Chapter 4
Atoms and Elements

Based on slides provided with Introductory Chemistry, Fifth Edition
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Atoms

Ancient Greek philosophers had to invent the idea of atoms

The only way to explain change:

The rearrangement of unchanging pieces, with space between them

- Usually identified with the modern idea of atom
- But it's deeper than that: it's about particles of "stuff"

The idea that matter has to exist as particles in order to change

- A structureless "blob" cannot change, not even its shape.
- Physicists don't even attempt to imagine matter without referring to particles
- They are looking for Dark Matter "particles" even though they know nothing about Dark Matter (most of the universe)

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.

Until early 20th century (1900s), atoms were still only a "working assumption" for many scientists.

Physicists would say "matter behaves as if it is made of atoms". They weren't sure they existed.

But chemists did regard the atoms as real already.

Nevertheless, nobody knew what they were made of, or their structure.

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

- Postulated the existence of negatively charged particles, that we now call **electrons**, using cathode-ray tubes.
  - The mysterious "rays" he observed were bending in response to a magnetic field. Light doesn't do that. It had to be composed of (negatively) charged particles.
- Those negative particles had to be coming from atoms, if all matter is made of atoms.
- He determined the **charge-to-mass ratio of an electron**. He didn't have the right experiment to determine the charge and the mass separately.

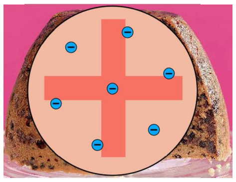
J.J. Thomson also reasoned:

- If the atom had negatively charged particles, it must have had a positively charged part too.
- Atoms must be electrically neutral because materials are normally neutral.

J. J. Thomson (cont.)

- 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.

But before we come to Rutherford …

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

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

Ernest Rutherford (1911)

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

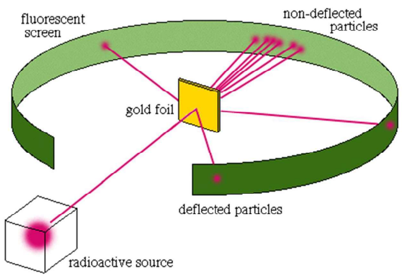
- Alpha (α)
  -- a particle with a +2 charge (least penetrating)
  -- much heavier than other types of radiation
- Beta (β)
  a high speed electron (–1 charge)
- Gamma (ϒ)
  high energy light (neutral) (most penetrating)

There are other kinds of radiation too

Ernest Rutherford (1911)

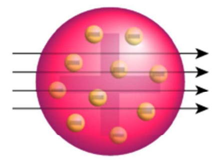
- Showed "plum pudding" model to be wrong
- Found that:
  — The atom has a very dense center with (+) charge: **nucleus**
  — Electrons travel around the nucleus, at large distances compared with the size of the nucleus
- And here is how he did that:

Rutherford's gold foil experiment (~1911)

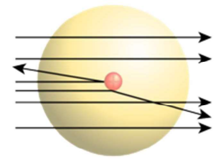


Alpha particles:
- much heavier than electrons
- positively charged

Most alpha particles went through with little or no deflection, **<u>but some bounced off in random directions</u>**



<u>Plum-pudding model:</u> Thompson
- Mass spread uniformly
- Predicts little or no deflections of alpha particles
- Does not match the experiment!



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

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

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

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

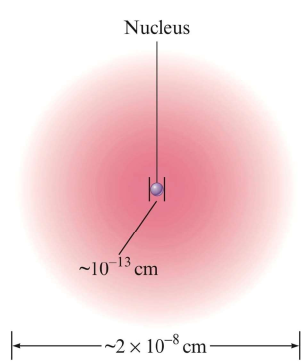
-- we normally ignore the mass of electrons

-- 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

Nuclear Atom Viewed in Cross Section



Actually this picture exaggerates the size of the nucleus.

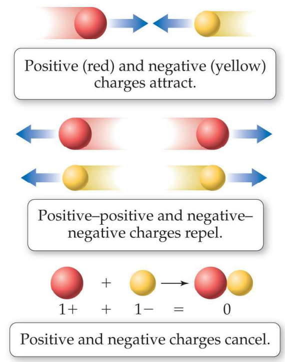
The nucleus would be invisible in this picture of an atom!

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

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

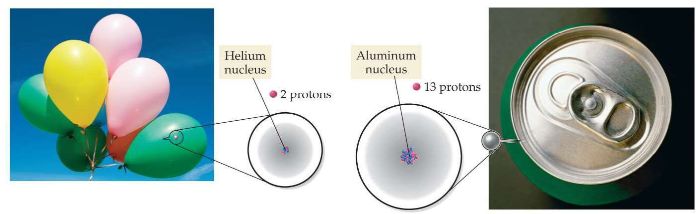
Electrical Charge Is a Property of Protons and Electrons

- Electrical charge is a fundamental property of protons and electrons.
- Positive and negative electrical charges attract each other.
- Positive–positive and negative–negative charges repel each other.
- Positive and negative charges cancel each other so that a proton and an electron, when paired, are charge-neutral.



How Atoms of the Elements Differ from One Another

- Elements are defined by their numbers of protons.
- It is the number of protons in the nucleus of an atom that identifies the atom as a particular element.
  -- because it determines the number of electrons in a neutral atom
- If an atom had a different number of protons, it would be a different element.
- The number of protons in the nucleus of an atom is its **atomic number** and is given the symbol **Z.**



Looking for Patterns: Recurring Properties

1	2	3	4	5	6	7	8	9	10	11	12	13	14	15	16	17	18	19	20
H	He	Li	Be	B	C	N	O	F	Ne	Na	Mg	Al	Si	P	S	Cl	Ar	K	Ca

- The color of each element box here represents its general features.
- We arrange them in rows so that similar properties align in the same vertical columns. This figure is similar to Mendeleev's first periodic table.

1 H							2 He
3 Li	4 Be	5 B	6 C	7 N	8 O	9 F	10 Ne
11 Na	12 Mg	13 Al	14 Si	15 P	16 S	17 Cl	18 Ar
19 K	20 Ca						

Looking for Patterns: Alkali Metals

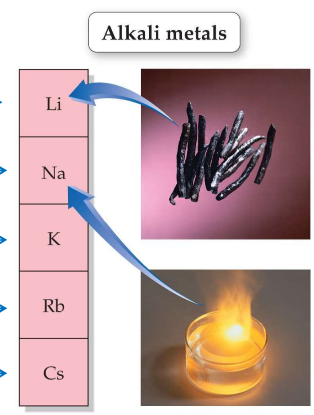
lithium

sodium

potassium

rubidium

cesium



Alkali metals are very reactive

For example, they react explosively with water

Looking for Patterns: Alkaline Earth Metals

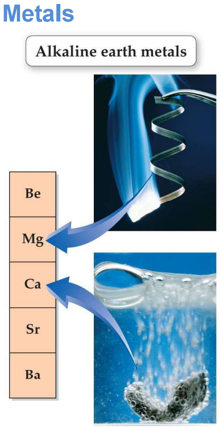
beryllium

magnesium

calcium

strontium

barium



Less reactive than alkali metals

Chemistry of alkaline earth metals share common features

Looking for Patterns: Halogens

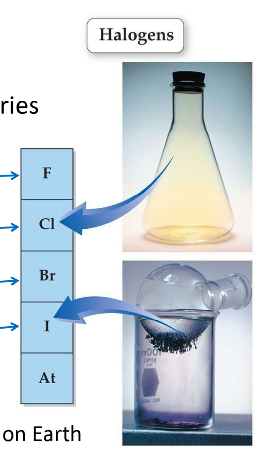
Halogens have different physical properties, but have similar chemistries

fluorine

chlorine

bromine

iodine



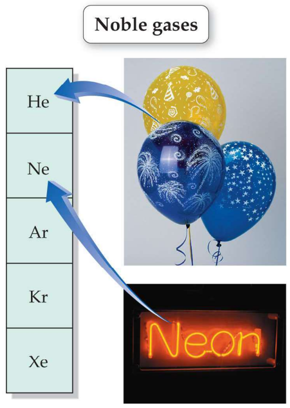
only a trace amounts of astatine on Earth at any time, produced by nuclear decay

Looking for Patterns: Noble Gases

The noble gases include helium (used in balloons), neon (used in neon signs), argon, krypton, and xenon.

Noble gases are very inert

They are not reactive
-- but heavier members, Kr and especially Xe do form a few compounds



The Periodic Table

- The idea developed around the same time (1869) by several scientists:

  When ordered by their atomic masses, the properties of the elements repeat (not exactly of course)

  – Modern periodic table orders the elements by atomic number, which is the true identifying property of an element
- Mendeleyev's version became famous, and he became known as the inventor, because he used his table to predict the existence of yet-unknown elements and their properties, extrapolating from known elements and the fact that their properties repeat periodically

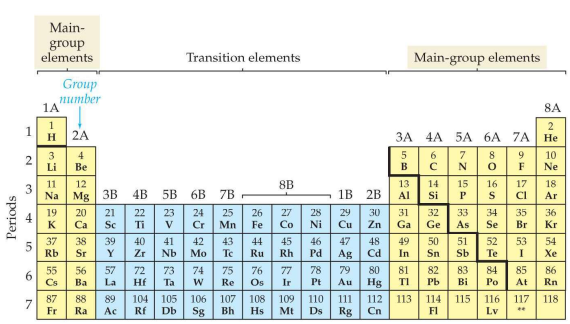
- Mendeleyev's version became famous, and he became known as the inventor, because he used his table to predict the existence of yet-unknown elements and their properties, extrapolating from known elements and the fact that their properties repeat periodically

Periodic Law Summarizes Many Observations

- Periodic law was based on observation.
- Like all scientific laws, the periodic law summarized many observations but did not give the underlying reason for the observation—only theories do that.
- For now, we accept the periodic law as it is, but in Chapter 9 we will examine a powerful theory that explains the law and gives the underlying reasons for it.

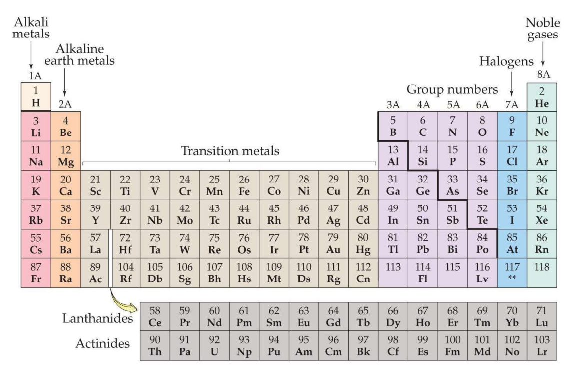
Main Group Elements and the Transition Series of Elements

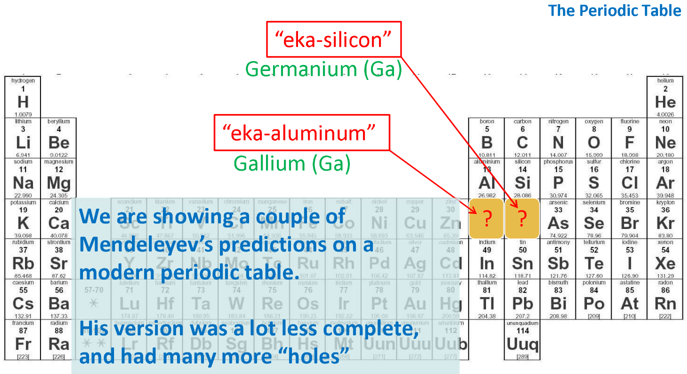
In main group elements, properties can generally be predicted based on the element's position.
In transition elements, properties tend to be less predictable based on the element's position.



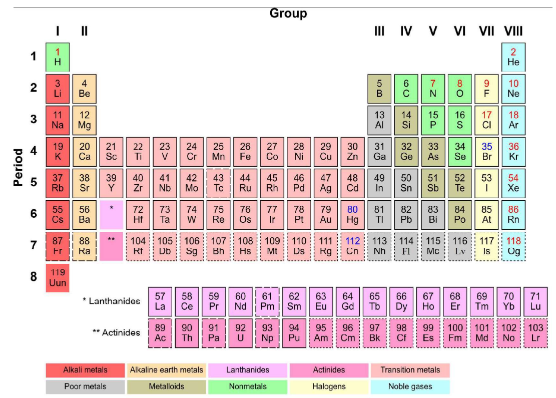
The Periodic Table with Groups 1A, 2A, 7A, 8A Highlighted

Groups: 1A, alkali metals; 2A, alkaline earth metals; 7A, halogens; 8A, noble gases





An example of a modern periodic table



The Periodic Table

Periods

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

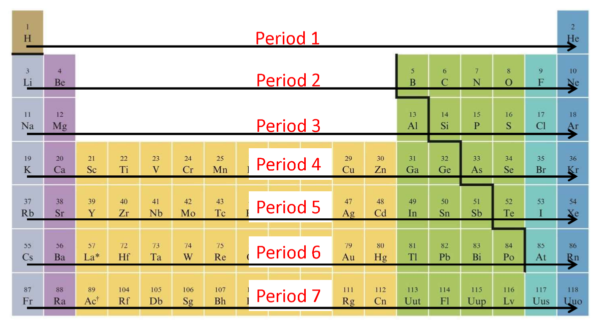
Groups or Families

- elements in the same **<u>vertical</u>** columns
- have similar chemical properties

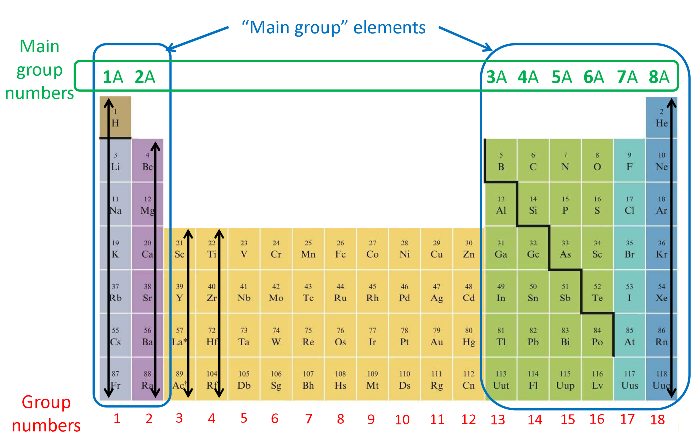
Most elements are metals

Nonmetals are huddled towards the top-right corner

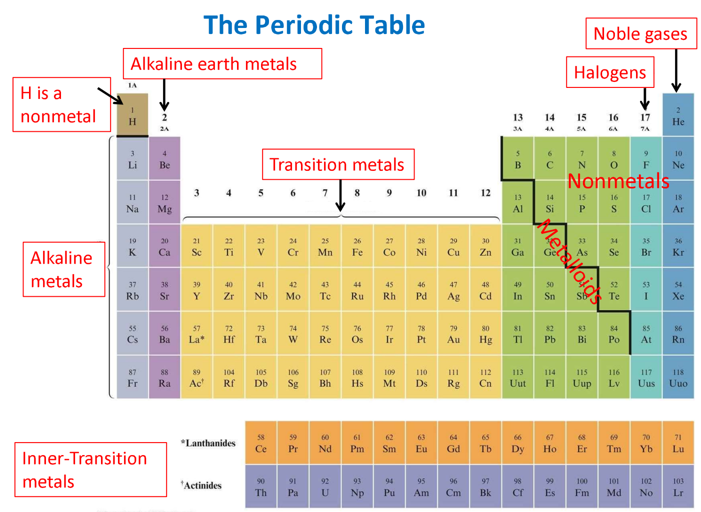
Periods are horizontal

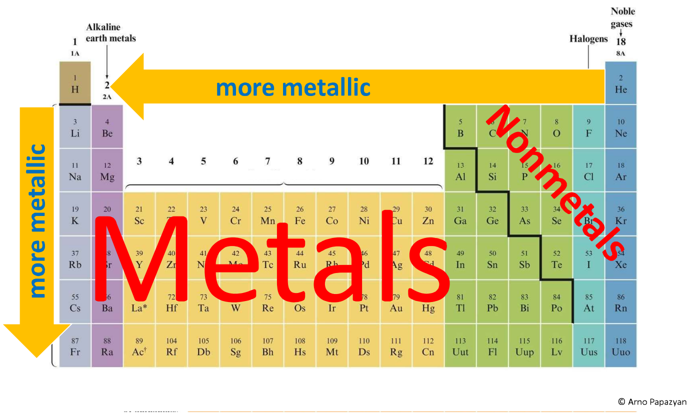


Groups are vertical



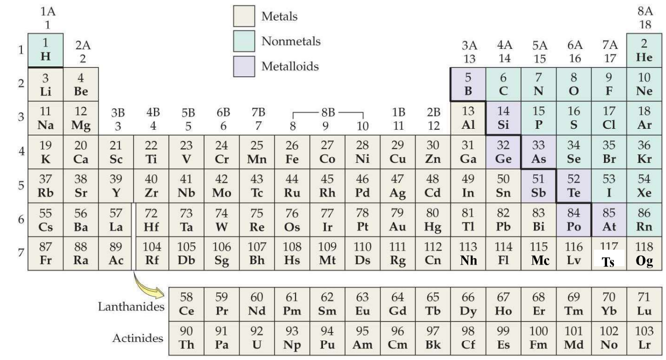
The Periodic Table





Metals, Nonmetals, Metalloids on the Periodic Table

The elements in the periodic table can be broadly classified as metals, nonmetals, and metalloids.



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Typical Properties of Metals

- **Metals** occupy the left side of the periodic table and have similar properties.
- Metals are good conductors of heat and electricity.
- Metals can be pounded into flat sheets (malleable).
- Metals can be drawn into wires (ductile).
- Metals are often shiny (lustrous).
- Metals tend to lose electrons when they undergo chemical changes.
- Good examples of metals are iron, magnesium, chromium, and sodium.

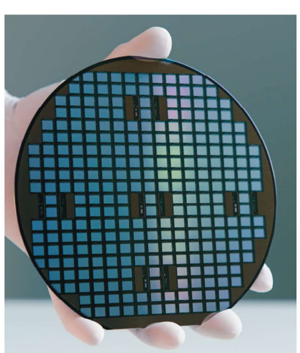
Properties of Nonmetals Vary

- **Nonmetals** occupy the upper right side of the periodic table.
- The dividing line between metals and nonmetals is the zigzag diagonal line running from boron to astatine.
- Nonmetals have more varied properties; some are solids at room temperature, while others are gases.
- As a whole, nonmetals tend to be poor conductors of heat and electricity.
- Nonmetals tend to gain electrons when they undergo chemical changes.
- Good examples of nonmetals are oxygen, nitrogen, chlorine, and iodine.

Properties of Metalloids

Metalloids lie along the zigzag diagonal line dividing metals and nonmetals. Metalloids, also called semimetals, display mixed properties.

- Metalloids are also called **semiconductors** because of their intermediate electrical conductivity, which can be changed and controlled.
- This property makes semiconductors useful in the manufacture of electronic devices that are central to computers, cell phones, and other modern gadgets.
- Silicon, arsenic, and germanium are good examples of metalloids.
- Silicon is shown here.



Atoms Lose or Gain Electrons to Form Ions

- In chemical reactions, atoms often lose or gain electrons to form charged particles called **ions.**
- Positive ions are called **cations.**
- Negative ions are called **anions.**
- The charge of an ion is shown in the upper right corner of the symbol.
- Ion charges are usually written with the magnitude of the charge first, followed by the sign of the charge.
- Examples: Mg^{2+}, O^{2-}

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)

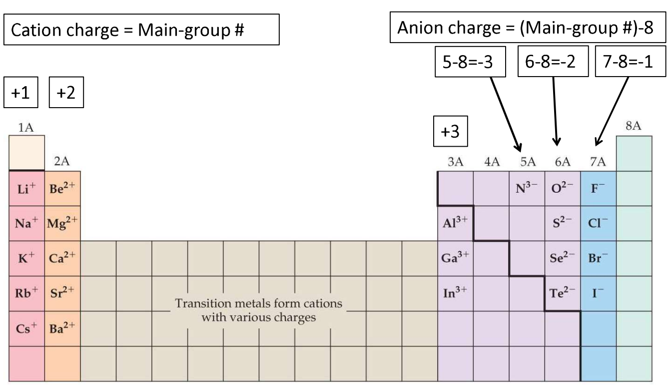
  anion charge = {main group #} - 8

Groups or "Families" of elements and their ions

Group or Family	Charge of ion	
Alkali Metals (1A)	1+	Main Group #
Alkaline Earth Metals (2A)	2+	
Aluminum *(not a group)* (3A)	3+	Main Group # / Main Group # of noble gases
Group 5 nonmetals (5A)	3–	5–8 = –3
Chalcogens (6A)	2–	6–8 = –2
Halogens (7A)	1–	7–8 = –1
Noble Gases (8A)	0	(they don't ionize)

Ions and the Periodic Table

Ions with charge predicted by the group number:



By the way:

These charges are taken on only when they are stabilized by nearby charges of the opposite sign.

-- such as in an ionic compound

-- or when that ionic compound is dissolved in a solvent with "polar" molecules, like water

Some parts are electron-rich (–), some parts are electron poor (+)

If you ignore this little point, you might think those elements always exist as ions!

A common misconception!

In chemistry, ions are made by losing or gaining electrons

Making Ions by Losing Electrons

In reactions, lithium atoms lose one electron (e^-) to form Li^+ ions.

$$\mathrm{Li} \longrightarrow \mathrm{Li}^+ + e^-$$

The charge of an ion depends on how many electrons were gained or lost and is given by the formula

$$\text{Ion charge} = \text{number of protons} - \text{number of electrons}$$
$$= \#p^+ - \#e^-$$

where p^+ stands for *proton* and e^- stands for *electron.*

For the Li^+ ion with 3 protons and 2 electrons, the charge is

Ion charge = 3 − 2 = 1+

Making Ions by Gaining Electrons

In reactions, fluorine atoms gain one electron (e^-) to form F^- ions:

$$F + e^- \longrightarrow F^-$$

The charge of an ion depends on how many electrons were gained or lost and is given by the formula

$$\text{Ion charge} = \text{number of protons} - \text{number of electrons}$$
$$= \#p^+ - \#e^-$$

where p^+ stands for *proton* and e^- stands for *electron.*

For the F^- ion with 9 protons and 10 electrons, the charge is

Ion charge = 9 – 10 = 1–

Isotopes: When the Number of Neutrons Varies

- All atoms of a given element have the same number of protons.
- They do not necessarily have the same number of neutrons.
- Atoms with the same number of protons but different numbers of neutrons are called **isotopes.**
- All elements have their own unique percent **natural abundance** of isotopes.

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

Atomic Number (**Z**)

Number of protons in the nucleus

Atomic Mass Number (**A**)

{number of protons} + {number of neutrons}

The number of neutrons in an isotope is the difference between the mass number and the atomic number:

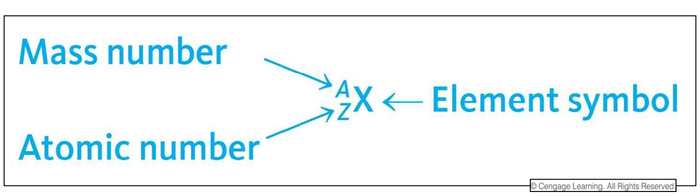
No. of neutrons = A – Z

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



Example -- Isotopes of Magnesium:

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

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

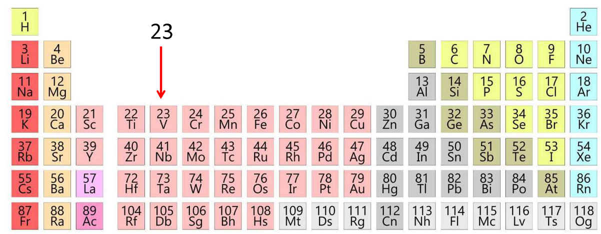
Isotope name: {element name}-{mass number}

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

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)



Practice:
Check your solution on next page

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

Practice 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

Check the periodic table Re's atomic number is 75.
Therefore it has 75 protons.
If the mass number is 185, #n = 185-75 = 110
since (mass #) = #p + #n

Practice:
Check your solution on next page

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.

Practice:

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

Practice: Complete the table

Check your solution on next page

Protons	Neutrons	Electrons	Atomic Number	Mass Number	Atomic Symbol
6	7				
		42		96	
					$^{27}_{13}Al$
			55	133	

Practice: Complete the table

Protons	Neutrons	Electrons	Atomic Number	Mass Number	Atomic Symbol
6	7	6	6	13	
42	54	42	42	96	
13	14	13	13	27	$^{27}_{13}Al$
55	78	55	55	133	

Practice:

Check your solution on next page

What is the number of electrons, protons and neutrons in the following?

$^{32}S^{2-}$

$^{40}Ca^{2+}$

$^{6}Li^{+}$

$^{27}Al^{3+}$

$^{15}N^{3-}$

Practice

What is the number of electrons, protons and neutrons in the following?

$^{32}S^{2-}$ 16p, 18e⁻, 16n

$^{40}Ca^{2+}$ 20p, 18e⁻, 20n

$^{6}Li^{+}$ 3p, 2e⁻, 3n

$^{27}Al^{3+}$ 13p, 10e⁻, 14n

$^{15}N^{3-}$ 7p, 10e⁻, 8n

Atomic mass unit
-- a mass unit appropriate for atoms

<u>Almost</u> the same value as the atomic mass number

but, not quite …

Atomic Mass Unit

Atomic mass unit is meant to approximate the atomic mass number. In general, it can't be exactly equal to it.

Things would be simple if:

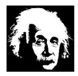
- protons and neutrons had the same exact mass
- and their masses did not depend on the kind of nucleus they are in
- Then, knowing the proton and neutron mass, we could calculate the mass of any isotope

  ^{13}C would have a mass of exactly 13 atomic mass units
  ^{35}Cl would have a mass of exactly 35 atomic mass units
  ^{37}Cl would have a mass of exactly 37 atomic mass units
- *But, alas …*

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 $E=mc^2$



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*

*The obsolete symbol "amu" you might see used in many sources including the textbook refers to an obsolete calibration based on ^{16}O, but is more descriptive. I will often use a.m.u. as a compromise acronym for "atomic mass unit"

So,

- ^{12}C is the standard for atomic mass, with a mass of exactly 12 atomic mass units (u) (that is, "a.m.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 = **47.948 a.m.u**

In other words,

Other atoms' masses are found relative to the mass of ^{12}C

If the mass of ^{35}Cl atom is measured to be 2.9141 times larger than that of ^{12}C, its mass in atomic mass units is

2.9141 x 12 = 34.9692 u

Note that this is pretty close to the atomic mass number (35) but not exactly the same.

So,

Atomic mass of an isotope (isotopic mass) in a.m.u. is not a whole number because the masses of protons and neutrons vary slightly.

The term "**atomic mass**", (for an element) when used by itself, refers to the weighted average of the isotopic masses of its isotopes, in a.m.u.

And it is also never a whole number because it is a weighted average. Even if isotopic masses were whole numbers (which they are not), atomic mass of the element would still not be a whole number.

Atomic Mass

In general, atomic mass is calculated according to the following equation:

Atomic mass =

(Fraction of isotope 1 × Mass of isotope 1) +
(Fraction of isotope 2 × Mass of isotope 2) +
(Fraction of isotope 3 × Mass of isotope 3) + …

where the fractions of each isotope are the percent natural abundances converted to their decimal values.

Atomic Mass

Refers to the entire element, as it exists as a mixture of isotopes.

For example, for Carbon:

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

Atomic mass of an element (as opposed to an isotope) is a **weighted average** of isotopic masses

Average Atomic Mass for Carbon:

98.89% of 12 a.m.u. + 1.11% of 13.0034 a.m.u.

Exactly equal to mass number (12)

not exactly 13

(0.9889)(12 a.m.u.) + (0.0111)(13.0034 a.m.u.) = $\boxed{\text{12.01 a.m.u.}}$

To think about what we just did in the previous slide …

- In calculations, percents are used in fractional form.
  -- 62.60% becomes 0.6260, 37.40% becomes 0.3740, etc
- Isotopic abundances are used in fractional form
  -- to calculate the weighted average of the isotopic masses
- Also note that fractional abundances add up to 1
  Just as the percent abundances add up to 100
  Can come in handy

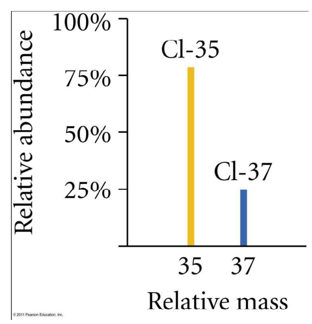
So, to calculate the atomic mass of carbon, we did:

(0.9889)(12 u) + (0.0111)(13.0034 u) = $\boxed{\text{12.01 u}}$

Mass Spectrum

17
Cl
35.45
chlorine

- A mass spectrum is a graph that gives the relative mass and relative abundance of each particle
- Relative mass of the particle is plotted in the *x*-axis
- Relative abundance of the particle is plotted in the *y*-axis



Example: Ga-69 with mass 68.9256 amu and abundance of 60.11% and Ga-71 with mass 70.9247 amu and abundance of 39.89%. Calculate the atomic mass of gallium.

Given:	Ga-69 = 60.11%, 68.9256 amu Ga-71 = 39.89%, 70.9247 amu
Find:	atomic mass, amu
Conceptual Plan:	isotope masses, isotope fractions → avg. atomic mass
Relationships:	$\text{Atomic Mass} = \sum (\text{fractional abundance of isotope})_n \times (\text{mass of isotope})_n$
Solution:	$\text{Atomic Mass} = (0.6011)(68.9256\text{ amu}) + (0.3989)(70.9247\text{ amu})$ $\text{Atomic Mass} = 69.7\underline{2}3041 = 69.72\text{ amu}$
Check:	the average is between the two masses, closer to the major isotope

Practice:
Check your solution on next page

- If copper is 69.17% Cu-63 with a mass of 62.9396 amu and the rest Cu-65 with a mass of 64.9278 amu, find copper's atomic mass

Practice: If copper is 69.17% Cu-63 with a mass of 62.9396 amu and the rest Cu-65 with a mass of 64.9278 amu, find copper's atomic mass

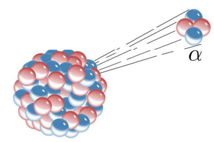
Given:	Cu-63 = 69.17%, 62.9396 amu Cu-65 = 100-69.17%, 64.9278 amu
Find:	atomic mass, amu
Conceptual Plan:	isotope masses, isotope fractions ⟹ avg. atomic mass
Relationships:	$\text{Atomic Mass} = \sum (\text{fractional abundance of isotope})_n \times (\text{mass of isotope})_n$
Solution:	$\text{Atomic Mass} = (0.6917)(62.9396\text{ amu}) + (0.3083)(64.9278\text{ amu})$ $\text{Atomic Mass} = 63.5525 = 63.55\text{ amu}$
Check:	the average is between the two masses, closer to the major isotope

Now that we learned about nuclei and particles …

Let's re-think what an alpha particle is

Remember? It was used in Rutherford's gold-foil experiment?

Alpha radiation



- **Nucleus of a Helium atom** ; has **+2 charge**
- Has **2 protons** and **2 neutrons**
  -- **Atomic number of the atom left behind changes** (because it has 2 less protons)
- Heaviest of the common radiation types

Origins of the Names and Symbols of the Elements

- Most chemical symbols are based on the English name of the element.
- Some symbols are based on Latin names.
- The symbol for potassium is K, from the Latin *kalium,* and the symbol for sodium is Na, from the Latin *natrium.*
- Additional elements with symbols based on their Greek or Latin names include the following:

lead	Pb	*plumbum*
mercury	Hg	*hydrargyrum*
iron	Fe	*ferrum*
silver	Ag	*argentum*
tin	Sn	*stannum*
copper	Cu	*cuprum*

Origins of the Names of the Elements

Early scientists gave newly discovered elements names that reflected their properties:

- *Argon,* from the Greek *argos*, means "inactive."

Other elements were named after countries. For example,

- *Polonium* after Poland
- *Francium* after France
- *Americium* after the United States of America.

Other elements were named after scientists.

Every element's name, symbol, and atomic number are included in the periodic table (inside the front cover) and in an alphabetical listing (inside the back cover) in this book.

Origin of the Names of the Elements

Curium is named after Marie Curie, a chemist who helped discover radioactivity and also discovered two new elements. Curie won two Nobel Prizes for her work.

Bromine originates from the Greek word *bromos*, meaning "stench."

Bromine vapor, seen as the red-brown gas in this photograph, has a strong odor.

Curium
96
Cm
(247)