Chemistry

Chemical Quantities
Lesson 6
Lesson Plan
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The Mole: A Measurement of Matter

Objectives: Describe how Avogadro's number is related to a mole of any substance; Calculate the mass of a mole of any substance

- What is a Mole?
  - We live in a quantitative world – numbers describe everything around us.
    - Grade on last test
    - Amount of money in our pockets
    - How many times you heard your favorite song on the radio yesterday
    - How far is it from your house to the school
    - How much do you weigh
  - In chemistry, chemist ask questions that also involve numbers
    - How many kg of iron can be obtained from a kg of iron ore
    - How many grams of hydrogen and nitrogen must be combined to make 200g of the fertilizer ammonia (NH_3)
  - As we move forward in our chemistry, we will be involved in analyzing the composition of samples of matter
    - Including calculations relating quantities of reactants and products to chemical equations
    - To do this we must be able to measure matter
  - How do you measure matter?
    - Count the number of CD's in your collection
    - Count the number of stuffed bears in your collection
    - How do you buy beans? Buy counting how many you are buying? Typically by weight.
    - Gold is sold by weight
    - Potatoes are sold by weight
    - How do you buy milk? By volume
    - Gasoline is by volume
    - Shampoo is by volume
    - There are numerous ways to measure matter
  - Common units of measurement
    - Socks are sold in ____ (pairs) meaning 2
    - Shoes are sold in pairs
    - Eggs are sold by the dozen – 12
      - A dozen of anything is 12
      - Except a baker's dozen which is 13
    - Apples are sold in 3 different ways
      - By count (meaning 5)

- By weight ($2/pound)
- Or by the volume – bushel ($12/bushel)
- We could set up an equivalency of count, mass and volume
  - By count: 1 dozen apples = 12 apples
  - By mass: 1 dozen apples = 2.0 kg apples
  - By volume: 1 dozen apples = 0.20 bushel apples

- The problem in measuring atoms or molecules is they are too small to be seen and too light to be weighed individually.
- The Mole (**mol**) – the mole is like a dozen – where a dozen is 12 of anything, a mole is 6.02×10^{23} of anything's.
  - A mole of apples would be 6.02×10^{23} apples
  - A mole of doughnuts would be 6.02×10^{23} doughnuts
  - 6.02×10^{23} is called Avogadro's number
    - In honor of Amedo Avogadro di Quadrenga
    - He did not determine the number, rather he clarified the difference between atoms and molecules
- Avogadro's number is used with the term "representative particles"
  - This could be atoms, molecules or formula units
  - The representative particle for most elements is the atom
  - There are 7 elements that normally occur as diatomic molecules (H_2, N_2, O_2, F_2, Cl_2, Br_2, and I_2)
    - The representative particle for these elements and all molecular compounds is the molecule
    - H_2O, or SO_2
  - The representative particle for ionic compounds is the formula unit
    - NaCl (1:1) or $CaCl_2$ (1:2)

Representative Particles and Moles

Substance	**Representative Particle**	**Chemical Formula**	**Representative particles in 1.00 mol**
Atomic nitrogen	Atom	N	6.02×10^{23}
Nitrogen gas	Molecule	N_2	6.02×10^{23}
Water	Molecule	H_2O	6.02×10^{23}
Calcium ion	Ion	Ca^{2+}	6.02×10^{23}
Calcium fluoride	Formula Unit	CaF_2	6.02×10^{23}
Sucrose	Molecule	$C_{12}H_{22}O_{11}$	6.02×10^{23}

Example Problem #1

How many moles of magnesium is 1.25×10^{23} atoms of magnesium?

Known	*Unknown*
Number of atoms = 1.25×10^{23} atoms of Mg	moles = ? mol Mg
1 mol Mg = 6.02×10^{23} atoms of Mg	

The desired conversion is atoms $\longrightarrow$ moles

The conversion factor is $\dfrac{1\,mol\ Mg}{6.02\ x\,10^{23}\ atoms\ Mg}$

Multiplying the known number of atoms of Mg by the conversion yields:

$$1.25\ x\,10^{23}\,atoms\ Mg \bullet \frac{1\,mol\ Mg}{6.02\ x\,10^{23}\ atoms\ Mg}$$

$$1.25\ x\,10^{23}\ \cancel{atoms\ Mg} \bullet \frac{1\,mol\ Mg}{6.02\ x\,10^{23}\ \cancel{atoms\ Mg}} = .208\,mol\ Mg\ =$$

$$= 2.08\ x\,10^{-1}\ mol\ Mg$$

Does the answer make sense?

Because the given number of atoms is less than ¼ of Avogadro's number, the answer should be less than ¼ mole of atoms. The answer should have three significant digits.

- Determining the number of atoms in a mole of a compound
    - First you must know how many atoms are in a representative particle of the compound – this is determined by the chemical formula
        - CO_2 has three atoms (1 carbon and 2 oxygen)
        - A mole of carbon dioxide contains Avogadro's number of CO_2 molecules
        - Thus a mole of carbon dioxide contains three times Avogadro's number of atoms
        - If we were talking about CO, then since it consists of two atom per molecule, there would be twice Avogadro's number of atoms in 1 mole

Sample Problem #2

How many atoms are in 2.12 mol of propane (C_3H_8)?

Known	*Unknown*
Number of moles = 2.12 mol (C_3H_8)?	Number of atoms = ?

atoms

1 mol C_3H_8 = 6.02 x 10^{23} molecules C_3H_8

1 molecule C_3H_8 = 11 atoms

(3 carbon and 8 hydrogen)

The desired conversion is moles $\longrightarrow$ molecules $\longrightarrow$ atoms

The first conversion factor is $\dfrac{6.02\ x\,10^{23}\ molecules\ C_3H_8}{1\,mol\ C_3H_8}$

The second conversion factor is $\dfrac{11\,atoms}{1\,molecule\ C_3H_8}$

$$\text{So } 2.12\ \cancel{mol\ C_3H_8}\ \text{x}\ \frac{6.02\ x\,10^{23}\ \cancel{molecules\ C_3H_8}}{\cancel{1\,mol\ C_3H_8}}\ \text{x}\ \frac{11\,atoms}{1\,\cancel{molecule\ C_3H_8}}$$

= 140.39 x 10^{23} atoms = 1.40 x 10^{25} atoms

Does the result make sense?

Because there are 11 atoms in each propane molecule and more than 2 molecules of propane, then the answer should be at least 2*10 times more than Avogadro's number . The answer has 3 significant digits based on the given measurements

Practice Problems:

1. How many atoms are in 1.00 mole of sucrose ($C_{12}H_{22}O_{11}$)?

$$1.00\ \cancel{mol\ C_{12}H_{22}O_{11}}\ \text{x}\ \frac{6.02\ x\ 10^{23}\ \cancel{molecules\ C_{12}H_{22}O_{11}}}{\cancel{1\ mol\ C_{12}H_{22}O_{11}}}\ \text{x}\ \frac{45\ atoms}{\cancel{1\ molecule\ C_{12}H_{22}O_{11}}} =$$

$270.90 \text{ x } 10^{23}$ atoms $= 2.71 \text{ x } 10^{25}$ atoms

2. How many atoms of C are in 2.0 moles of $C_{12}H_{22}O_{11}$?

$$2.00\ \cancel{mol\ C_{12}H_{22}O_{11}}\ \text{x}\ \frac{6.02\ x\ 10^{23}\ \cancel{molecules\ C_{12}H_{22}O_{11}}}{\cancel{1\ mol\ C_{12}H_{22}O_{11}}}\ \text{x}\ \frac{12\ atoms}{\cancel{1\ molecule\ C_{12}H_{22}O_{11}}} =$$

$144.48 \text{ x } 10^{23}$ atoms $= 1.44 \text{ x } 10^{25}$ atoms

3. How many atoms of O are in 2.0 moles of $C_{12}H_{22}O_{11}$?

$$3.65\ \cancel{mol\ C_{12}H_{22}O_{11}}\ \text{x}\ \frac{6.02\ x\ 10^{23}\ \cancel{molecules\ C_{12}H_{22}O_{11}}}{\cancel{1\ mol\ C_{12}H_{22}O_{11}}}\ \text{x}\ \frac{11\ atoms}{\cancel{1\ molecule\ C_{12}H_{22}O_{11}}} =$$

$241.70 \text{ x } 10^{23}$ atoms $= 2.42 \text{ x } 10^{25}$ atoms

4. How many atoms of H are in 2.0 moles of $C_{12}H_{22}O_{11}$?

$$2.00\ \cancel{mol\ C_{12}H_{22}O_{11}}\ \text{x}\ \frac{6.02\ x\ 10^{23}\ \cancel{molecules\ C_{12}H_{22}O_{11}}}{\cancel{1\ mol\ C_{12}H_{22}O_{11}}}\ \text{x}\ \frac{22\ atoms}{\cancel{1\ molecule\ C_{12}H_{22}O_{11}}} =$$

$264.88 \text{ x } 10^{23}$ atoms $= 2.65 \text{ x } 10^{25}$ atoms

5. How many atoms are there in 1.14 mol of SO_3?

$$1.14\ \cancel{mol\ SO_3}\ \text{x}\ \frac{6.02\ x\ 10^{23}\ \cancel{molecules\ SO_3}}{1\ \cancel{mol\ SO_3}}\ \text{x}\ \frac{4\ atoms}{\cancel{1\ molecule\ SO_3}} =$$

$27.45 \text{ x } 10^{23}$ atoms $= 2.75 \text{ x } 10^{24}$ atoms

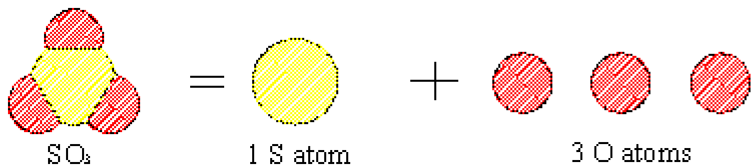
6. How many moles are there in $4.65 \text{ x } 10^{24}$ molecules of NO_2?

$$4.65 \text{ x } 10^{24}\ \cancel{\text{molecules } NO_2}\ \text{x}\ \frac{1\ mol\ NO_2}{6.02\ x\ 10^{23}\ \cancel{molecules\ NO_2}} = .7724 \text{ x } 10^{1} \text{ mol} = 7.72 \text{ mol}$$

- How big is Avogadro's number?
  - Use the animal mole for a basis
    - Typical mole is 15 cm long, 5 cm high, 150g mass
    - $6.02 \text{ x } 10^{23}$ animal-mole x $\frac{150g}{1\ animal-mole}$ $=9.03 \text{ x } 10^{22}$ kg

            - more than 1% the mass of the Earth
            - 1.3 times the mass of the moon
            - 60 times the mass of the Earth's oceans
            - If spread over the Earth, it would create a layer 8 million animals thick
        - 6.02×10^{23} ~~animal-mole~~ x $\frac{15\ cm}{1\ \cancel{animal-mole}}$ $= 9.03 \times 10^{19}$ km
            - long enough that if lined up end to end the animals would stretch from Earth to the nearest star, Alpha Centauri and back to earth again
    - If you make 6 billion stacks of animals and set one person to counting one stack at the rate of 1000 animals per second, it would take more than 3000 years to count 6.02×10^{23} animals.
    - Avogadro's number is really big!
- The Mass of a Mole of an element
    - Working with individual atoms, or even a billion atoms would be impossible since there would be too little of the substance to handle.
    - Chemists get around this by working with grams of a substance
    - By using the atomic mass of an element, expressed in grams, we have the **gram atomic mass**
        - The gam of carbon is 12.0 g
        - The gam of Hydrogen is 1.0 g
        - The gam of Oxygen is 16.0 g
        - What is the gam of Hg and Fe? (200.5 g, 55.8 g)
    - The gram atomic mass of any element contains 6.02×10^{23} atoms of that element
    - Remember that the atomic mass unit was based on carbon-12 and the there were 12 amu in one cabon-12 atom?
        - And that hydrogen atom had 1 amu?
        - Then just as 1 atom of carbon weighs 12 times 1 atom of hydrogen, the 100 atoms of carbon weighs 12 times 100 atoms of hydrogen.
        - In the same way, any number of carbon atoms will weigh 12 times the same number of hydrogen atoms.
        - Or if we were to compare carbon and oxygen, we would find that 12 g of carbon atoms would weigh as much as 16 g of oxygen atoms – and both would contain the same number of atoms – Avogadro's number of atoms.
    - So now the mole can be defined as the amount of substance that contains as many representative particles as the number of atoms in 12 g of carbon-12
        - 12 g of carbon is 1 mole of carbon and contains 6.02×10^{23} atoms
        - 1 g of hydrogen is 1 mole of hydrogen and contains 6.02×10^{23} atoms

- 16 g of oxygen is 1 mole of oxygen and contains 6.02 x 10^{23} atoms
- The gram atomic mass is the mass of 1 mole of atoms of any element

- The Mass of a Mole of a compound
  - Like a atom, a molecule is far too small to be weighed, so again we use the mole
  - First we must know the formula of the compound – this tells us the number of atoms of each element in a representative particle.
  - 

    - SO_3 – has 1 sulfur and 3 oxygen atoms
    - Add the atomic mass of all the atoms making up one molecule (1 * 32.1 amu + 3 * 16.0 amu = 32.1 amu + 48 amu = 80.1 amu (the molecular mass of sulfur trioxide)
    - When working with atoms, we used the gram atomic mass, with compounds we use **gram molecular mass** – same concept, but for compounds rather than atoms. That is 1 gmm of a compound is 1 mol of that compound
    - 1 amu = 1gmm
    - So 80.1 amu = 80.1 gmm = 1mol of SO_3

Examples:

1. What is the gram molecular mass of hydrogen peroxide (H_2O_2)?

Known	*Unknown*
Molecular formula = H_2O_2	gmm = ?g

1 gam H = 1 mol H = 1.0 g H
1 gam O = 1 mol O = 16.0 g O
Since we know that there is 2 mol of H and 2 mol of O in 1 mol of hydrogen peroxide, we must convert moles of atoms into grams by using the conversion factors (g/mol) based on the gram atomic mass of each element. The sum of the masses of the elements gives the gram molecular weight.

$$2\,\cancel{mol\ H}\ x\ \frac{1.0\ g\ H}{1\,\cancel{mol\ H}} = 2.0g\ H$$

$$2\,\cancel{mol\ O}\ x\ \frac{16.0\ g\ O}{1\,\cancel{mol\ O}} = 32.0g\ O$$

2.0 g H + 32.0 g O = 34.0 g = gmm of H_2O_2

2. 1 mol of glucose molecules (blood sugar) $C_6H_{12}O_6$
6 carbons, 12 hydrogens, and 6 oxygens
6 * 12.0 + 12 * 1.0+6 * 16.0 = 180.0 g $C_6H_{12}O_6$ = 1 gmm $C_6H_{12}O_6$

3. 1 mol of H_2O (water)
   2 hydrogen and 1 oxygen
   $2 * 1.0 + 1 * 16.0 = 18.0$ g H_2O = 1 gmm H_2O

4. 1 mol of paradichlorobenzene molecules (moth crystals) $C_6H_4Cl_2$
   6 carbons, 4 hydrogens, and 2 chlorines
   $6 * 12.0 + 4 * 1.00 + 2 * 35.5 = 147.0$ g $C_6H_4Cl_2$ = 1 gmm $C_6H_4Cl_2$

5. What is the gram molecular mass of C_2H_6?
   2 carbons and 6 hydrogens
   $2 * 12.0 + 6 * 1.0 = 30.0$ g

6. What is the gram molecular mass of PCl_3?
   1 phosphorous and 3 chlorine atoms
   $1 * 31.0 + 3 * 35.5 = 137.5$ g

7. What is the gram molecular mass of C_3H_7OH?
   3 carbon, 8 hydrogen, and 1oxygen
   $3 * 12.0 + 8 * 1.0 + 1 * 16.0 = 60.0$ g

8. What is the gram molecular mass of N_2O_5?
   2 nitrogen and 5 oxygen
   $2 * 14.0 + 5 * 16.0 = 108.0$ g

9. What is the mass of 1.00 mole of chlorine?
   Cl occurs as a diatomic atom, so it is Cl_2
   $2* 35.5 = 71.0$g

10. What is the mass of 1.00 mole of nitrogen dioxide?
    NO_2
    1 nitrogen and 2 chlorine
    $1 * 14.0 + 2 * 16.0 = 46.0$ g

11. What is the mass of 1.00 mole of carbon tetrabromide?
    CBr_4
    1 carbon and 4 bromine
    $1 * 12.0 + 4 * 79.9 = 331.6$ g

12. What is the mass of 1.00 mole of silicon dioxide?
    SiO_2
    1 silicon and 2 oxygen
    $1 * 28.1 + 2 * 16.0 = 60.1$ g

- How about ionic compounds?
  - Remember the representative unit for a molecular compound is the molecule, but for an ionic compound it is the formula unit

- When calculating the mass of a mole of an ionic compound we use the **gram formula mass**.
- gfm is just like the gmm, in that the gfm equals te formula mass expressed in grams – simply sum the atomic masses of the ions in the formula unit of the compound

Examples:

1. What is the gram formula mass of ammonium carbonate ($(NH_4)_2CO_3$)?

Knowns | *Unknowns*
Formula unit = $(NO_4)_2CO_3$ | gfm = ? g
1 gam N = 1 mol N – 14.0g N
1 gam H = 1 mol H = 1.0 g H
1 gam C = 1 mol C = 12.0 g C
1 gam O = 1 mol O = 16.0 g O

The formula shows that a mole of this ionic compound is composed of 2 mol of nitrogen, 8 mol of hydrogen, 1 mol of carbon and 3 mol of oxygen. Moles of atoms are converted to grams using the conversion factor based on the gram atomic masses. The sum of the masses of the elements gives the gram formula mass.

$$2\,\cancel{mol\ N}\ x\ \frac{14.0\ g\ N}{1\,\cancel{mol\ N}} = 28.0\ g\ N$$

$$8\,\cancel{mol\ N}\ x\ \frac{1.0\ g\ H}{1\,\cancel{mol\ H}} = 8.0\ g\ H$$

$$1\,\cancel{mol\ C}\ x\ \frac{12.0\ g\ C}{1\,\cancel{mol\ C}} = 12.0\ g\ C$$

$$3\,\cancel{mol\ N}\ x\ \frac{16.0\ g\ O}{1\,\cancel{mol\ O}} = 48.0\ g\ O$$

so 28.0 g N + 8.0 g H + 12.0 g C + 48.0 g O = 96.0 g as the gram formula mass of $(NO_4)_2CO_3$

2. Calculate the gram formula mass of K_2O.
2 potassium, 1 oxide
2 * 39.1 g + 1 * 16.0 g O = 94.2 g

3. Calculate the gram formula mass of $CaSO_4$.
1 calcium, 1 sulfur and 4 oxygen
1 * 40.1 g + 1 * 32.1 g + 4 * 16.0 g = 136.2 g

4. Calculate the gram formula mass of CuI_2.
1 copper and 2 iodide
1 * 63.6 g + 2 * 126.9 g = 317.4 g

5. Find the gram formula mass of barium fluoride.
BaF_2 = 1 barium and 2 fluoride

$1 * 137.3\ g + 2 * 19.0\ g = 175.3\ g$

6. Find the gram formula mass of strontium cyanide.
   $Sr(CN)_2$ = 1 strontium, 2 carbon and 2 nitrogen
   $1 * 87.6\ g + 2 * 12.0\ g + 2 * 14.0\ g = 139.6\ g$

6. Find the gram formula mass of sodium hydrogen carbonate.
   $NaHCO_3$ = 1 sodium, 1 hydrogen, 1 carbon, and 3 oxygen
   $1 * 23.0\ g + 1 * 1.0\ g + 1 * 12.0\ g + 3 * 16.0\ g = 84.0\ g$

7. Find the gram formula mass of aluminum sulfite.
   $Al_2(SO_3)_3$ = 2 aluminum, 3 sulfur and 9 oxygen
   $2 * 27.0\ g + 3 * 32.1\ g + 9 * 16.0\ g = 294.3\ g$

Section Review Problems

Mole-Mass and Mole-Volume Relationships

Objectives: Use the molar mass to convert between mass and moles of a substance; Use the mole to convert among measurements of mass, volume, and number of particles

- The Molar Mass of a Substance
  - Just learned three new terms
    - Gram atomic mass (for elements)
    - Gram molecular mass (for molecular compounds)
    - Gram formula mass (for ionic compounds)
    - Each represents a mole for a particular substance
- To represent all of the above we use **molar mass** to represent the mass to make one mole of any substance.
  - This can cause problems – for example if we talk about a molar mass of oxygen, are we talking about atoms of oxygen (O), then the molar mass is 16.0 g. If we mean the oxygen molecule (O_2), then the molar mass is 32.0 g (2 x 16.0 g).
    - Chemist use more specific terms when questions may be raised.

Sample Problems:

1. How many grams are in 9.45 mol of dinitrogen trioxide (N_2O_3)?

Known	*Unknown*
9.45 mol N_2O_3	mass = ? g N_2O_3
1 mol N_2O_3 = 76.0 g	

moles $\longrightarrow$ grams

$$9.45\ \cancel{mol\ N_2O_3}\ x\ \frac{76.0\ g\ N_2O_3}{1.00\ \cancel{mol\ N_2O_3}} = 718.2\ g\ N_2O_3 = 718\ g\ N_2O_3$$

Does the result make sense?

Since 1 mole = 76 g, then we have just under 10 mol so the answer should be about 700. Be sure to round your answer off to the proper number of significant digits.

2. Find the mass, in grams of 3.32 mol K.

$$3.32\ mol\ K\ x\ \frac{39.1\ g\ K}{1\ mol\ K} = 129.8\ g\ K = 130\ g\ K$$

3. Find the mass, in grams of 4.52 x 10^{-3} mol $C_{20}H_{42}$.

$$4.52\ x\ 10^{-3}\ mol\ C_{20}H_{42} = \frac{282.0\ g\ C_{20}H_{42}}{1\ mol\ C_{20}H_{42}} = 1.274\ g\ C_{20}H_{42} = 1.27\ g\ C_{20}H_{42}$$

4. Find the mass, in grams of 0.0112 mol K_2CO_3.

$$0.0112\ mol\ K_2CO_3 = \frac{138.2\ g\ K_2CO_3}{1\ mol\ K_2CO_3} = 1.547\ g\ K_2CO_3 = 1.55\ g\ K_2CO_3$$

5. Calculate the mass, in grams of 2.50 mol sodium sulfate.
Sodium Sulfate is Na_2SO_4

$$2.50\ mol\ Na_2SO_4 = \frac{142.1\ g\ Na_2SO_4}{1\ mol\ Na_2SO_4} = 355.25\ g\ Na_2SO_4 = 355\ g\ Na_2SO_4$$

6. Calculate the mass, in grams of 2.50 mol iron(II) hydroxide.
Iron(II) hydroxide is $Fe(OH)_2$

$$2.50\ mol\ Fe(OH)_2 = \frac{89.9\ g\ Fe(OH)_2}{1\ mol\ Fe(OH)_2} = 224.75\ g\ Fe(OH)_2 = 225\ g\ Fe(OH)_2$$

7. Find the number of moles in 92.2 g of iron (III) oxide ($Fe_2\ O_3$).

Known	*Unknown*
Mass – 92.2 g $Fe_2\ O_3$	Number of moles = ? mol $Fe_2\ O_3$
1 mol $Fe_2\ O_3$ = 159.6 $Fe_2\ O_3$	
conversion is grams $\rightarrow$ moles	

$$92.2\ \cancel{g\ Fe_2O_3}\ x\ \frac{1.00\ mol\ Fe_2O_3}{159.6\ \cancel{g\ Fe_2O_3}} = 0.5776\ mol\ Fe_2O_3 = 0.578\ mol\ Fe_2O_3$$

Does the result make sense?
Because the mass (about 90 g) is slightly larger than the mass of ½ mol of Fe_2O_3 (about 160 g), the answer should be slightly larger than ½ mole.

8. Find the number of moles in 3.70 x 10^{-1} g B.

$$3.70\ x\ 10^{-1}\ \cancel{g\ B}\ x\ \frac{1.00\ mol\ B}{10.8\ \cancel{g\ B}} = \frac{3.70\ x\ 10^{-1}\ mol\ B}{10.8} = 3.426\ x\ 10^{-2}\ mol\ B = 3.43\ x\ 10^{-2}\ mol\ B$$

9. Find the number of moles in 2.74 g TiO_2.

$$2.74\ \cancel{g\ TiO_2}\ x\ \frac{1.00\ mol\ TiO_2}{79.9\ \cancel{g\ TiO_2}} = \frac{2.74\ mol\ TiO_2}{79.9} = 3.429\ x\ 10^{-2}\ mol\ TiO_2 = 3.45\ x\ 10^{-2}\ mol\ TiO_2$$

10. Find the number of moles in 847 g $(NH_4)_2CO_3$.

$$847\ \cancel{g\ (NH_4)_2CO_3}\ x\ \frac{1.00\ mol\ (NH_4)_2CO_3}{96.0\ \cancel{g\ (NH_4)_2CO_3}} = \frac{847\ mol\ (NH_4)_2CO_3}{96.0} =$$

$$8.822\ mol\ (NH_4)_2CO_3 = 8.82\ mol\ (NH_4)_2CO_3$$

10. Calculate the number of moles in 75.0 g of dinitrogen trioxide (N_2O_3).

$$75.0\ g\ N_2O_3\ x \frac{1\ mol\ N_2O_3}{76.0\ g\ N_2O_3} = .9868\ mol\ N_2O_3 = 9.87\ x\ 10^{-1} mol\ N_2O_3$$

11. Calculate the number of moles in 75.0 g of nitrogen gas (N_2).

$$75.0\ g\ N_2\ x \frac{1\ mol\ N_2}{28.0\ g\ N_2} = 2.678\ mol\ N_2 = 2.68\ mol\ N_2$$

12. Calculate the number of moles in 75.0 g of sodium oxide (Na_2O).

$$75.0\ g\ Na_2O\ x \frac{1\ mol\ N_2O}{62.0\ g\ N_2O} = 1.209\ mol\ Na_2O = 1.21\ mol\ Na_2O$$

- The Volume of a Mole of Gas
  - From examples we did earlier, we found that 1 mol of glucose had a gmm of 180.0 g, 1 mol of water had a gmm of 18.0 g and 1mole of paradichlorobenzene had a gmm of 147.0 g. The volume that these compounds occupy vary according to density.
  - Gases are more predictable under the same physical conditions.
  - When talking about physical characteristics of the states of matter (solid, liquid and gas), that a gas expanded with heat – so the volume of a gas will change with an increase of heat. If a gas is in a restricted volume, then heating the gas will raise the pressure of the gas.
    - Because of this variation, the volume of a gas is usually measured at a **standard temperature and pressure** – abbreviated as **STP**.
    - Standard temperature is 0ºC
    - Standard pressure is 101.3 kPa, or 1 atmosphere (atm)
  - At STP, 1 mol of gas occupies a volume of 22.4 L
  - 22.4 L is known as the **molar volume** of a gas
  - Because 1 mol of anything contains Avogadro's number of representative particles, then 22.4 L of gas at STP contains 6.02×10^{23} representative particles of that gas
    - Would this differ for gaseous elements compared with gaseous compounds? (no – a mole one gas contains just as many representative particles as another gas – think of molecules and atoms)
    - Would 22.4 L of one gas have the same mass as 22.4 L of another gas at STP?
      - Probably not – since gases are composed of different elements, just as solid compounds are composed different elements and have different mass
      - Only gasses with the same molar mass at STP would have equal masses for equal volumes at STP

Sample Problems

1. Determine the volume, in liters, of 0.60 mol SO_2 gas at STP.

Known | *Unknown*
0.60 mol SO_2 | volume = ? L SO_2
1 mol SO_2 = 22.4 L SO_2
Conversion is moles $\longrightarrow$ liters

$$0.60\ \cancel{mol\ SO_2}\ x\ \frac{22.4\ L\ SO_2}{1\ \cancel{mol\ SO_2}} = 13.44\ L\ SO_2 = 13\ L\ SO_2$$

2. What is the volume at STP of 3.20×10^{-3} mol CO_2?

$$3.20 \times 10^{-3}\ \cancel{mol\ CO_2}\ x\ \frac{22.4\ L\ CO_2}{1\ \cancel{mol\ CO_2}} = 71.68\ x\ 10^{-3}\ L\ CO_2 = 7.17\ x\ 10^{-2}\ L\ CO_2$$

3. What is the volume at STP of 0.960 mol CH_4?

$$0.960\ \cancel{mol\ CH_4}\ x\ \frac{22.4\ L\ CH_4}{1\ \cancel{mol}\ CH_4} = 21.50\ L\ CH_4 = 21.5\ L\ CH_4$$

4. What is the volume at STP of 3.70 mol N_2?

$$3.70\ \cancel{mol\ N_2}\ x\ \frac{22.4\ L\ N_2}{1\ \cancel{mol\ N_2}} = 82.88\ L\ N_2 = 8.29\ L\ N_2$$

5. Assuming STP, how many moles are in 67.2 L SO_4?

$$67.2\ \cancel{L\ SO_2}\ x\ \frac{1\ mol\ SO_2}{22.4\ \cancel{L\ SO_2}} = 3.00\ L\ SO_2$$

6. Assuming STP, how many moles are in 0.880 L He?

$$0.880\ \cancel{L\ He}\ x\ \frac{1\ mol\ He}{22.4\ \cancel{L\ He}} = 3.928 \times 10^{-2}\ L\ He = 3.93 \times 10^{-2}\ L\ He$$

7. Assuming STP, how many moles are in 1.00×10^3 L C_2H_6?

$$1.00 \times 10^{3}\ \cancel{L\ C_2H_6}\ x\ \frac{1\ mol\ C_2H_6}{22.4\ \cancel{L\ C_2H_6}} = 44.64\ L\ C_2H_6 = 4.46 \times 10^{1}\ L\ C_2H_6$$

- Remember density? For a solid it was the $\frac{mass}{volume}$ which for our lab was measured in g/ml. For gases, density is usually measured in g/L
  - The experimentally determined density of a gas at STP is sued to calculate the molar mass of that gas
    - The gas can be an element or a compound
    - What makes a balloon float or sink?
      - Density of gas in the balloon compared to the surrounding air
        - Balloons filled with by mouth contain the same air as they are surrounded by plus the weight of the balloon
        - Balloons filled with helium tend to float in air as long as the density difference is great enough to include the weight of the balloon
        - Would a helium balloon float in an atmosphere of helium? (no)

Example Problems:

1. The density of a gaseous compound containing carbon and oxygen is 1.964 g/L at STP. Determine the molar mass of the compound.

Known	*Unknown*
Density = 1.964 g/L	molar mass = ? g/mol
1 mol (gas at STP) = 22.4 L	

Conversion $\frac{g}{L} \rightarrow \frac{g}{mol}$

$$\frac{1.964\ g}{L}\ x\ \frac{22.4\ L}{1\ mol} = 43.99\ g/mol = 44.0\ g/mol$$

Does the result make sense?
The ratio of the calculated mass to the volume is about 2, which is about what the density was.

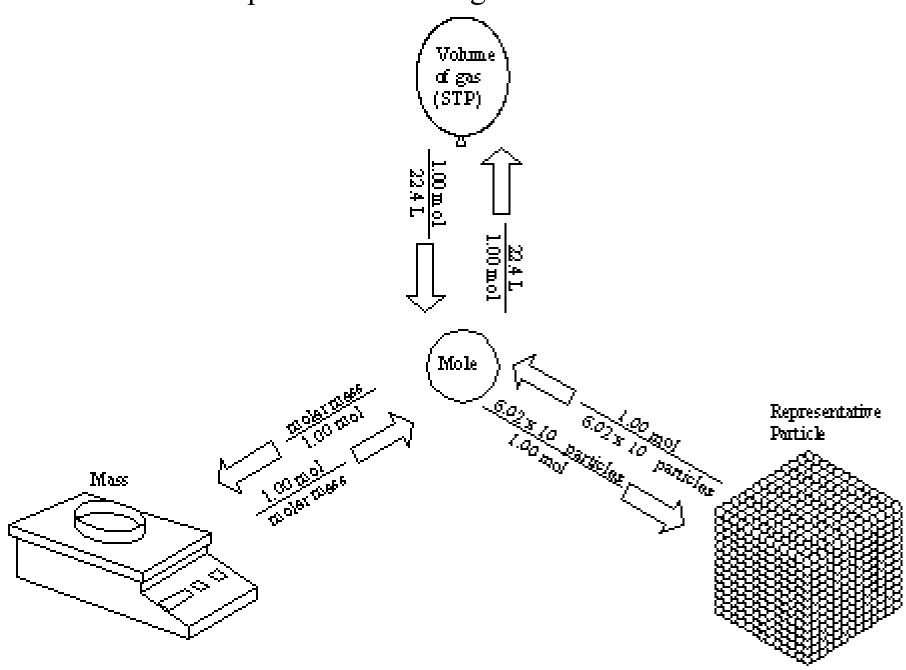
2. A gaseous compound composed of sulfur and oxygen that is linked to the formation of acid rain has a density of 3.58 g/L at STP. What is the molar mass of this gas?

$$\frac{3.58\ g}{\cancel{L}}\ x\ \frac{22.4\ \cancel{L}}{1\ mol} = 80.192\ g/mol = 80.2\ g/mol$$

3. What is the density of krypton gas at STP?

$$\frac{83.8g}{1\cancel{mol}}\ x\ \frac{1\ \cancel{mol}}{22.4\ L} = 3.741\ g/mol = 3.74\ g/mol$$

- The Mole Road Map
  - We have now examined the mole in terms of particles, mass, and volume of gases at STP. The below figure shows the relationship of these terms – note that we must pass through the mole as an intermediate step when converting from one to another.



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Percent Composition and Chemical Formulas

Objectives: calculate the percent composition of a substance from its chemical formula or experimental data; Derive the empirical formula and the molecular formula of a compound from experimental data.

- Calculating the Percent Composition of a Compound
  - When taking care of a yard, it is important to feed the yard what it needs to grow – typically this would be a mixture of nitrogen, phosphorus and potassium. Most fertilizers are labeled with 3 numbers, such as 15-10-15, indicating the amount of each element in the mixture.
  - The relative amount of each element in a compound is expressed in **percent composition** or **percent mass**
  - The percent composition or percent mass of a compound will contain as many values as there are elements in a compound, and they must total to 100.0%
  - % mass of element E = $\dfrac{\text{grams of element E}}{\text{grams of compound}}\, x\, 100\%$

Example

If an 8.20 g of magnesium combines completely with 5.40 g of oxygen to form a compound, what is the percent composition of this compound?

Known	*Unknown*
Mass of magnesium = 8.20 g	Percent Mg = ?% Mg
Mass of oxygen = 5.40 g	Percent O = ?% O

Mass of compound = 8.20 g + 5.40 g = 13.60 g

$$\%Mg = \frac{mass\ of\ Mg}{grams\ of\ compound}\, x\, 100\%$$

$$\%Mg = \frac{8.2\ g}{13.6\ g}\, x\, 100\% = 60.29\% = 60.3\%$$

$$\%O = \frac{mass\ of\ O}{Grams\ of\ compound}\, x\, 100\%$$

$$\%O = \frac{5.40\ g}{13.60\ g}\, x\, 100\% = 39.70\% = 39.7\%$$

Checking our results: 60.3% + 39.7% = 100.0%

Practice Problems

If 9.03 g of Mg combines completely with 3.48 g N to form a compound, what is the percent composition?

9.03 g + 3.48 g = 12.51 g total mass of compound

$$\%Mg = \frac{9.03\ g}{12.51\ g}\ x\ 100\% = 72.2\%$$

$$\%N = \frac{3.48\ g}{12.51\ g}\ x\ 100\% = 27.8\%$$

If 29.0 g of Ag combines completely with 4.30 g S to from a compound, then what is the percent composition?

$$\%\ Ag = \frac{29.0\ g}{33.30\ g}\ x\ 100\% = 87.1\%$$

$$\%\ S = \frac{4.30\ g}{33.30\ g}\ x\ 100\% = 12.9\%$$

When a 14.2 g sample of mercury(II) oxide is decomposed into its elements by heating, 13.2 g Hg is obtained. What is the percent composition of this of the compound?

$$\%Hg = \frac{13.2\ g}{14.2\ g}\ x\ 100\% = 93.0\%$$

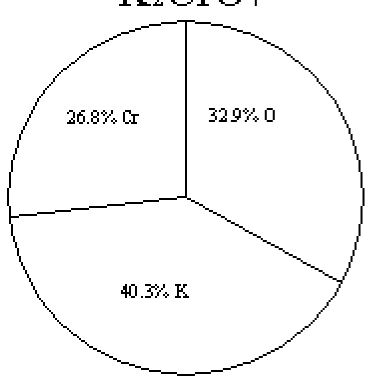
$$\%O = \frac{1.0\ g}{14.2\ g}\ x\ 100\% = 7.0\%$$

- So far we have calculated the percent by mass of an element in a compound – now we learn to calculate the % composition of a compound
    - With % mass we used the values of a known amount of a compound and the known amount of the substances in that compound
    - With % composition, we are not concerned with being given a known value of a compound, rather we use the molar mass of the compound and the grams of an element in mol of the compound.

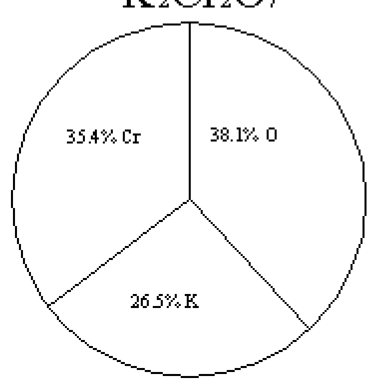
$$\%\ mass = \frac{grams\ of\ element\ in\ 1\ mol\ compound}{molar\ mass\ of\ compound}\ x\ 100\%$$

    - Compare two compounds with the same elements, but with different chemical formulas and how the percent composition differs

K_2CrO_4



$K_2Cr_2O_7$



Sample Problems

Calculate the percent composition of propane (C_3H_8)

Known	*Unknown*
molar mass of C_3H_8 = 44.0g/mol	percent C = ?%C
Mass of C in 1 mol C_3H_8 = 36.0 g	percent H = ?%H
Mass of H in 1 mol C_3H_8 = 8.0 g	

$$\%C = \frac{36.0\ g}{44.0\ g} x\, 100\% = 81.8\%$$

$$\%H = \frac{8.0\ g}{44.0\ g} x\, 100\% = 18.2\%$$

Calculate the percent composition of the following compounds

Ethane – C_2H_6 (80.0% C, 20.0% H)

Sodium bisulfate – $NaHSO_4$ (19.2 % Na, 0.83% H, 26.7% S, 53.3% O)

Ammonium chloride – NH_4Cl (26.2% N, 7.5% H, 66.4% Cl)

Calculate the percent nitrogen in these common fertilizers

$CO(NH_2)_2$ (46.7% N)

NH_3 (82.4% N)

NH_4NO_3 (35.0% N)

- Using Percent as a Conversion Factor
  - We can use % composition as a conversion factor in determining the mass of one element in a given mass of a compound.

Example Problems

Calculate the mass of carbon in 82.0 g of propane (C_3H_8).

Known	*Unknown*
Mass of C_3H_8 = 82.0 g	mass of carbon = ? g C

First we must calculate the % composition of carbon in C_3H_8

$$\%C = \frac{36.0\ g}{44.0\ g} x\, 100\% = 81.8\%$$

Next we use the % composition in relationship to 100%

$$mass\ of\ compound\ x \frac{\%\ of\ element\ in\ grams}{100\%\ in\ grams} = mass\ of\ element\ in\ compound$$

$$82.0\ \cancel{g\ C_3H_8} = \frac{81.8\ g\ C}{100.0\ \cancel{g\ C_3H_8}} = 67.1\ g\ C$$

Calculate the mass of hydrogen in each of the following:

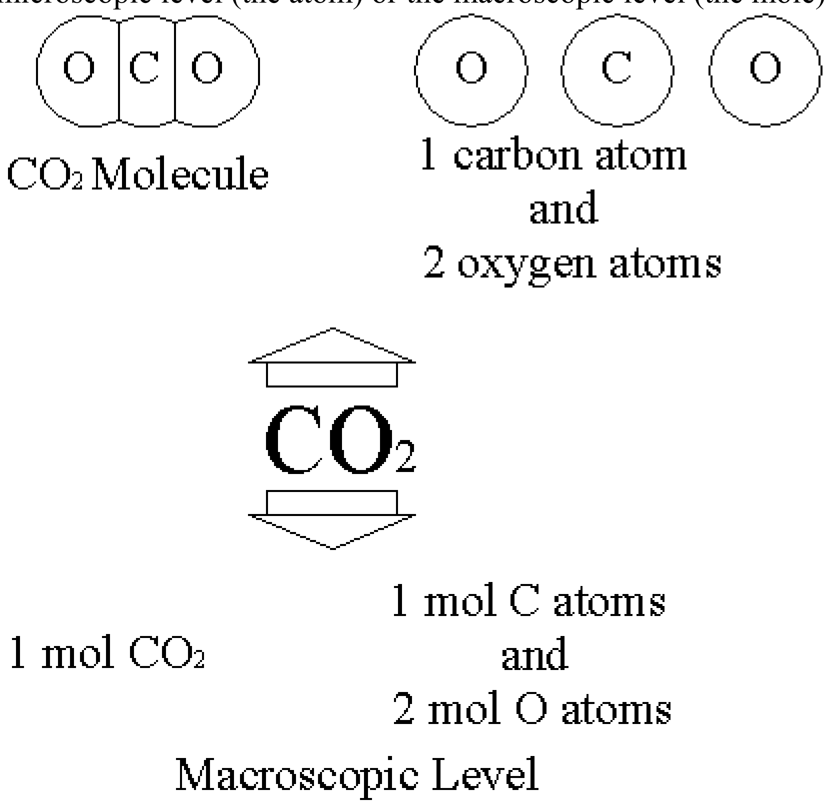
350 g C_2H_6 - 70 g H

20.2 g $NaHSO_4$ – 0.17 g H

2.14 g NH_4Cl – 0.16 g H

Calculate the grams of nitrogen in the 125 g of each fertilizer:
$CO(NH_2)_2$ – 58.4 g N
NH_3 – 103g N
NH_4NO_3 – 43.8 g N

- Calculating Empirical Formulas
  - The empirical formula gives the lowest whole number ratio of atoms in a compound – the empirical formula for H_2O_2 would be HO – As seen, the empirical formula may or may not be the same as the molecular formula.
  - The empirical formula tells us about the kinds of atoms and the relative count of atoms, or moles of atoms in a molecule or formula unit
  - As noted above, empirical formula can be interpreted at the microscopic level (the atom) or the macroscopic level (the mole).



  - If you use the mole, the empirical formula is the lowest whole number ratio of moles that combine to form a compound
  - Again, the empirical formula may or may not be the same as the molecular formula – if is not, then the molecular formula is a simple multiple of the empirical formula
  - So how do you calculate an empirical formula – you use percent composition!

Sample Problem

What is the empirical formula of a compound that is 25.9% nitrogen and 74.1% oxygen?

Known	*Unknown*
% of nitrogen = 25.9% N	empirical formula = $N_?O_?$
% of oxygen = 74.1% O	

Strategy:
The lowest whole number ration of moles of nitrogen atoms to oxygen atoms, or empirical formula is to be calculated. The percent composition tells the ration of masses of nitrogen atoms to oxygen atoms in the compound. The ratio of masses is changed to a ratio of moles by using the conversion factors based on the molar mass of each element. This mole ratio is then reduced to the lowest whole number.

According to the percent composition, in 100.0 g of the compound there is 25.9 g of nitrogen and 74.1 g of oxygen. These values are used to convert to moles:

$$mass\ of\ element\ x\ \frac{1\ mol\ element}{gam\ of\ element} = mol\ of\ element$$

$$25.9\ g\ N = \frac{1\ mol\ N}{14.0\ g\ N} = 1.85\ mol\ N$$

$$74.1\ g\ O = \frac{1\ mol\ O}{16.0\ g\ O} = 4.63\ mol\ O$$

So the mole ratio of nitrogen to oxygen is $N_{1.85}O_{4.63}$ – which is not a lowest whole number ratio. So we have a little more work to do. To get closer, we should divide each of the values obtained by the smaller of the two values:

$$\frac{1.85\ mol\ N}{1.85} = 1\ mol\ N$$

$$\frac{4.63\ mol\ O}{1.85} = 2.50\ mol\ O$$

So we are closer, but we still do not have a lowest whole number ratio – to obtain the lowest whole number ratio, we multiply both number by the smallest number that will give a result as a whole number ratio – start with 2 and work you way up!

$$1.00\ mol\ N\ x\ 2 = 2.00\ mol\ N = 2\ mol\ N$$

$$2.50\ mol\ O\ x\ 2 = 5.00\ mol\ O = 5\ mol\ O$$

So the empirical formula is N_2O_5.

Calculate the empirical formula of each compound
94.1% O, 5.9% H – OH
79.8% C, 20.2% H – CH_3
67.6% Hg, 10.8% S, 21.6% O – $HgSO_4$
27.59% C, 1.15% H, 16.09% N, 55.17% O – C_2HNO_3

1,6-diaminohexane is used to make nylon. What is the empirical formula given 62.1% C, 13.8% H, and 24.1% N? C_3H_8N

- Calculating Molecular Formulas

Formula (name)	Classification of formula	Molar mass
CH	Empirical	13
C_2H_2 (ethyne)	Molecular	26 (2 x 13)
C_6H_6 (benzene)	Molecular	78 (6 x 13)
CH_2O (methanal)	Emperical and molecular	30
$C_2H_4O_2$ (ethanonic acid)	Molecular	60 (2 x 30)
$C_6H_{12}O_6$ (glucose)	Molecular	180 (6 x 30)

- From the above table we see that we have two groups of 3 compounds. The first group contains carbon and hydrogen, the second group contains carbon, hydrogen and oxygen
- Also note that each compound in a group has a number of atoms that is a whole number ratio, thus a different molar mass
- Just as we could determine the empirical formula of a compound, we can also determine the molecular formula of a compound if we know the empirical formula and the molar mass of the molecular compound
    - From the empirical formula, we can calculate the empirical formula mass (efm) – just the molar mass of the empirical formula
    - The known molar mass of the compound is then divided by the empirical molar mass – this gives the number of empirical formula units in a molecule of the compound, and is the multiplier to convert the empirical formula to a molecular formula

Sample Problems

Calculate the molecular formula of the compound whose molar mass is 60.0 g and has an empirical formula of CH_4N.

Known	*Unknown*
Empirical formula = CH_4N	molecular formula = ?
Molar mass = 60.0g	
EFM = 12.0 + 4.0 + 14.0 = 30.0g	

$$multiplier = \frac{molar\ mass\ of\ compound}{efm} = \frac{60.0\ g}{30.0\ g} = 2$$

so 2 x C, 2 x H_4, and 2 x N = $C_2H_8N_2$

Find the molecular formula of each compound give its empirical formula and molar mass

Ethylene glycol (CH_3O), used in anti-freeze, molar mass = 62 g/mol – $C_2H_6O_2$

p-dichlorobenzene (C_3H_2Cl) (mothballs), molar mass = 147 g/mol – $C_6H_4Cl_2$

Which pairs of molecules have the same empirical formula?

a. $C_2H_4O_2$, $C_6H_{12}O_6$ - yes
b. $NaCrO_4$, $Na_2Cr_2O_7$ - no