Lecture 2 Notes: Number in Chemistry -- Stoichiometry

PRS quiz on reading

Logistics

- Show chem31a.stanford.edu, syllabus & calendar
- Office hours start today after 2nd lecture (2:15-3:30p). Come see the instructors
- Sections start Th & F – discuss section rules.
- Laboratory dress code will be enforced in section
- Contact Marissa Caldwell (macaldwe@stanford.edu) for section assignment issues.
- Reading and reading problems due F before lecture
- Problem Set 1 due Monday by 1:15pm. Submit in boxes opposite Mudd 235 by 1:15pm

Group work in section is just like in Science

- Be demanding of yourself
- Be kind to others
- Engage in argument. Not the mean kind, the thoughtful kind.
- No where to run, no where to hide.
- Be proud of each other – thinking is hard work.

The development of natural science

- **Galileo's** strict adherence to the **experimental method 400 years ago**
- **Newton**'s development of the first **general scientific theory 330 years** ago.
- Natural science is **very young** compared with most branches of knowledge.

- Chemistry is even younger
- **Lavoisier's identification of the element oxygen** 230 years ago http://www.chemheritage.org/classroom/chemach/forerunners/lavoisier.html

Lavoisier in many ways represents the epitome of what we mean by the discipline of chemistry. After studying law in college to satisfy his lawyer father and practicing geology for several years after graduation, he was soon pushing the envelope of the emerging and eminently practical science of chemistry. At age 32 he was appointed head of what might be the first national lab, part of the **Royal Gunpowder and Saltpeter Administration** in Paris.

Lavoisier perfected the defining method of chemistry -- a repeated cycle of:

- making matter
- measuring matter
- modeling matter

Goal: as general a theory of matter and its transformations as experiment will sustain.

Lavoisier's key paper establishing oxygen as an element in air is available in translation: http://web.lemoyne.edu/~giunta/lavoisier.html

Using the methods perfected by Lavoisier and the atomic theory advocated a couple decades later by **Dalton**, chemists continued to struggle to understand chemistry throughout the 19th century.

- As late at 1849, during the gold rush that created the state of California, chemists remained unsure whether water was HO or H_2O.
- Evidence for atoms was still indirect and controversial.
- Even late in the 19th century, many scientists considered the atomic theory just a hypothesis.
- Lack of easily observable evidence of atoms made chemistry a very late bloomer.

The breakthrough came with the discovery of the **free electron by J. J. Thomson in 1897, just over a century ago**. I want to tell the story of the electron and what else it unearthed to give a sense of how observations and measurements led to the modern science of chemistry.

Electrons

- Electricity was well known
- Particles of electricity were not yet known.
- Thomson found that a stream of particles comes out of a negatively charged piece of hot metal in a vacuum.
- The particles then accelerate away from the negatively charged metal.
- This told him that the particles are negatively charged – named them "electrons".
- The hotter the metal, the more frequently electrons break away.
- The more negatively charge the hot metal, the more rapidly the electrons accelerate away from the metal.

- When these electrons hit a piece of matter, visible light is often emitted, which provides an easy way to see where the stream of electrons is going.
- This electron-stimulated emission is the basis of fluorescent and neon lamps as well as old-fashion televisions.
- Thomson measured the ratio of the mass of the electron, $\boldsymbol{m_e}$, to its electrical charge, $\boldsymbol{e}$, from the acceleration caused by the negatively charge metal:

$$\boldsymbol{m_e/e = 5.7\text{x}10^{-12}\text{kg/C}}$$

- He could not measure either m_e or e separately.
- Millikan published a definitive measurement of $\boldsymbol{e}$ in 1913. He collected small numbers of electrons on small oil droplets and carefully measured the charge on the oil droplets.
- The charge came in integer multiples of a small constant charge: $\boldsymbol{e = 1.6\text{x}10^{-19}\text{C}}$
- With m_e/e and e accurately known, the mass of the electron is obtained: $\boldsymbol{m_e = 9.1\text{x}10^{-31}\text{kg}}$.

Millikan's 1923 Nobel Address is a great example of the power of careful experiments to unambiguously establish physical phenomena that are well beyond normal observation: http://nobelprize.org/physics/laureates/1923/millikan-lecture.html.

Having separated electrons from a chunk of metal and accelerated them, the next step was to study what, in addition to giving off light, happens when matter is hit by electrons.

Mass Spectrometry

- When an accelerated electron runs through a gas, other electrons can be knocked out of the gas – these are called "secondary electrons".
- The accelerated electron is like a queue ball knocking pool balls out of a neatly racked set of balls – except that typically only one secondary electron is knocked out per collision.
- The positively charged particles formed are called ions.
- Thomson and many others measured the mass-to-charge ratio of the positive ions.
- The charge of the ions is a positive multiple of ***e,*** usually +***e***
- The lightest ion that is ever formed is called a proton.
- From the proton's mass-to-charge ratio and the value of *e*, the **mass of the proton =** $\boldsymbol{m_p}$ **= 1.67x10^{-24}g** (*on board*)
- The proton is about **2,000** times heavier than the electron.
- Positive ions are enough heavier than the electron that the mass of the electron can often be ignored.
- The measurement of the mass of ions this way is called "mass spectrometry"

<u>Masses of ions are integer multiples of mass of proton</u>

- Ions from water vapor have masses of 1.0, 2.0, 16.0, 17.0 and 18.0 times the mass of the proton.
- Ions from air have masses of 14.0, 16.0, 28.0 and 32.0 times the mass of the proton.
- Fractional multiples, such as 16.5, are never observed.

Relation to "combining masses" of Dalton's atomic theory

Ions from water vapor

mass	*identity*
1.0 m_p	H^+ (proton or hydrogen ion)
2.0 m_p	H_2^+ (2 protons and an electron)
16.0 m_p	O^+ (the oxygen ion)
17.0 m_p	HO^+
18.0 m_p	H_2O^+

Ions from air

mass	*identity*
14.0 m_p	N^+
16.0 m_p	O^+
28.0 m_p	N_2^+
32.0 m_p	O_2^+

- Dalton had to make difficult, indirect measurements
- Great confusion about Dalton's combining masses
- Thomson measured the exact masses of the ions
- Add in the small mass of the missing electron to get the masses of atoms and molecules.

How can we use these masses?

How many water molecules are in one liter ($1.0dm^3$) *of liquid water?*

(number H_2O) = (mass of all H_2O in sample)/(mass of one H_2O)

$= (\text{volume sample})(\text{density } H_2O)/(18.0\ m_p)$

$= (1.0dm^3)(1.0g/cm^3)/[18.0(1.67x10^{-24}g)]$

$= (1.0x10^3\cancel{cm^3})(1.0\cancel{g}/\cancel{cm^3})/(3.00x10^{-23}\cancel{g})$

$= 10^{26}/3.0 = \mathbf{3.3x10^{25}}$

That is a lot of water molecules: 33 septillion molecules!

Isotopes

- Chlorine is not as well behaved as hydrogen, oxygen and nitrogen.
- Mass spectroscopy shows that three quarters of chlorine atoms weigh 35.0 m_p and the other fourth weigh 37.0 m_p
- Dalton's followers had found that chlorine had a combining mass of 35.5 times hydrogen's combining mass.
- 35.5 m_p is the weighted average of 35.0 m_p and 35.5 m_p

  $35.5\ m_p = (0.75)(35.0\ m_p) + (0.25)(37.0\ m_p)$
- Many elements have more than one mass of atom – called "isotopes"
- Isotopes of the same element have almost identical chemical properties
- Which is why they weren't identified by Lavoisier, Dalton or their followers.

Mass and Number in Chemistry:

We will take a more systematic approach to the branch of chemistry called "Stoichiometry" than the book uses. That is why I have not assigned all of Chapters 1 and 2. Please be sure you understand this week's activity in section. Really grapple with the material to gain a conceptual understanding of the role of **number** in chemistry. It will serve you very well when we get to the harder parts of the course.

<u>Atomic Mass Unit</u>

- Internationally recognized unit of mass for atoms, molecules and subatomic particles is the atomic mass unit
- Standard abbreviation is "u", but we will to use "amu" as our book does

- 1 atomic mass unit is approximately the mass of the proton or the hydrogen atom. More precisely, it is defined to be exactly one twelfth the mass of a ^{12}C atom:

$$1\text{amu} \equiv \frac{\text{mass } ^{12}\text{C}}{12} = 1.6605402\text{x}10^{-27}\text{kg} = 1.6605402\text{x}10^{-24}\text{g}$$

("≡" means "defined to be")

The Mole

The mole (mol) is a very useful unit of number. It is *the number of atomic mass units in a gram*.

$$1\text{mol} \equiv \frac{1\text{g}}{1\text{amu}} = \frac{1g}{1.6605402 \times 10^{-24}\text{g}} = 6.0221367\text{x}10^{23}$$

The mole is a synonym for Avogadro's **Number**

To help you understand the definition better, let me define the dozen as "the number of inches in a foot":

$$1\,\text{dozen} \equiv \frac{1\,\text{foot}}{1\,\text{inch}} = \frac{12\,\text{inch}}{1\text{inch}} = 12$$

Now think of the mole as "the number of atomic mass units in a gram".

Teachers, textbooks and even some international organizations often treat the mole as something other than a number an provide algorithms for solving certain types of chemistry problems using the mole. These approaches can leave the student with a fragile understanding that makes it harder to tackle more difficult concepts later in a chemistry course. In this class, I want you to treat the mole as **JUST** a number. Please set aside any thing you may have learned previously about "chemical amount" or "dimension of chemical substance", or

anything you may read in the unassigned parts of our text. Just consider the mole to be a gigantic cousin of the dozen.

<u>A convenient relationship that makes calculations simpler</u>

Since: $$1\text{mol} \equiv \frac{1\text{g}}{1\text{amu}}$$

We can rearrange to: $\boxed{1\text{amu} = 1\text{g/mol}}$

- 1amu and 1g/mol are **exactly** the same thing.
- 1g/mol is a gram divided by Avogadro's Number – it's tiny!
- The mass of the proton is about 1g/mol

$$m_p \cong 1\text{amu} = 1\text{g/mol}$$

How many water molecules are in 1.0dm^3 *of liquid water?*

(number H_2O) = (volume sample)(density H_2O)/(mass of one H_2O)

$= (1.0\text{dm}^3)(1.0\text{g/cm}^3)/(18.0\text{amu})$

$= (1.0\text{x}10^3\cancel{\text{cm}^3})(1.0\cancel{\text{g}}/\cancel{\text{cm}^3})/(18.0\cancel{\text{g}}/\text{mol})$

$=$ **56mol** $= 3.3\text{x}10^{25}$ (from previous version of problem)