MASS RELATIONS and STOICHIOMETRY

(Mostly **Chapter 3** – Part II)

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Atomic Mass

The ***atomic mass*** of a hydrogen atom is **1.0079** (from the P.T.)

- *No unit?* What is the unit for atomic mass?
  - Kg? g? ng? (None of these)

- Atomic masses are unitless. It is because they are **relative masses** where the mass of one atom is compared to the mass of carbon-12 isotope

Atomic mass of C-12 atom = 12.00 amu
(assigned mass)

- *Conversely*

1 amu = 1/12 mass of C-12 atom = **mass of H-1 atom**

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Atomic Mass - *Cont.*

- *Atoms are very small:*
  - Hard to measure individual mass
  - Measured as a bunch. (Even the mass spectrometer relies on a bunch of atoms!)
- Use *same sized bunch* of two different atoms, say H and O, then compare their masses.

Mass of 1 million O atoms	26.58×10^{-18} g
Mass of 1 million H atoms	1.675×10^{-18} g

- Divide both masses by the mass of 1 million H atoms (smaller atom) --- gives ***relative mass*** of O to H

Answer: **Relative mass of O to H: 15.89 amu**

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Atomic Mass - *Cont.*

- Recall: **Isotopes** = forms of the same element with different number of neutrons (different mass numbers)
- ***Relative atomic mass*** of <u>isotopes</u> takes into account the *relative mass* and *abundance* of each isotope

Example: Naturally occurring Cl consists of atoms of *relative isotopic masses* of 34.97 (with 75.53 % abundance) and 36.97 (with 24.47 % abundance). What is the relative atomic mass of Cl?

Rel. atomic mass of Cl = (0.7553 x 34.97) + (0.2447 x 36.97)
= **35.46 amu**

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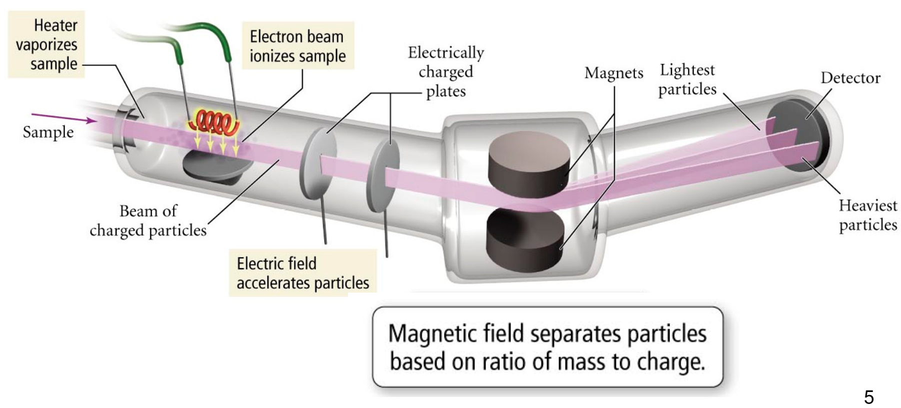
Atomic Mass - *Cont.*

How are atomic masses measured? (After all, atoms are extremely small)

- Using an instrument called ***mass spectrometer***



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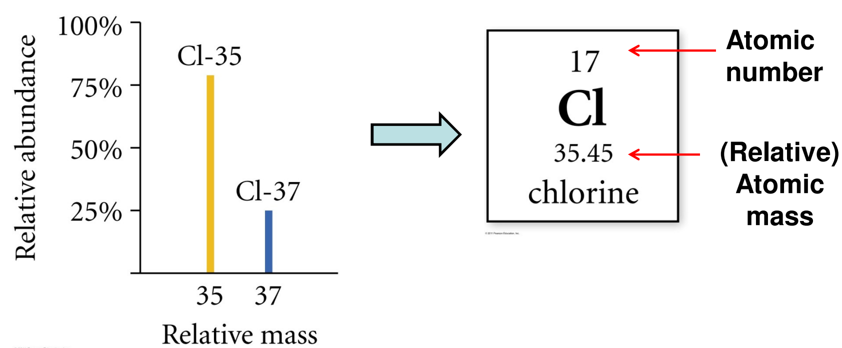
Mass Spectrometer - *Cont.*

In the mass spectrometer *vaporized atoms* are bombarded with fast-moving electrons, which knock out an electron from the atom. The cations formed pass through a magnet, which separate them according to their mass. The instrument provides a measure of the relative mass (compared to ^{12}C) and the relative number of each isotope.

The **mass spectrum** of naturally occurring chlorine.



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Atomic Mass vs. Atomic Weight

What is the difference between (relative) atomic mass and atomic weight?

- These terms are used interchangeably, although the former is a more accurate term to use. Why?
  - Because mass does not change regardless of location (whereas weight changes depending on gravity).

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Avogadro's Number

Fact: A sample of any element with a mass in grams equal to its atomic mass contains the same number of atoms, regardless of the element's identity

- In 1811 an Italian chemist, **Amedeo Avogadro** (1776-1856), determined that equal volumes of gases at the same temperature and pressure contain the same number of molecules regardless of their chemical nature and physical properties. This number (Avogadro's number) is **6.022 X 10^{23}**.

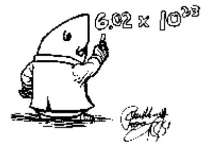
http://www.chemistry.co.nz/avogadro.htm

Avogadro's number (N_A) = the number of atoms present in a sample whose mass in grams is *numerically equal to the atomic mass* of the element, also called a mole.

http://gemini.tntech.edu/~tfurtsch/scihist/avogadro.htm

Avogadro's number:
6.022 x 10^{23} particles = 1 mole

The Mole



The term *mole* literally means a small mass.

- If the relative mass of a single ^{12}C atom is 12.000 amu, then *one mole of ^{12}C atoms would have a mass of 12.000 grams.*
- A **mole** (***n***) of any substance *contains the same number* of particles as there are atoms in exactly 12 grams of the ^{12}C isotope of carbon. In other words, *one mole of a substance contains an Avogadro's number of particles!*

1 mole, *n* = 6.022×10^{23} atoms or particles

Analogy: **1 dozen = 12** items of anything!

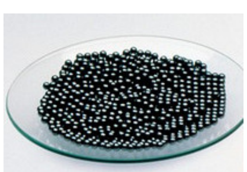
The mole is like a chemist's dozen!

The *mole*: A chemist's dozen

1 dozen eggs
12 eggs

1 dozen pears
12 pears

1 dozen apples
12 apples



1 mole lead (Pb) shots
6.022×10^{23} Pb atoms

1 mole mercury (Hg)
6.022×10^{23} Hg atoms

1 mole sulfur (S)
6.022×10^{23} S atoms

Molar mass of various substances. Image available at fphoto.photoshelter.com

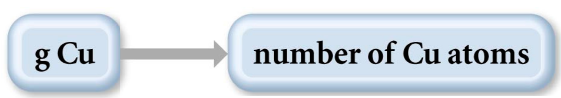
Number of Moles and Avogadro's Number

$$\frac{6.022 \times 10^{23}\ \text{Cu atoms}}{1\ \text{mol Cu}}$$

1 mole Cu = 6.022 x 10^{23} Cu atoms

Atomic Mass and Avogadro's Number

Q. How many atoms are in (a) 63.55 g of Cu and (b) 8.006 g of helium?



Answer: Cu's given mass equals its atomic mass, so it contains **6.022 x 10^{23} Cu atoms;** He's given mass represents **2** x 6.022 x 10^{23} or **1.044 x 10^{24} He atoms**

Work:

Atomic Mass and Avogadro's Number – *Cont.*

- One atomic mass in *grams* of an element contains an Avogadro's number of atoms, or 6.022 x 10^{23} atoms. This gram mass is also called ***molar mass***.

Example: How many atoms are in 10 g of sulfur?

Answer: 1.88 x 10^{23} S atoms

Work:

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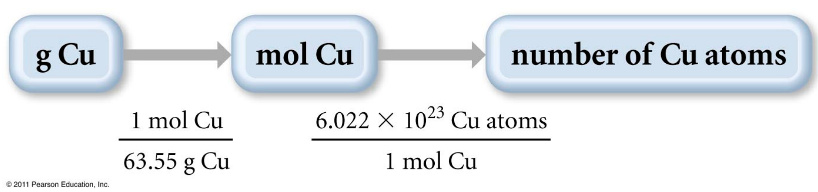
Mass, Mole and Avogadro's Number – *Cont.*



1 mole of Cu atoms = 6.022 x 10^{23} Cu atoms = 65.55 g Cu

2 moles of O atoms = 12.04 x 10^{23} O atoms = 32.00 g O

1 mole of H_2O molecules = 6.022 x 10^{23} H_2O molecules = 18.02 g* H_2O

1 mole of O_2 molecules = 6.022 x 10^{23} O_2 molecules = 32.02 g* O_2

**** Where do these masses come from?***

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Mole – *Cont.*

Thus far you have learned of the following relationships:

1 mole = 6.022 x 10^{23} atoms = 1 molar mass (in *g*) of the atom

However, a *mole* (much like a dozen) represents 6.022 x 10^{23} items of anything (atoms, ions, molecules)

Thus, for molecules, we can write the relationships:

1 mole = 6.022 x 10^{23} molecules = 1 molar mass (in g) of the compound

Note: Recall that covalent or molecular compounds consist of molecules

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1 dozen eggs
mass = 2 lbs

1 dozen pears
mass = 6 lbs

1 dozen apples
mass = 4 lbs

1 mole lead (Pb) shots
mass = 207.2 g

1 mole sulfur (S)
mass = 32.065 g

1 mole sugar ($C_{12}H_{22}O_{11}$)
mass = 342.3 g

1 mole mercury (Hg)
mass = 200.61 g

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Molar mass of various substances. Image available at fphoto.photoshelter.com

Molar Mass

The ***molar mass*** (or ***molecular weight***), **MM**, of a substance refers to the mass in grams of one mole of that substance (usually an atom or a molecule)

- *For atoms*: Molar mass equals its atomic mass in grams
- *For molecules*: Molar mass is determined by adding up the atomic masses of all the atoms in the chemical formula

Exercise: Calculate the MM of (a) N_2O_4 and (b) $Ca(NO_3)_2$.

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Mass and Moles

Converting Moles into Grams

Converting the moles of a substance, *n*, to mass in grams, *g*, and vice versa is often necessary. Why?

- Chemical reactions are based on moles, not on mass
- A mole, unlike mass, is not directly measurable
- Thus, we need to be able to convert mole to mass in grams (see earlier examples)

mass in grams: $g \text{ mass} = n \times \text{MM}$

# of moles (n) — **Molar mass (g/mole)** (MM)

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Converting Moles to Grams

Exercise: Calculate the mass in grams of:

(a) 2 moles of C_2H_2
(b) 0.52 mole of iron, and
(c) 1.5×10^{-2} mole Na_2CO_3

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Converting Grams to Moles

$$\text{moles} = \frac{\text{grams}}{\text{MM}}$$

Exercise: Calculate the number of moles in:

(a) 35.25 g C_2H_2
(b) 100 g of Fe

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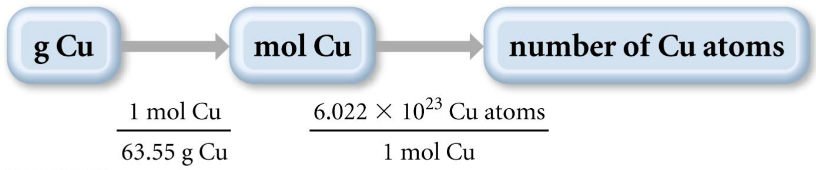
Summary: Mass, Mole and Avogadro's Number



g H_2O → **mol H_2O** → **Number of H_2O molecules**

$$\frac{1 \text{ mol } H_2O}{18.01 \text{ g } H_2O} \qquad \frac{6.022 \times 10^{23} \; H_2O \text{ molecules}}{1 \text{ mol } H_2O}$$

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Percent Composition & Molecular Formula

The mass ***percent composition*** of a compound is specified by citing the mass percents of the elements present.

- The mass % composition of water, H_2O, is 11.19 % H and 88.81 % O. *How is this determined?*

$$\text{Percentage} = \frac{\text{part}}{\text{whole}} \times 100\%$$

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Calculating % Composition

<u>Step 1</u>: Calculate the *molar mass* of the compound

<u>Step 2</u>: Calculate the *% of each element* from the number of atoms per chemical formula, atomic mass and the molar mass.

Exercise: Determine the mass percent composition of (a) $CaCl_2$ and (b) Na_2CO_3.

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Calculating % Composition - *Cont.*

- $CaCl_2$

$$\text{Mass \% Ca} = \frac{\text{molar mass Ca}}{\text{molar mass CaCl}_2} \times 100\%$$

$$\text{Mass \% Cl} = \frac{2 \times \text{molar mass Cl}}{\text{molar mass CaCl}_2} \times 100\%$$

$$2 \times \text{molar mass Cl} = 2(35.45\ \text{g/mol}) = 70.90\ \text{g/mol}$$

$$\text{molar mass CaCl}_2 = 1(40.08) + 2(35.45) = 110.98\ \text{g/mol}$$

$$\text{Mass \% Ca} = \frac{40.08\ \text{g/mol}}{110.98\ \text{g/mol}} \times 100\% = 36.11\%$$

$$\text{Mass \% Cl} = \frac{70.90\ \text{g/mol}}{110.98\ \text{g/mol}} \times 100\% = 63.88\%$$

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Mass Relations in Chemical Reactions: STOICHIOMETRY

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Stoichiometry

The word *stoichiometry* derives from two Greek words: *stoicheion* (meaning "element") and *metron* (meaning "measure")

Stoichiometry deals with calculations of the masses of reactants and products involved in a chemical reaction.

Importance *of Stoichiometry* - it allows chemists to determine:

- The amount of raw materials needed to produce a certain product
- Yield of a certain product
- The number of molecules are in a microgram of biomolecules, such as the neurotransmitter *serotonin*

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Stoichiometry and Chemical Reactions

Recall from Chapter 2 that *atoms are never created or destroyed in a chemical reaction*.

Consequence: The number of atoms which were present before the reaction must be present after the reaction.

A chemical equation which meets this criterion is said to be **balanced**.

Stoichiometry is often *used to balance chemical equations*. For example, the two diatomic gases hydrogen and oxygen can combine to form a liquid, water, in an exothermic reaction, as described by the equation:

$$H_2 + O_2 \rightarrow H_2O$$

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Wait ! The number of O atoms on the left side and the right side of the equation is not the same => the equation is not balanced!

Stoichiometry and Chemical Reactions – *Cont.*

Before you balance chemical equations, you have to understand how these equations are written.

Writing chemical equations:

1. Write the formulas of reactants on the left side and products on the right side of the equation.
2. Use an arrow to separate reactants from products (The arrow means "to produce" or "to yield")
3. Indicate the physical state of each reactant and product after the formula.
   - (g) for gases, (l) for liquids, (s) for solids and (aq) for ions or molecules dissolved in water, called *aqueous* solutions.

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Writing Chemical Reactions

Exercise: Write an equation for the reaction where:
(1) ethylene gas, C_2H_4, combines with oxygen gas to produce carbon dioxide gas and water vapor, *and*
(2) solid potassium chlorate decomposes to form solid potassium chloride and oxygen gas.

Is this what you got?

(1) $C_2H_{4\,(g)} + O_{2\,(g)} \longrightarrow CO_{2\,(g)} + H_2O$

(2) $KClO_{3\,(s)} \longrightarrow KCl_{(s)} + O_{2\,(g)}$

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Balancing Chemical Reactions

To balance a chemical reaction:

1. Count *how many of each type of atoms* are on each side of the equation

Example: $C_2H_{4\,(g)} + O_{2\,(g)} \rightarrow CO_{2\,(g)} + H_2O$

Reactants	Products
2 C	**1 C**
4 H	**2 H**
2 O	**3 O** (= 2 from CO_2 + 1 from H_2O)

2. *Place coefficients* in front of the appropriate reactant or product to balance the number of atoms on both sides of the equation.

- NOTE: A coefficient of 1 is understood but is not shown.
- In the example above, we can balance the # of C's first by placing a coefficient in front of CO_2

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Balancing Reactions – *Cont.*

$$C_2H_4\,_{(g)} + O_2\,_{(g)} \rightarrow \mathbf{2}\,CO_2\,_{(g)} + H_2O$$

2 C	2 C
4 H	2 H
2 O	? O

- Then we balance H by placing a coefficient before H_2O

$$C_2H_4\,_{(g)} + O_2\,_{(g)} \rightarrow \mathbf{2}\,CO_2\,_{(g)} + \mathbf{2}\,H_2O$$

3. *Update your inventory of atoms*. If all the numbers are balanced, then the equation is balanced. Otherwise, adjust the coefficients until all the atoms are balanced.

$$C_2H_4\,_{(g)} + O_2\,_{(g)} \rightarrow \mathbf{2}\,CO_2\,_{(g)} + \mathbf{2}\,H_2O$$

2 C	2 C
4 H	4 H
2 O	6 O = 4 from CO_2 and 2 from H2O

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Balancing Reactions – *Cont.*

- O is the only atom left to be balanced (6 O on the right), so we write:

$$C_2H_4\,_{(g)} + \mathbf{3}\,O_2\,_{(g)} \rightarrow \mathbf{2}\,CO_2\,_{(g)} + \mathbf{2}\,H_2O$$

3 x 2 = 6 O's!

4. *Check your answer*. Are all the atoms balanced? If so, you're done!

$$C_2H_4\,_{(g)} + \mathbf{3}\,O_2\,_{(g)} \rightarrow \mathbf{2}\,CO_2\,_{(g)} + \mathbf{2}\,H_2O$$

2 C	2 C
4 H	4 H
6 O	6 O

Balanced!

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Balancing Reactions – What Not to Do

The first NO-NO when balancing chemical equations is:

Never change the subscripts in a chemical formula in order to balance an equation.

$$H_2 + O_2 \longrightarrow H_2O_2$$

NOT ALLOWED!
When this subscript is added, we get a completely different reaction.

http://wps.prenhall.com/wps/media/objects/602/616516/Chapter_03.html

Changing water, **H_2O**, to **H_2O_2** changes the chemical species! H_2O_2, or hydrogen peroxide, is NOT the same as water.

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Balancing Chemical Reactions - *Cont.*

Another thing to remember about chemical equations is that:

The coefficients in a chemical equation represent ratios and <u>not</u> the actual amounts or reactants and/or products present.

Exercises:

Balance the following equations:

1. $NaCl + MgF_2 \rightarrow NaF + MgCl_2$
2. $FeCl_3 + Ca_3(PO_4)_2 \rightarrow CaCl_2 + FePO_4$
3. $AgNO_3 + Be(OH)_2 \rightarrow AgOH + Be(NO_3)_2$
4. $Mg + Mn_2O_3 \rightarrow MgO + Mn$

NOTE: When ***polyatomic ions*** are present, balance them as a group instead of individual atoms. *Example*: Balance C & O in $Al_2(CO_3)_3$ as "CO_3"

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Are these your answers?

1. $2\ NaCl + MgF_2 \rightarrow 2\ NaF + MgCl_2$
2. Step 1: $FeCl_3 + Ca_3(PO_4)_2 \rightarrow CaCl_2 + 2\ FePO_4$

   Step 2: $2\ FeCl_3 + Ca_3(PO_4)_2 \rightarrow CaCl_2 + 2\ FePO_4$

   Step 3: $2\ FeCl_3 + Ca_3(PO_4)_2 \rightarrow 3\ CaCl_2 + 2\ FePO_4$
3. $2\ AgNO_3 + Be(OH)_2 \rightarrow 2\ AgOH + Be(NO_3)_2$
4. $Mg + Mn_2O_3 \rightarrow MgO + 2\ Mn$

   $Mg + Mn_2O_3 \rightarrow 3\ MgO + 2\ Mn$

   $3\ Mg + Mn_2O_3 \rightarrow 3\ MgO + 2\ Mn$

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