CHEM 205 section 03

LECTURE #8 Tues. Jan.29, 2008

ASSIGNED READINGS:

TODAY'S CLASS: finish Ch.3

NEXT CLASS: **most of Ch.4**

(1)

3.5 Formulae & molar masses of compounds

- **Molar mass (MM)** = mass (g) of one mole of substance
  - For ionic compounds: MM also called formula weight, F_w
  - For molecular cmpds: MM also called molecular weight, M_w
- **To calculate MM**: must know chemical formula
  - Consider mole ratio of elements given in formula
  - Subscripts indicate #moles of element per mole of compound
  - Add up atomic masses (mass per mole of atoms) for each element…taking into account #moles of each in compound…

$CaCl_2$ *MM = (1 mol x 40.078 g/mol Ca) + (2 mol x 35.453 g/mol Cl)*
= 110.984 g/mol

Caffeine $C_8H_{10}N_4O_2$ →

MM = 194.201 g/mol

(2) **NEXT: How to experimentally determine formulas…**

3.6 Formulae of compounds: describing composition

$$\text{Mass \%} = \frac{\text{mass}_{element\text{-}A} \text{ in sample}}{\text{mass}_{total} \text{ of sample}} \times 100\%$$

experiment ⇩ ⇧ calculate

$$\text{Mole \%} = \frac{\text{moles}_{element\text{-}A} \text{ in sample}}{\text{moles}_{total\text{-}atoms} \text{ in sample}} \times 100\%$$

⇦ ⇨ Formula

If formula known: Iron in iron (III) oxide, Fe_2O_3

- MM = (2 x 55.845g/mol Fe)+ (3 x 15.999g/mol O) = 159.69 g/mol
- mass % Fe: $\frac{\text{111.69 g Fe per mole } Fe_2O_3}{\text{159.69 g total per mole } Fe_2O_3} \times 100\% = 69.94\%$
- mole % Fe: $\frac{\text{2 moles Fe per mole } Fe_2O_3}{\text{5 moles (Fe+O) per mole } Fe_2O_3} \times 100\% = 40\%$

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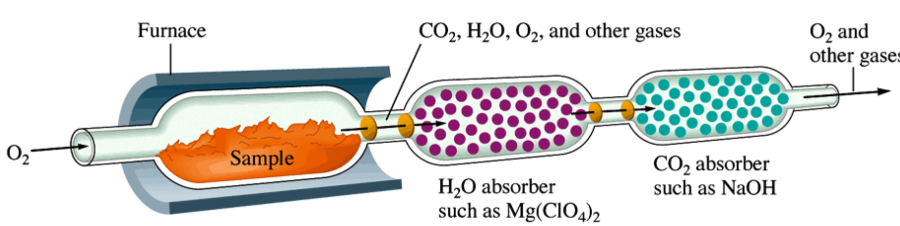
How do we find elemental composition by exp't?

Exploit a known reaction
- start with known mass of reactant
- measure mass of products formed
- deduce % mass of each element in reactant

See details in Ch.4

Commonly used for organic compounds (contain mostly C&H):

- Burn organic sample in O_2: converts C to CO_2, H to H_2O
- Determine mass ↑ of absorbers



Zumdahl's Fig. 3.5: A schematic diagram of a combustion analyzer

EXAMPLE: What is the formula for Aspirin?

Given: Elemental Analysis data: 60.0% C, 4.4% H, 35.6% O
Molar mass measurement (for expt, see Chem206): ~ 180 g/mol

Mass %	Mass in 100 g	MM (g/mol)	Moles	Ratio	Whole # ratio
60.0% C	60.0 g	12.01	5.00	2.24	8.96
4.4% H	4.4 g	1.00	4.4	1.97	7.88
35.6% O	35.6 g	16.00	2.23	1.00	4.00

Sample calculations, for carbon:

Mass of C in 100g sample:
60.0% x 100.0g
= 0.600 x 100.0g
= 60.0 g

Moles of C:
$\frac{60.0g}{12.01\ g/mol}$
= 5.00 mol

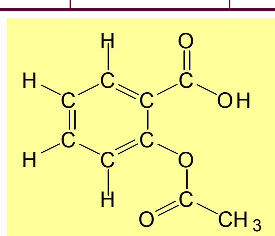
Ratio: Normalize to least abundant element
$\frac{\#mol\ C}{\#mol\ O} = \frac{5.00}{2.23}$
= 2.24

Whole # ratio: multiply all ratios by same **integer coefficient**, to convert all to whole #s

EXAMPLE: What is the formula for Aspirin?

Given: Elemental Analysis data: 60.0% C, 4.4% H, 35.6% O
Molar mass measurement: approx. 180 g/mol

Mass %	Mass in 100 g	MM (g/mol)	Moles	Ratio	Whole # ratio
60.0% C	60.0 g	12.01	5.00	2.24	8.96
4.4% H	4.4 g	1.00	4.4	1.97	7.88
35.6% O	35.6 g	16.00	2.23	1.00	4.00



Rounds to: $C_9H_8O_4$

Empirical MM = 180.17 g/mol
Matches experimental MM
∴ Molecular formula = $C_9H_8O_4$

Empirical Formula Determination: Strategy

Given: elemental composition data
Goal: find mole ratio of elements present

1. Determine mass of each element in your sample.
   - if given mass % data, but no sample mass: use 100 g
   - if given mass of each product formed: see Ch.4…
2. Determine #moles of each element in your sample.
3. Normalize the mole data:
   divide each mole value by the least abundant element's value
   ⇨ gives MOLE RATIO of elements, relative to one of them.
4. Scale to whole numbers:
   multiply each normalized mole value by the smallest integer (SAME FOR ALL!) that yields a whole number for each element.

⇨ RESULT: smallest whole number mole ratio of elements
= empirical formula.

EASIEST TO USE A TABLE.

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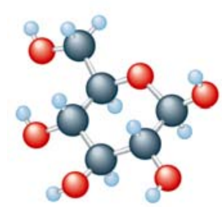
Empirical formula (= simplest) *vs.* Molecular formula

Glucose:
molecular formula: $C_6H_{12}O_6$ 180.158 g/mol
empirical formula: CH_2O 30.026 g/mol



molecular formula = (empirical formula)$_n$ *[n = an integer]*

- Experimental mass % data ⇨ empirical formula
  - Fine for ionic compounds… *BUT…*
  - Molecular formulae NOT ALWAYS same as empirical
- Need: experiment to measure MM *(see Chem 206)*
- Calculate ratio: $\dfrac{\text{MM (measured by experiment)}}{\text{MM (empirical formula weight)}}$
- Multiply empirical formula by ratio to find molecular formula

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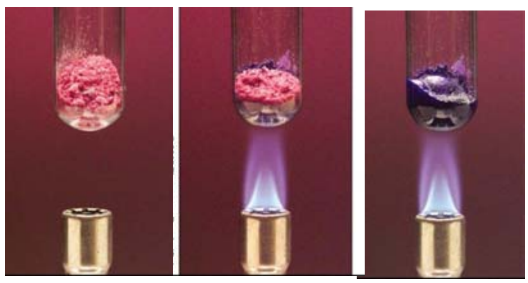
3.7 Hydrated compounds

- Many compounds can have water molecules become trapped inside their crystal lattice
  - Common for ionic solids isolated from aqueous solution
  - Some substances asborb H_2O from air: called "hygroscopic"
  - # of H_2O's trapped per mole of substance varies for different substances: represented by $M_mE_n \cdot xH_2O$ ← *a hydrated compound*
  - "Waters of hydration" can be driven off by heating (*i.e.,* can find x…)

RED/PINK **$CoCl_2 \cdot xH_2O$**
Hydrated cobalt(II) chloride



BLUE **$CoCl_2$**

"Anhydrous" cobalt(II) chloride

How many waters of hydration were there?
(must do expts…)

(9) Figure 3.17

EXAMPLE: What is x in $CoCl_2 \cdot xH_2O$?

Given: Elemental Analysis data: 24.77% Co, 29.80% Cl

Rest must be H_2O!

Mass %	Mass in 100 g	MM (g/mol)	Moles	Ratio	Whole # ratio
24.77% Co	24.77	58.93	0.4203	1.000	1
29.80% Cl	29.80	35.45	0.8406	2.000	2
45.43% H_2O	45.43	18.02	2.521	6.000	6

Formula: $CoCl_2 \cdot 6H_2O$
Cobalt(II) chloride hexahydrate

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Determining how many waters of hydration: by drying

Sample final exam question *work must be shown…*

A 4.450 g sample of hydrated lithium iodide, $LiI \cdot x\, H_2O$, is dried in an oven. When the anhydrous salt is removed from the oven, its mass is 3.170 g. What is the value of x?

Ans: 3

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ASSIGNED READINGS

- BEFORE NEXT CLASS:

  Read rest of Ch. 3 & Ch. 4 sections 4.1-4.2

  Master Ch.1-3 material & exercises

- Practice: naming/formulae of compounds
  determining mass % & formulae
  balancing chemical equations

(12)