Chapter 10

"Chemical Quantities"

Yes, **you <u>will</u> need a calculator for this chapter!**

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Section 10.1
The Mole: A Measurement of Matter

- <u>OBJECTIVES:</u>
  - <u>Describe</u> methods of *measuring* the amount of something.

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Section 10.1
The Mole: A Measurement of Matter

- <u>OBJECTIVES:</u>
  - <u>Define</u> *Avogadro's number* as it relates to a mole of a substance.

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Section 10.1
The Mole: A Measurement of Matter

- <u>OBJECTIVES:</u>
  - <u>Distinguish</u> between the *atomic mass* of an element and its *molar mass.*

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Section 10.1
The Mole: A Measurement of Matter

- <u>OBJECTIVES:</u>
  - <u>Describe</u> how the mass of a mole of a compound is *calculated*.

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How do we measure items?

- You can measure *mass*,
- or *volume*,
- or you can *count pieces*.
- We measure mass in <u>grams</u>.
- We measure volume in <u>liters</u>.
- We count pieces in <u>MOLES</u>.

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What is the mole?

We're not talking about this kind of mole!

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Moles (is abbreviated: mol)

- **It is <u>an amount</u>**, defined as the number of carbon atoms in exactly 12 grams of carbon-12.
- **1 mole = 6.02 x 10^{23}** of the representative particles.
- Treat it like a very large dozen
- 6.02 x 10^{23} is called:

8 Avogadro's number.

Similar Words for an amount

- **<u>Pair</u>: 1 pair of shoelaces**
  = 2 shoelaces
- **<u>Dozen</u>: 1 dozen oranges**
  = 12 oranges
- **<u>Gross</u>: 1 gross of pencils**
  = 144 pencils
- **<u>Ream</u>: 1 ream of paper**
  = 500 sheets of paper

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What are Representative Particles?

- **The smallest pieces of a substance:**
  1) **For a molecular compound: it is the <u>molecule</u>.**
  2) **For an ionic compound: it is the <u>formula unit</u> (made of ions).**
  3) **For an element: it is the <u>atom</u>.**
     - » **Remember the 7 diatomic elements? (made of molecules)**

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Types of questions

- How many *oxygen atoms* in the following?
  - $CaCO_3$ — 3 atoms of oxygen
  - $Al_2(SO_4)_3$ — 12 (3 x 4) atoms of oxygen
- How many *ions* in the following?
  - $CaCl_2$ — 3 total ions (1 Ca^{2+} ion and 2 Cl^{1-} ions)
  - NaOH — 2 total ions (1 Na^{1+} ion and 1 OH^{1-} ion)
  - $Al_2(SO_4)_3$ — 5 total ions (2 Al^{3+} + 3 SO_4^{2-} ions)

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Practice problems (<u>round to 3 sig. figs</u>.)

- How many molecules of CO_2 are in 4.56 moles of CO_2? 2.75 x 10^{24} molecules
- How many moles of water is 5.87 x 10^{22} molecules? 0.0975 mol (or 9.75 x 10^{-2})
- How many <u>atoms</u> of carbon are in 1.23 moles of $C_6H_{12}O_6$? 4.44 x 10^{24} atoms C
- How many moles is 7.78 x 10^{24} formula units of $MgCl_2$? 12.9 moles

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Measuring Moles

- Remember relative atomic mass?
  - The amu was one twelfth the mass of a carbon-12 atom.
- Since the mole is the number of atoms in 12 grams of carbon-12,
- the <u>decimal number</u> on the periodic table is also the mass of 1 mole of those atoms in grams.

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Gram Atomic Mass (gam)

- Equals the mass of 1 mole of an element in grams (from periodic table)
- 12.01 grams of C has the same number of pieces as 1.008 grams of H and 55.85 grams of iron.
- We can write this as: 12.01 g C = 1 mole C (this is also the molar mass)
- We can count things by weighing them.

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Examples

- How much would 2.34 moles of carbon weigh? 28.1 grams C
- How many moles of magnesium is 24.31 g of Mg? 1 mol Mg
- How many atoms of lithium is 1.00 g of Li? 8.72×10^{22} atoms Li
- How much would 3.45×10^{22} atoms of U weigh? 13.6 grams U

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What about compounds?

- in 1 mole of H_2O molecules there are two moles of H atoms and 1 mole of O atoms (think of a compound as a **<u>molar ratio</u>**)
- To find the mass of one mole of a compound
  - –determine the number of moles of the elements present
  - –Multiply the number times their mass (from the periodic table)
  - –add them up for the total mass

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<u>Calculating Formula Mass</u>

Calculate the formula mass of magnesium carbonate, $MgCO_3$.

12 Mg 24.3050	6 C 12.0107	8 O 15.9994

24.3 g + 12 g + 3 x (16.00 g) = 84.3 g

Thus, 84.3 grams is the formula mass for **$MgCO_3$**.

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Section 10.2 Mole-Mass and Mole-Volume Relationships

- <u>OBJECTIVES:</u>
  - –<u>Describe</u> how to *convert* the mass of a substance to the number of moles of a substance, and moles to mass.

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Section 10.2 Mole-Mass and Mole-Volume Relationships

- OBJECTIVES:
  - Identify the *volume* of a quantity of gas at STP.

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Molar Mass

- **Molar mass** is the generic term for the mass of one mole of *any* substance (expressed in grams/mol)
- The same as:
  1) Gram Molecular Mass (for molecules)
  2) Gram Formula Mass (ionic compounds)
  3) Gram Atomic Mass (for elements)
  - molar mass is just a much broader term than these other specific masses

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Examples

- Calculate the molar mass of the following and tell what type it is:

Na_2S = 78 g/mol gram formula mass

N_2O_4 = 92 g/mol gram molecular mass

C = 12 g/mol gram atomic mass

$Ca(NO_3)_2$ = 164 g/mol gram formula mass

$C_6H_{12}O_6$ = 180 g/mol gram molecular mass

$(NH_4)_3PO_4$ = 149 g/mol gram formula mass

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Since Molar Mass is…

- The number of grams in 1 mole of atoms, ions, or molecules,
- We can make *conversion factors* from these.
  - To change between grams of a compound and moles of a compound.

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For example

- How many moles is 5.69 g of NaOH?

(Solution on next slides)

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For example

- **How many moles is 5.69 g of NaOH?**

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For example

- How many moles is 5.69 g of NaOH?

- We need to change 5.69 grams NaOH to moles

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For example

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- 1mole Na = 23 g 1 mol O = 16 g 1 mole of H = 1 g

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For example

- How many moles is 5.69 g of NaOH?

- We need to change 5.69 grams NaOH to moles
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The Mole-Volume Relationship

- Many of the chemicals we deal with are in the physical state as: **gases.**
  - They are difficult to *weigh (or mass)*.
- But, we may still need to know how many moles of gas we have.
- Two things effect the volume of a gas:
  a) Temperature and b) Pressure
- We need to compare all gases at the same temperature and pressure.

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Standard Temperature and Pressure

- 0ºC and 1 atm pressure
  - is abbreviated "STP"
- At STP, 1 mole of *any gas* occupies a volume of 22.4 L
  - Called the molar volume
- This is our fourth equality:
  1 mole of *any gas at STP* = 22.4 L

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Practice Examples

- What is the volume of 4.59 mole of CO_2 gas at STP? = 103 L
- How many moles is 5.67 L of O_2 at STP? = 0.253 mol
- What is the volume of 8.8 g of CH_4 gas at STP? = 12.3 L

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Density of a gas

- D = m / V (density = mass/volume)
  - for a gas the units will be: g / L
- We can determine the density of any gas at STP if we know its formula.
- To find the density we need: 1) mass and 2) volume.
- If you assume you have 1 mole, then the mass is the molar mass (from periodic table)
- And, at STP the volume is 22.4 L.

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Practice Examples (D=m/V)

- Find the density of CO_2 at STP.
  D = 44g/22.4L = 1.96 g/L
- Find the density of CH_4 at STP.
  D = 16g/22.4L = 0.714 g/L

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Another way:

- *If given the density*, we can find the molar mass of the gas.
- Again, pretend you have 1 mole at STP, so V = 22.4 L.
  modify: D = m/V to show: **m = D x V**
- "m" will be the mass of 1 mole, since you have 22.4 L of the stuff.
- What is the molar mass of a gas with a density of 1.964 g/L? **= 44.0 g/mol**
- How about a density of 2.86 g/L? = 64.0 g/mol

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Summary

- **<u>These four items are all equal:</u>**
  - a) 1 mole
  - b) molar mass (in grams/mol)
  - c) 6.02×10^{23} representative particles (atoms, molecules, or formula units)
  - d) 22.4 L of gas at STP

  Thus, we can make conversion factors from these 4 values!

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Section 10.3 Percent Composition and Chemical Formulas

- OBJECTIVES:
  - Describe how to *calculate* the percent by mass of an element in a compound.

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Section 10.3 Percent Composition and Chemical Formulas

- OBJECTIVES:
  - Interpret an *empirical formula.*

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Section 10.3 Percent Composition and Chemical Formulas

- OBJECTIVES:
  - Distinguish between *empirical* and *molecular formulas*.

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Calculating Percent Composition of a Compound

- **Like all percent problems:**

$$\left[\frac{\textbf{part}}{\textbf{whole}} \textbf{ x 100 \%}\right] \textbf{= percent}$$

1) Find the mass of each of the components (the elements),
2) Next, divide by the total mass of the compound; then x 100

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Example

- Calculate the percent composition of a compound that is made of 29.0 grams of Ag with 4.30 grams of S.

$$\frac{29.0 \text{ g Ag}}{33.3 \text{ g total}} \text{ X 100 = 87.1 \% Ag}$$

$$\frac{4.30 \text{ g S}}{33.3 \text{ g total}} \text{ X 100 = 12.9 \% S}$$

Total = 100 %

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Getting it from the formula

- If we know the formula, assume you have 1 mole,
- then you know the mass of the elements and the whole compound (these values come from the periodic table!).

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Examples

- Calculate the percent composition of C_2H_4? 85.7% C, 14.3 % H
- How about Aluminum carbonate? 23.1% Al, 15.4% C, and 61.5 % O
- Sample Problem 10.10, p.307
- We can also use the percent as a *conversion factor*
- Sample Problem page 308

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Formulas

***Empirical formula*: the <u>lowest whole number ratio</u> of atoms in a compound.**

***Molecular formula*: the <u>true number of atoms of each element</u> in the formula of a compound.**

- **Example: molecular formula for benzene is C_6H_6 (note that everything is divisible by 6)**
  - **Therefore, the empirical formula = CH (the lowest *whole* number ratio)**

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Formulas (continued)

Formulas for *ionic compounds* are <u>ALWAYS</u> empirical (the lowest whole number ratio = <u>cannot</u> be reduced).

Examples:

NaCl MgCl$_2$ Al$_2$(SO$_4$)$_3$ K$_2$CO$_3$

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Formulas (continued)

Formulas for *molecular compounds* <u>MIGHT</u> be empirical (lowest whole number ratio).

Molecular: (Correct formula)	H_2O	$C_6H_{12}O_6$	$C_{12}H_{22}O_{11}$
	↓	↓	↓
Empirical: (Lowest whole number ratio)	H_2O	CH_2O	$C_{12}H_{22}O_{11}$

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Calculating Empirical

- Just find the lowest whole number ratio
  - $C_6H_{12}O_6$ = CH_2O
  - CH_4N = this is already the lowest ratio.
- A formula is not just the ratio of *atoms*, it is also the ratio of ***<u>moles</u>***.
- In 1 <u>mole</u> of CO_2 there is <u>1 mole of carbon</u> and <u>2 moles of oxygen</u>.
- In one molecule of CO_2 there is 1 atom of C and 2 atoms of O.

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Calculating Empirical

- We can get a ratio *<u>from the percent composition.</u>*

1) Assume you have a 100 g sample
   - the percentage become grams (75.1% = 75.1 grams)
2) Convert grams to moles.
3) Find lowest whole number ratio by dividing each number of moles by the smallest value.

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Example

- Calculate the empirical formula of a compound composed of 38.67 % C, 16.22 % H, and 45.11 %N.
- Assume 100 g sample, so
- **38.67 g C x $\frac{\text{1mol C}}{\text{12.0 g C}}$ = 3.22 mole C**
- **16.22 g H x $\frac{\text{1mol H}}{\text{1.0 g H}}$ = 16.22 mole H**
- **45.11 g N x $\frac{\text{1mol N}}{\text{14.0 g N}}$ = 3.22 mole N**

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Example

- **The ratio is $\frac{\text{3.22 mol C}}{\text{3.22 mol N}} = \frac{\text{1 mol C}}{\text{1 mol N}}$**
- **The ratio is $\frac{\text{16.22 mol H}}{\text{3.22 mol N}} = \frac{\text{5 mol H}}{\text{1 mol N}}$**

= $C_1H_5N_1$ which is = CH_5N

- A compound is 43.64 % P and 56.36 % O. What is the empirical formula? **= P_2O_5**
- Caffeine is 49.48% C, 5.15% H, 28.87% N and 16.49% O. What is its empirical formula? **= $C_4H_5N_2O$**

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Empirical to molecular

- Since the empirical formula is the *lowest ratio*, the actual molecule would weigh more.
  - By a whole number multiple.
- Divide the actual molar mass by the empirical formula mass – you get a whole number to *increase each coefficient* in the empirical formula
- Caffeine has a molar mass of 194 g. what is its molecular formula? = $C_8H_{10}N_4O_2$

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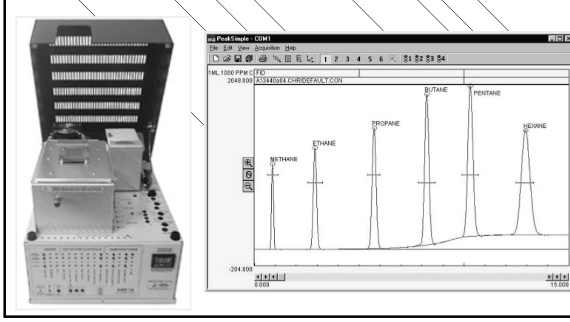
Note page 313 – Gas Chromatography used for chemical analysis



End of Chapter 10

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