Chapter 1 Fundamental Concepts

Fundamental definitions:

- **Atoms, elements, compounds, molecules, chemical formulas…**

Scientific measurements

- **SI base units, SI prefixes, conversion factors, precision versus accuracy, sig figs (covered in appendices A, B and G in hard copy of text)**

Chemical problem solving (aka the scientific method)

First DGDs:

Sept 11th, 13th

Review of sig figs, unit conversions, chem problem solving strategies and nomenclature rules

Note: Prep for Success (Sapling) can help with review

Système International (SI) Base Units

Table 1.2	Base SI Units	
Quantity	**Unit**	**Symbol**
Mass	kilogram	kg
Length	metre	m
Time	second	s
Temperature	kelvin	K
Amount	mole	mol
Electric current	**ampere**	A

Frequently Used SI Prefixes

Table 1.1 Frequently Used Prefixes for Magnitudes

Prefix	Symbol	Number	Exponential Notation
atto	a	0.000 000 000 000 000 001	10^{-18}
femto	f	0.000 000 000 000 001	10^{-15}
pico	p	0.000 000 000 001	10^{-12}
nano	n	0.000 000 001	10^{-9}
micro	μ	0.000 001	10^{-6}
milli	m	0.001	10^{-3}
centi	c	0.01	10^{-2}
kilo	k	1000	10^{3}
mega	M	1 000 000	10^{6}
giga	G	1 000 000 000	10^{9}
tera	T	1 000 000 000 000	10^{12}

Conversion Factors

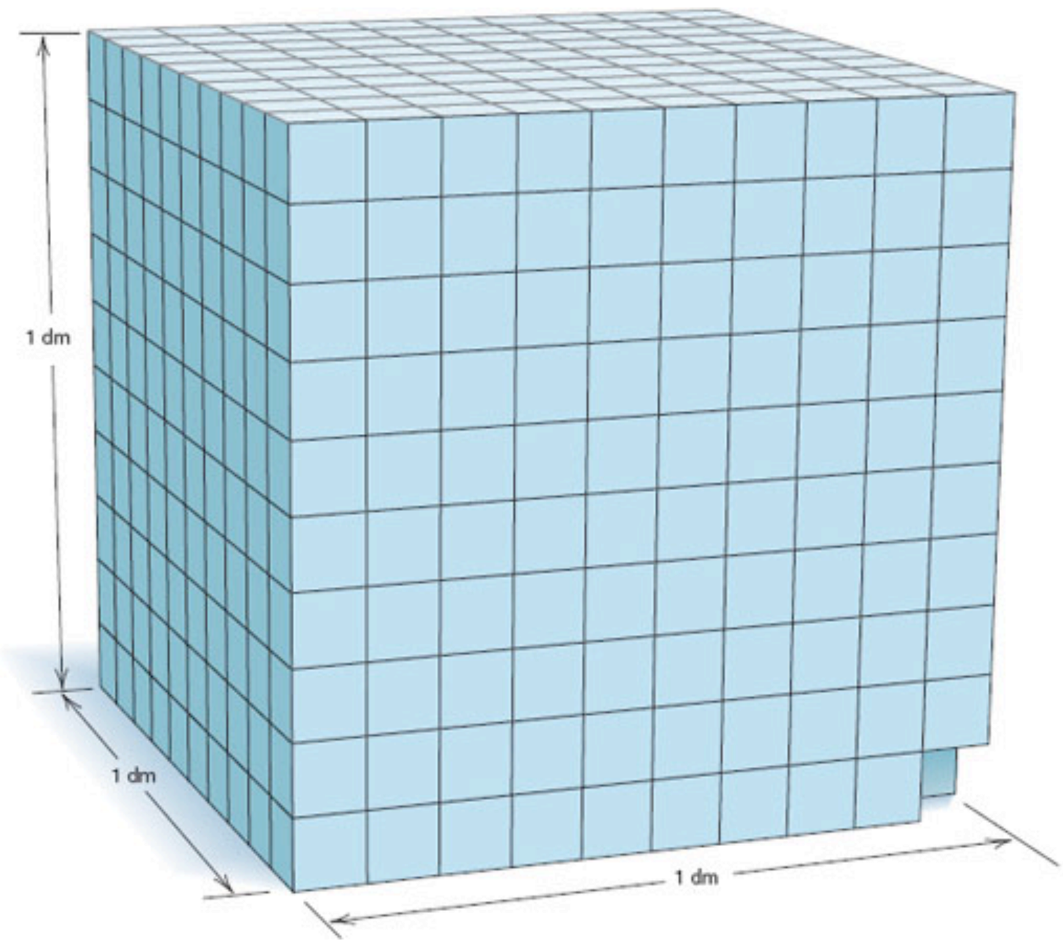
Converts a value from one unit to another in…

Dimensional Analysis

Quantity with desired units	=	Quantity with given units	x	Conversion Factor

For example, how many ps are in 1 min?

Units of Volume



Some volume equivalents:

$1\ m^3 = 1000\ dm^3$

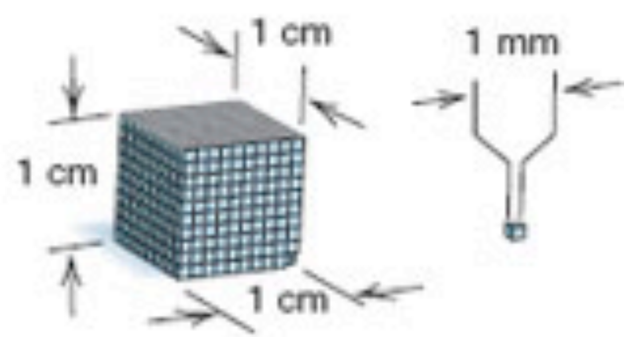
$1\ dm^3 = 1000\ cm^3$

$= 1\ L = 1000\ mL$

$1\ cm^3 = 1000\ mm^3$

$= 1\ mL = 1000\ \mu L$

$1\ mm^3 = 1\ \mu L$



Converting Units of Volume

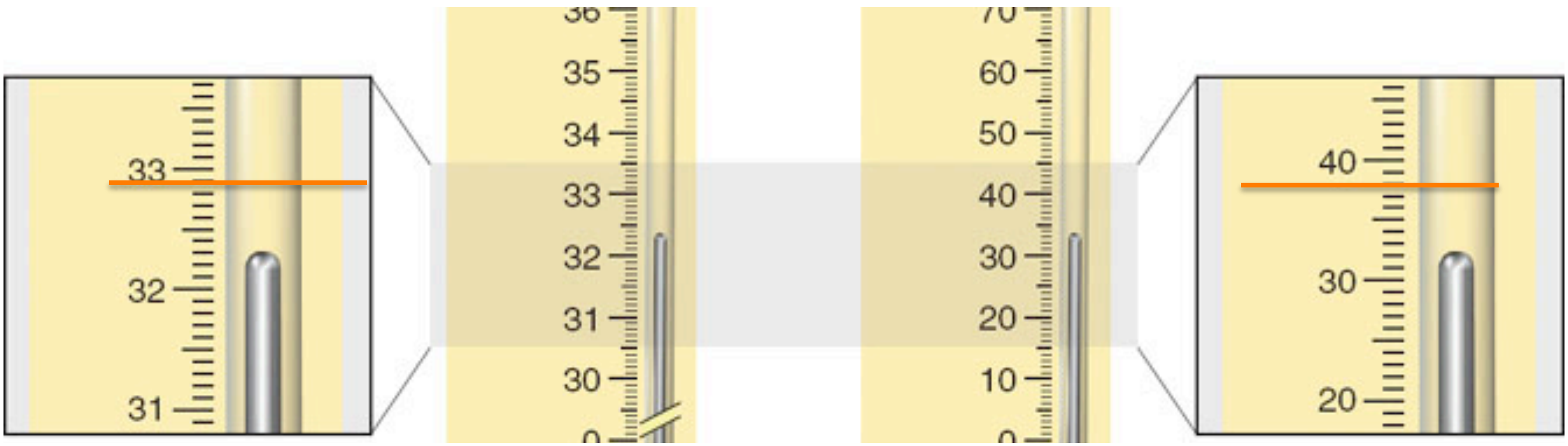
Within a cell, proteins are synthesized on particles called ribosomes. Assuming ribosomes are spherical, what is the volume (in L) of a ribosome whose average radius is 10.7 nm? The equation for the volume of a sphere is:

$$V = \frac{4}{3}\pi r^3$$

Uncertainty

All measurements have uncertainty
- the last digit is the uncertain digit
- the number of sig figs in a measurement includes both certain and uncertain digits!

What are these 2 thermometer readings? # of sig figs?



Significant Digits in the Lab

The specifications for a 100 mL volumetric flask state that it is guaranteed to contain the specified volume to within 0.1 mL. How many significant figures should you use to record the volume contained in the flask?

a) 1
b) 2
c) 3
d) 4
e) The measurement would be an exact number because a volumetric flask is used.

Significant Figures in Calculations

Addition and Subtraction:

The answer has the same number of decimal places as there are in the measurement with the <u>fewest decimal places.</u>

Multiplication and Division:

The answer contains the same number of significant figures as there are in the measurement <u>with the fewest significant figures.</u>

$$\begin{array}{r} 2.06 \text{ mL} \\ -\ 1.1 \text{ mL} \\ \hline 0.96 \text{ mL} \end{array}$$

$$2.2 \times 3.7845 = 8.32590$$

Sig. Figs and Rounding

In calculations carry through one or two extra significant digits. Round up if first digit to be eliminated is 5 or higher, round down if this digit is less than 5.

For example:

$$\frac{25.65 \text{ mL} + 37.4 \text{ mL}}{73.55 \text{ s}\left(\frac{1 \text{ min}}{60 \text{ s}}\right)}$$

Echo360 Question:

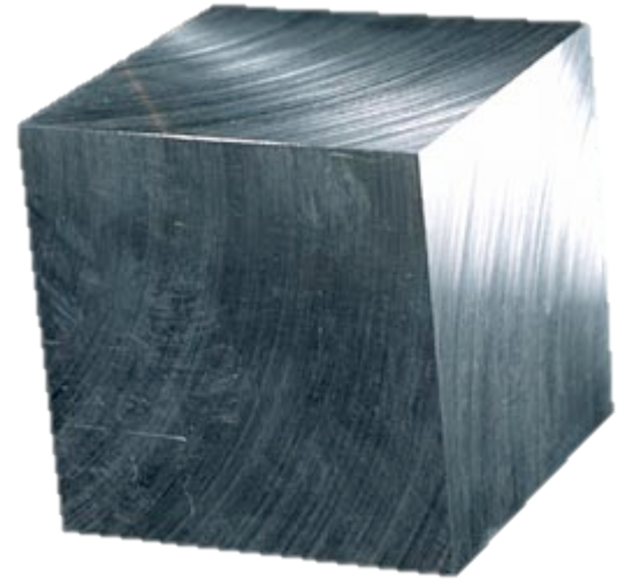
The result of (3.8621 x 1.5630) - 5.98 is properly written as:

a) 0.06 b) 0.056 c) 0.0565 d) 0.05646 e) 0.056462

Chemical Problem Solving

1. What are you being asked to find?
2. Visualize the problem.
3. Organize the data.
    - What data is given and how is related to the value you are trying to calculate?
4. Identify a process to solve the problem.
    - Which equations do you need?
5. Manipulate the equations.
6. Substitute and calculate.
    - Keep track of units and sig figs.
7. Check it – does the result make sense?

A small cube of aluminum measures 15.6 mm on a side and has a mass of 10.25 g. What is the density of aluminum in g/cm^3?



a) 2.70×10^{-3} g/cm^3

b) 0.657 g/cm^3

c) 2.70 g/cm^3

d) 657 g/cm^3

e) 6.57×10^{-4} g/cm^3

Atom Accounting

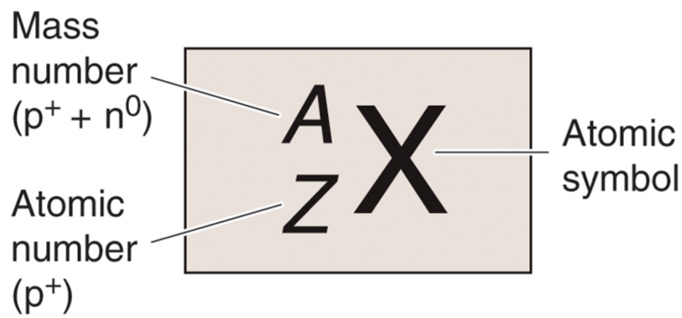
1 mole = the number of atoms in 12 g of pure ^{12}C

Avagadro's number = N_A = 6.022×10^{23} units in one mole

Molar mass = the mass of 1 mole of that substance

- To calculate molar mass, we must consider <u>isotopes</u>
- Each isotope of an element will have a slightly different mass.

Atomic Isotopes (Chapter 21.1)



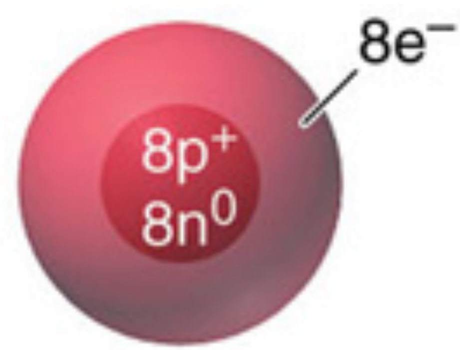
X = **Atomic symbol of the element**

A **= mass number;** $A = Z + N$

Z **= atomic number**

N **= number of neutrons in the nucleus**

For example: $^{16}_{8}\mathrm{O}$



An atom of oxygen-16

Elemental Molar Mass

- **In nature the elements exist as a mixture of isotopes**
- **To calculate the elemental (or average) molar mass we must account for the relative proportions of each isotope**
- **Elemental molar mass = weighted average of the respective molar masses of the natural isotopes of a given element**
- **ex. carbon has two naturally occurring isotopes:**

	^{12}C	^{13}C
Mass	12 g/mol	13.00335 g/mol
Abundance	98.892 %	1.108 %

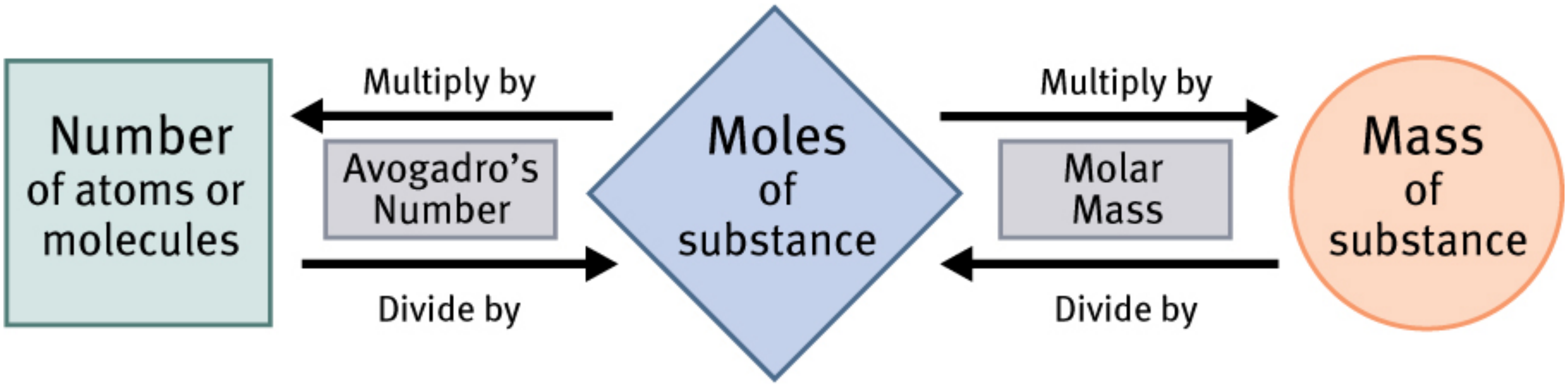
The elemental molar mass of carbon is:

Elemental Molar Mass Calculation Example

Boron (B; Z = 5) has two naturally occurring isotopes. Find the percent abundances of ^{10}B and ^{11}B, given the following data: elemental molar mass of B = 10.81 u, isotopic mass of ^{10}B = 10.0129 g/mol, and isotopic mass of ^{11}B = 11.0093 g/mol.

Conversion Between Amount, Mass and Number

Molar mass and Avagadro's number are essentially conversion factors:



Sample Calculation with Molar Mass

Potassium-40 is the main source of natural radioactivity in animals. Its percent natural abundance among K isotopes is 0.012%. An average human being contains about 160 g of potassium. How many atoms of radioactive ^{40}K does each person contain?

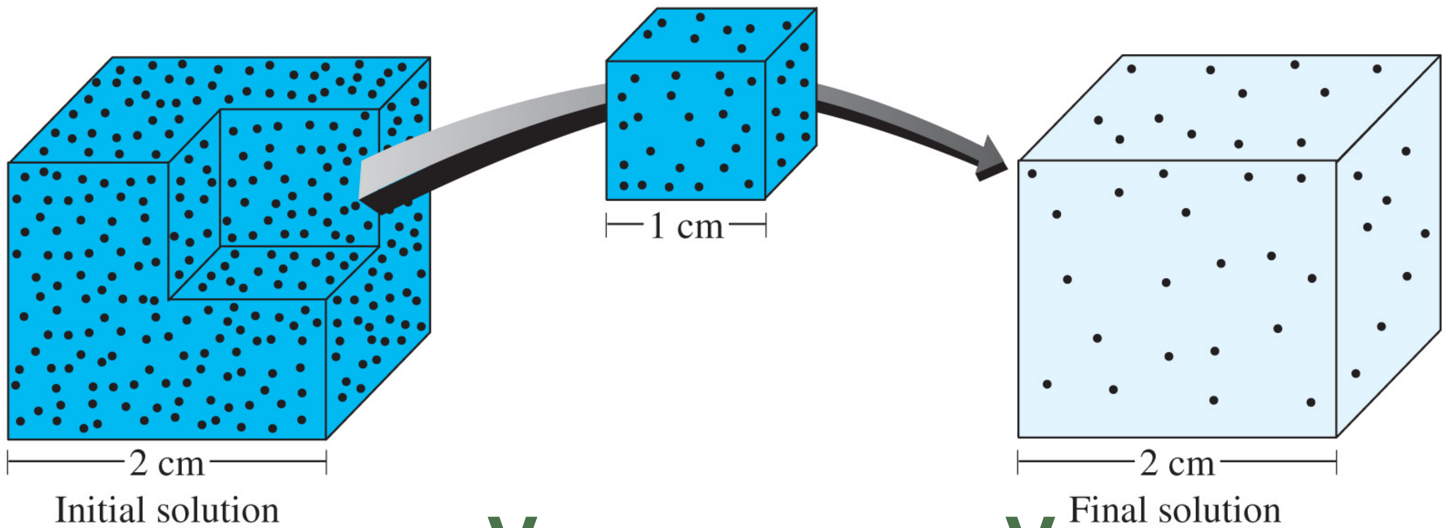
Solutions

- Lots of chemistry is done in solution

Concentration	Symbol	Equation
Molarity	*c*	$\frac{n_{solute}}{V_{solution}}$
Mass %	m/m%	$\frac{m_{solute}}{m_{solution}} = \frac{\text{mass of solute in g}}{\text{100 g solution}}$
Mole fraction	*X*	$X_A = \frac{n_A}{n_{Total}}$
Molality	*m*	$\frac{m_{solute}}{V_{solution}} = \frac{n_{solute}}{\text{1 kg solvent}}$

Described in more detail in Chapter 9.1

Preparing Solutions by Dilution



$$c_{initial} \cdot V_{initial} = c_{final} \cdot V_{final}$$

e.g. You have 50.0 mL of 3.0 M NaOH and you want 0.50 M NaOH. What do you do?

A Dilution Calculation For You…

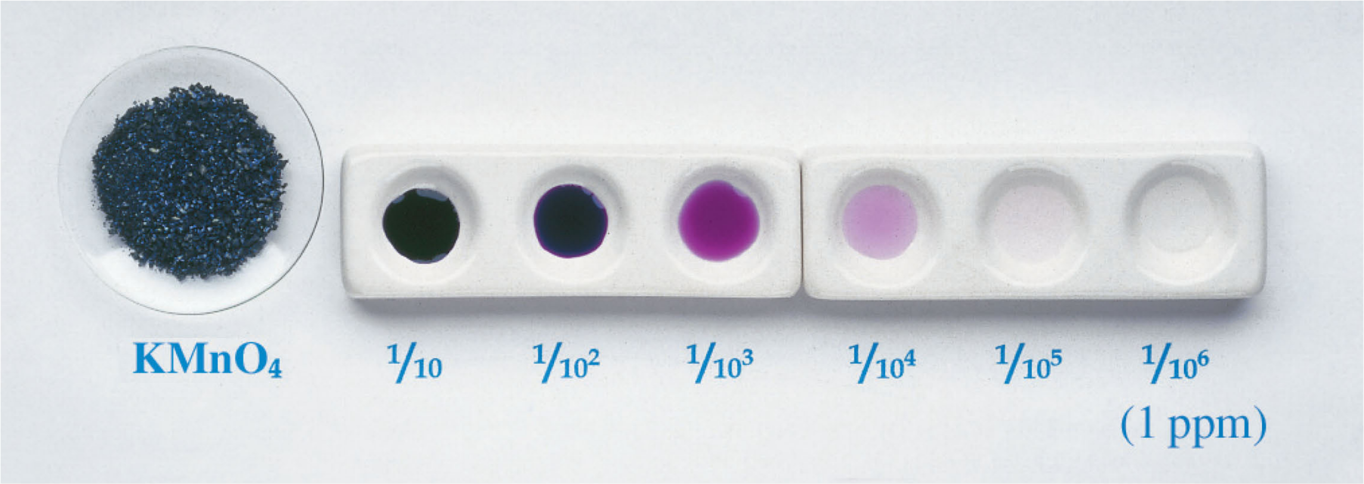
You have a standard solution of 1.00 mol/L $KMnO_4$. To prepare a solution of 100 mL of 0.100 mol/L $KMnO_4$, how many mL of the standard solution will you need?

A. 0.10 mL

B. 1 mL

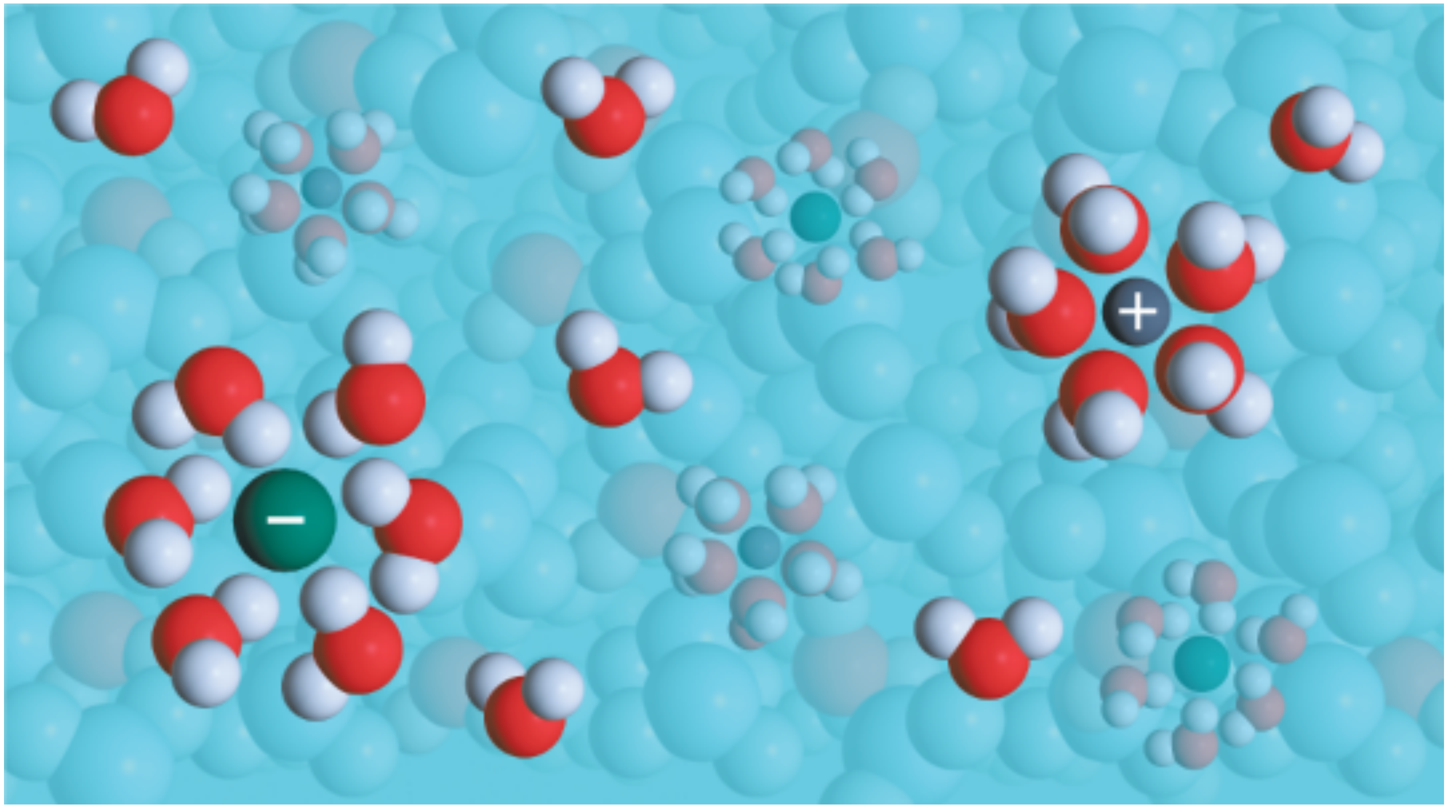
C. 10 mL

D. 100 mL



To text in your answer, send 'ZVI A,B,etc' to (613) 777-0647

Ions in Solution



e.g. $KMnO_4(aq) \rightarrow K^+ (aq) + MnO_4^- (aq)$

If you have a 0.30 M $KMnO_4$, this means that

$[K^+] = [MnO_4^-] = 0.30$ M

What happens if Na_2CO_3 is dissolved in water?

- This is a strong electrolyte
- Complete dissociation occurs when dissolved in water

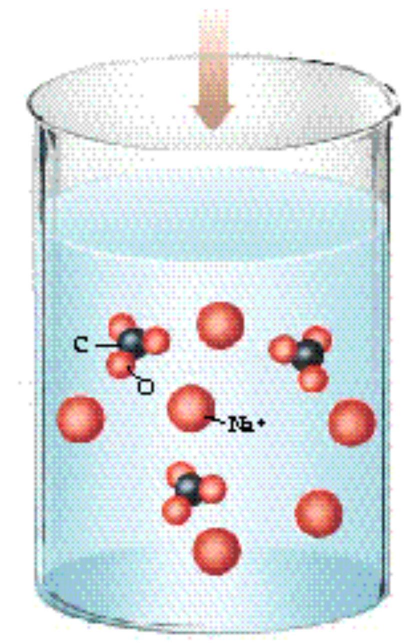
$Na_2CO_3(aq) \rightarrow 2\ Na^+(aq) + CO_3^{2-}(aq)$

If $[Na_2CO_3] = 0.100$ M, then

$[Na^+] =$

$[CO_3^{2-}] =$

3 Na_2CO_3



Equations for Aqueous Ionic Reactions

Chemical equations can be written in the form of a:

1) **Molecular Equation – Ionic species written in non-dissociated form**

$$Ba(NO_3)_2\,(aq) + Na_2CO_3\,(aq) \rightarrow BaCO_3\,(s) + 2\,NaNO_3\,(aq)$$

2) **Net Ionic Equation – cancel out aqueous ions appearing on both sides of equation**

Naming Compounds

- We can see that some ions are made up of single atoms, other are polyatomic
- You will need to recognize common polyatomic ions
- From Appendix G of the hardcover text:

Cations		Anions	
NH_4^+	Ammonium	OH^-	Hydroxide
H_3O^+	Hydronium	ClO_3^-	Chlorate
Hg_2^{2+}	Mercury (I)	ClO_4^-	Perchlorate
		NO_3^-	Nitrate
		NO_2^-	Nitrite
		MnO_4^-	Permanganate
		CO_3^{2-}	Carbonate
		HCO_3^-	Bicarbonate
		$Cr_2O_7^{2-}$	Dichromate
		PO_4^{3-}	Phosphate
		SO_4^{2-}	Sulfate

Can you answer these nomenclature questions?

a) Name the ionic compound formed from iodine and cadmium

b) Write the empirical formula for aluminum sulfide

c) Write the formula for ferric oxide.

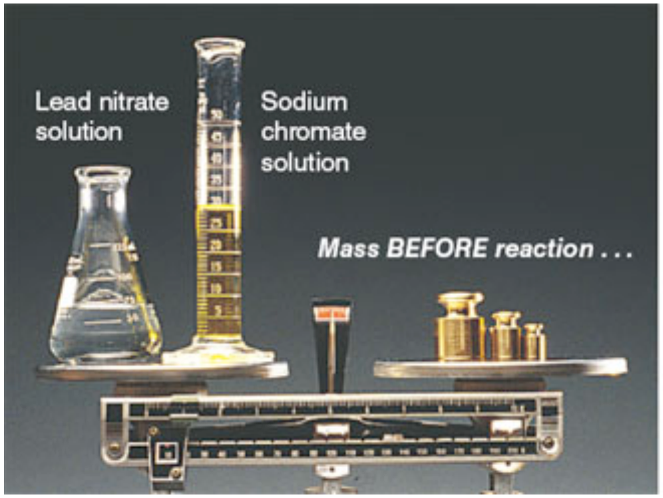
d) Give the formula for cupric nitrate trihydrate

e) What is the name of Br^- and its acid?

f) What is the name of NO_2^- and its acid?

g) What is the name for SF_4?

The Law of Mass Conservation



Total mass before reaction = Total mass after reaction

- **This reflects the fact that the number of atoms of each element is conserved in a chemical reaction.**
- **Chemical equations must be balanced to reflect this!**

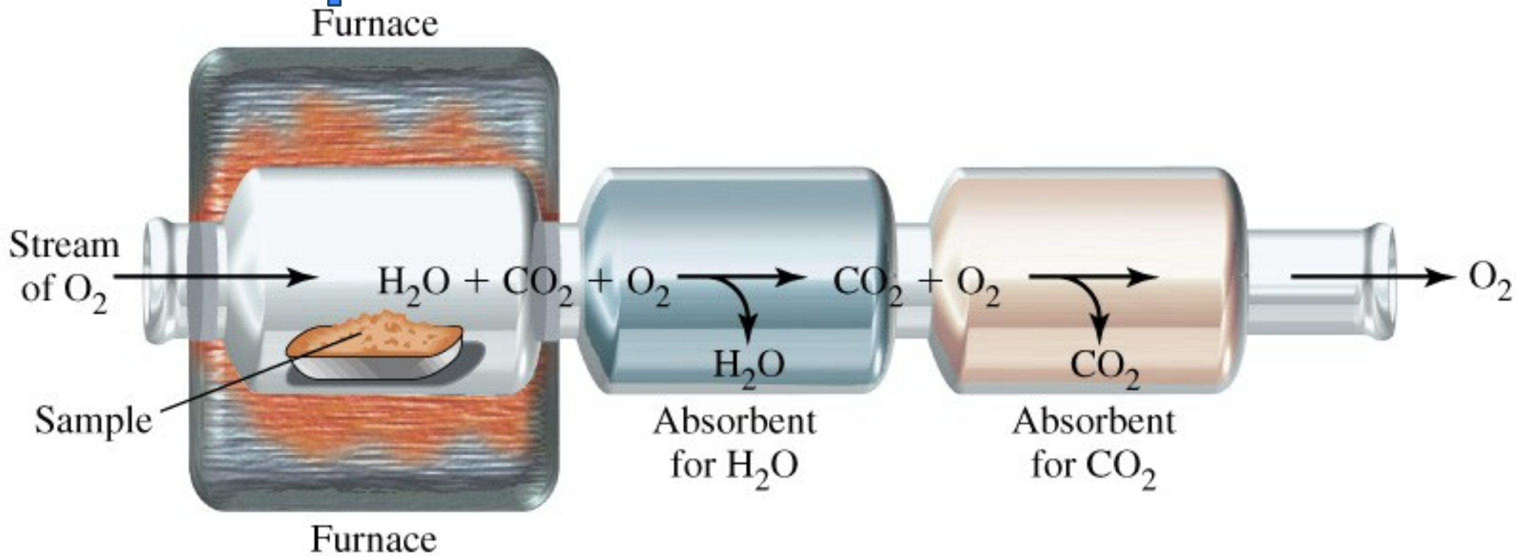
Strategy for Balancing Equations

- **Balance elements that occur in only one compound on each side first. Start with most complex compounds.**
- **Balance free elements last.**
- **Do not alter chemical formulas of compounds.**
- **Fractional coefficients are acceptable and can be cleared at the end by multiplication.**

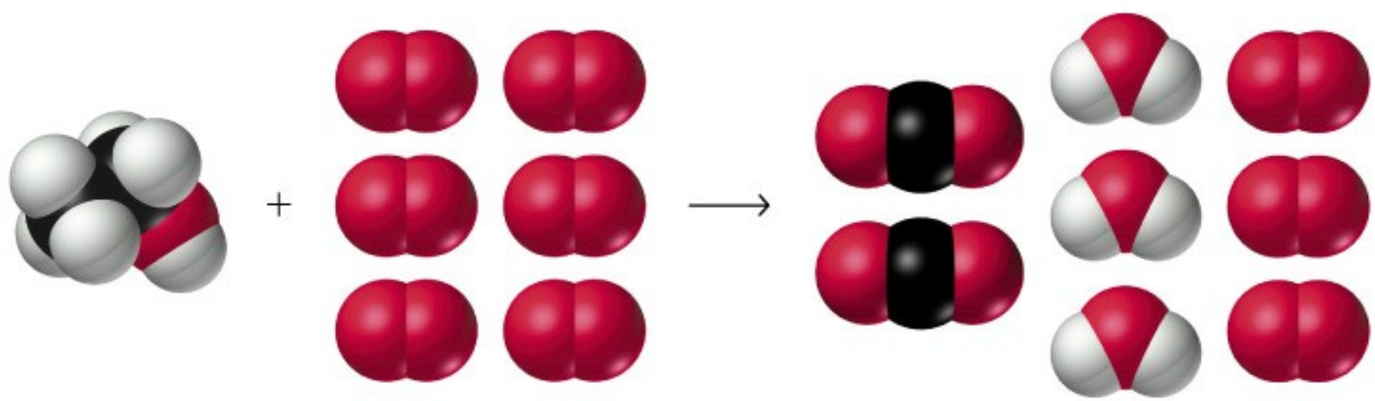
e.g. Balance:

$$Fe_3O_4 + HCl + Cl_2 \longrightarrow FeCl_3 + H_2O + O_2$$

Balanced Equations for Combustion Reactions



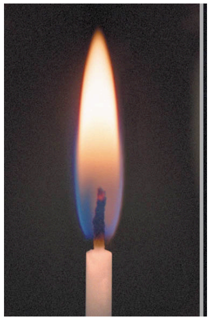
(a)



(b)

Example of Calculation from Combustion Data

When 129.48 g of an unknown hydrocarbon with $\mathcal{M}$ = 114.22 g/mol undergoes combustion, 88.94 g CO_2, and 40.54 g H_2O is produced. What is the molecular formula for the unknown?



Step 1: Determine the number of moles of each element from the products of combustion.

Combustion Data Calculation Continued…

Step 2: **Write a tentative formula.**

Step 3: **Divide by the smallest value.**

Step 4: **Determine the mass of the empirical formula.**

These questions can always be done by following the same 4 steps!

Your Turn…

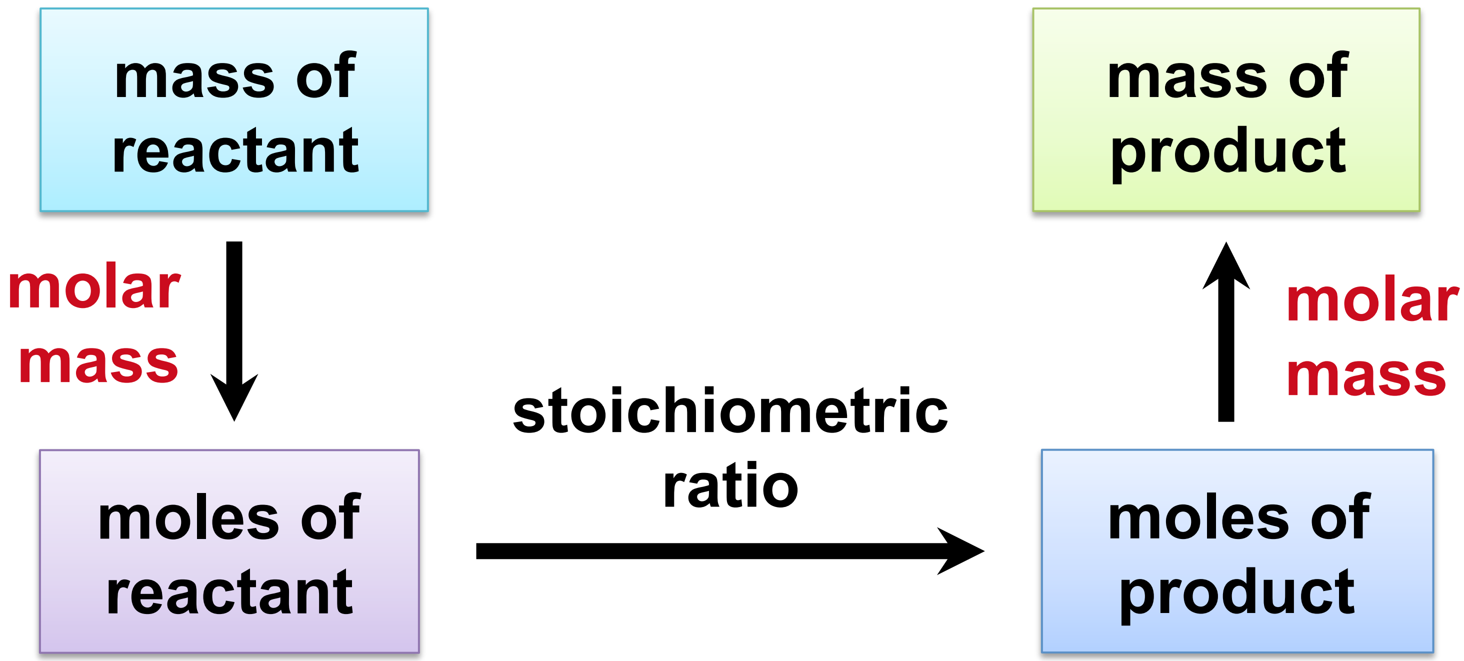
An oxide of nitrogen contains 30.45 % N by mass. What is the empirical formula of the oxide? (MW for N is 14.01 g/mol, O is 16.00 g/mol).

A) NO_2
B) N_2O
C) NO
D) NO_3

After you are done, consider this question:
What is the molecular formula of this oxide if M = 90 g/mol?

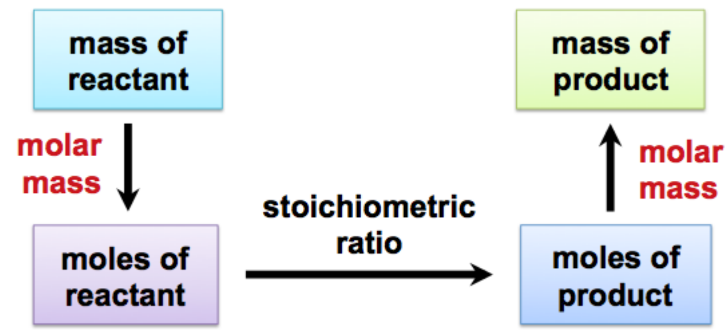
Stoichiometry

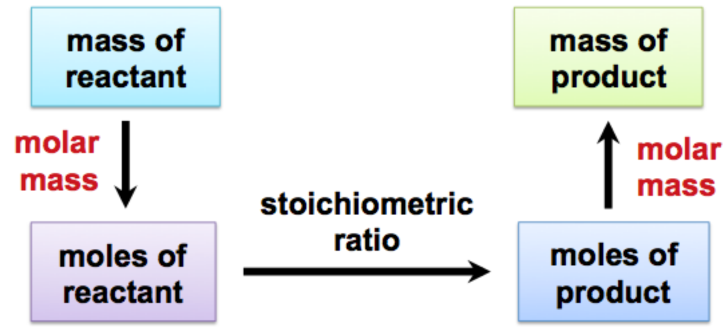
- The study of the quantitative aspects of a chemical reaction.
- if the quantity of reactant is known, it is possible to find the quantity of product that will be formed (or vice-versa)
- to simplify the process, we use the mole method:



Example of a Mole Method Calculation

If 454 g of NH_4NO_3 decomposes, how much N_2O and H_2O (in grams) are formed?




mass of reactant
molar mass
moles of reactant
stoichiometric ratio
moles of product
molar mass
mass of product

Reaction Yield Calculation

In many reactions, not all the reactants are converted into the desired product.

This is calculated using:

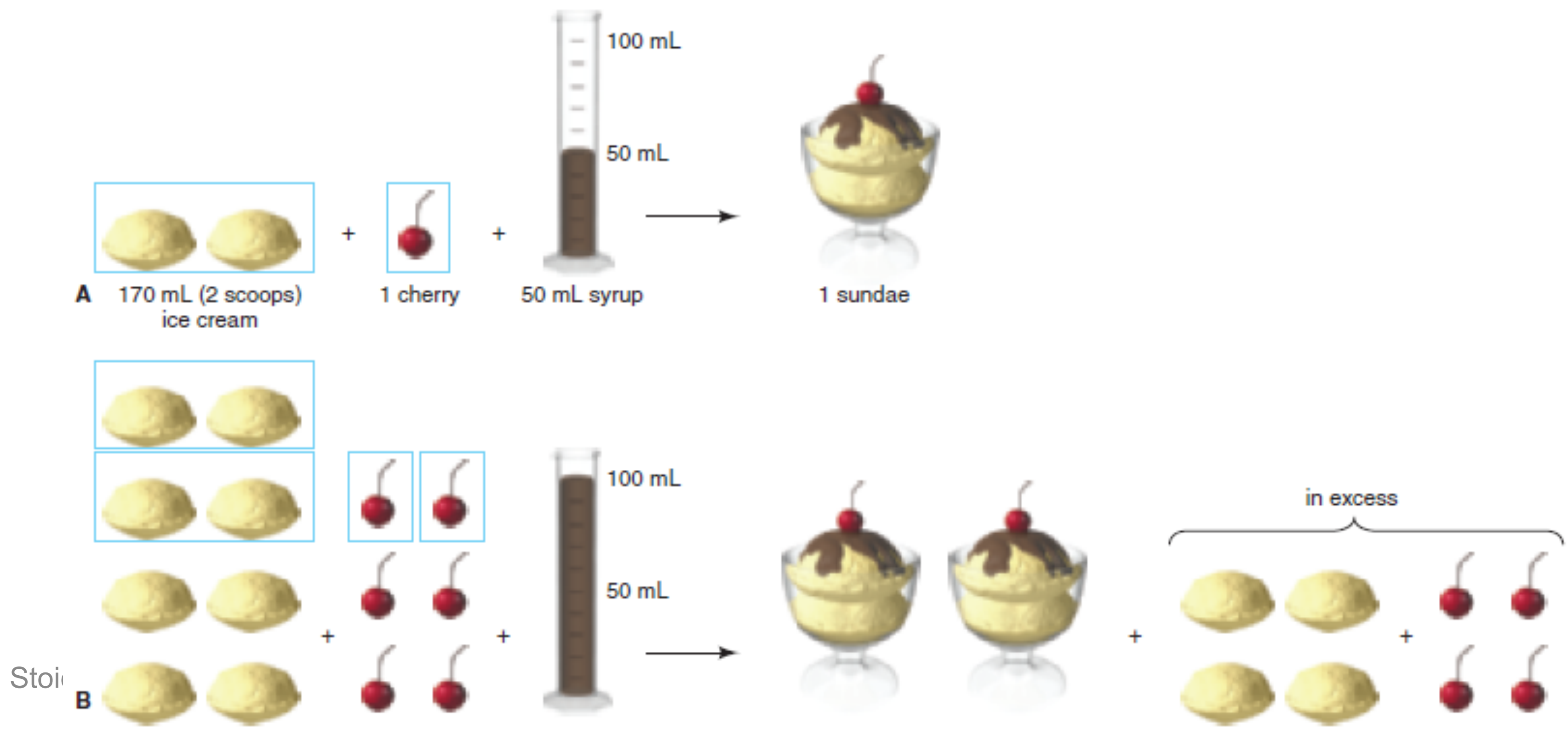
$$\%\ \text{yield} = \frac{\text{actual amount of product}}{\text{theoretical amount of product}} \times 100\%$$

How do we calculate the yield for a multi-step reaction?

e.g. The antidepressant sertraline is synthesized in 6 steps with the following yields: 80%, 80%, 50%, 100%, 48%, and 30%.

Limiting Reagents

- In real reactions, one reactant can limit the amount of product that can form.
- The limiting reactant will be completely used up in the reaction.
- The reactant that is *not* limiting is in *excess* – some of this reactant will be left over.



Solving Limiting Reactant Problems

1) Write the balanced chemical equation.
2) Convert reagent quantities into moles.
3) Use reaction stoichiometry to find the amount of product that would be produced by each reagent.
4) The reagent that produces the smallest amount of product is limiting.

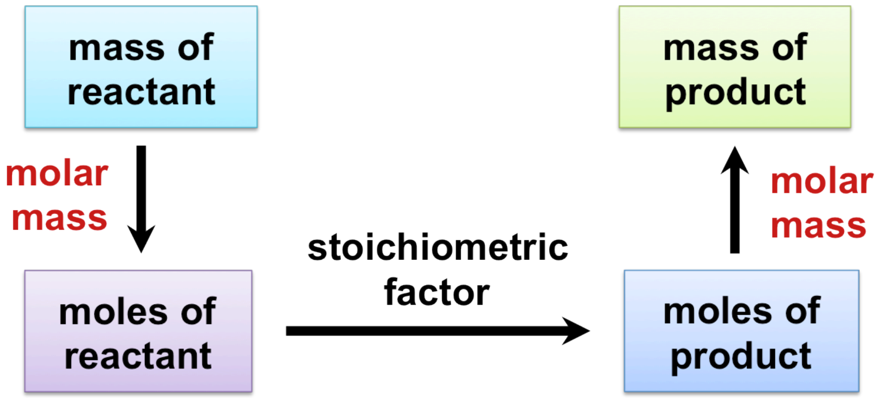
Example: What mass of solid aluminum sulfide can be prepared from the reaction of 10.0 g of aluminum and 15.0 g of sulfur? What mass of the non-limiting reagent is in excess?

1) Write the balanced chemical equation.

2) Convert reagent quantities into moles.

3) Find the amount of product that would be produced by each reagent

4) Calculate the mass of product from the amount of limiting reagent.



5) What mass of the non-limiting reagent is in excess?

Calculate amount required for reaction, and subtract that from amount added to reaction.

Reaction Tables

- used to keep track of the quantities in a limiting-reactant problem
- top row shows balanced equation as the column headings.

The table contains rows for reactant and product:
Initial quantities (usually moles)
Change in quantities during the reaction
Final quantities remaining after the reaction

So for the previous example:

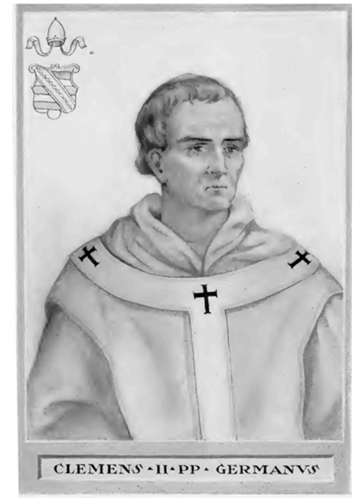
$$2\ Al(s) + \quad 3\ S(s) \quad \rightarrow \quad Al_2S_3(s)$$

Initial

Change

Final

Limiting Reactants in Solution



Example Problem:

Lead acetate was used as a sweetener before its toxic effects were known. It is possible to distinguish it from sugar by adding it to a NaCl solution, since it will form an insoluble chloride salt.

When 268 mL of 1.50 mol/L lead(II) acetate ($Pb(OAc)_2$) reacts with 130. mL of 3.40 mol/L sodium chloride, what mass (g) of solid lead(II) chloride can form? (Sodium acetate solution also forms.)

What is the balanced equation?

What is the limiting reagent?

Mass of $PbCl_2$ ($\mathcal{M}$ = 278.1 g/mol)?

To try on your own:

Can you write the reaction table for this example?

An ionic stoichiometry question for you…

In an environmental test for lead (207.2 g/mol) in soil, the lead is oxidized using permanganate ions in acidic solution, according to the following (balanced) reaction:

$$5\ Pb\ (s) + 2\ MnO_4^{-}\ (aq) + 16\ H^{+}(aq) \rightarrow 5\ Pb^{2+}\ (aq) + 2\ Mn^{2+}(aq) + 8\ H_2O\ (l)$$

A 25.0 g sample of soil was found to react with 2.24×10^{-3} mol of $KMnO_4$. What was the percent composition by mass of lead in the soil sample?

A. 4.6%
B. 1.2%
C. 7.4%
D. 0.46%
E. I don't know

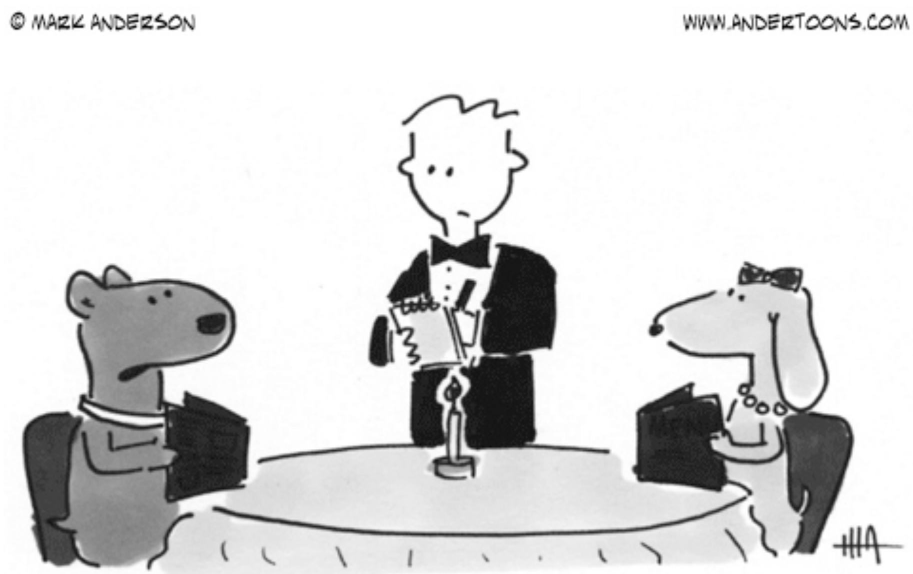
Recommended Exercises

Chemistry Fundamentals

Ch. 1: 11, 13, 14, 15, 16, 20, 22, 23, 25, 77, 78, 81, 85, 87, 101, 102
Ch. 21: 1, 2

Stoichiometry and Equations:

Ch. 1: 27, 31, 35, 40, 41, 43, 45, 57, 60, 64, 68, 72, 75, 89, 92, 94, 97, 103, 104, 106, 109, 114, 115, 117
Ch. 9: 6, 8, 9, 11, 54



"We'll both have the homework."