Lecture Notes

Chemistry

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Jasper Runco

Instructed by
Catherine Drennan
Assisted by

Contents

Chapter 1

Average atomic mass

Date: 2020-05-19
Anouncements:

make inkscape always open in workspace 3 ✓
get inkscape shortcut manager configured

1.1 Introduction

The periodic table we take for granted is the result of thousands of years of work trying to sort out the complexity of the physical world. How do you make sense of it?

At the end of the day, we realized that the building blocks, elements, that make up matter are of a limited number, and they are disproportionately present. They have different properties and are themselves made up of other building blocks.

With chemistry, we can start to make use of the math and physics to understand the world. On top of that, we build biology, based on molecular interactions.

Elements and atoms

Observe that different substances have different properties and they react with each other differently. Certain types of air particles. . . it leads to a question . If we break down carbon to the smallest possible size that it would still have the properties of carbon. We call this an atom, the most basic unit of any element. They are unimaginably small. One hair is one million carbon atoms wide.

The elements are related to each other

A carbon atom is made up of the same subatomic particles as a gold atom.

Proton

The elements are written in order of atomic number, which is the number of protons in it's nucleus. Protons define the element.

Neutrons

Neutrons are in the nucleus and their number can change without changing the element. Example: ${}^{12}_{6}\mathrm{C}$ and ${}^{14}_{6}\mathrm{C}$.

Electrons

These buzz around the nucleus and determine the net charge of the atom.

To operate at such a small scale, we need to have some units of measurement. How to we define mass at the atomic scale. Historically, scientists have used the **unified atomic mass unit (amu)** and is defined as

$$u = 1.660540 \times 10^{-27}\mathrm{kg}.$$

This seems like a strange number. In fact, this deffinition makes it cleaner when talking about the mass of atoms. The mass of a proton is approximately $1u$ and a neutron is approximately $1u$

The electron mass is almost $\frac{1}{2000}$ the mass of the proton and neutron.

1.2 Periodic table

The number on top is the elements atomic number, i.e. the number of protons in the nucleus. But what would the mass of hydrogen be? There are different **isotopes** of hydrogen, the most common form has 1 proton, 0 neutrons, and 1 electron. It's mass is $\approx 1.008u$. But the periodic table takes the weighted average of all isotopes to give the **average atomic mass**. Because they don't put units in, it's really the relative atomic mass.

1.3 Worked example: Atomic weight calculation

If we look at this table, we see carbon-12 is the most common isotope on earth.

Table 1.1: Atomic mass

^{12}C 98.89%	12.0000 amu
^{13}C 1.110%	13.0034 amu

The periodic table reports the atomic weight of carbon as 12.01. It is callculated by taking the weighted average of these two isotopes:

$$0.9889 \cdot 12 + 0.0111 \cdot 13.0034 = 12.0111374 \approx 12.01.$$

The difference in atomic mass is from the number of neutrons the isotope has.

1.4 The mole and Avogadro's number

Here we connect the avg. atomic mass to the masses we will likely handle in a lab. The chemistry community has come up with the following tool, take Lithium for example: it has $6.94\frac{u}{atomLi}$. How many atoms would a sample have it it had the mass $6.94g$Li. This number, named Avogadro's number, is

$$\boxed{\approx 6.022 \times 10^{23}} \tag{1.1}$$

Constant: Avogadro's number

One **mole** of an element is this many atoms of the element. In chemistry practice finding the number of moles is a useful thing, and finding the number of atoms is simply multiplying the number of moles by Avogadro's number. So to go from mass to moles, devide by molar mass. Then multiply by Avogadr's number if you want to know the number of atoms.

1.5 Atomic number, mass number, and isotopes

Definition 1 (atomic number (Z))**.** *The number of protons in a nucleus, the number of electrons (in a neutral atom)*

Definition 2 (isotope)**.** *atoms of a single element (same atomic number) that differ in the number of neutrons in their nuclei (different masses)*

Chapter 2

Atomic Structure

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2.1 History: discovery of the electron

J. J. Thompson was interested in cathonde rays. He had a hydrogen gas cylinder and put a charge to it, and got these rays. He was curious to see what would happen to the rays as they passed a charged deflection plate, and saw the cathode ray being deflected towards the positive plate, which told him the cathode rays were made of **negatively** charged particles.

$\Delta x_{(-)} \propto \frac{e_{(-)}}{m_{(-)}} \frac{charge}{mass}$.

This deflection was directly proportional to charge and inversely to mass. When he applied a greater charge, he saw a small deflection towards the negatively charged particle, which told him there were also **positively** charged particles.

$$\Delta x_{(+)} \propto \frac{e_{(-)}}{m_{(-)}} \frac{charge}{mass}.$$

The relative deflection told him something about the masses of these particles

$$\frac{\left|\Delta x_{(-)}\right|}{\left|\Delta x_{(+)}\right|} = \frac{\left|\frac{e_{(-)}}{m_{(-)}}\right|}{\left|\frac{e_{(+)}}{m_{(+)}}\right|} = \frac{m_{(+)}}{m_{(-)}}.$$

This negatively charged particle was later named the electron and its mass was very small $m = 9.11 \times 10^{-31} kg$.

2.2 Rutheford Discovery of the Nucleus (1911)

Rutheford and his students did an experiment with a sample Curie sent them. They had a detector that counted the number of α particles comming off the radioactive sample. They counted and then put thin gold foil in the path. When they counted, they seemed to be the same. They moved the detector behind the foil to detect a small number of backscatter.

$$P = \frac{\text{count rate backscattered}}{\text{count rate of incident particles}} = \frac{20}{132,000} = 2 \times 10^{-4}.$$

2.2.1 Interpretation:

The Au atoms are mostly **empty.**

The majority of each atom's mass is concntrated in a very small volume, we call now the **nucleus**

$$\boxed{Z = \text{ atomic number}} \tag{2.1}$$

$$Z$$

$$\boxed{e = \text{ the absolute value of an electron's charge}} \tag{2.2}$$

$$e$$

Definition 3. *Charge of the electron in an atom* $= -Ze$

Definition 4. *Charge of the nucleus in an atom* $= +Ze$

Rutheford used backscattering to measure the diameter of the nucleus

Definition 5 (Diameter of the nucleus)**.** $10^{-14}m$

2.3 (Failure of) the classical description of the atom

Definition 6 (Columb's Force law)**.** *Columb Force* $F(r) = \frac{Q_1 Q_2}{4\pi\epsilon_0 r^2}$

$$\boxed{Q_x = \text{ charge on particle x}} \tag{2.3}$$

$$\boxed{r = \text{ distance between two charges}} \tag{2.4}$$

Definition 7 (Permittivity constant of a vacuum)**.** $\epsilon_0 = 8.854 \times 10^{-12} C^2 J^{-1} m^{-1}$

The electron should plummet into the nucleus, the equations don't work at small scales.

Chapter 3

Wave-Particle Duality of Light

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3.1 Properties of Waves

Waves have a periodic variation of some quntity.

Definition 8 (Amplitude). *The deviation from the average level*

Definition 9 (Wavelength). *(λ)The distance between succesive maxima*

Definition 10 (Frequency). *(ν) Number of cycles per unit time. Units of frequency are # of cycles / second = $Hz = s^{-1}$.*

Definition 11 (Period). $\frac{1}{\nu}$ *The time it takes for one cycle to occure*

Definition 12 (Intensity). $= a^2$

Definition 13 (Speed). $\lambda\nu$ *(m/sec)*

Constant 1 (Speed of light). $c = \lambda\nu = 2.9979 \times 10^8\ m/s\ c \approx 670,000,000\ mph \approx 186,000\ miles/s$

Note: ROY G. BIV

3.1.1 Superposition

Waves can interfere constructively when they are superpositioned in phase. Distructive when out of phase.

3.1.2 Light as a particle

Photoelectric effect

A beam of UV light hitting a metal surface can eject electrons if the frequency ν is greater or equal to a threshold frequency ν_0

They changed the frequency with a constant intensity: below threshold frequency no electrons, at threshold there were a certain number of electrons, but above the threshold there was no change in number of ejections.

They looked at the KE of the light as a function of the frequency, and found they increased proportionally.

They then looked at how the intensity increased the KE, but it stayed constant.

Looking at the number of electrons as a function of intensity, they were directly proportional.

These results were all unexpected, untill Einstein looked into the data. Different metals had different threshold frequencies, but they all had the same slope. $m = 6.626 \times 10^{-34}$ Js

Planck came up with the number when studying black body radiation.

Constant 2 (Planck's constant). $h = 6.626 \times 10^{-34}\ Js$

Rewriting the equation of the line, Einstein came up with the frequency being proportional to the energy

$$\boxed{K.E. = h\nu - h\nu_0} \tag{3.1}$$

Photoelectric effect

ν **-** frequency of incident light

$h\nu$ **-** the **energy of the incident light** $= E_i$

ν_0 **-** threshold frequency

$h\nu_0$ **-** threshold energy or **workfunction (ϕ)**

Ultimately, he realized energy of light is proportional to frequency.

$$\boxed{E = h\nu} \tag{3.2}$$

Photon Energy - Einstein (1905)

Note 1 (Units:). $J = (Js)(s^{-1})$

Light is made up of energy "packets" called photons, where the energy of the photon depends on its frequency.

Definition 14 (Intensity). *number of photons per second (Units: Watt, J/s)*