

CHEM 115

Course Review, First Half

Lecture Slides
May 15, 2007
Prof. Sevian



The final exam will be in
HLL-3507

The final exam is scheduled
for Friday, May 25,
8:00-11:00am

Macro-Particle-Symbolic

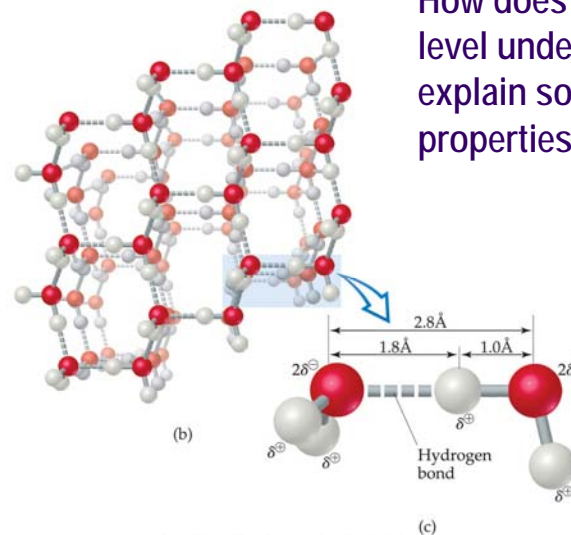
Each level provides information

- Macroscopic
 - Properties of matter
- Particle Level
 - Interactions between particles that give rise to matter's properties
- Symbolic
 - Representation of matter

Agenda

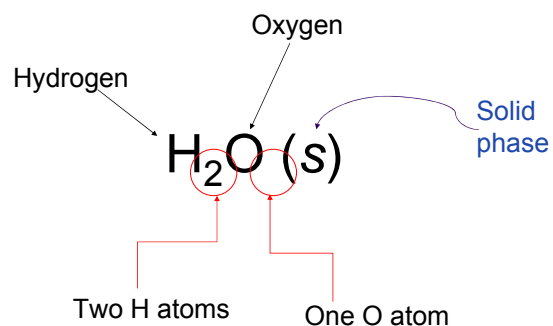
- Matter and measurement (ch. 1)
- Atoms, molecules and ions (ch. 2)
- Basic stoichiometry (ch. 3)
- Solution stoichiometry and reaction basics (ch. 4)

How does a particle level understanding explain some properties of ice?

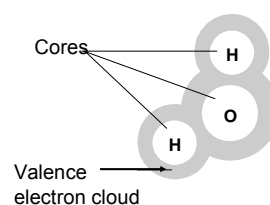


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What information does a chemical formula contain?

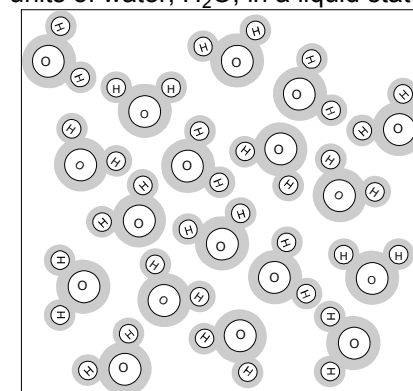


$H_2O(s)$ Representation



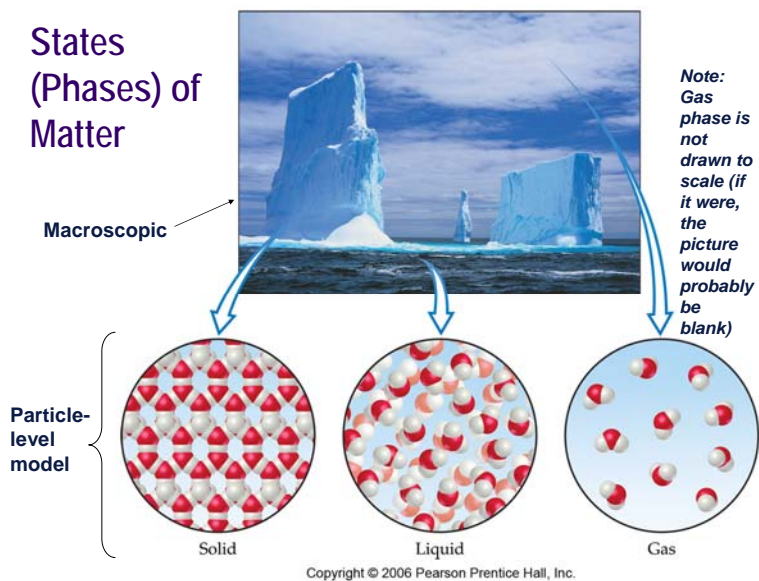
A single molecule (unit) of water, H_2O

A two-dimensional picture of several units of water, H_2O , in a liquid state

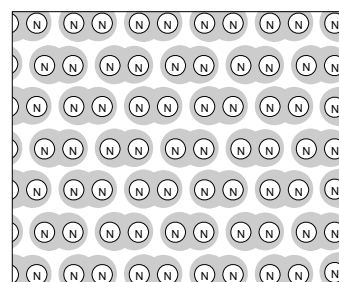


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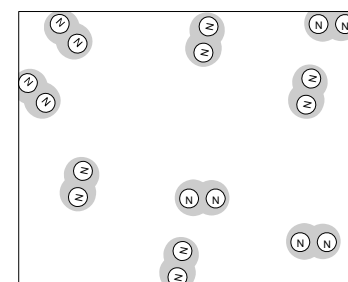
States (Phases) of Matter



Solid vs. Gaseous N_2



$N_2(s)$



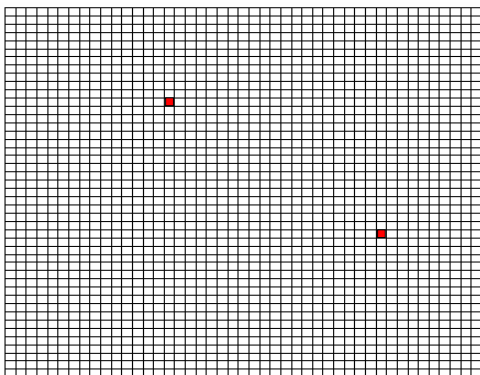
$N_2(g)$

Note: Not drawn to scale. True density of a gas is about 1/1000th of solid.

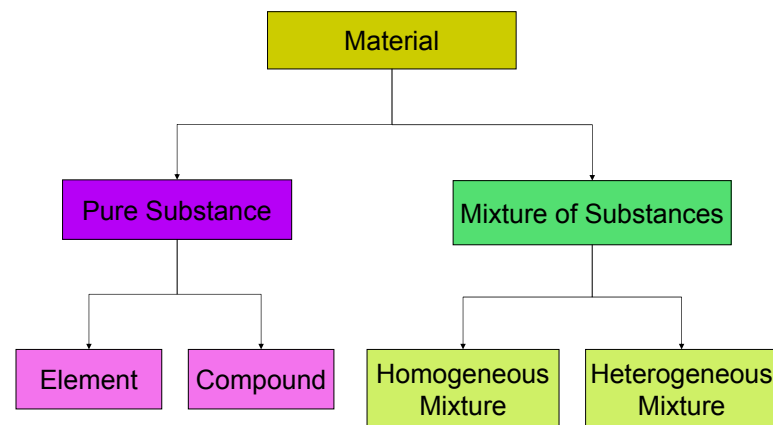
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Gas particles are very far apart from each other

A more accurate depiction: 1 in 1000 at room temperature



Categorizations of Matter



What goes into any calculation?

Three kinds of information

1. Measured values
2. Exact values
3. Derived (calculated) values

The least reliable number determines the reliability of the final calculated result. Usually it is a measured value that determines the significant digits in the result.

Measured Values

Measurement	Units most often used	Instrument often used
Length Wavelength	m, cm nm	Ruler (spectrometer)
Mass	g, kg	Balance
Volume	L, mL, m ³ , cm ³	Ruler (to measure dimensions) Graduated cylinder (for liquids)
Time	s	Clock
Temperature	°C, °F, K	Thermometer
pH	No units	pH gauge
Pressure	kPa, atm, mmHg	manometer

Exact Values

- Integer-based values
 - Fractions: $\frac{1}{2}$, $\frac{3}{4}$
 - Counting numbers: 2 electrons
 - Metric conversions: 1 meter = 100 cm
 - Two important exact conversions you need to remember
 - Distance: 1 inch = 2.54 cm
 - Energy: 1 calorie = 4.184 Joules
- Constants of nature are usually treated as exact values
 - Speed of light in vacuum, $c = 2.99792458 \times 10^8$ m/s
 - Pi, $\pi = 3.141592654\dots$ (no units)
 - Planck's constant, $h = 6.62617636 \times 10^{-34}$ J s
 - Gas constant, $R = 8.3144126$ J/mol·K

Derived (Calculated) Values

Value	Units most often used
Amount of a substance	mol
Energy	Joules, cal, kcal
Solution concentration	Molarity (M) = mol/L
Density	g/mL or g/cm ³
Velocity	m/s
Molar mass	g/mol

Rules for Sig Figs

- Generally, count the digits
- Zeroes written to the left don't count
0.00056 has 2 sig figs
- Zeroes written to the right do count
81.00 and 0.0008100 both have 4 sig figs
- Convention for numbers not containing a decimal point
7200 has 2 sig figs, 7200. has 4 sig figs
- See pp. 22-23 in the text for rules

Rules for Sig Figs

- Multiplication/division rule
 - The measurement with the least total sig figs wins
- Addition/subtraction rule
 - The measurement with the least decimal places (compared to the decimal point) wins
- Other rules can wait until you need to do more complicated calculations

Example of Addition Rule

Problem

Find the sum: $28.6 + 8.289 + 0.003 + 1007.56$

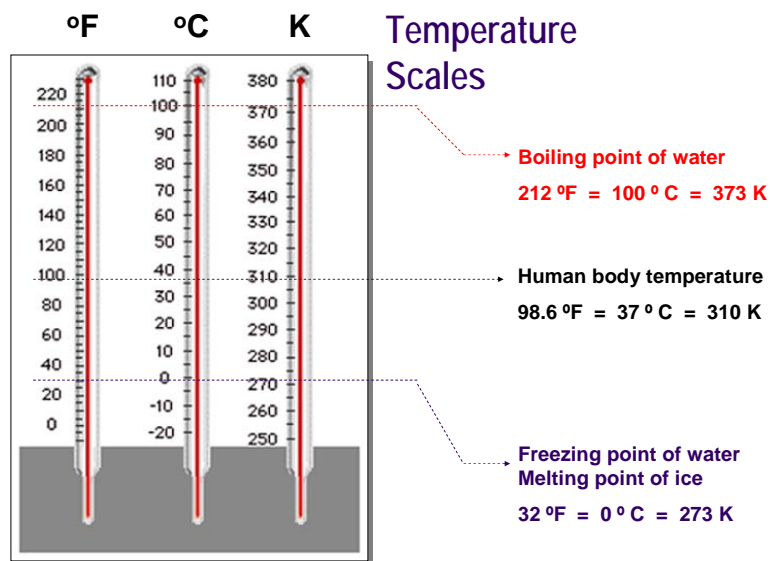
Solution

Line up the numbers at the decimal point, compare, and cut off at least significant (compared to decimal point)

$$\begin{array}{r}
 28.6 \\
 8.289 \\
 0.003 \\
 + 1007.56 \\
 \hline
 1044.452 \quad \text{|| rounds to} \quad 1044.5
 \end{array}$$

Metric Prefix Meanings for Conversions

Prefix	Meaning	Example
Centi (c)	1/100 th of	1 cm = 0.01 m 100 cm = 1 m
Milli (m)	1/1000 th of	1 mL = 0.001 L 1000 mL = 1 L
Kilo (k)	1000 of	1 kg = 1000 g
Micro (μ)	10 ⁻⁶ of	1 μ mol = 10 ⁻⁶ mol 1,000,000 μ mol = 1 mol
Nano (n)	10 ⁻⁹ of	1 nm = 10 ⁻⁹ m



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Temperature Conversions Require Equations

Celsius ($^{\circ}\text{C}$) \leftrightarrow Kelvin (K)

You need to memorize this conversion!!!

$$K = ^{\circ}\text{C} + 273.15 \quad \text{or} \quad ^{\circ}\text{C} = K - 273.15$$

Example: A gas has a temperature of $25.8\text{ }^{\circ}\text{C}$. What is the temperature in Kelvin?

$$\begin{aligned}
 K &= ^{\circ}\text{C} + 273.15 \\
 &= 25.8 + 273.15 = 298.95\text{ K} \quad \text{Rounds to } 299.0\text{ K}
 \end{aligned}$$

Sig figs:
One place beyond decimal pt

Really Big Numbers

In a 22.4 liter sample of air at standard conditions, there are approximately this many particles present:

$$602,204,531,000,000,000,000,000$$

2
3
3
3
3
3
3
3

$$= 6.02204531 \times 10^{23} \text{ particles}$$

Really Small Numbers

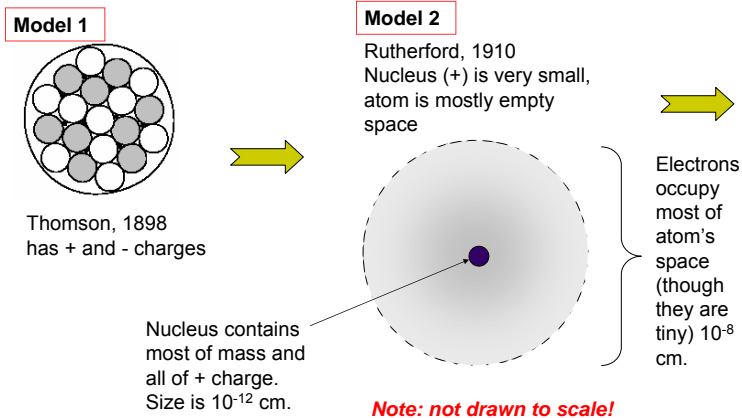
A single snowflake has a mass of approximately

$$0.000\ 0030 \text{ kg}$$

3
3

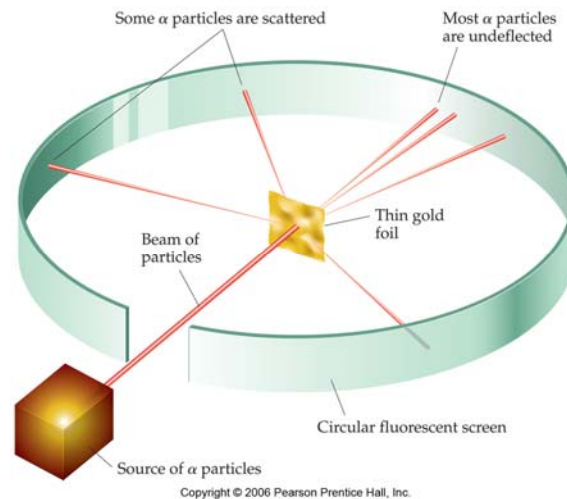
$$= 3.0 \times 10^{-6} \text{ kg}$$

What does an atom look like?



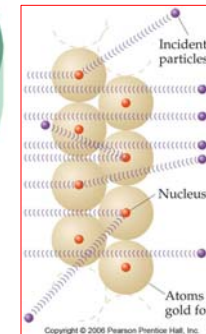
<http://micro.magnet.fsu.edu/primer/java/scienceopticsu/powersof10/>

Rutherford's Experiments

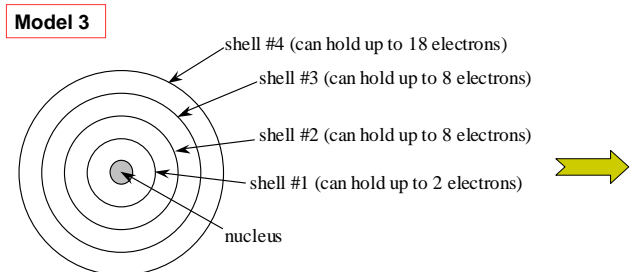


"It was almost as if you fired a 15-inch shell into a piece of tissue paper and it came back and hit you."

— Ernest Rutherford



What does an atom look like?



Bohr, 1912

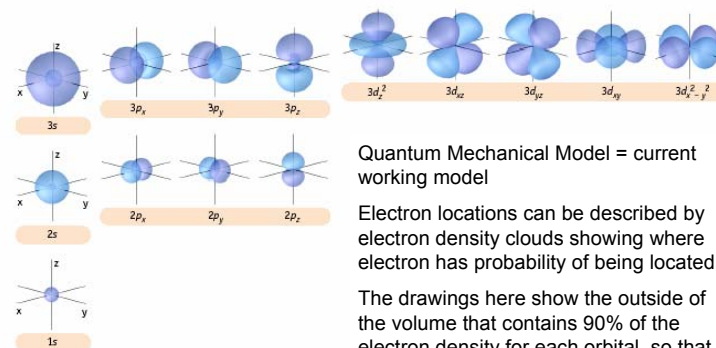
Note: not drawn to scale!

This model explains the hydrogen spectrum (but nothing else)

What does an atom look like?

Model 4

Schrodinger, Heisenberg, Bohr, de Broglie and some others, 1912-1924



Quantum Mechanical Model = current working model

Electron locations can be described by electron density clouds showing where electron has probability of being located.

The drawings here show the outside of the volume that contains 90% of the electron density for each orbital, so that you can see the 3-dimensional shapes of different orbitals.

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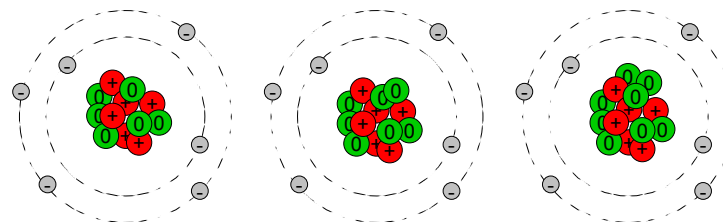
What's in an Atom?

	Location	Charge	Mass
Proton	Nucleus	+	~1 a.m.u.
Neutron	Nucleus	0	~1 a.m.u.
Electron	Most of atom's space	-	1/1836 th of an a.m.u.

- ⊖ electron
- 0 neutron
- + proton

How Can the Nucleus Vary?

Three different isotopes of carbon atoms
 (Note: Bohr model is incorrect, but very useful)



Same	All have 6 electrons	All have 6 protons	All neutral
Different	6 neutrons	7 neutrons	8 neutrons
Symbol	$^{12}_6\text{C}$	$^{13}_6\text{C}$	$^{14}_6\text{C}$

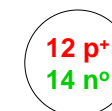
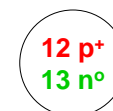
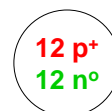
Specific isotopes

Fill in the missing information				
Symbol	Protons	Neutrons	Mass Number	Electrons
$^{11}_5\text{B}$	5	6	11	5
$^{37}_{17}\text{Cl}^-$	17	20	37	18

Isotopes and Natural Abundances

The mass of a typical sample of an element is a weighted average of the masses of the isotopes

Isotopes of magnesium



This is the atomic weight on the periodic table

Natural Abundance

78.99%

10.00%

11.01%

Mass of Isotope (amu)

23.9850

24.9858

25.9826

$$18.95 + 2.499 + 2.861 = 24.31$$

Reading the Periodic Table

12	Atomic number
Mg	Symbol
24.31	Atomic weight

Regions of the Periodic Table

See p. 50 in text

Transition metals

Lanthanides and actinides

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Organization of the Periodic Table

Terminology

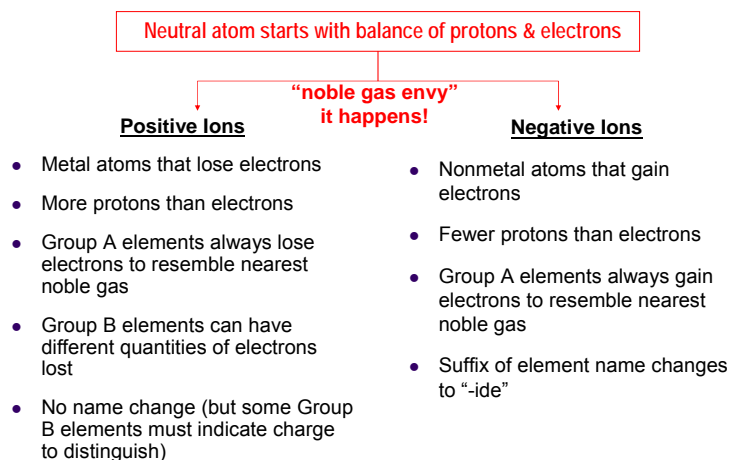
- Period = row across
- Group = column down
- Several common groups
 - Group 1A: Alkali metals
 - Group 2A: Alkaline earth metals
 - Group 7A: Halogens
 - Group 8A: Noble gases
 - Groups B: Transition metals
- Early chemists (Mendeleev, Moseley) organized the Periodic Table according to properties of elements
- There are reasons why the Periodic Table is organized the way it is (quantum mechanics explains periodicity of elements)

Main points about atoms & isotopes

- Over the years, various models of atomic structure have been developed
- All models are incorrect, some models are useful under certain circumstances
- The majority of the space that an atom takes up is the domain of electrons in the atom (the nucleus is very small)
- Nucleus is positive (due to protons), electrons outside/around nucleus are negative
- Symbolic notation for isotopes (atomic number, mass number, symbol)
- Atomic mass (weight) is a weighted average of all the naturally occurring isotopes of an element
- Organization of the periodic table (which is another model)

Ions and Their Names

When Protons \neq Electrons



Names of Common Ions

Find these elements on the Periodic Table and convince yourself why they take the charges they do.

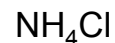
<u>Ion</u>	<u>Name</u>	
Cl ⁻	chloride	<div style="font-size: 3em; line-height: 1; display: inline-block; vertical-align: middle;">}</div> <p style="margin: 0;">Rule If nonmetal, change ending to “-ide”</p>
Na ⁺	sodium	
Mg ²⁺	magnesium	
O ²⁻	oxide	
N ³⁻	nitride	
Sr ²⁺	strontium	
Al ³⁺	aluminum	

Ionic vs. Molecular Compounds

- For simplicity, let's compare some ionic and molecular compounds in the solid state
- Ionic compounds:
NaCl = table salt, also called sodium chloride
NH₄Cl = ammonium chloride
- Molecular compounds:
H₂O (s) = ice
C₁₂H₂₂O₁₁ = sucrose

How do their macroscopic properties compare?

An Ionic Compound (in solid state)

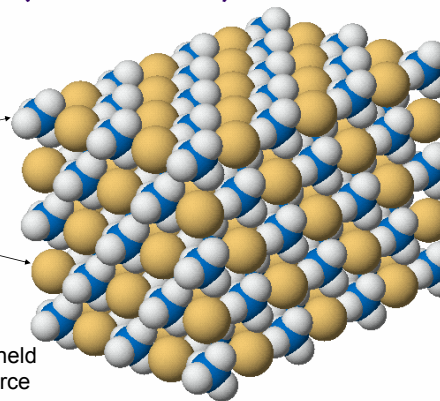


made of

NH₄⁺ ions

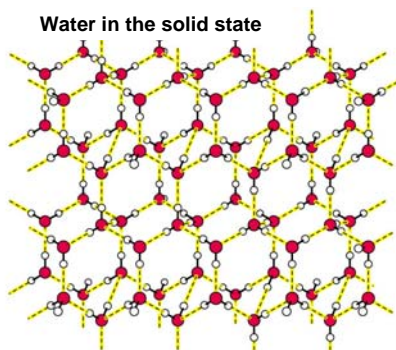
and

Cl⁻ ions



- Regular, repeating lattice structure
- Positive and negative ions held by attractive electrostatic force
- Every + ion surrounded by - ions
- Every - ion surrounded by + ions

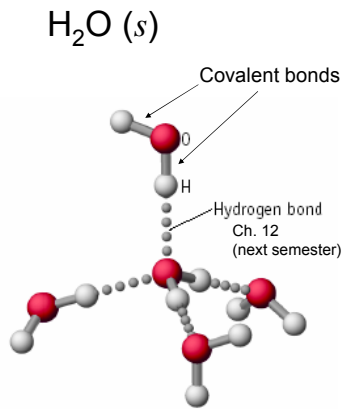
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A molecular compound (in solid state)

- Regular, repeating lattice structure
- Molecules at lattice site (not ions)
- Every lattice site has the same molecule at it
- Molecules held in lattice by attractions to each other that are weaker than covalent bonds

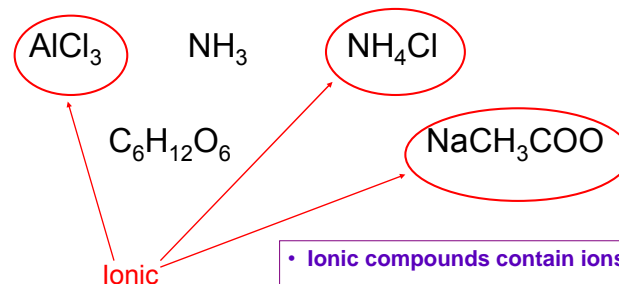
Ice crystal structure picture from http://cwx.prenhall.com/horton/media/lib/media_portfolio/text_images/E_G02_05.JPG



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You need to be able to...

... distinguish between ionic and molecular compounds so that you can name them.



- Ionic compounds contain ions
- Molecular compounds have only nonmetals in them

Common Ions

- Monatomic
 - Group A elements have only one possible charge
 - Group B elements (transition metals) usually have more than one possible charge
- Polyatomic
 - See pp. 62 and 64 for lists of ions you need to memorize (name, formula, charge)

Families of polyatomic ions

Chlorine family

- ClO^- (a.k.a. OCl^-)
- ClO_2^-
- ClO_3^-
- ClO_4^-

Chlorine family

- hypochlorite
- chlorite
- chlorate
- perchlorate

Bromine family

- BrO^- (a.k.a. OBr^-)
- BrO_2^-
- BrO_3^-
- BrO_4^-

Bromine family

- hypobromite
- bromite
- bromate
- perbromate

Naming Conventions

1. Ionic compounds - binary

NaCl sodium chloride

FeCl_3 iron (III) chloride

FeCl_2 iron (II) chloride

AgI silver iodide

Naming Conventions

1. Ionic compounds – contains polyatomic

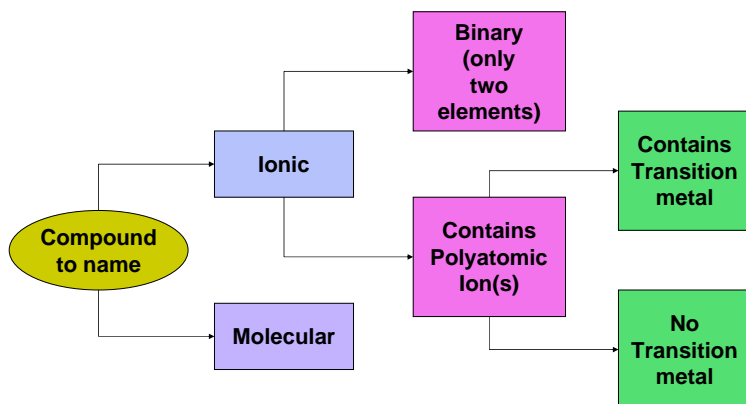
Na_2CO_3 sodium carbonate

NH_4Br ammonium bromide

$\text{Mg}(\text{C}_2\text{H}_3\text{O}_2)_2$ magnesium acetate

CuSO_4 copper (II) sulfate

Naming Rules So Far



Naming Conventions

2. Molecular compounds

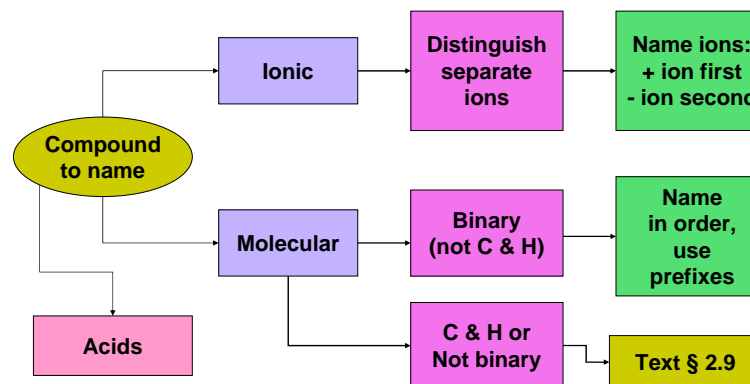
NO_2	nitrogen dioxide
NO_3	nitrogen trioxide
N_2O_4	dinitrogen tetroxide
CO	carbon monoxide
CO_2	carbon dioxide
P_2O_5	diphosphorus pentoxide
CH_4	methane
C_2H_6	ethane

Be sure to read section 2.9 in your text: you will be responsible for knowing how to name simple organic compounds

Counting to 10 in Greek to name binary molecular compounds

1. Mono
2. Di
3. Tri
4. Tetr(a)-
5. Pent(a)-
6. Hex(a)-
7. Hept(a)-
8. Oct(a)-
9. Non(a)-
10. Dec(a)-

Naming Rules So Far

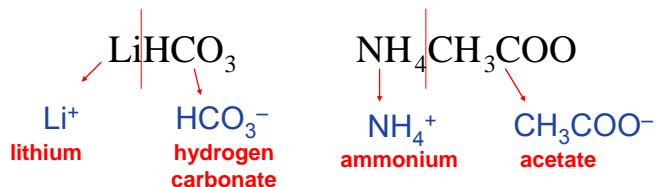


Common Mistakes in Naming

- Look for ions vs. no ions



- If ionic compound, regardless of how many total atoms, it has only a first name (+ ion) and a last name (- ion)



Naming Conventions

3. Acids

- HCl
- H_2CO_3
- HBr
- HOCl
- HOClO
- $\text{HC}_2\text{H}_3\text{O}_2$
- $\text{C}_2\text{H}_3\text{O}_2\text{H}$
- HNO_2
- HNO_3

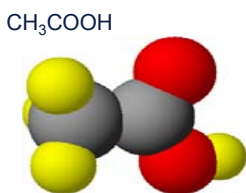
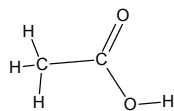
They look like ionic compounds, except that the positive ion is always H^+

What would you name them if they were ionic compounds?

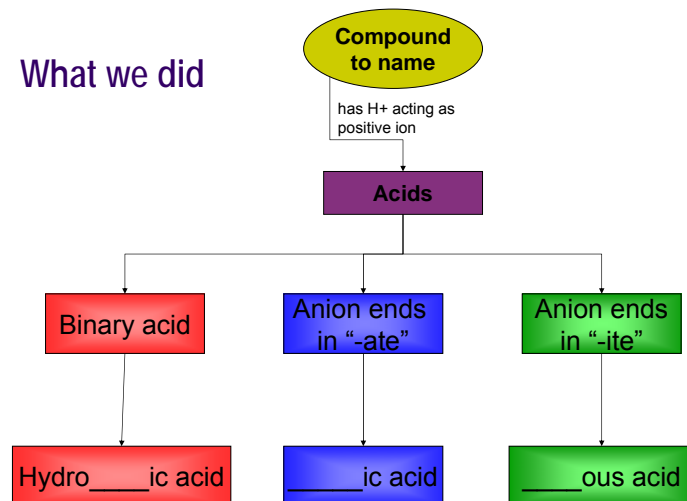
Naming acids properly

3. Acids

- HCl → Hydrochloric acid
- H_2CO_3 → Carbonic acid
- HBr → Hydrobromic acid
- HOCl → Hypochlorous acid
- HOClO → Chlorous acid
- $\text{HC}_2\text{H}_3\text{O}_2$ → Acetic acid
- $\text{C}_2\text{H}_3\text{O}_2\text{H}$ → Acetic acid
- HNO_2 → Nitrous acid
- HNO_3 → Nitric acid



What we did



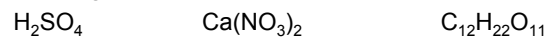
Naming Conventions Summary

First determine if the compound is ionic, molecular, or acid.

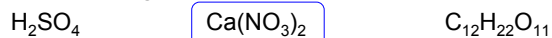
1. Ionic compounds
 - a) Binary or contain polyatomic ion(s)?
 - b) Can the metal cation have more than one oxidation state?
2. Molecular compounds
 - a) Binary (except not C and H)
 - b) Hydrocarbons (contains C and H)
3. Acids
 - a) Binary
 - b) Anion ends in "-ate"
 - c) Anion ends in "-ite"

Moles scale from atoms

- How many atoms of O are in one unit of each of the following compounds?



- How many moles of O atoms are in 1 mol of each of the following compounds?



Possible conversions are:

$$\frac{6 \text{ mol O atoms}}{1 \text{ mol Ca}(\text{NO}_3)_2} \quad \text{or} \quad \frac{1 \text{ mol Ca}(\text{NO}_3)_2}{6 \text{ mol O atoms}}$$

What is the Mass of a Mole?

- Different kinds of particles have different masses
- Since the quantity of particles in a mole is the same for any particle, the mass of a mole of particles varies depending on the identity of the particles
- By definition,

$$1 \text{ mol of } {}^{12}_6\text{C particles} = 12 \text{ g exactly}$$

- Other molar masses always in reference to definition.

Molar Mass

Molar mass means the mass of one mole

- For monatomic elements, use atomic weight on the Periodic Table (do you remember what "atomic weight" means?)
- Examples:

What is the mass of 1.000 mol of Ar?

$$1.000 \text{ mol Ar} \times \frac{39.95 \text{ g Ar}}{1 \text{ mol Ar}} = 39.95 \text{ g Ar}$$

What is the mass of 2.5 mol of Xe?

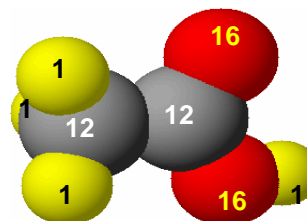
$$2.5 \text{ mol Xe} \times \frac{131.3 \text{ g Xe}}{1 \text{ mol Xe}} = 328.25 \text{ g Xe} \quad \rightarrow \quad 330 \text{ g Xe}$$

What is the amount of money here?



How did you figure this out?

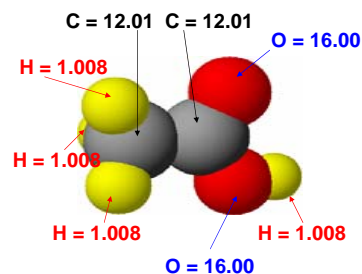
What is the amount of amu's here?



Answer:
60 amu

How did you figure this out?

Actually, the masses aren't integers



Mass of one molecule of CH_3COOH is
60.05 amu

Mass of 1 mole of CH_3COOH molecules is
60.05 grams

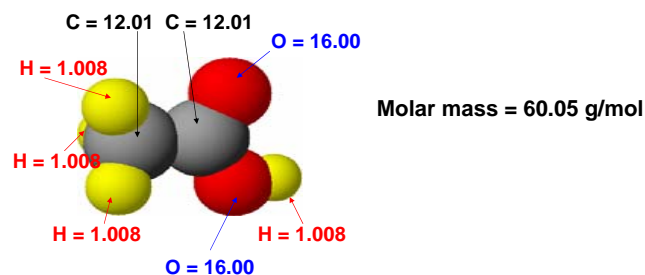
Important note: Since there is always the same number of particles in a mole, when you determine the amu's of a unit, you are determining the mass in grams of a mole of that unit.

Moles and masses, so far

- Mass (in grams) of a mole is the same number as the mass (in amu) of one unit in the mole
- Use the average atomic weights on the periodic table as masses of individual atoms
- There are many names for this quantity:
 - Molar mass
 - Molecular weight (when the unit is a molecule)
 - Formula weight (when the unit is the simplest ratio of ions in an ionic crystal)

More on Molar Mass

What is the mass of 0.00497 mol of CH₃COOH?



$$0.00497 \text{ mol CH}_3\text{COOH} \times \frac{60.05 \text{ g}}{1 \text{ mol}} = 0.298 \text{ g CH}_3\text{COOH}$$

Conversions possible so far

(look for opportunities to use these in dimensional analysis)

- Moles ↔ mass (grams)
- Moles of atoms ↔ moles of a unit (e.g., molecule, neutral ionic crystal formula)
- Moles of something ↔ particles in the something

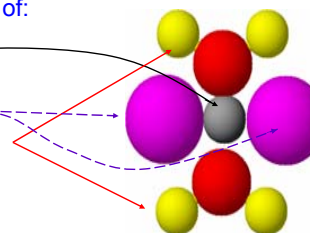
Hydrated Ionic Compounds

How many moles are in 5.55 g of BaCl₂•2H₂O?

One unit of ionic formula consists of:

One Ba²⁺ ion = 137.3
 Two Cl⁻ ions = 2 × 35.45
 Two H₂O molecules = 2 × 18.02

Formula weight = 244.2 g/mol



$$5.55 \text{ g BaCl}_2 \cdot 2\text{H}_2\text{O} \times \frac{1 \text{ mol}}{244.2 \text{ g}} = 0.0227 \text{ mol BaCl}_2 \cdot 2\text{H}_2\text{O}$$

How Many Molecules?

Too many to count!

How many molecules of water are in a 2.00-L bottle that is filled with water? Water has a density of 1.000 g/mL.



Start

$$\frac{2.00 \text{ L of H}_2\text{O}}{1} \times \frac{1000 \text{ mL}}{1 \text{ L}} \times \frac{1.000 \text{ g}}{1 \text{ mL}} \times \frac{1 \text{ mol}}{18.02 \text{ g}} \times \frac{6.022 \times 10^{23} \text{ molecules}}{1 \text{ mol}} =$$

End

6.68 × 10²⁵ molecules of H₂O

1

Composition of a ...

- Usually composition is given in mass percent (also called weight %)
- Use the chemistry definition of percent

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

- Know the difference between fraction and percent – a fraction can be expressed as a percent by multiplying by 100 (“percent” means part out of 100)

Fraction 0.7585 is the same as 75.85%

Composition of a Compound

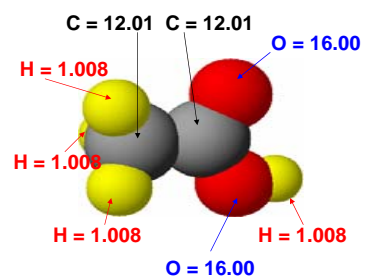
What is the composition by mass of acetic acid (CH_3COOH)?

$$\% \text{C} = \frac{24.02}{60.05} \times 100\% = 40.00\%$$

$$\% \text{H} = \frac{4.032}{60.05} \times 100\% = 6.714\%$$

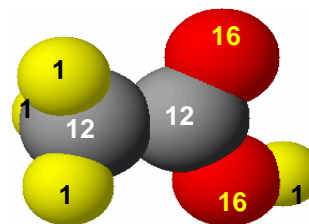
$$\% \text{O} = \frac{32.00}{60.05} \times 100\% = 53.29\%$$

$$\text{Check: } 40.00 + 6.714 + 53.29 = 100.00\%$$



Recall from earlier:
Molar mass = 60.05 g/mol

What does % composition mean?



Note: These are approximate atomic masses, for the purpose of demonstrating % composition. When actually calculating % composition, use the values from the Periodic Table.

Approximate
% Composition by Mass

$$\text{fraction C} = \frac{2 \times 12}{60} = 40\%$$

$$\text{fraction H} = \frac{4 \times 1}{60} = 6.7\%$$

$$\text{fraction O} = \frac{2 \times 16}{60} = 53\%$$

Composition of a Hydrated Compound

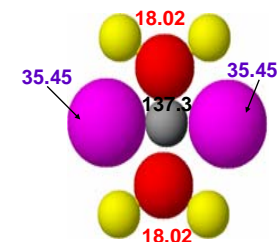
Heating barium chloride dihydrate ($\text{BaCl}_2 \cdot 2\text{H}_2\text{O}$) drives off the water, leaving the anhydrous compound (BaCl_2). The chemical reaction is



If you begin with a 10.0 g sample of the hydrated compound, what mass of water will be lost?

$$\% \text{H}_2\text{O} = \frac{2 \times 18.02}{244.2} \times 100\% = 7.379\%$$

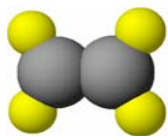
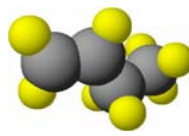
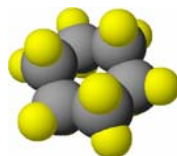
$$\begin{aligned} \text{mass of H}_2\text{O in sample} &= 7.379\% \text{ of } 10.0 \text{ g} \\ &= 0.07379 \times 10.0 \text{ g} \\ &= 0.738 \text{ g} \end{aligned}$$



Formula weight = 244.2 g/mol

Chemical Compounds and Mass

- Chemical formula to percent composition
 - Need to determine parts and whole
 - Use definition of percent
- Going the other direction
 - Percent composition alone is not enough information to determine molecular formula

C₂H₆C₄H₁₀C₆H₁₄

All three of these have 14.37% H and 85.63% C by mass

Example

% Composition to Empirical Formula

Analysis of a particular compound shows that it is composed of 73.14% carbon, 7.37% hydrogen, and the remainder is oxygen. What is the compound's empirical formula?

$$\left. \begin{array}{l} \frac{73.14 \text{ g C}}{12.01 \text{ g}} \times \frac{1 \text{ mol C}}{12.01 \text{ g}} = 6.089 \text{ mol C} \\ \frac{7.37 \text{ g H}}{1.008 \text{ g}} \times \frac{1 \text{ mol H}}{1.008 \text{ g}} = 7.31 \text{ mol H} \\ \frac{19.49 \text{ g O}}{16.00 \text{ g}} \times \frac{1 \text{ mol O}}{16.00 \text{ g}} = 1.218 \text{ mol O} \end{array} \right\} \begin{array}{l} \text{specific whole number ratio of} \\ \text{moles C : moles H : moles O} \\ \text{is } 6.089 : 7.31 : 1.218 \\ \\ = \frac{6.089}{1.218} : \frac{7.31}{1.218} : \frac{1.218}{1.218} \\ = 5 : 6 : 1 \\ \\ \text{so, empirical formula is C}_5\text{H}_6\text{O}_1 \end{array}$$

General Idea

% Composition to Empirical Formula

Analysis of a particular compound shows that it is composed of X% carbon, Y% hydrogen, and Z% oxygen.

What is the compound's empirical formula?

MASS (grams) ⇒ MOLES (mol)

$$\frac{X \text{ grams}}{100 \text{ grams}} \text{ are carbon}$$

$$\frac{Y \text{ grams}}{100 \text{ grams}} \text{ are hydrogen}$$

$$\frac{Z \text{ grams}}{100 \text{ grams}} \text{ are oxygen}$$

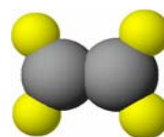
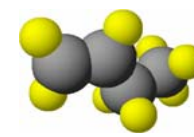
⇒ specific whole number ratio of
moles C : moles H : moles O
is a : b : c
empirical formula is C_aH_bO_c

Next Step: How to get to Molecular Formula

- A hydrocarbon has 85.63% carbon by mass. What is its empirical formula?

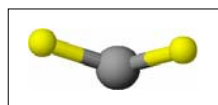


- What else do you need to know to determine molecular formula?

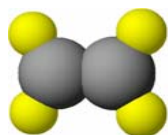
C₂H₆C₄H₁₀C₆H₁₄

(All three have 14.37% H and 85.63% C by mass)

How many empirical formulas?

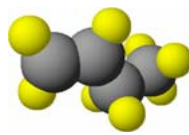


Empirical Unit CH₂
Mass is 12.01 + 2(1.008) = 14.03 g/mol



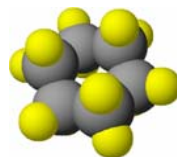
C₂H₆

$$\begin{aligned} & 2 \times (\text{CH}_2) \\ & = 2 \times (14.03) \\ & = \mathbf{28.06 \text{ g/mol}} \end{aligned}$$



C₄H₁₀

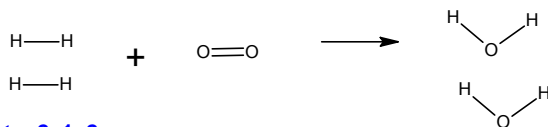
$$\begin{aligned} & 4 \times (\text{CH}_2) \\ & = 4 \times (14.03) \\ & = \mathbf{56.12 \text{ g/mol}} \end{aligned}$$



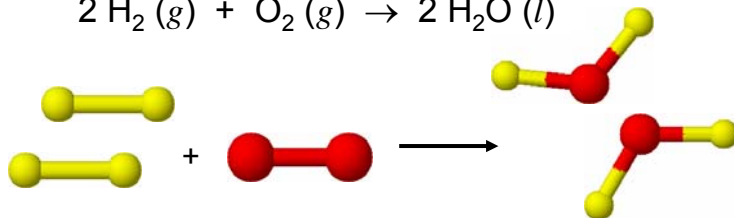
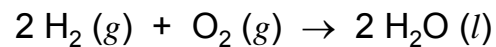
C₆H₁₄

$$\begin{aligned} & 6 \times (\text{CH}_2) \\ & = 6 \times (14.03) \\ & = \mathbf{84.18 \text{ g/mol}} \end{aligned}$$

Formation of Water



Coefficients: 2, 1, 2



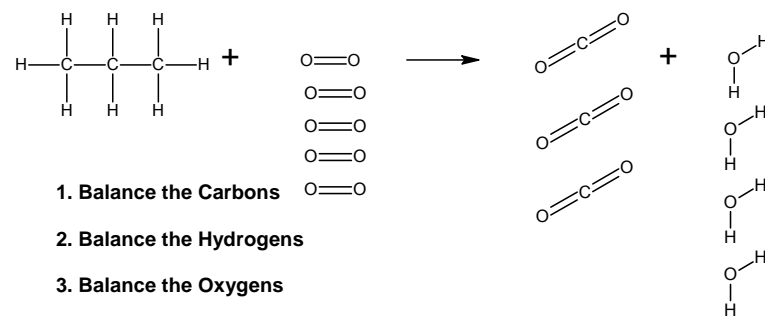
Chemical Equations

- Represent a chemical change of matter
- Reactants (starting materials) on left
- Products (ending materials) on right

Reactants → Products

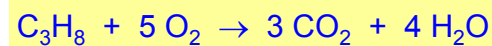
- What goes in must come out, just connected (bonded) differently

Combustion of propane



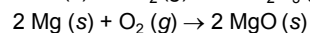
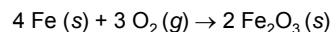
1. Balance the Carbons
2. Balance the Hydrogens
3. Balance the Oxygens
4. Is it balanced?

Balanced chemical equation:

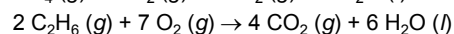
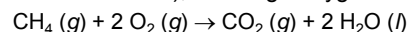


Combustion Reactions

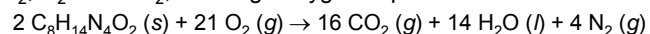
- Flame-producing reaction of a chemical with oxygen to produce oxides



- Hydrocarbons (C and H) burn to produce CO₂ and H₂O (the oxides of C and H), if enough oxygen is present



- Organic compounds (containing C, H, N and O) burn to produce CO₂, H₂O and N₂, if enough oxygen is present



Balancing chemical equations

Step 2: The balancing act

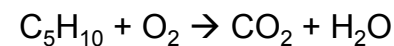
- Two methods
 - Accounting method
 - Pictures method
- Both methods work
- Whichever one you use just depends on your mathematical inclination
- Method 2 is more time consuming

Balancing chemical equations

Step 1: What species to balance?

- The important question is: What are the smallest units (species) that both go in and come out?
 - If the chemical reaction involves only ions, chances are you should balance ions instead of atoms
 - If there are ionic compounds involved, knowing how to name them will assist greatly in balancing, because the names are the ions
- Practice: what species should be balanced?
 - $\text{C}_5\text{H}_{10} + \text{O}_2 \rightarrow \text{CO}_2 + \text{H}_2\text{O}$
 - $\text{CaSO}_4 + \text{AgNO}_3 \rightarrow \text{Ag}_2\text{SO}_4 + \text{Ca}(\text{NO}_3)_2$
 - $\text{Mg} + \text{O}_2 \rightarrow \text{MgO}$
 - $\text{Al} + \text{CuSO}_4 \rightarrow \text{Al}_2(\text{SO}_4)_3 + \text{Cu}$

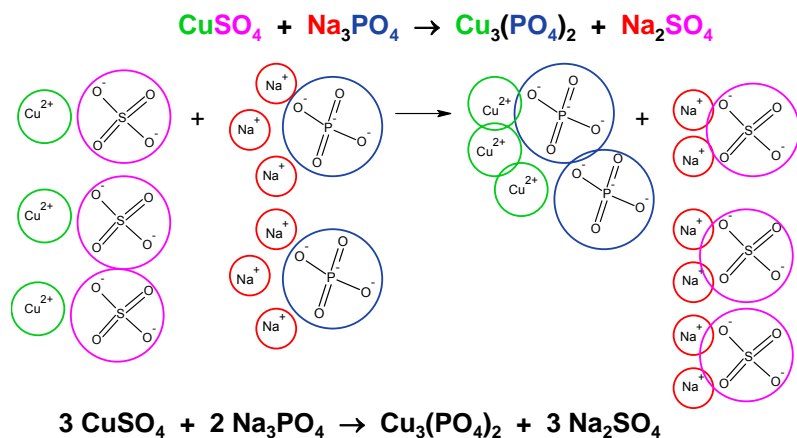
Accounting method of balancing



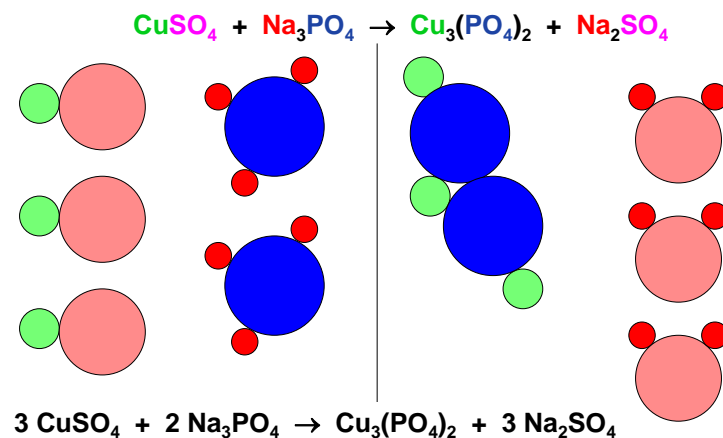
- Balancing units are C, H and O
- Balance the O's last, since they are present in two different products
- 5 C in \rightarrow 5 C must come out, so CO₂ has coefficient 5
- 10 H in \rightarrow 10 H must come out, so H₂O has coefficient 5
- Now, 5 CO₂ + 5 H₂O produced means (10 + 5) or 15 O's come out, so 15 O must go in
- The only way to get 15 O to go in is to have the coefficient of O₂ be 7.5
- So far, we have $\text{C}_5\text{H}_{10} + 7.5 \text{O}_2 \rightarrow 5 \text{CO}_2 + 5 \text{H}_2\text{O}$
- Double all the coefficients to get

$$2 \text{C}_5\text{H}_{10} + 15 \text{O}_2 \rightarrow 10 \text{CO}_2 + 10 \text{H}_2\text{O}$$

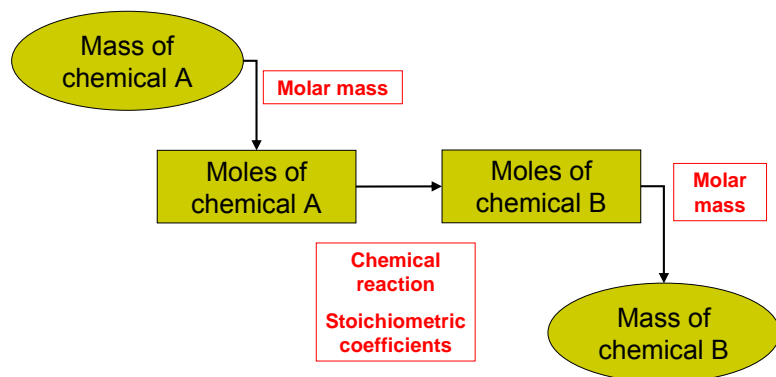
Pictures method of balancing



Let's see that again

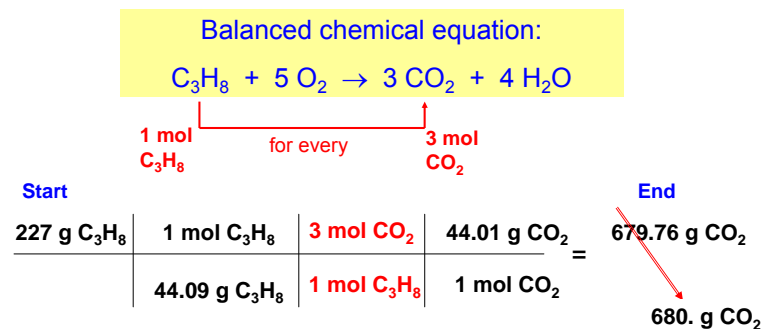


General Strategy for Simple Stoichiometry Problems



Simple Stoichiometry

What mass of carbon dioxide gas is produced when 227 grams of propane (C_3H_8) combust completely?



Types of Stoichiometry Problems

- Simple stoichiometry: Mass of one chemical (reactant or product) is specified. Find out mass of another chemical required or produced in the reaction (assuming just enough of each reactant is present). Must use a balanced chemical equation.
- Chemical analysis: Known and unknown chemicals or quantities given. Figure out unknown chemicals or quantities.
- Limiting reagent: Masses of two different reactants are specified. One of the reactants limits the reaction (it gets used up first). Figure out the maximum mass of a product that could be formed if all of the limiting reactant is used up. Must use a balanced chemical equation.

Limiting Reagent Problems

- If you use an excess of one reactant, it is more likely that all of the other reactant will be used up in the chemical reaction.
- The reactant that is in excess will be present in the final products, since it didn't all get used up.

Chemical Analysis

1. Unknown quantity of a known substance
 - Two known chemicals react to exactly the point of completion
 - Quantity of one of them is known
 - Determine quantity that must have been present of the other one
 - Example: titration
2. Known quantity of an unknown substance
 - A specific amount of an unknown substance reacts
 - Known products are formed, and their quantities are measured
 - Determine empirical formula of the unknown reactant from this information
 - Examples: gravimetric analysis, combustion analysis

Solutions

- Almost nothing in nature occurs naturally as a pure substance
- Most chemical reactions of interest take place in solution, between chemicals that are dissolved in water
- Medium in which something is dissolved is called a solvent
- Water is often called the "universal solvent"
 - It is possible to have solvents other than water (e.g., tincture of iodine is solid iodine crystals dissolved in ethanol)
 - It is possible to have solutions that are not liquids (e.g., air is a bunch of different gases dissolved in N₂)
- The chemical that dissolves is called the solute
- In this course, we will focus on aqueous solutions

Strategies for figuring out what is in a solution

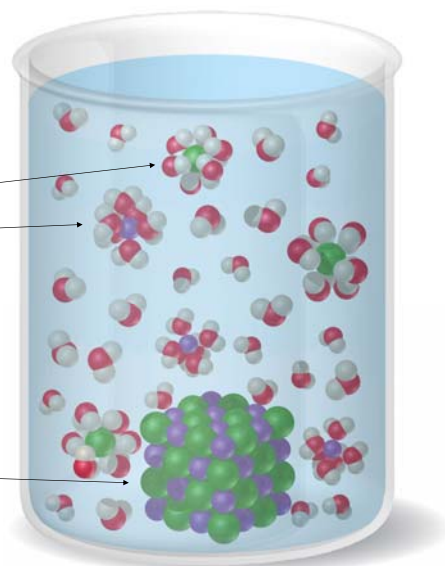
- A solution is a mixture
- An aqueous solution is when a solute is mixed into water (i.e., water is the solvent)
- Usually, you cannot tell by looking at the solution whether it is pure water or an aqueous solution
- Tests you can do in the laboratory:
 1. Evaporation: after water evaporates off, the solute will remain behind as a solid
 2. Electrical conductivity
 - An aqueous solution that contains an ionic compound as the solute will conduct electricity
 - An aqueous solution that contains a molecular compound as the solute will not conduct electricity
 3. Qualitative analysis of ions: use deductive reasoning, trying to react the unknown solution with known solutions containing specific ions

A saturated solution

Cl⁻ ions surrounded by H₂O molecules

Na⁺ ions surrounded by H₂O molecules

Some NaCl (s) remains undissolved because the solvent is holding the maximum amount of Na⁺ and Cl⁻ ions that it can



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What is electrical conductivity?

- Electrical = charged parts due to electrons not balanced with protons on particles
- Conductivity = parts can circulate so that complete circuit can form

Two conditions exist for something to be electrically conductive:

1. Must have charged parts
2. The parts must be mobile

Some Ions Dissolve in Water Better Than Others

Definitions are fuzzy

Very **soluble**

- More than 0.10 mol can dissolve in 1 L of water

Moderately soluble

- Something in between

Pretty much **insoluble**

- Less than 0.01 mol can dissolve in 1 L of water

Understanding Solubility

- “Solubility” refers to the ability of an ionic compound to dissolve in water. More soluble means more dissolves.
- There is no such thing as “completely insoluble.” Solubility is relative. Rather than applying the dualism of soluble vs. insoluble, it is more accurate to talk about *degree of solubility*.
- At the macroscopic level, a common threshold for determining solubility is 0.1 to 1 gram per 100 mL of water.
- At the particle level, when an ionic solid dissolves in water, it breaks apart into single ions which then become surrounded by water molecules.

Classifying Chemical Reactions Based on Evidence Observed

Consider only reactions that occur in aqueous solutions

1. Precipitation reactions
 - Formation of a precipitate (solid) where there wasn't one before
2. Acid-base reactions
 - pH of product solution is different from the pH's of the reactant solution
3. Oxidation-reduction (redox) reactions
 - If the reactions are separated in a special way, electricity can be generated

*Categories are not unique, in part because 1 and 2 are based on evidence, but 3 is based on particle-level model

Solubility guidelines

If one ion from the “Soluble Compounds” list is in an ionic compound, then the ionic compound will dissolve in water

TABLE 4.1 Solubility Guidelines for Common Ionic Compounds in Water

Soluble Ionic Compounds	Important Exceptions
Compounds containing NO_3^-	None
$\text{C}_2\text{H}_3\text{O}_2^-$	None
Cl^-	Compounds of Ag^+ , Hg_2^{2+} , and Pb^{2+}
Br^-	Compounds of Ag^+ , Hg_2^{2+} , and Pb^{2+}
I^-	Compounds of Ag^+ , Hg_2^{2+} , and Pb^{2+}
SO_4^{2-}	Compounds of Sr^{2+} , Ba^{2+} , Hg_2^{2+} , and Pb^{2+}
Insoluble Ionic Compounds	Important Exceptions
Compounds containing S^{2-}	Compounds of NH_4^+ , the alkali metal cations, and Ca^{2+} , Sr^{2+} , and Ba^{2+}
CO_3^{2-}	Compounds of NH_4^+ and the alkali metal cations
PO_4^{3-}	Compounds of NH_4^+ and the alkali metal cations
OH^-	Compounds of the alkali metal cations, and NH_4^+ , Ca^{2+} , Sr^{2+} , and Ba^{2+}

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History of Classes of Reactions

- Names of “classes” of chemical reactions reflects the history of understanding them
- Macroscopic = properties observed in laboratory
 - Precipitation
 - Acid-base
 - Combustion
- Particle level = model of underlying process
 - Redox

Particle Level

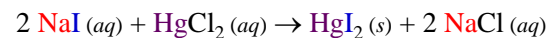
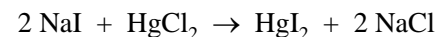
Chemical Reactions

- It is called a “chemical reaction” only if a chemical change occurs
- Chemical changes are ALWAYS about ELECTRONS
- Four kinds of processes involved
 - Transfer of a “proton” (a naked hydrogen atom) from one chemical species to another
 - Sharing of electron pairs between chemical species
 - Transfer of an electron from one chemical species to another
 - Sharing of single electrons between chemical species

Precipitation Rxns

Symbolic Representation

Sodium iodide + mercury (II) chloride



sodium iodide + mercury (II) chloride → mercury (II) iodide + sodium chloride

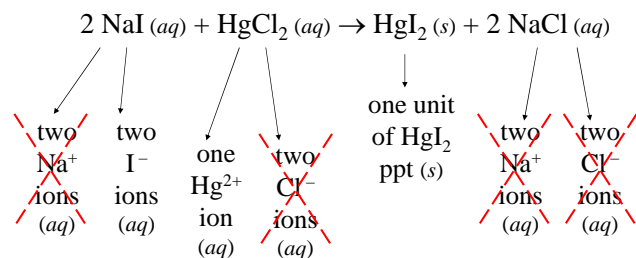
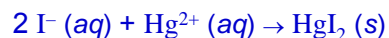
Many names for this pattern (synonyms):

- Exchange reaction
- Metathetical reaction
- Double displacement reaction
- Double replacement reaction

Precipitation Rxns

Symbolic Representation

Sodium iodide + mercury (II) chloride

Net reaction: two I^- ions + one Hg^{2+} ion → one unit of HgI_2 ppt

Writing Net Ionic Equations

1. Start with the balanced reaction, written with phases
2. Identify ions in aqueous solution, ionic solids that precipitate, and any molecules on both sides of the arrow
3. Cross out any spectator ions
4. What's left is the net ionic equation

Particle Level View

Acids vs. Bases

Acids

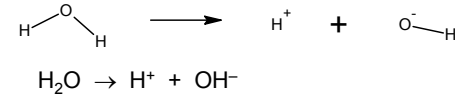
- When a chemical that is an acid is added to water, it increases the H^+ concentration in the solution
- Registers less than 7 on pH scale because $[H^+]$ is greater than neutral water, which is $1.0 \times 10^{-7} \text{ mol/L}$

Bases

- When a chemical that is a base is added to water, it increases the OH^- concentration in the solution
- Registers more than 7 on pH scale because $[OH^-]$ is greater than neutral water, which is $1.0 \times 10^{-7} \text{ mol/L}$

Why is Water So Special?

Water itself breaks into ions:



In a sample of pure water, equal quantities of these ions are present:

$$[H^+] = [OH^-] = 1.0 \times 10^{-7} \text{ mol/L}$$

But when you add a solute that is an acid or base to water, the balance changes.

Strong vs. Weak Acids

- Strong acids (pH 1 to 4) are strong electrolytes
 - Only a few common acids are strong
 - HCl, HBr, HI, HNO₃, H₂SO₄, HClO₄** (*memorize these*)
 - A few other less common acids are strong (you do not need to memorize these)
 - e.g.*, **HClO₃, HBrO₃** and some others
- Weak acids (pH 5-7) are weak electrolytes
 - It is safe to assume that all other acids are weak
 - Most common weak acids:*
 - HF, HNO₂, HClO₂, HSO₃, CH₃COOH**

Another Way to Form Acids

Acid anhydrides

- Nonmetal oxides
- When dissolved in water, react with water to form acid solution
- No change in oxidation number* of the nonmetal
- Examples:
 - $\text{SO}_3 + \text{H}_2\text{O} \rightarrow \text{H}_2\text{SO}_4$
 - $\text{SO}_2 + \text{H}_2\text{O} \rightarrow \text{H}_2\text{SO}_3$
 - $\text{N}_2\text{O}_5 + \text{H}_2\text{O} \rightarrow 2 \text{HNO}_3$
 - $\text{P}_2\text{O}_5 + 3 \text{H}_2\text{O} \rightarrow 2 \text{H}_3\text{PO}_4$

*We will study oxidation numbers soon...

Strong vs. Weak Bases

- Strong bases (pH 11 to 14) are strong electrolytes
 - All bases containing the hydroxide ion (OH⁻) are strong

Most common strong bases:

NaOH, KOH, LiOH, Ca(OH)₂, Mg(OH)₂
- Weak bases (pH 7-10) are weak electrolytes
 - Bases that do not contain the hydroxide ion (OH⁻) are weak

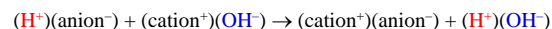
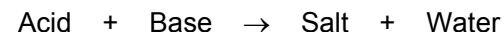
Most common weak bases:

NH₃, CH₃NH₂, C₆H₅NH₂
- Organic alcohols are not bases
 - Organic molecules that contain an alcohol functional group (-OH) do not ionize to form hydroxide ions

Grouping Reactions

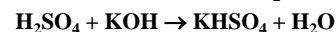
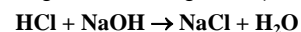
- Exchange rxns: no oxidation number change
 - Precipitation rxns
 - Acid-base rxns
 - Acid + Base → Salt + Water
 - Gas-forming rxns
- Redox rxns: oxidation numbers change
 - Single replacement (*e.g.*, many electrochemistry reactions)
 - Formation of compound from elements (*e.g.*, synthesis of sulfuric acid anhydride)
 - Decomposition of compound into elements (*e.g.*, producing pure metals from metal ores)
 - Combustion

What happens to pH when you mix acid and base solutions? **Neutralization rxn**

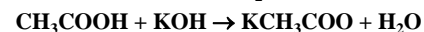
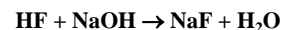


This reaction occurs if the Base contains OH⁻ (in other words, only strong bases)

Strong acid + Strong base (strong acid ionizes nearly completely):



Weak acid + Strong base (weak acid is mostly present as whole molecule):



Important difference: Net ionic equations with strong acids are different than net ionic equations with weak acids!

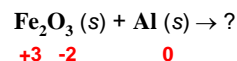
What are oxidation numbers?

- A charge (sometimes real, sometimes fake) that can be assigned to each atom in a compound
 - In charged particles (ions within ionic solid, ions dissolved in water), the charge is real
 - In molecules (which are neutral), atoms don't have real charges
- Measures electron density that resides on the individual atom
 - Caveat: if the atom is in a molecule, and there is more than one of a particular kind of atom, in reality each one doesn't have the same electron density
- A model that provides an accounting method for explaining electron transfer in redox reactions

Redox Reactions

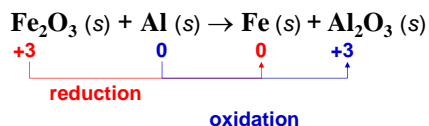
Symbolic Representation

Thermite reaction: Iron (III) oxide + aluminum



<http://www.jce.divched.org/JCESoft/CCA/samples/cca7thermite.html>

What are the products?



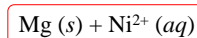
- Fe_2O_3 gets reduced to Fe metal. It gets reduced by Al. Therefore, Al is the reducing agent.
- Al metal gets oxidized to Al_2O_3 . It gets oxidized by Fe_2O_3 . Therefore, Fe_2O_3 is the oxidizing agent.

Solution Concentration

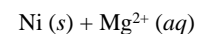
- Many reactions of interest occur in aqueous solution
- To be able to quantify how much of a chemical reacts, and how much product is made, it is necessary to know quantities in moles
- In pure materials, the concentration of particles can be presented as density
- In mixtures that are aqueous solutions, the concentration of the particles of interest (solute) is usually presented as molarity
- Concentration just means how many particles (of interest) are present in a given amount of space
- Since Molarity = moles solute / Liter of solution, if you know the molarity and the volume of solution, you can find the moles of solute

Ranking reactivity of metals with other metal cations in aqueous solution

- Called the activity series
- Metals that are higher are more reactive, *i.e.*, will react more readily than those below them
- Example: which reaction will occur?



or

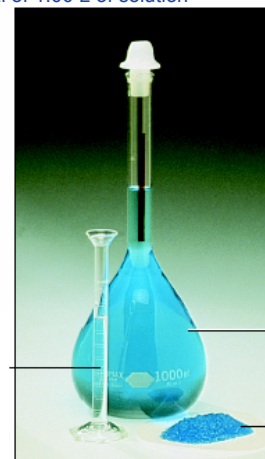


Metal	Oxidation Reaction
Lithium	$\text{Li}(s) \rightarrow \text{Li}^+(aq) + e^-$
Potassium	$\text{K}(s) \rightarrow \text{K}^+(aq) + e^-$
Barium	$\text{Ba}(s) \rightarrow \text{Ba}^{2+}(aq) + 2e^-$
Calcium	$\text{Ca}(s) \rightarrow \text{Ca}^{2+}(aq) + 2e^-$
Sodium	$\text{Na}(s) \rightarrow \text{Na}^+(aq) + e^-$
Magnesium	$\text{Mg}(s) \rightarrow \text{Mg}^{2+}(aq) + 2e^-$
Aluminum	$\text{Al}(s) \rightarrow \text{Al}^{3+}(aq) + 3e^-$
Manganese	$\text{Mn}(s) \rightarrow \text{Mn}^{2+}(aq) + 2e^-$
Zinc	$\text{Zn}(s) \rightarrow \text{Zn}^{2+}(aq) + 2e^-$
Chromium	$\text{Cr}(s) \rightarrow \text{Cr}^{3+}(aq) + 3e^-$
Iron	$\text{Fe}(s) \rightarrow \text{Fe}^{2+}(aq) + 2e^-$
Cobalt	$\text{Co}(s) \rightarrow \text{Co}^{2+}(aq) + 2e^-$
Nickel	$\text{Ni}(s) \rightarrow \text{Ni}^{2+}(aq) + 2e^-$
Tin	$\text{Sn}(s) \rightarrow \text{Sn}^{2+}(aq) + 2e^-$
Lead	$\text{Pb}(s) \rightarrow \text{Pb}^{2+}(aq) + 2e^-$
Hydrogen	$\text{H}_2(g) \rightarrow 2\text{H}^+(aq) + 2e^-$
Copper	$\text{Cu}(s) \rightarrow \text{Cu}^{2+}(aq) + 2e^-$
Silver	$\text{Ag}(s) \rightarrow \text{Ag}^+(aq) + e^-$
Mercury	$\text{Hg}(l) \rightarrow \text{Hg}^{2+}(aq) + 2e^-$
Platinum	$\text{Pt}(s) \rightarrow \text{Pt}^{2+}(aq) + 2e^-$
Gold	$\text{Au}(s) \rightarrow \text{Au}^{3+}(aq) + 3e^-$

↑ Ease of oxidation increases

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Copper (II) sulfate solution has 25.0 g of $\text{CuSO}_4 \cdot 5\text{H}_2\text{O}$ (FW 250. g/mol) included in a total of 1.00 L of solution



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How to Calculate Concentration of a Solution

Need to know:

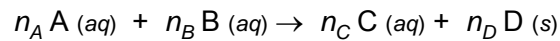
- Amount of solute (in mol)
- Amount of solution (in L)

$$\begin{aligned} \text{Molarity} &= \frac{\text{mol of solute}}{\text{L of solution}} \\ &= \frac{\left(25.0 \text{ g} \times \frac{1 \text{ mol}}{250. \text{ g}} \right)}{1.00 \text{ L}} \\ &= 0.100 \frac{\text{mol}}{\text{L}} = 0.100 \text{ M} \end{aligned}$$

Variations on the Theme

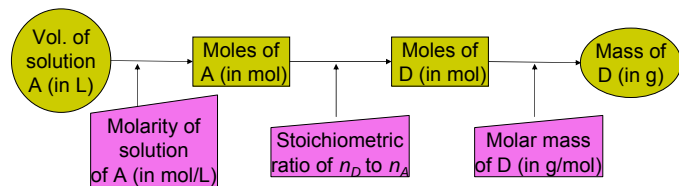
- Three variables:
 - Concentration of solution (molarity)
 - Amount of solute (moles or grams)
 - Volume of solution (liters)
- Given any two, you can always calculate the third
 - What is the concentration of a solution made by mixing ... (solute amount) into water to make a (certain number of liters) of solution?
 - How much solute (grams?) would be required to make (certain number of liters) of a (specify concentration) molar solution?
 - What volume of a (specify concentration) molar solution must be used to obtain (solute amount)?

Solution Stoichiometry



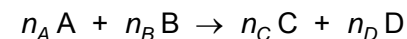
A typical scenario

Given **volume** of a **certain molar** solution of chemical A, find mass that could be produced of chemical D.



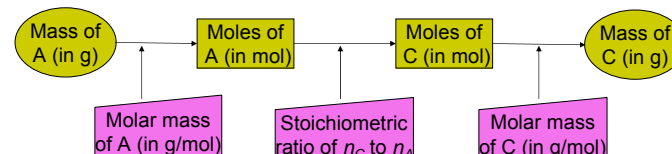
$$\text{volume of solution of A (in L)} \times \frac{\text{molarity of solution A (mol)}}{1 L} \times \frac{n_D \text{ mol D}}{n_A \text{ mol A}} \times \frac{\text{molar mass of D (g)}}{1 \text{ mol D}} = \text{mass of D (in g)}$$

Reminder: Stoichiometry



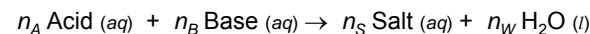
A typical scenario

Given mass of chemical A, find mass that could be produced of chemical C.



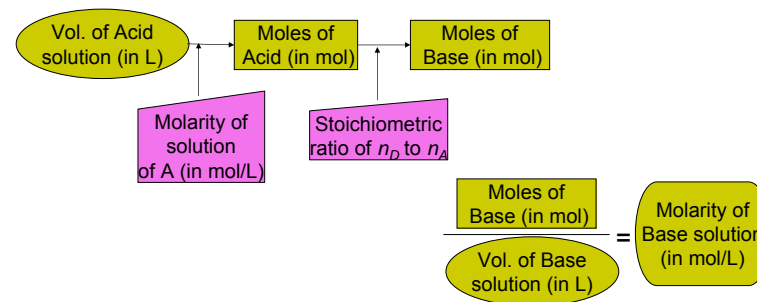
$$\text{mass of A (in g)} \times \frac{1 \text{ mol A}}{\text{molar mass of A (g)}} \times \frac{n_C \text{ mol C}}{n_A \text{ mol A}} \times \frac{\text{molar mass of C (g)}}{1 \text{ mol C}} = \text{mass of C (in g)}$$

Acid-Base Titration



A typical scenario

Given **volume** of a **certain molar** solution of Acid, find concentration of the Base solution if **certain volume** of Base used.



$$\frac{\text{Moles of Base (in mol)}}{\text{Vol. of Base solution (in L)}} = \text{Molarity of Base solution (in mol/L)}$$

Summary of reactions, mostly in aqueous solutions

- Precipitation reactions are exchange reactions
 - One (or more) of the products is insoluble in water
- Acid base reactions are exchange reactions
 - One reactant (the acid) has a loosely bound H^+ ion (can be a weak or strong acid), the other reactant (the base) has an OH^- ion
- Redox reactions involve the transfer of electrons from one reactant to another
 - Involve changes of oxidation number
 - Combustion reactions (which are not aqueous reactions) are redox reactions in which oxygen is always a reactant, and the products are always the oxides of every element in the other reactant which is being burned
 - Oxidation number of one element **increases** in the reactant species (molecule or ion) being **oxidized**
 - Oxidation number of one element **decreases** in the reactant species (molecule or ion) being **reduced**
- Net ionic equation can be written for any reaction
- Solution stoichiometry
 - Involves calculations using molarity of solution
 - Often carried out with technique called titration