Chapter 2

Atoms and the Atomic Theory

1

Origin of the idea of atoms

Starts with the ponderings about Nature and existence in pre-Socratic Greek philosophy (600 BC – 425 BC)

Their work was called "Natural Philosophy"

In fact, until the 19th century the term "Natural Philosophy" was the term used for "Science", or the practice of studying Nature.

What you normally hear/read about how ancient Greek philosophers arrived at the idea of atoms is <u>almost certainly wrong</u>.

2

Pre-Socratic Greek philosophers arrived at the concept of atoms because there was no other way to explain the existence of change!

Parmenides: Change is not real; it's an illusion!

A→B means that A is disappearing into nothing, and B is appearing out of nothing

A→ nothing

nothing→B

If pure substances are blobs with no internal structure, there is no mechanism to explain transformations

"Something" cannot turn into "Nothing", and "Nothing" cannot turn into "Something"
Therefore, change cannot be real!

- Solid reasoning. Pretty bold. But kind of unacceptable.
- But these well-reasoned conclusions had to be proven wrong. That led to the idea of atoms.

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Democritus and the "atomists":

- Parmenides is correct in stating that parts of the universe can't just disappear, or appear out of nothing.
- But we <u>can</u> have changes and transformations, if things are made up of tiny, invisible particles that rearrange and recombine.
- **Those fundamental, permanent, indivisible particles (atoms) themselves don't change (just like Parmenides concluded),** <u>but</u> their rearrangements and re-combinations are the cause for observed change.
- Different atoms can come together to form a new substance.
- Each substance has its characteristic particle, and the attributes of these particles determine the attributes of the substance they make up.

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Atomists:
Oh, by the way, there has to be a void (empty space) for atoms to move around and do their rearranging.

Paraphrasing Democritus:
"Reality is nothing but the atoms moving around in the void"

The first mechanical view of universe
-- originally expressed much more poetically (literally in poem form)

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So, the ancient (460-400 BC) Greek concept of atom was not simply an intuitive guess or idle speculation

It was the inescapable result of impressively disciplined thinking by a few generations of philosophers.

<u>They had to</u> invent the idea of atoms to explain the existence of change

Usually identified with the modern idea of atom

But it's deeper than that: it's about **<u>particles of "stuff"</u>**

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Very famous physicist Richard Feynman (1918-1988)

From Feynman Lectures:

"If, in some cataclysm, all of scientific knowledge were to be destroyed, and only one sentence passed on to the next generations of creatures, <u>what statement would contain the most information in the fewest words?</u>

I believe it is the *atomic hypothesis* (or the atomic *fact*, or whatever you wish to call it) that *all things are made of atoms—little particles that move around in perpetual motion, attracting each other when they are a little distance apart, but repelling upon being squeezed into one another*. In that one sentence, you will see, there is an *enormous* amount of information about the world, **<u>if just a little imagination and thinking are applied</u>**. "

http://www.feynmanlectures.caltech.edu/I_01.html

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Universe is made up of ~5% ordinary matter, ~27% Dark Matter and ~68% Dark Energy
We don't know what the latter two really are, but …
Whether through intuition, or by thinking along the lines of ancient Greek philosophers:

Dark Matter density varies across the universe, and …
Physicists are looking for Dark Matter "particles", even though they know nothing about Dark Matter

"Dark Energy" appears to be a constant, and …
They <u>aren't</u> looking for particles of Dark Energy (the mysterious "thing" that makes the universe expand faster and faster)

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Now, back to chemistry and on to the (early) modern era …

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Robert Boyle was the first "chemist" (1661)

- Used the "scientific method"
- Performed quantitative experiments on the pressure and volume of gases
- Developed the first experimental definition of an element:

A substance is an element unless it can be broken down into two or more simpler substances

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Three Important Laws:

<u>Law of conservation of mass</u> (Lavoisier):

Mass is neither created nor destroyed in a chemical reaction.

<u>Law of definite proportion</u> (Proust):

A given compound always contains exactly the same proportion of elements by mass.

<u>Law of multiple proportions</u> (Dalton):

When two elements form a series of compounds, the ratios of the masses of the second element that combine with a fixed amount of the first element will be ratios of small whole numbers.

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<u>Law of conservation of mass</u> (Lavoisier):

Mass is neither created nor destroyed in a chemical reaction.

A plant grows from a tiny seed up to a huge tree.
Does this violate the Law of Conservation of Mass? Explain.

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<u>**Law of definite proportion**</u> (Proust):

A given compound always contains exactly the same proportion of elements by mass.

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Question:

According to the Law of Definite Proportions:

a) If the same two elements form two different compounds, they do so in the same ratio
b) It is not possible for the same two elements to form more than one compound
c) The ratio of the masses of the elements in a compound is always the same
d) The total mass after a chemical change is the same as before the change.

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A sample of chemical X is found to contain 5.0 grams of oxygen, 10.0 grams of carbon, and 20.0 grams of nitrogen. The law of definite proportion would predict that a 70 gram sample of chemical X should contain how many grams of carbon?

a) 5.0 grams
b) 7.0 grams
c) 10. grams
d) 15 grams
e) 20 grams

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<u>**Law of multiple proportions**</u> (Dalton):

When **two** elements form a **series of compounds**:

For a <u>fixed amount of one element</u>,
<u>the amounts of the **other**</u> element in those compounds
will be in **ratios of "small" whole numbers**.

1:2, 1:3, 2:3, 5:2 etc.

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<u>**Law of multiple proportions**</u> (Dalton):

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will be in **ratios of "small"* whole numbers**.

1:2, 1:3, 2:3, 5:2 etc.

Compounds known at that time had at most a few atoms of any element in their formulas. Quantitative methods were not good enough to deal with larger numbers anyway. But the principle is still valid, even when the numbers are large, and our methods are accurate and precise enough to tell apart 133:100 from 4:3

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Example

- Carbon and oxygen make two compounds. Their molecules are composed of:

1 carbon with 1 oxygen for "Carbon monoxide"
1 carbon with 2 oxygens for "Carbon dioxide"

Without knowing anything about the mass of carbon and oxygen atoms, we can say that *"carbon dioxide" has twice as much oxygen for a given amount of carbon as "carbon monoxide"*.

How about the amount of carbon per oxygen?

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Example (cont.)

1 carbon with 1 oxygen for "Carbon monoxide"

1 carbon with 2 oxygens for "Carbon dioxide"

Now we are comparing the amount of carbon per oxygen in the two compounds. It's a bit trickier.

We can make it easier to think about by making the number of oxygens the same. So we take two carbon monoxide molecules, corresponding to 2 oxygens.

2 carbons per 2 oxygens for "Carbon monoxide"

1 carbon with 2 oxygens for "Carbon dioxide"

Carbon dioxide has half as much carbon for a given amount of oxygen as carbon monoxide

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A compound of chromium (symbol: Cr) and oxygen (symbol: O) has the formula CrO_2. Using the Law of Multiple Proportions, guess the formulas for the other compounds of Cr and O, using the data provided in the table below.

Compound	Mass of Cr (g)	Mass of O (g)
CrO_2	52	32
Compound A	39	12
Compound B	26	24

In CrO_2, there is 32 g of O per 52 g of Cr → $\frac{32\text{ g O}}{52\text{ g Cr}} = \frac{0.615\text{ g O}}{1\text{ g Cr}}$

In A, there is 12 g of O per 39 g of Cr → $\frac{12\text{ g O}}{39\text{ g Cr}} = \frac{0.308\text{ g O}}{1\text{ g Cr}}$

→ In A, there is 0.308 g of O per gram of Cr
Compared to 0.615 g O per gram of Cr in CrO_2 → $\frac{0.308\text{ g}}{0.615\text{ g}} \approx \frac{1}{2}$

→ In A, there is half as much O **per gram of Cr** as in CrO_2

So ….

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Is it ok to also say "In A, there is half as much O per each atom of Cr"?

Yes. For a given amount of Cr, A has half as much O as CrO_2.

The sample size doesn't matter.

The amount of Cr we are considering doesn't change the **relative** amounts of oxygen in the two compounds.

If one compound is half as rich in oxygen as the other compound, that fact won't change with sample size.

For any given amount of Cr, compound A has half as much O as CrO_2.

After all, we could consider the amount of O for a mass of Cr equal to the mass of 1 Cr atom. It's still "for a given amount of Cr".

So the **relative** amounts applies equally to the number of atoms contained in the two compounds.

1 O atom instead of 2 O atoms per Cr atom

So the formula for A can be: **CrO**

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Now guess the formula for Compound B using the same table and the same reasoning

Compound	Mass of Cr (grams)	Mass of O (grams)
CrO_2	52	32
Compound A	39	12
Compound B	26	24

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Which of the following pairs of compounds can be used to illustrate the law of multiple proportions?

A) NH_4 and NH_4Cl
B) ZnO_2 and $ZnCl_2$
C) H_2O and HCl
D) NO and NO_2

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- **Law of Definite Proportions** did give a hint that matter was acting "as if" it were made of discrete particles (atoms)

  -- at least to those who liked the idea of atoms.

  -- Otherwise why else would two elements come together in a definite ratio of masses?

  -- Think Lego pieces that have different masses

- On the other hand, it was reasonable to think that there might be some other explanation.

- After all, the mass ratios were nothing special. The numbers didn't look like "counts" (or rather, ratios of counts of particles.)

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- **Law of Multiple Proportions** all of a sudden produced these simple ratios of **small integers** relating different compounds of the same two elements.
- That was a much stronger hint that different elements were combining using discrete particles (atoms)
- One had to be irrationally opposed to the idea of atoms to ignore the hint.
- In fact it was stated by Dalton as **one of the predictions by his atomic theory**.

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Dalton's Atomic Theory (1808)

- *Each element is made up of tiny particles called atoms.*
- *The atoms of a given element are identical;* the atoms of different elements are different in some fundamental way or ways. Not quite, we now know there are "isotopes"
- *Chemical compounds are formed when atoms of different elements combine with each other.* A given compound always has the same relative numbers and types of atoms.
- *Chemical reactions involve reorganization of the atoms*—changes in the way they are bound together.
- The atoms themselves are not changed in a chemical reaction.

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Gay-Lussac's Observations (1808)

- Measured (under same conditions of T and P) the volumes of gases that reacted with each other.
- Ratio of volumes of gases used in a reaction are simple integers.

For example:

1 liter of oxygen reacts with 2 liters of hydrogen to form water.

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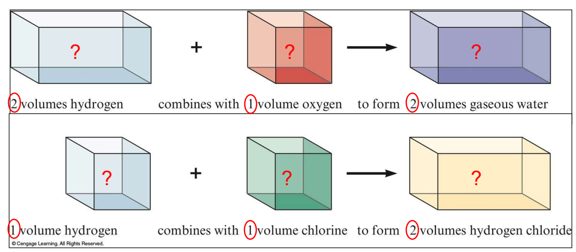
Gay—Lussac's Results



Ratio of volumes of gases used in a reaction are ratios of simple integers.

- Has nothing to say about molecules or atoms
- It's just a law with no microscopic insight

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Avogadro's ~~Hypothesis~~ Theory (1811)

Gay-Lussac's results got Avogadro thinking:

There is no reason for those simple ratios of reactant volumes used, unless …

At the same T and P, equal volumes of different gases contain the same number of particles.

In other words:

Volume of a gas is determined by the number, not the size or anything else of the gas particles

Explained Gay-Lussac's observations

Not accepted until 1860!

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Avogadro's theory was our first true connection to the atomic world

Chemists did not suspect that gases like hydrogen, oxygen, chlorine, etc. were made of diatomic molecules

- If you don't know how heavy each oxygen atom and each hydrogen atom is,
- And you see that 1 g of hydrogen combines with 8 grams of oxygen to form water,
- Assuming the simplest formula for water (HO), you conclude that oxygen atoms are 8 times as heavy as hydrogen atoms

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Likewise, you would also think

H + Cl → HCl

Hydrogen + Chlorine → Hydrogen Chloride

But when you get

Hydrogen + Chlorine → Hydrogen Chloride

You figure hydrogen and chlorine must be diatomic:

$H_2 + Cl_2 \rightarrow 2\ HCl$

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And when you find

Water → Hydrogen + Oxygen

instead of the following expected from HO molecules

Water → Hydrogen + Oxygen

You figure water molecules must be H_2O instead of HO

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So,

Avogadro's "hypothesis", for the first time, made a connection between macroscopic (human scale) measurements and the number of particles in a sample.

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Until early 20th century (1900s), atoms were still a "working assumption", as far as physicists were concerned.

- Chemists were sure atoms actually existed
- Physicists insisted on more direct evidence
  - All they could say was that "matter behaves as if it is made of atoms"

Either way, nobody knew what they were made of, or their internal structure, if any.

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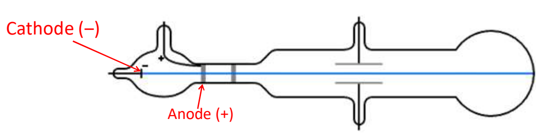
J. J. Thomson (Around 1900; i.e. beginning of 20th century)



J. J. Thompson studied the "cathode rays" and determined:

- They were made of "matter"
  - not "immaterial" like light
- They were made of negatively charged particles
  - Later called "**electrons**"

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- J. J. Thomson also calculated the **mass-to-charge ratio** (m/e) of the electron from the amount of deflection of the rays by electric and magnetic fields
- Those negative particles had to be coming from atoms, if all matter were made of atoms
- Being negatively charged, they could not be the entire atom
  - **Atoms must be electrically neutral** because materials are normally neutral.
- If the atom had negatively charged particles, it must have had a **positively charged part too**.

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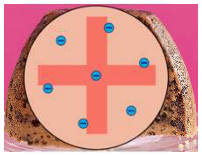
J. J. Thomson came up with the "plum pudding" model:

(–) electrons dispersed in a ball of (+) charge



Pro:
- Only has what he knew experimentally
- Assumed the simplest possible distribution for the positive charge, about which he knew that he knew nothing

Con:
Had little chance of being right and violated fairly basic physics because:
- (–) charges touching the (+) "blob" would lead to an energy of negative infinity

Plum pudding model turned out to be wrong.
Rutherford's "gold foil" experiment gave the first hint.

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But before we come to Rutherford …

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Robert Millikan (1909)

- Performed experiments involving charged oil drops.
  - Made microscopic oil droplets; determined their size
  - Calculated their weight from their size and density
  - Charged microscopic oil droplets with random charges
  - The electrical force needed to balance the droplets against gravity let him calculate the charges on droplets
  - Charges turned out to be multiples of a certain small number
  - That had to be the smallest possible charge: i.e. the charge of an electron
- Charge on a single electron = 1.6×10^{-19} Coulombs
- Calculated the mass of the electron (9.11×10^{-31} kg) from the charge-to-mass ratio measured by Thomson.

Don't memorize numbers like that just because they are on the slides

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Henri Becquerel (1896)

- Discovered radioactivity by observing the spontaneous emission of radiation by uranium.
- Using a magnetic field he found that there were three kinds of radiation:
  - negatively charged
  - positively charged
  - neutral

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Ernest Rutherford (1911)

Classified three types of radioactive emission based on their penetrating power (instead of charge, as Becquerel did)

- Alpha (α)
  -- a particle with a +2 charge (least penetrating)
  -- much heavier than other types of radiation
- Beta (β)
  a high speed electron (–1 charge)
- Gamma (ϒ)
  high energy light (neutral) (most penetrating)

There are other kinds of radiation too

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Ernest Rutherford (1911)

- Showed "plum pudding" model to be wrong
- Found that:
  — The atom has a very dense center with (+) charge: **nucleus**
  — Electrons travel around the nucleus, at large distances compared with the size of the nucleus
- And here is how he did that:

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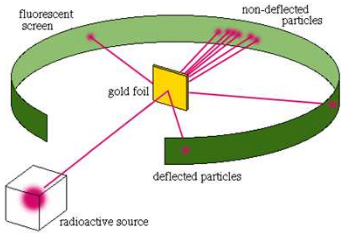
Rutherford's gold foil experiment (~1911)



Alpha particles:
- much heavier than electrons
- positively charged

Most alpha particles went through non-deflected,
but some bounced off in all directions

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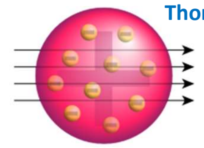
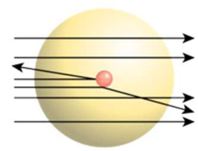
Thompson



Plum-pudding model:
Mass spread uniformly
No (or very small) deflections
Does not match the experiment!

Rutherford



Nuclear model:
Mass is concentrated at center
Very light electrons fill the volume
Mostly no deflection
Occasional big deflection
Matches the gold-foil experiment

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Neutrons

- hypothesized by Rutherford in 1920
- discovered experimentally in 1932 by James Chadwick
- have no electric charge (neutral)
- mass and size similar to proton; very slightly heavier

- Act as glue holding the positively charged protons together
- There is a special nuclear force, the "strong nuclear force" between neutrons and protons that can overcome the electrostatic repulsion between positively charged protons.
- Without neutrons the repulsion between the (+) charged protons would break the nucleus apart

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So, the atom contains:

Electrons:
- outside the nucleus
- negatively charged
- much lighter than protons and neutrons
- spread over a much larger volume than the nucleus

Protons:
- in the nucleus
- positive charge equal in magnitude to the electron's negative charge.
- tiny but heavy (dense)

Neutrons:
- in the nucleus
- no charge
- density similar to proton; very slightly heavier

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The nucleus is:
- **Small** compared with the overall size of the atom.
- Extremely **dense**; accounts for almost all of the atom's mass

Each proton or neutron is ~1850 times heavier than an electron

-- we normally ignore the mass of electrons
-- proton and neutron masses are slightly different
-- and their mass changes slightly depending on nucleus they are in

Very strong nuclear forces → Very large nuclear energies → Large variations in very large nuclear energies → $E = mc^2$ noticeable variation in particle mass

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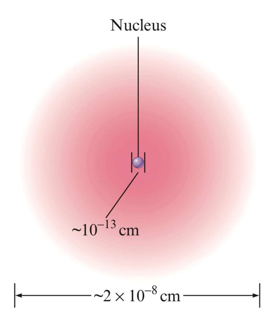
Nuclear Atom Viewed in Cross Section



Actually this picture exaggerates the size of the nucleus.

The nucleus would be invisible in this picture of an atom!

$$\frac{2 \times 10^{-8}\ \text{cm}}{10^{-13}\ \text{cm}} = 200{,}000$$

The atom is 200,000 times larger than the nucleus.
200,000 pixel-wide screen needed for the nucleus to occupy 1 pixel.

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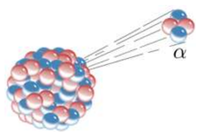
Now that we learned about nuclei and particles …

Let's make sure we know what an alpha particle is

- **Nucleus of a Helium atom** ; has **+2 charge**
- Heaviest of the common radiation types
- Has **2 protons** and **2 neutrons**

-- **Atomic number of the atom left behind changes** (because it has 2 less protons)



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Isotopes

- Atoms with the same number of protons ⇨ same element
  - but different numbers of neutrons.
- Isotopes show virtually identical chemical properties because chemistry is done by the electrons.
  (the isotopes of the lightest elements like H or Li have measurable chemical differences, but the reasons for that is beyond the scope of the course)
- In nature most elements are mixtures of isotopes. The relative abundances of isotopes on Earth are fairly well fixed

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Isotopes are identified by:

Atomic Number (Z) = number of protons (p)

Mass Number (A) = number of protons (p) + number of neutrons (n)

Z = p   A = p + n

Mass number → $^{A}_{Z}X$ ← Element symbol

Atomic number → Z

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Example -- Isotopes of Magnesium:

		Natural abundance
$^{24}_{12}\mathrm{Mg}$	12 protons, 12 neutrons	79%
$^{25}_{12}\mathrm{Mg}$	12 protons, 13 neutrons	10%
$^{26}_{12}\mathrm{Mg}$	12 protons, 14 neutrons	11%

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Isotope symbol: $^{24}_{12}\mathrm{Mg}$

Showing the Z value is redundant, since its value is fixed for the isotopes of a given element.
If it's a magnesium isotope, Z is always 12

Isotope name: {element name}– {mass number}

$^{24}_{12}\mathrm{Mg}$ is magnesium–24

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Atomic masses ("atomic weight")

"Atomic mass" of an element is an average quantity

- It is a weighted average of isotopic masses
- because elements almost always have more than one isotope with significant abundance

But before we deal with averaging isotopic masses,
let's understand the "atomic mass unit"
(used to be denoted by ***a.m.u***, nowadays just ***u***)

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Atomic Mass Unit

Atomic mass unit is meant to approximate the atomic mass number.

But in general it can't be exactly equal to it.

Things would be simple if:

- protons and neutrons had the same exact mass
- and their masses did not depend on the kind of nucleus they are in
- Then, knowing the proton and neutron mass, we could calculate the mass of any isotope
  - ^{13}C would have a mass of exactly 13*u*
  - ^{35}Cl would have a mass of exactly 35*u*

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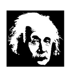
Alas,

- protons and neutrons have slightly different masses,
- and their exact masses depend on the kind of nucleus they are in

  nuclear energies involved in binding these particles are large enough to show up as measurable mass, because $E=mc^2$



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So, the mass of an atom, measured in atomic mass units *(u)*,

- instead of simply corresponding to the number of protons and neutrons in a nucleus (i.e. "mass number"),
- needs to be "calibrated" on a particular nucleus
  - ➢ ^{12}C was chosen
- It is almost equal to the mass number, but not quite, for all elements and isotopes, except the calibration isotope ^{12}C, whose mass is defined to be 12*u*

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- So,
- ^{12}C is the standard for atomic mass, with a mass of exactly 12 atomic mass units (u).
- The masses of all other atoms are measured relative to this standard.

- So, if ^{48}**Ti** is measured to be 3.9957 times heavier than ^{12}C,

➢ then its atomic mass is 3.9957x12 = ***47.948u***

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Atomic Mass

Elements occur in nature as mixtures of isotopes

Atomic mass of an element

(as opposed to an isotope)

is a weighted average of isotopic masses

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Atomic Mass Unit

Atomic Mass

For example, for Carbon:

% Abundance	Isotope
98.89%	^{12}C
1.11%	^{13}C

Simply "atomic mass" for an element

Average Atomic Mass for Carbon:

98.89% of 12 u + 1.11% of 13.0034 u

Exactly equal to mass number (12)

not exactly 13

(0.9889)(12 u) + (0.0111)(13.0034 u) = **12.01 u**

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Atomic Mass

Even though natural carbon does not contain even a single atom with mass 12.01, for stoichiometric purposes, we can consider carbon to be composed of only one type of atom with a mass of 12.01 (like our average jellybean)

This enables us to count atoms of natural carbon by weighing a sample of carbon.

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Atomic Mass

Practice

An element consists of 62.60% of an isotope with mass 186.956 u and 37.40% of an isotope with mass 184.953 u. Calculate the average atomic mass and identify the element.

Average Atomic Mass =
(0.6260)(186.956 u) + (0.3740)(184.953 u) = 186.2 u

Rhenium (Re)

We can identify an element by its average atomic mass (the "atomic mass" reported in the periodic table)

It's not the same as identifying the element from the isotopic mass (which we can't)

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Atomic Mass

Practice

Indium has atomic number 49 and atomic mass 114.8 u. Naturally occurring indium contains a mixture of indium-112 and indium-115, respectively. What is the approximate ratio of the indium-112 abundance to that of the indium-115?

Remember: **Isotopic mass numbers** can be used to *approximate* **isotopic masses**.

a) 7:93
b) 25:75
c) 50:50
d) 75:25
e) 93:7

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Atomic Mass

Concept question

If a sample of zinc (Zn) and a sample of aluminum (Al) contain the same number of atoms, which of the following claims is true?

Zn: 65.38 a.m.u.
Al: 26.98 a.m.u.

✓a) The mass of Zn sample is more than twice that of the Al sample.
b) The mass of Zn sample is more than that of the Al sample, but it is less than twice as much.
c) The mass of Al sample is more than twice the mass of the Zn sample.
d) The mass of Al sample is more than that of the Zn sample, but it is less than twice as much.
e) The two samples have equal masses.

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The concept of "mole"

In chemistry it's the count of atoms that matter.

2 hydrogen atoms combined with 1 oxygen atom form a water molecule.

We can directly measure masses, but chemistry doesn't care about masses. It cares about individual atoms and their counts.

We need a concept that keeps track of the number of atoms (or the entities they form, like molecules):

1 mole = $6.02214076 \times 10^{23}$ entities

Avogadro's Number
(defined to be exact in 2019)

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Avogadro's number is now defined exactly as $6.02214076 \times 10^{23}$, but is usually used with 3 or 4 significant figures:

6.022×10^{23} entities/mole

↑

- "elementary entities", or "characteristic units"
- Not just atoms or molecules
- Can be atoms, ions, molecules, formula units, electrons, protons, neutrons, photons, etc.

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The unit "mole" represents a count.
Just like "dozen" represents a count.

So, in principle, we can use "mole" to count things other than atoms or molecules:

1 mole of marbles, that is 6.022×10^{23} of them, can cover the surface of the Earth to a depth of 50 miles

If you had 1 mole of dollars, you could spend 1 billion dollars every second and you wouldn't run out of money for 19 million years

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But why 6.022×10^{23} particles per mole?

Why not a simpler, smaller number?

We need it to be of the order of 10^{23} because atoms are extremely light, and it takes that many of them to have the kind of masses we normally deal with (i.e. human-scale quantities)

Why 6.022…, and not 1?

It allows us to calculate the mass of 1 mole (molar mass) directly from atomic masses. We use the same numbers but use grams instead of a.m.u!

69

1 carbon-12 atom has a mass of **12 a.m.u.**

6.022×10^{23} carbon-12 atoms (i.e. 1 mole of ^{12}C) have a mass of **12 grams**

The atomic mass of iron (Fe) is **55.845 a.m.u**

(the average mass of iron atoms is 55.845 a.m.u.)

6.022×10^{23} iron atoms (i.e. 1 mole of iron) have a mass of **55.845 grams**

70

The symbol for "mole" is "mol"

- Not much shorter than "mole"
- But it allows us to use the singular form just as with the symbols for meter or gram.

1.5 mole or 1.5 moles?
0.75 mole or moles?

We just say:
1.5 mol
0.75 mol

71

Practice

Calculate the number of iron atoms in a 4.48 mol sample of iron.

72

Practice

Which of the following is closest to the average mass of one atom of copper?

a) 63.55 g
b) 52.00 g
c) 58.93 g
d) 65.38 g
e) 1.055×10^{-22} g

73

Practice

Which of the following 100.0-g samples contains the greatest number of atoms?

a) Magnesium b) Zinc c) Silver d) Calcium
e) Same for all

74

How many C atoms in 1 mole of Carbon?

How many O_2 molecules in 1 mole of Oxygen gas?

How many O atoms in 1 mole of Oxygen gas?

How many formula units in 1 mole of $CaCl_2$?

How many Cl atoms in 1 mole of $CaCl_2$?

What is the mass of 1 mole of Carbon?

What is the mass of 1 mole of Oxygen gas?

What is the mass of 1 mole of $CaCl_2$?

75

Practice

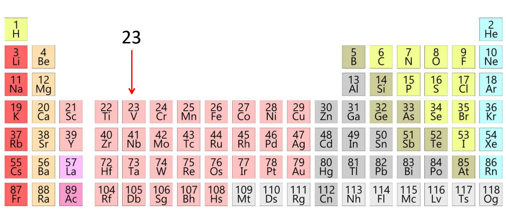
Calculate the number of copper atoms in a 63.55 g sample of copper.

76

A certain isotope X contains 23 protons and 28 neutrons.

- What is the mass number of this isotope?
- Identify the element.

- The identity of an atom is determined by its number of **protons**
- It’s Vanadium (V) because of the number of protons (23), **not** the mass number (51)



77

The element rhenium (Re) exists as 2 stable isotopes and 18 unstable isotopes. The nucleus of rhenium-185 contains:

a) 75 protons and 75 neutrons
b) 75 protons and 130 neutrons
c) 130 protons and 75 neutrons
d) 75 protons and 110 neutrons

78

Which of the following statements are true?

I. The number of protons is the same for all neutral atoms of an element.

II. The number of electrons is the same for all neutral atoms of an element.

III. The number of neutrons is the same for all neutral atoms of an element.

a) I, II, and III are true
b) Only I and II are true
c) Only II and III are true
d) Only I and III are true
e) I, II, and III are false.

79

Ions

Cations

- Formed when an atom loses one or more electrons

$Li \rightarrow Li^+ + e^-$

$Ba \rightarrow Ba^{2+} + 2e^-$

Anions

- Formed when an atom gains one or more electrons

$F + e^- \rightarrow F^-$

$N + 3e^- \rightarrow N^{3-}$

Note that we put the size of the charge before the sign:
2+ instead of +2
3– instead of –3

80

Ions

Unless the atom is undergoing a nuclear reaction (which is not chemistry), ions are formed only by **gaining or losing electrons**

81

Ions are denoted by a superscript on the right side of the entity, indicating the charge.

Examples:

Cl^- Ca^{2+} S^{2-}

82

How is an ion formed in a chemical process, starting with a neutral atom?

a) By adding or removing protons
b) By adding or removing electrons
c) By adding or removing neutrons
d) All of these are true
e) Two of these are true.

83

A certain isotope X^+ contains 54 electrons and 78 neutrons.

What is the mass number of this isotope?

84

Which of the following statements regarding Dalton's atomic theory are still considered true?

I. Elements are made of tiny particles called atoms.
II. All atoms of a given element are identical.
III. A given compound always has the same relative numbers and types of atoms.
IV. Atoms are indestructible.

By the way:
This is the part where the "**Law** of Definite Proportions" is **explained** (if we read his whole theory), in Dalton's Atomic **Theory**.

85

The Periodic Table

Periods

– **horizontal** rows of elements
– properties change in a similar way in each period
– with each new period, the trend repeats (or more like "rhymes")

Groups or Families

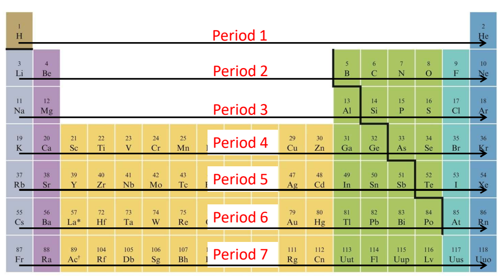
– elements in the same **vertical** columns
– have similar chemical properties

Most elements are metals

Nonmetals are huddled towards the top-right corner

86

Periods are horizontal



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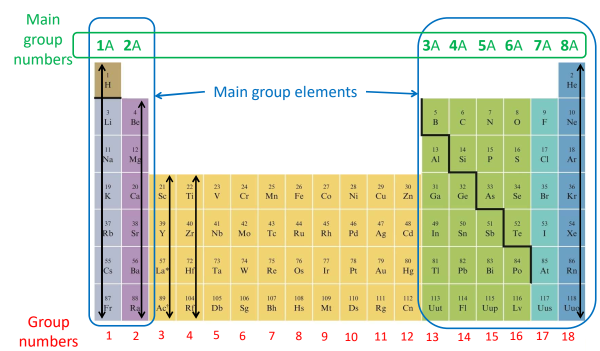
Groups are vertical



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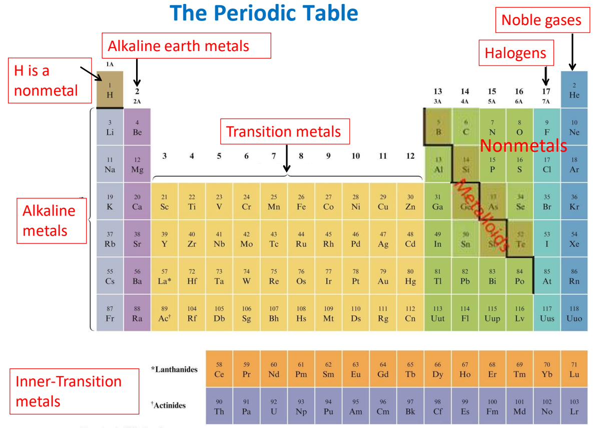
The Periodic Table



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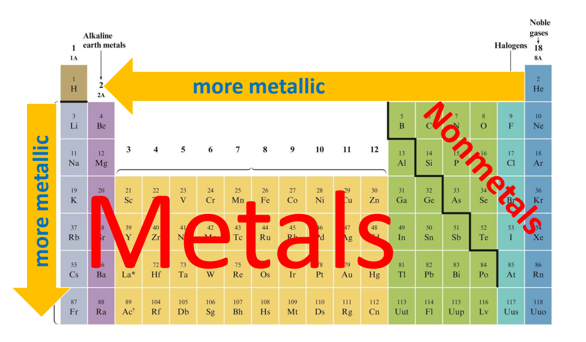

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Groups or "Families" of elements and their ions

- Metals form cations by losing electrons
- Main group metals lose as many electrons as their main group number

  cation charge = main group #

- Nonmetals form anions by gaining electrons
- Nonmetal anion charge size is the difference between main group number and 8 (main group number of noble gases at the very right)

  anion charge = {main group #} - 8

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