Introduction and Fundamental Concepts

Chemistry 35

Fall 2000

Introductory Stuff

- Chemistry is an *experimental* science
- How is science done?
  - *Observation*
  - *Hypothesis*
  - *EXPERIMENT*
    - Observation
    - Revise Hypothesis?
    - NEW Experiment

- Eventually: *hypothesis* -> Theory -> ***LAW***

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Experiment:

- Fill three balloons with different gas mixtures:

  - Balloon I: N_2 and O_2
  - Balloon II: H_2 and O_2
  - Balloon III: H_2 and N_2

Question: "What will happen if we touch a lit candle to the balloons? Will a *CHEMICAL REACTION* occur?"

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Balloon I

- essentially, AIR (80% N_2, 20% O_2)

  - **O_2:** *quite reactive*
  - **N_2:** *generally, pretty inert* (used as a preservative atmosphere for food, chemicals)

-**Prediction:** ***NO REACTION***
(this is our CONTROL)

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Balloon II

- Mixture is:

  ROCKET FUEL!

- *Significant* reaction:

  - audible acoustic retort
  - flames
  - screams from the dead and wounded

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Balloon III

- H_2 is highly reactive with O_2, but not so much with N_2

  Prediction: no significant reaction, unless there is some source of O_2

- ***Result:***

  SUBSTANTIAL explosion!

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How do we write a reaction?

- In words:

  oxygen + hydrogen -> (product)

- In symbols:

  O_2 *(g)* + $2H_2$ *(g)* -> $2H_2O$ *(g)* (+heat)

  *- called a **combustion reaction***

- Symbols indicate *structure:*

  O=O + 2H-H -> 2H-O-H

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Chemical Change

- Bonds were broken (H-H and O=O)
- Bonds were formed (2H-O bonds)
- The *IDENTITIES* of the reactant compounds were **CHANGED by the process**
- *How does this differ from a purely **physical** change?*

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Reaction Energetics

- Energy is *consumed* to break bonds
- Energy is ***released*** **when making bonds**

If $E_{released} > E_{consumed}$,
HEAT is given off

(*Thermodynamics!*)

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Quantifying Chemical Change

- H_2 and O_2 react in a 2:1 ratio to make water

-reaction must be balanced
(Conservation of Mass)

-for a compound, elements always present in *definite proportions*
(*Stoichiometry*)

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Quantifying Chemical Change

- In our experiment, we had a 1:1 mixture of gases:

H_2	O_2
H_2	O_2
Both React	Only *ONE* Reacts
(Limiting Reagent)	(EXCESS Reagent)

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Quantifying Chemical Change

- Energy release:

8.0×10^{-19} Joules

(for each molecule of O_2)

In our balloon: 15,000 Joules

- about 3.5 KCalories (Kcals)
- Calories does not = “Calories”

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Gas Properties

- Why did the balloon EXPLODE?
  - -rapidly heated gases
  - -rapid *EXPANSION* of gases
  - -***explosive shockwave!***
- How can we make a *bigger* explosion?
- Where did the product (water) go?
  - -2 grams of reactant gases make less than a thimbleful of water
  - -*vaporized!*

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Other Questions:

- Can the *REVERSE reaction occur?*

  $2H_2O \rightarrow 2H_2 + O_2$

  How?
- Why doesn't the H_2/O_2 mixture *spontaneously* explode (without being ignited)?

  KINETICS!

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Still *More* Questions!

- What happened with Balloon III?

  Two possible reactions:

  1) $3H_2 + N_2 \rightarrow 2NH_3$
  2) $2H_2 + O_2 \rightarrow 2H_2O$

  (#1 is too slow; #2 got O_2 from the air)

- How do we *know* that atoms and molecules exist?

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Some **History**

- Democritus

  -Greek Philosopher (5th Century B.C.)

  Asked: *"Is there a limit to which something can be divided and yet still remain the same material?"*

  YES!

  Atomos ("indivisible")

  BUT: *there was no experimental evidence*

  "Show us!" – Aristotle and Plato

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More History

- <u>Lavoisier</u>

  -18th Century Frenchman

  -Wrote the 1st Chemistry text

  -Considered the “Father of Chemistry”

  Rigorously quantified masses **before and after** a chemical reaction (in a closed system):

  <u>Law of Conservation of Mass</u>

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Still more History

- <u>Proust</u>

  -also an 18th Century Frenchman

  Found that: *regardless of how prepared*, a compound will always have the same composition (same elements present in the same proportion)

  Law of Definite Proportions

<u>BUT:</u> still didn’t know ***WHY*** these laws were followed

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along comes: John Dalton

- 19th Century British Schoolteacher
- He found: elements can combine to form *different* compounds when they combine in *different proportions:*

Law of Multiple Proportions

- Based on these laws, he proposed an:

Atomic Theory of Matter

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Dalton's Atomic Theory (1808)

- Matter is composed of **Atoms**
- Atoms are *indivisible particles*
- All atoms of an element are identical
- Different elements -> Different atoms
- Atoms retain their identities in chemical reactions
- Compounds are formed by *combining atoms* of different elements in simple whole-number ratios

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Quantitative Measurements

"If you can't measure it, you can't really understand it."

- Measured quantities have ***two parts*** :
  1) ***number***
  2) ***unit***

Example: **2.367** *quarts*

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Units

- ***SI*** Units (***S***ystème ***I***nternational) are the commonly used units for scientific measurements

- ***Base*** SI Units:

  Length: **meter, m**

  Mass: **kilogram, kg**

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Base SI Units

TABLE 1.4 SI Base Units

Physical Quantity	Name of Unit	Abbreviation
Mass	Kilogram	kg
Length	Meter	m
Time	Second	s[a]
Electric current	Ampere	A
Temperature	Kelvin	K
Luminous intensity	Candela	cd
Amount of substance	Mole	mol

[a] The abbreviation sec is frequently used.

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Prefixes

TABLE 1.5 Selected Prefixes Used in the Metric System

Prefix	Abbreviation	Meaning	Example
Giga-	G	10^9	1 gigameter (Gm) = 1×10^9 m
Mega-	M	10^6	1 megameter (Mm) = 1×10^6 m
Kilo-	k	10^3	1 kilometer (km) = 1×10^3 m
Deci-	d	10^{-1}	1 decimeter (dm) = 0.1 m
Centi-	c	10^{-2}	1 centimeter (cm) = 0.01 m
Milli-	m	10^{-3}	1 millimeter (mm) = 0.001 m
Micro-	m [a]	10^{-6}	1 micrometer (m m) = 1×10^{-6} m
Nano-	n	10^{-9}	1 nanometer (nm) = 1×10^{-9} m
Pico- m	p	10^{-12}	1 picometer (pm) = 1×10^{-12}
Femto- m	f	10^{-15}	1 femtometer (fm) = 1×10^{-15}

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More Units

- *Derived* SI Units:

Volume: **m^3**

-more commonly: **cm^3** (cc or mL)

-even more commonly: **Liter, L** (NOT an SI unit!)

Density: **g/cm^3**

-originally used to define mass (1 gram = mass of water in 1 cm^3)

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Energy Units

Not quite so obvious: *$kg\text{-}m^2\text{-}s^{-2}$*

-*Gravitational Potential Energy* =

mass x acceleration x length =

kg x m/s^2 x m = ***Joules***

-*Kinetic Energy* = mass x $(\text{velocity})^2$

= kg x m^2/s^2 = ***Joules***

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Dimensional Analysis

- Use *conversion factors* to change units:

1 inch = 2.54 cm
12 inches = 1 foot
1 m = 1000 mm
1 Å = 10^{-10} m

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Dimensional Analysis Example

-How many cm are there in 1.7 miles?

We know: 2.54 cm = 1 inch
12 inches = 1 foot
1 mile = 5280 feet

miles -> feet -> inches -> cm

$$1.7 \text{ miles} \times \frac{5280 \text{ feet}}{1 \text{ mile}} \times \frac{12 \text{ inches}}{1 \text{ foot}} \times \frac{2.54 \text{ cm}}{1 \text{ inch}} = \mathbf{273{,}588.48 \text{ cm}}$$

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Significant Figures (digits)

- There is *measurement uncertainty* associated with **every** measured quantity:

1.7 miles – uncertainty is in last digit

273,588.48 cm **–** *where is the uncertainty?*

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Sig Figs: Rules

- Measurement uncertainty expressed in last digit
- In *addition and subtraction*: uncertainty in result is limited by value with the greatest ***absolute uncertainty***
- In *multiplication and division*: uncertainty in result is limited by value with the greatest ***relative uncertainty***

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Rounding and Zeros

Rounding

If <5, *round down*

If >5, *round up*

If =5, ***round to nearest EVEN number***

Only round at the END of a calculation!

Zeros

All zeros are significant ***EXCEPT*** those that **only** locate a decimal point

Not certain? Use **Scientific Notation**

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Example: Final Rounded Answer

$1.7 \text{ miles} \times \frac{5280 \text{ feet}}{1 \text{ mile}} \times \frac{12 \text{ inches}}{1 \text{ foot}} \times \frac{2.54 \text{ cm}}{1 \text{ inch}} =$ ***273,588.48 cm***

-limited to TWO sig figs in result

- ***273,588.48 cm rounds to:***

270,000 cm

or

2.7×10^5 cm (best!)

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The "Munowitz Perspective"

- Overview

  -Matter behaves according to the basic ***physical laws*** of nature, so we can use this knowledge to understand *Chemistry*

  -What ***FORCES*** predominate at the atomic and molecular levels?

  -**Gravitational?**

  -**Electromagnetic?**

  -Gosh, this seems a lot like PHYSICS!

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Force, Work and Energy

- **Force:** depends on **mass** or **charge**

  -UNIT: Newtons (N = kg-m-s^{-2})

- **Work**

  - *the application of* ***force*** *over a distance*

- **Energy**

  - *the ability to* ***do work***

  -UNIT: Joules (J = N-m = kg-m^{2}-s^{-2})

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Quantifying Matter: Moles

- **An SI Unit (N_o):**
  1 mol = 6.022137 x 10^{23} species

  Avogadro's Number

  -mole: Latin ("heap, pile")

  Amedeo Avogadro's Hypothesis (1811):
  Equal volumes of different gases contain equal numbers of particles

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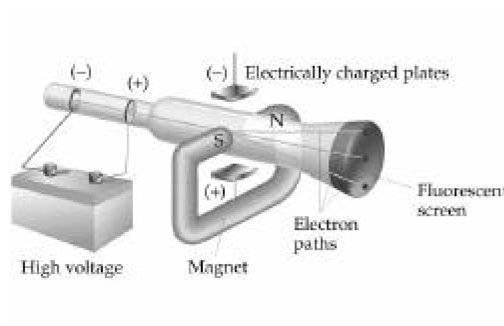
Dalton was *WRONG!*

- Well . . . not *entirely* wrong, but

  Atoms ***ARE*** divisible

- **1897:**
  J.J. Thomson



Cathode Rays:
charge/mass ratio

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Millikan Oil Drop Experiment

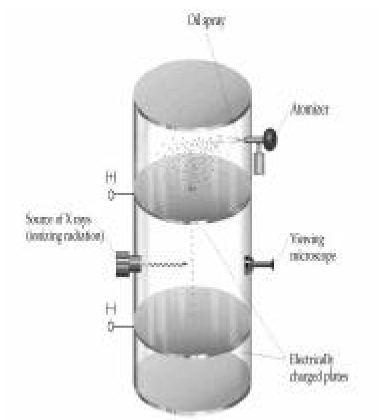
- **1909:**

  ->measured charge on an electron:

  1.6 x 10^{-19} C

  -> **calc (using Thomson's data) the MASS of an electron:**

  9.11 x 10^{-28} g



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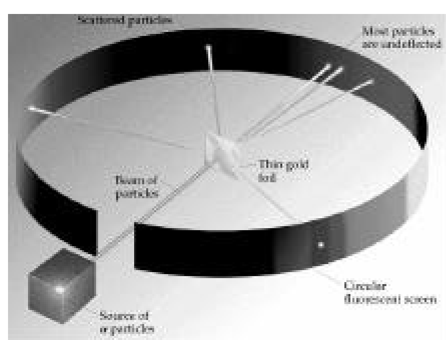
Plum Pudding and Rutherford

"The atom has a structure like plum-pudding" - *J.J. "Boom-Boom" Thomson*

1910: Ernest Rutherford

-shot α-particles at a thin foil of gold

-led to the **Nuclear** model of the atom



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The Nuclear Atom

- **Atoms are composed of:**

  electrons (e^-): neg charge, light
  protons (p^+): pos charge, **heavy**
  neutrons (n^0): NO charge, **heavy**

$$\#p^+ = \#e^-$$

$$n^0 \text{ (mass)} \approx p^+ \text{ (mass)} \approx 2000 \times e^- \text{ (mass)}$$

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How do Elements Differ?

- Elements classified by $\#p^+$:

  Atomic Number (Z) = $\#p^+$

- What about $\#n^0$?

  $\#n^0$ affects *mass number* only:

  Isotopes!

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Isotopes of Hydrogen

^{1}**<u>H</u>**: 1 proton, 0 neutrons, 1 electron

^{2}**<u>H</u>**: 1 proton, *1 neutron,* 1 electron
- Deuterium

^{3}**<u>H</u>**: 1 proton, *2 neutrons*, 1 electron
-Tritium

-Behave the same, chemically

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Quantifying Atomic Mass

- New Unit: ***amu*** (atomic mass unit)

  By definition:

  mass of ^{12}C nucleus = 12.000 amu

- For a mole of an element, we use the ***molar mass*** (= g/mol) which is *numerically* identical to the #amu for the element:

  1 mol ^{12}C = 12.000 grams (exactly)

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Why Fractional Molar Masses?

- Need to consider the natural abundances of *isotopes*

Example: Chlorine

$75.5\%\ ^{35}Cl + 24.5\%\ ^{37}Cl$

$(0.755)(34.97) + (0.245)(36.97) =$ ***35.45 g/mol***

-This is a *weighted average*; 1 mol of Cl will have a mass of 35.45 grams

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The Same or Not the Same?

- Are all Cl_2 molecules the same?
  -3 possible combos *(isotopomers):*
  $^{35}Cl^{35}Cl$ or $^{35}Cl^{37}Cl$ or $^{37}Cl^{37}Cl$
  (56.5%) (37.3%) (6.2%)
- Ok, what about Hemoglobin?
  $C_{2954}H_{4516}N_{780}O_{806}S_{12}Fe_{4}$ - a BIG molecule!
  3 3 2 3 4 4 <- *isotopes (nat'l)*

-The chances of any two hemoglobin molecules in a drop of blood being isotopically IDENTICAL, is VERY VERY SMALL!

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