

A dozen is a convenient way for bakers to talk of buns, whereas a mole is a convenient way for chemists to talk of atoms, molecules, or particles, as one mole of particles would fit in the palm of your hand.

2. Molar Mass



The Mole

- **By definition, 1 mole is the number of atoms in 12.0 g of Carbon-12**
- **so 12.0 g of carbon contains approx. $6.02214179... \times 10^{23}$ carbon atoms and is defined as 1 mole of carbon**
- **1 mole is further approximated to three sig figs: 6.02×10^{23}**

Mass Numbers on the Periodic Table take on two meanings

- **Meaning #1:**
- **Atomic Mass: found by adding the number of _____ with the number of _____ – units are atomic mass units (a.m.u.)**
- **A carbon-12 atom has _____ protons and _____ neutrons and a mass of _____ a.m.u.**
- *Recall that masses on the p.t. are not whole numbers due to weighted _____ of isotopes*

- **Meaning #2:**
- **Molar Mass:** the mass in grams of 1 mole of that element (units are)
- Avogadro figured out how many atoms were in 12.0g of carbon, and called it '1 mole' of carbon
- Then, the mass of one mole of hydrogen atoms is 1.0g because each hydrogen atom has 1/12 the mass of a carbon atom, and so on
- So, the number for **atomic** and **molar** masses is the same, but their definitions and their units are different

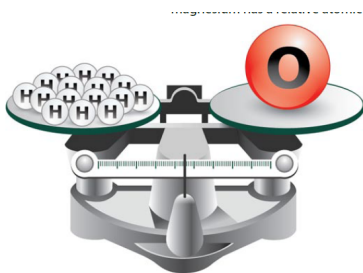


Figure 3.1c The mass of an oxygen atom is equal to the mass of 16 hydrogen atoms.

Molar Mass of Elements

- 1 mol of C is and 1 mol of Fe is
- this is because a C atom has mass than an Fe atom
- a C atom has 12p and 12-14n, whereas an Fe atom has 26p, and 28-34n, so an Fe atom is much heavier
- thus a mole of Fe atoms would be much heavier than a mole of C atoms

Molar Mass of a Compound

- The total mass of all atoms in the molecule
 - Multiply the mass of each element by the subscript on that element in the formula
 - E.g. Molar mass for H₂O
-
- so, 1 mole of water (6.02 x 10²³ molecules) has a mass of

Molar Mass of a Compound

- The same rules apply for diatomic molecules such as
- 1 mole of oxygen gas:

- H_2SO_4
-
- $\text{Cu}(\text{NO}_3)_2$
-
- Iron II sulfate trihydrate
-

Calculate the Molar Mass

- Ne
- just look at periodic table g/mol
- **Round all molar masses to the nearest TENTH (one decimal place)**
- SF_6
-
- $\text{C}_2\text{H}_5\text{OH}$
-

HOMEWORK:

Mole Problems #0
Worksheet - Part 1



3. MOLE CALCULATIONS

Moles to Mass

- Find the mass of one mole of NaCl
-
- Find the mass of two moles of NaCl
-
- How did you get the answer?

- What is the mass of 0.500 moles of NaCl?

•

- Find the mass of 0.35 moles of NaCl.

•



- Notice the mole unit cancels (one on top, one on the bottom), leaving only the gram unit!

- Using a table:

- | | |
|-----------------------|---------------------------------|
| <u>starting value</u> | <u>unit we want for answer</u> |
| | unit we have for starting value |
-

- Top left and bottom right units cancel, so we are left with unit we want in answer (the top right unit)

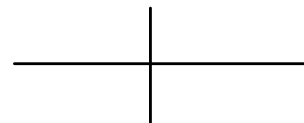
- Do all calculations for moles using a table as it will be very helpful for more complicated problems.
- Calculate the mass of 0.25mol of MgBr_2 .
- Molar mass $\text{MgBr}_2 =$



Calculate the mass of each

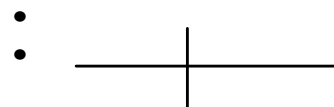
- 1.20 mol of HNO_3
- 0.60 mol of acetic acid

- Calculate the mass of 2.40 mol of ammonium carbonate.



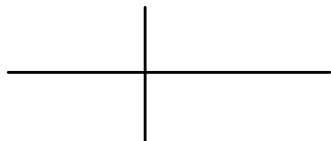
Mass to Moles

- How many moles are in 100.0g of CaS ?
-



- Note: Molar mass conversion can be written as 72.2g in one mole (72.2g/mol) or that one mole has a mass of 72.2g (1 mol/72.2g)

- How many moles are in 22.5g of aluminum oxide?



HOMEWORK: Mole Problems #0 Worksheet

- Calculate the number of moles in each of the following:
 - 50.0g of copper (II) oxide
 - 24mg of NaCl (Hint: change to g first)

**Finding Molar Mass Using
Experimental Data**

- If you do not know the chemical formula of a substance, its molar mass can be found if you know the mass of a sample and the number of moles it contains.

Example

- If a 0.35 mole sample of a compound has a mass of 44.4 g, what is its molar mass?
- Molar mass = $\frac{\text{grams}}{\text{moles}}$
- If you know this compound is either FeCl_3 or FeCl_2 you can compare molar masses with these compounds and find a match.
 - $\text{FeCl}_3 =$
 - $\text{FeCl}_2 =$
 -
- The unknown compound is

- How many molecules are in 0.500 mole of NaCl ?
-
- How many molecules are in 0.500 mole of CuS ?
-
-
- Keep in mind that Avogadro's number is **independent** of the type of molecule or object we are talking about.
 - Mass is **dependent** because every type of atom has a different mass!

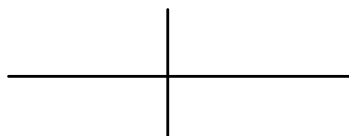
Moles to Molecules

- We must use Avogadro's number
- There are 6.02×10^{23} particles in 1 mole
- How many molecules are in 1.00 mole of NaCl ?
-
- How many molecules are in 2.00 mole of NaCl ?

Name	Equivalence Statement	Conversion Factors	
Avogadro's number	$1 \text{ mol} = 6.02 \times 10^{23} \text{ items}$	$\frac{6.02 \times 10^{23} \text{ items}}{1 \text{ mol}}$	$\frac{1 \text{ mol}}{6.02 \times 10^{23} \text{ items}}$

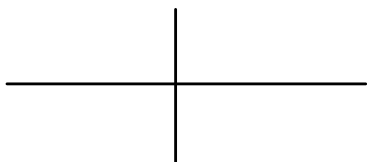
- Use the TABLE for all questions!
- How many molecules are in 1.20 mol of PbSO_4 ?
-
-

- How many molecules in 1.30 mol of potassium chloride?



Molecules to Moles

- How many moles in 5.42×10^{22} molecules of Na_2S ?



Exercises

- Calculate the number of molecules in each of the following:
 - 0.75 mol H_2SO_4
 - 2.35×10^{-1} mol $\text{Ca}(\text{OH})_2$

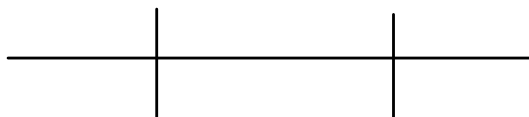
Exercise

- How many moles in 3.21×10^{23} molecules of KBr ?

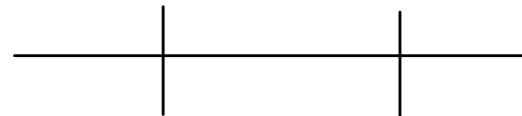
Moles to Atoms

- If we are dealing with an element, do moles to atoms just as you did moles to molecules, because an element is made up of singular atoms (except the 7 diatomics)
- If we are dealing with compounds, we must do a 'moles to molecules' step and then a 'molecules to atoms' step

- How many atoms are in 1.50 mol of calcium carbonate?



- How many atoms are in 0.10mol of H_2SO_4
- 1st -How many molecules are in 0.10 mol of H_2SO_4 ?
- 2nd -How many atoms in one molecule of H_2SO_4 ?

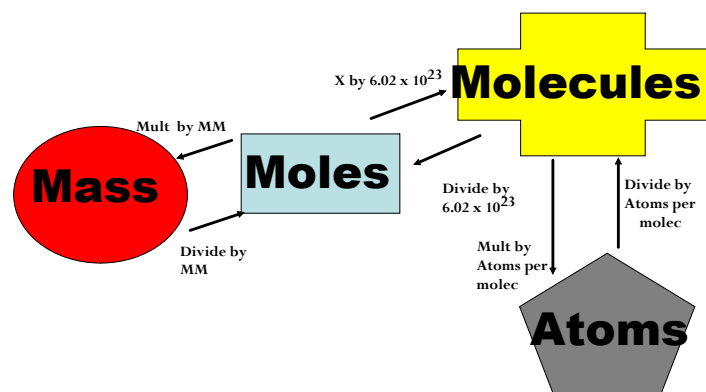


- Note: the number of atoms in a molecule is a counting number, not a measurement, so sig figs don't apply!

Exercise

- Determine the number of oxygen atoms in 0.65 mol $\text{Mg}(\text{OH})_2$.

MOLE MAP



HOMEWORK:

Mole Problems #1 - Parts 1 & 2

Multi Step Mole Problems

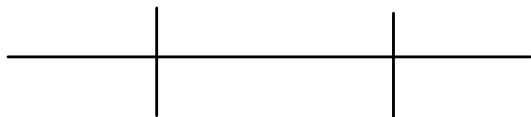
- Many times problems will involve more than one conversion in the same question.
- These problems are solved by using multiple conversion factors.
- Pay attention to cancelling units to aid in putting the numbers in the correct positions

Mass to Molecules

- How many molecules are present in 60.0g of NaOH?
- Find the molar mass of NaOH and set up your table

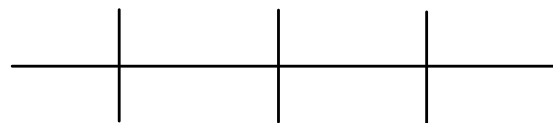
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- How many molecules are present in 5.00g of magnesium sulphate?



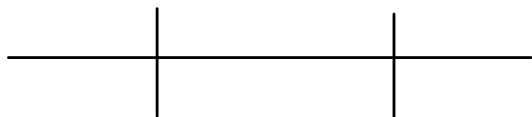
Mass to Atoms

- How many atoms would be present in 4.50g of copper (II) chloride
- Find the formula and MM for copper (II) chloride



Molecules to Mass

- What is the mass of 1.2×10^4 molecules of copper (II) chloride?



Exercise

- How many carbon atoms would be present in 795.0mg of acetic acid?

HOMEWORK:

Mole Problems #1

- Parts 3 & 4

Moles and Volumes of Gases

- 1.00 mole of any gas will occupy a volume of 22.4 L at STP
-
- STP = Standard Temperature and Pressure
- Standard Temperature = 0 degrees C
- Standard Pressure = 1 atmosphere

Moles to Volume

- How much space does 0.563 moles of H₂ gas take up at STP?

Volume to Moles

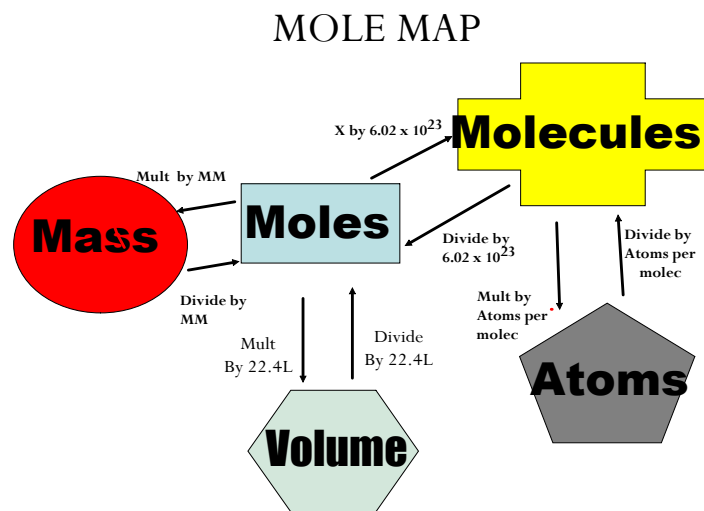
- How many moles in 18.7L of CH₄ gas (methane) at STP?

Mass to Volume

- What would be the volume of 90.0g of H₂ gas at STP?

Molecules to Volume

- What volume (in litres) is occupied by 8.14×10^{22} molecules of CO₂ gas (at STP)?

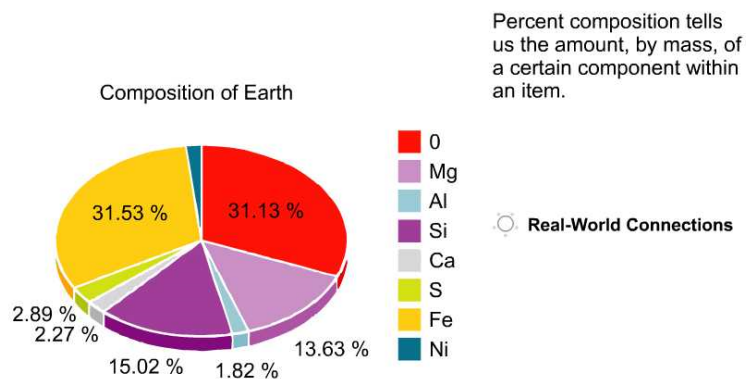


HOMEWORK:

Mole Problems #1 - Part 5

Mole Problems #2 Worksheet

4. Percent Composition



What is it?

- The percentage (by mass) of each element in a chemical formula
- AND
- The percentage (by mass) of each component in a mixture



From a Chemical Formula:

$$\% \text{ Composition of X} = \frac{\text{Mass of X}}{\text{Total mass of sample}}$$

- It doesn't matter how much of a substance we have, the % composition stays the same so we can assume we have 1 mole.
- % composition is an intensive property.

Find the % composition of Fe_2O_3

- Molar mass of Fe =
- Molar mass of O =
- Molar mass of Fe_2O_3 =

- Mass Fe =
-
-
- % Fe =

- Mass of O =
-
- % O =
-

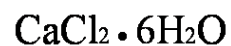
Fe₂O₃ is iron and oxygen
by mass

What is the % composition of C₆H₁₂O₆

From experimental data:

- A sample of a gas compound is analyzed and found to contain 23.5 g nitrogen and 53.9 g oxygen. What is the % composition?

What is the percent water in calcium chloride hexahydrate ?



- Total mass of mixture

% Composition of a Mixture

- Brass is made by melting together copper, zinc and other metals
- What is the percent composition of the alloy if the starting materials are 350 kg copper, 180 kg zinc and 25 kg tin?

• % copper =

• % zinc =

• % tin =

HOMEWORK:

Percent Composition
Worksheet

5. Empirical Formulas



What is Empirical Formula?

- The formula of a compound determined from experimental data (mass or percent)
- Gives whole number ratio of elements
- Simplest or “reduced” form of the molecular formula

Examples

- Molecular Formula: C_2H_6
- Empirical Formula: CH_3
-
- Molecular Formula: $C_6H_{12}O_6$
- Empirical Formula:
-
- Molecular Formula: $C_4H_9O_3$
- Empirical Formula:

1. Find Mass of component elements

- Data may be in mass or as a percentage
- If data is in percent, assume 100 g total and use percent as mass
- 25.3 % element A and 74.7 % element B become
g element A and g element B if we work
with 100 g total

2. Convert Mass to Moles

- Convert each element to moles by dividing by molar mass of the element

3. Find Mole Ratio

- Divide each number of moles by the smallest number

4. Find Whole Number Ratio

- Multiply to eliminate all decimals (or fractions)

Example # 1

- A 43.2 g lump of charcoal (carbon) burns with oxygen and produces 159.0 g of a new compound. What is the empirical formula of this compound?
- Step 1:

2. Find moles of each element (divide by molar mass)

3. Divide to find mole ratio (always divide by smallest # of moles)

-

Step 4 unnecessary for this question!

- moles of oxygen for each mole of carbon
- Empirical Formula is . . .

Multiply to find Whole Numbers

• Element	mass in 100.0 g	mol	mol ratio	simplest ratio
•				
•				
•				
•				
•				
•				

Empirical Formula is . . .

Example #2

- Analysis of a compound showed that it contained 70.0 % iron and 30.0 % oxygen. What is the formula of this compound?
- Assume we have 100 g

Example #3: Find Empirical Formula if a compound has 20.2 % Al, 79.8 % Cl

• Element	mass in 100 g	mol	mol ratio	simplest ratio
•				
• Al	20.2 g			
• Cl	79.8 g			

HOMEWORK:

Percent Composition & Empirical
Formulas Worksheet



6. Molecular Formula

Actual Formula of a Compound

Molecular Formula

- Experimental data can only give the simplest whole number ratios for a formula
- Some compounds may have the same empirical formula, but have different molecular formulas
- E.g. NO_2 and N_2O_4
- C_2H_2 and C_6H_6
- Molecular formula is always a whole number multiple of the empirical formula
- To find the multiplier, divide the molar mass by the empirical mass (mass of empirical formula)

Example #1

- Many hydrocarbons (compounds containing only hydrogen and carbon) have the same empirical formula.
- Two compounds both have an empirical formula of CH_2 . The molar mass of compound A is 28.0 g/mol. The molar mass of compound B is 70.0 g/mol. What is the molecular formula of each?

Compound B

- Empirical mass =
-
- Multiplier = $\frac{\text{Molar Mass B}}{\text{Empirical Mass}} =$

- Molecular formula of B =

Compound A

- Empirical mass =

$$\frac{\text{Molar Mass A}}{\text{Empirical Mass}} =$$

- Molecular formula of A =

Example # 2

- The empirical formula for a compound is NO_2 . If 0.405 moles of the substance has a mass of 37.2 g. What is the molecular formula?

HOMEWORK:
Empirical & Molecular
Formula Worksheet