relatively large head

$$\% \text{ element} = \frac{\text{(number of atoms of element)}(\text{atomic mass of element})}{\text{Formula weight of molecule}} \times 100\%$$

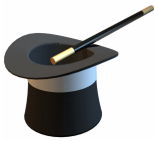
Question: If one molecule of water has a mass of 18.02 amu, then what fraction of a water molecule's mass is due to oxygen? What is this value as a percent?

Task: Determine the following:

1. The % mass of C and O in carbon dioxide

2. The % mass of C, H and O in formaldehyde (CH₂O).

The Mole



Just like the dozen, *the MOLE is just a number that represents a bigger number*. Since atoms and/or molecules are very small (i.e. to see a collection of atoms, say in your hand, you need a lot of them), the mole is a **VERY large number** :

1 dozen = 12 things (eggs) 1 mole = 6.02×10^{23} things (e.g. atoms)

Task: To get an idea about how many atoms there are in a mole of atoms, write 6.02×10^{23} as a regular number:

1 mole = _____

Note: The mole is sometimes called **Avogadro's number (N_A)**, so:



1 mole = $N_A = 6.02 \times 10^{23}$ things

Nerd stuff: When do you think *some* chemists celebrate mole day?

Discussion: Why is the mole ($N_A = 6.02 \times 10^{23}$) such a 'strange' number? Why not just 1×10^{20} or something?



&





Proof: Determine how many carbon atoms (12.011 amu/atom) would weigh 12.011 grams. This number is, by definition, a mole (N_A) of C atoms ($1 \text{ amu} = 1.661 \times 10^{-24} \text{ g}$).



Convert between grams, moles and number of atoms SEPARATELY using atomic mass values (Periodic table) and Avogadro's number respectively.

Grams ↔ Moles ↔ Number of atoms

Task: Use the periodic table and the above information discussed above, determine the following quantities:

1. The mass of 2.0 moles of carbon atoms
2. The number of moles of lead (Pb) atoms in 35.5 grams of lead.
3. The number of Pb atoms in 35.5 grams of lead



Recall that a molecule is the ‘sum of its parts’ (atoms). Thus, simply add the masses of all the atoms in a molecule to find the molecule’s FW in amu, *OR* the mass of 1 mole of molecules (MOLAR MASS, \mathcal{M}) in grams. THIS IS THE POWER OF THE MOLE(!)

Proof: What is the mass in amu of 1 molecule of water (H_2O). What is the mass of one *mole* of water molecules in grams? ($1 \text{ amu} = 1.661 \times 10^{-24} \text{ g}$)

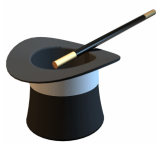
Note: The units of Molar mass (\mathcal{M}) are grams/mole (i.e. the number of grams in 1 mole of material).

Task: Calculate the Molar masses (\mathcal{M}) of the following compounds:

Carbon dioxide:

Diphosphorous pentoxide:

Calcium chloride:



Just like with atomic masses, Molar masses can be used to convert between grams, moles and number of molecules.

Grams ↔ Moles ↔ Number of molecules

Worked Example: How many molecules of sugar ($C_6H_{12}O_6$) are there in a 2.15 gram packet of sugar?

Plan: Write down what you are given and what you can immediately figure out:

Mass sugar = 2.15 g

$$\begin{aligned} M_{C_6H_{12}O_6} &= 6C + 12H + 6O \\ &= 6(12.011 \text{ g/mol}) + 12(1.01 \text{ g/mol}) + 6(16.00 \text{ g/mol}) \\ &= 180.2 \text{ g/mol} \end{aligned}$$

Use the conversion factor of $1 \text{ mol } C_6H_{12}O_6 = 180.2 \text{ g}$ to find # moles sugar in 2.15 g of sugar:

$$2.15 \text{ g} \times \frac{1 \text{ mol}}{180.2 \text{ g}} = \underline{0.0119 \text{ mols}}$$

Use the conversion factor of $1 \text{ mol} = 6.02 \times 10^{23} \text{ molecules}$ to find # molecules of sugar in 0.0119 mols. (2.15 g) of sugar:

$$0.0119 \text{ mol} \times \frac{6.02 \times 10^{23} \text{ molecules}}{1 \text{ mol}} = \underline{7.16 \times 10^{21} \text{ molecules}}$$



Remember: You CANNOT convert directly from grams to number of molecules. i.e. YOU MUST ALWAYS GO THROUGH MOLES:

Grams ↔ Moles ↔ Number of molecules

Things to remember:



Key relationships:

FW = sum of all atomic masses in a SINGLE molecular formula (amu/molecule)

\mathcal{M} = sum of all atomic masses in any SINGLE molecular or ionic formula (grams/mole)

$$\mathcal{M} = \frac{\text{number grams material}}{\text{number moles of material}}$$

1 mole = 6.02×10^{23} particles

FW and \mathcal{M} have identical numerical values but DIFFERENT units

You can write a conversion pyramid showing the relationship between \mathcal{M} , g and mols!



The relationship between \mathcal{M} , #grams and #moles is THE most frequently used equations in chemistry? Why?



Answer:



Observation(!):

Task: Determine the following quantities:

1. The number of molecules in 5.0 g of oxygen gas
2. The number of molecules in 12.5 grams of P_2O_5 (s)
3. The mass in grams of 2.76×10^{23} molecules of H_2O (l)

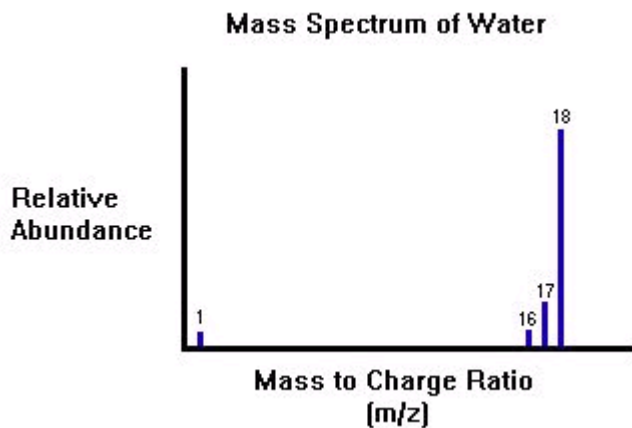
Empirical & Molecular Formulas from Analyses

Recall: Mass Spectrometry

Scientists are able to determine the mass of *individual* atoms and/or molecules using a *mass spectrometer*. The resulting mass spectrum gives a parent peak (M^+), which gives (is) the *formula weight* (and, therefore, *molecular weight*) of the specie under study.

Ever wondered why the masses of individual isotopes for any stable *element* or *molecule* are known to a large number of sig. figs? They are measured using a very accurate / precise mass spectrometer.

Extra Credit (to be completed outside of class): Review the mass spec. section on page 68 of your text *once more*. Use this information to determine the identities of the molecular fragments giving rise to each peak in the 'mass spec' of water.



The Mass Spectrum recorded for water

Hand in a concise, typed report next lecture for 2 pts e.c.

Determining Empirical and Molecular Formulas 'by hand'

Recall the definitions of *molecular formula* and *empirical formula*:



Molecular Formula: the actual number and type of atoms in a compound, e.g. hydrogen peroxide = H_2O_2

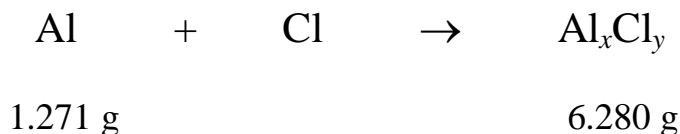
Empirical Formula: the lowest whole number ratio of each type of atom in a compound e.g. hydrogen peroxide = HO



Recall that the Empirical formula is the lowest *whole number* ratio of each type of atom in a compound \Rightarrow **find the moles of each type of atom and then find their ratio.**

Worked Example: A 1.271 g sample of Al(s) was allowed to react with chlorine. The mass of aluminum chloride produced was 6.280 g. Determine the empirical formula of aluminum chloride.

Step 1. Write an *unbalanced* chemical equation (do not assume balancing numbers or formulas for these problems). Find the mass of the missing reactant by applying the *conservation of mass law*.



Step 2. Find the *moles* of each reactant using the atomic masses from the periodic table.

Moles Al =

Moles Cl =

Step 3. Substitute the # moles determined for each type of atom in the product's empirical formula.

Al Cl

Step 4. Find the lowest whole number ratio of each type of atom in the empirical formula. This is the final answer.

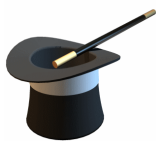
Al Cl

Finding the Molecular Formula from the Empirical Formula



Recall that the molecular formula is some whole number of times larger than the empirical formula (e.g. H_2O_2 compared to HO (x2)).

\Rightarrow the molecular formula will be 'heavier' than the empirical formula by the same factor (x2)



Mass spectroscopy data (supplied) will give the Formula Weight (FW) of the element or compound under study. This value is then compared with the empirical formula weight (EW) to find the molecular formula

Task: Work out the molecular masses of H_2O_2 and HO . What is their ratio?

$M_{\text{H}_2\text{O}_2} =$

$M_{\text{HO}} =$

Ratio =

Find the ratio of the molecular formula to the empirical formula – this information tells you how much 'bigger' the molecular formula is than the empirical formula.



Worked Example: Ethylene glycol contains 38.7 % C, 9.7 % H and 51.6 % O. Calculate the empirical and molecular formula of ethylene glycol given its molar mass = 60 g/mol (60 amu/molecule)



When given the % by mass values of each atom in a compound assume a 100 g sample – the % and g values are then the same.

‘Slides and Ladders’ – Finding Theoretical Yield and % Yield

Important Definitions

Theoretical Yield: The amount of product, in grams, expected (calculated) for a reaction.

Actual Yield: The amount of product, in grams, recovered (weighed) for a reaction.

% Yield :
$$\% \text{ Yield} = \frac{\text{Actual Yield}}{\text{Theoretical Yield}} \times 100\%$$

Discussion: Are theoretical and actual yields ever the same (i.e. does % yield = 100%) in practice? What factors influence the % yield?

Finding the Theoretical Yield using ‘Slides and Ladders’



Worked Example: If 5.00 g of Propane ($\text{C}_3\text{H}_8(\text{l})$) is combusted in excess oxygen gas, what mass of water is expected to be formed? What mass and volume of CO_2 (g) (at STP) would you expect to collect?

Step 1. Write a balanced Chemical Equation *and*

Step 2. Set up g, \mathcal{M} and mole 'ladder' grid

Step 3. Fill in the ladder grid with as much information as possible – this is typically supplied gram weights and molar mass data.

Step 4. Convert g \rightarrow moles by 'climbing' down ladder(s) ($\text{g} / \mathcal{M} = \text{moles}$).



Remember: Molar masses are calculated for ONE molecular formula only. I.E. ignore any balancing numbers when figuring out \mathcal{M} values

Step 5. Convert moles reactant \rightarrow moles product(s) by comparing balancing numbers and 'sliding' across.

Step 6. Convert moles product(s) \rightarrow grams product(s) by 'climbing' up ladder(s) ($\text{moles} \times \mathcal{M} = \text{grams}$).

Note: moles of gas can be converted to Liters of gas using:

1 mole any gas = 22.4 L at STP

Group Task: Determine the following quantities

1. What mass of dissolved HCl is needed to completely react 5.00g of CaCO₃(s), according to the following *unbalanced* reaction? What volume of CO₂(g) is generated at STP?

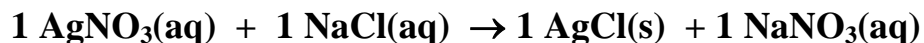


2. What mass of magnesium oxide is recovered when 1.56 g of Mg(s) is burnt in air to give MgO(s)? What volume of oxygen gas is consumed during this process (assume STP).
3. A student recovers 1.59 g of CaCO₃ (s) from an experiment when they should have produced 2.32 g. What is the student's % yield for their reaction?
4. Let's talk about your 'Alum' lab.....

'Slides and Ladders' – Limiting Reactant Problems



Recall: Reactants ALWAYS combine in the ratio defined by their respective balancing numbers:



i.e. 1 mole of $\text{AgNO}_3(\text{aq})$ will react exactly with 1 mole of $\text{NaCl}(\text{aq})$

Problem: It is VERY difficult to add exactly the right ratio of reactants in the lab \Rightarrow

There will be too *much* of one reactant (the excess (XS) reactant)

There will be too *little* of one reactant (the limiting reactant)

Discussion questions: If 1.0 moles of $\text{AgNO}_3(\text{aq})$ is added to 1.5 moles of $\text{NaCl}(\text{aq})$, then:

1. Which reactant is INXS?
2. Which reactant is limiting?
3. How many moles of $\text{AgCl}(\text{s})$ would be formed?
4. How many moles of $\text{NaCl}(\text{aq})$ would remain unreacted?



Compare the 'Ideal' (from the balanced equation) and the 'Real' (given) ratio of reactants to determine which is the limiting reactant ($\text{AgNO}_3(\text{aq})$)



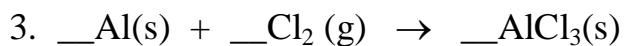
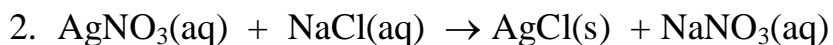
Since the LIMITING reactant will run out first, it determines the amount of product that can be formed (as well as the amount of XS reactant that is left behind)

Questions: Work out the mass of each product formed in the following reactions, assuming 10.0 grams of each reactant are initially mixed together.



Use a regular slides and ladders approach, but ‘slide’ across (to find moles of product) using the molar ratio determined by the limiting reactant.

Use conversions factors to find out the number of moles required to react with one of the reactants in ‘NON 1:1 problems’





Try the following questions (taken from the first and second practice midterms).

“% Mass”

Calculate the % mass of *each* type of atom in the following materials:



Ba = 58.8%

S = 13.8 %

O = 27.4%

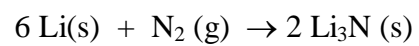


N = 46.7%

O = 53.3%

“Limiting”

Lithium and nitrogen react to produce lithium nitride:



If 7.00 grams of each reactant undergo a reaction with 85.0 % yield, how many grams of Li_3N are obtained from the reaction?

[10.0g]

“Take out?”

Monosodium glutamate (MSG) is a popular flavor enhancer used in many fast foods. Assuming MSG contains 35.51% C, 4.77% H, 37.85% O, 8.29% N, and 13.6% Na, and has a molar mass of 169 g/mol determine:

The Empirical formula of MSG



The Molecular formula of MSG

