Chapter 2
Atoms, Molecules, and Ions

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Cem 151

Atomic Theory of Matter

The theory of atoms:

Original to the Greeks

Leuccipus, Democritus and Lucretius

(Aristotle thought they were nuts)

He believed that one could divide up a piece of matter an infinite number of times, that is, one never came up with a piece of matter that could not be further divided. He suggested that everything in the world was made up of some combination of four elements: earth, fire, water, and air. The elements were acted upon by the two forces of gravity and levity. Gravity was the tendency for earth and water to sink, and levity the tendency for air and fire to rise.

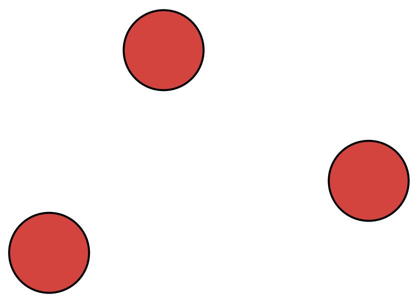
John Dalton (1805-1808)

Revived the idea and made it **science** by measuring the atomic weights of 21 elements.

That's the key thing because then you can see how elements combine.

Dalton’s Postulates

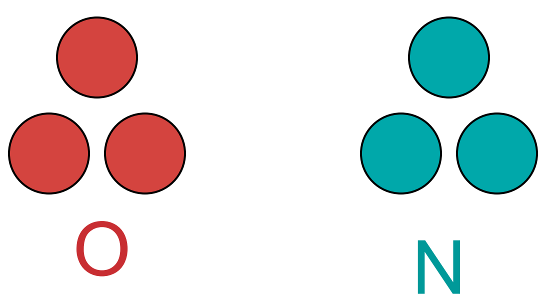
Each element is composed of extremely small particles called atoms.



Tiny balls make up the world

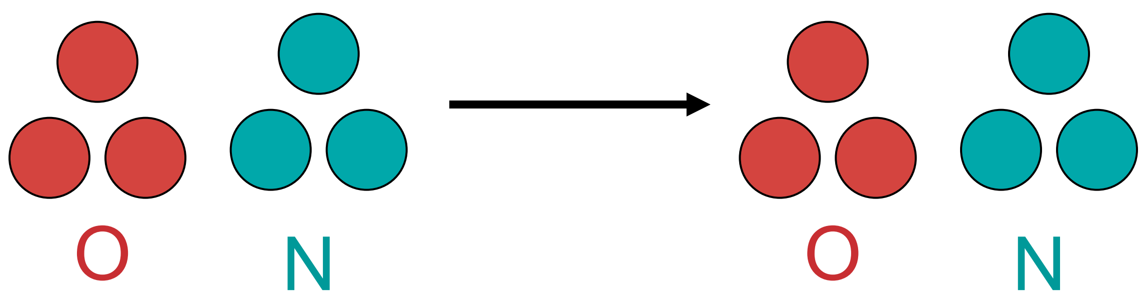
Dalton’s Postulates

All atoms of a given element are identical to one another in mass and other properties, but the atoms of one element are different from the atoms of all other elements.



Dalton's Postulates

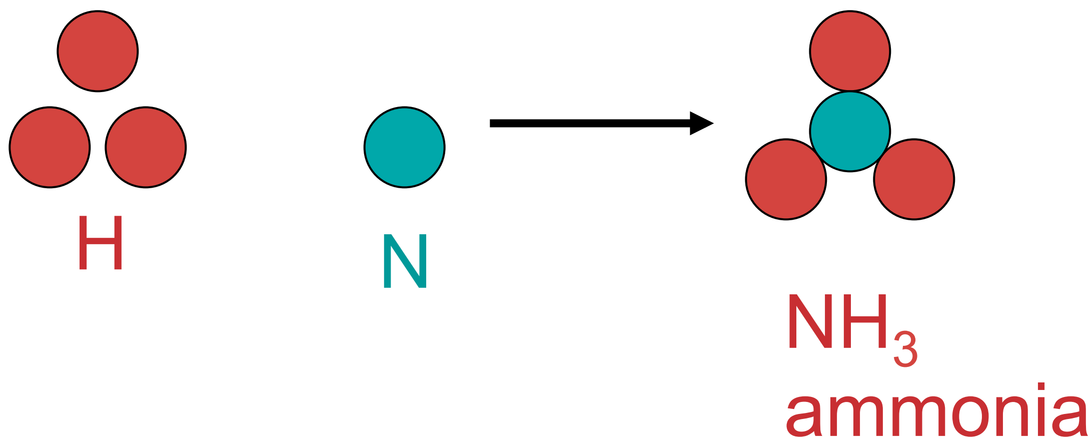
Atoms of an element are not changed into atoms of a different element by chemical reactions; atoms are neither created nor destroyed in chemical reactions. (As far as Dalton knew, they couldn't be changed at all).



Red O's stay Os and aqua N's stay N's.

Dalton's Postulates

Compounds are formed when atoms of more than one element combine; a given compound always has the same relative number and kind of atoms.

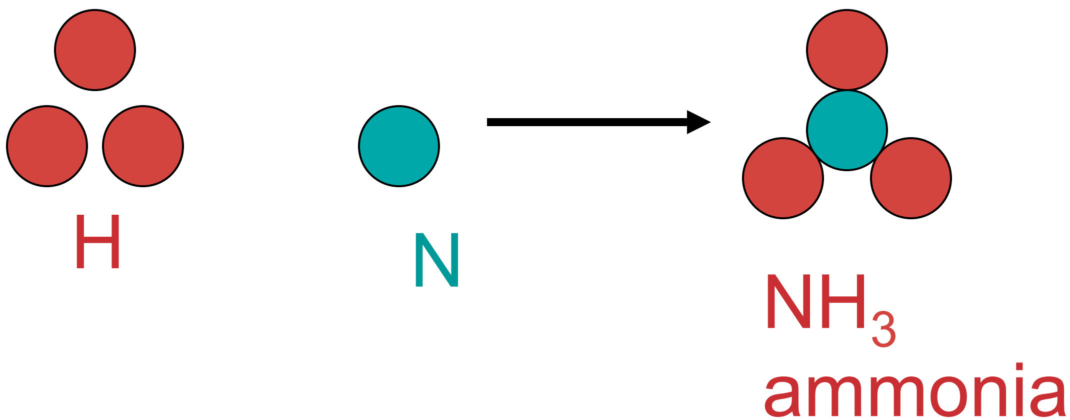


Chemistry happens when the balls rearrange

Law of Constant Composition

Joseph Proust (1754–1826)

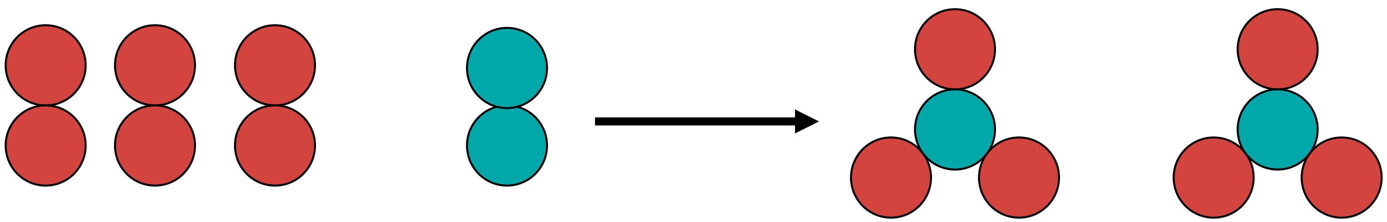
- Also known as the law of definite proportions.
- The elemental composition of a pure substance never varies.
- The relative amounts of each element in a compound doesn't vary.



ammonia always has 3 H and 1 N.

Law of Conservation of Mass

The total mass of substances present at the end of a chemical process is the same as the mass of substances present before the process took place.



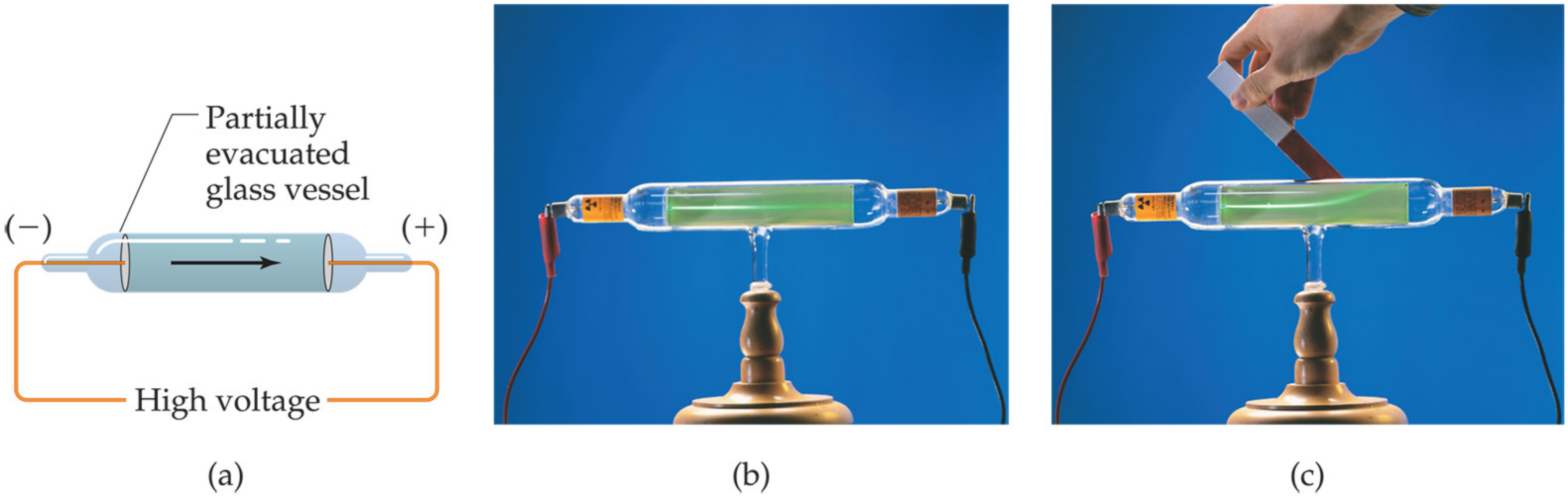
$3H_2$ + N_2 → $2NH_3$

ammonia

The atoms on the left all appear on the right

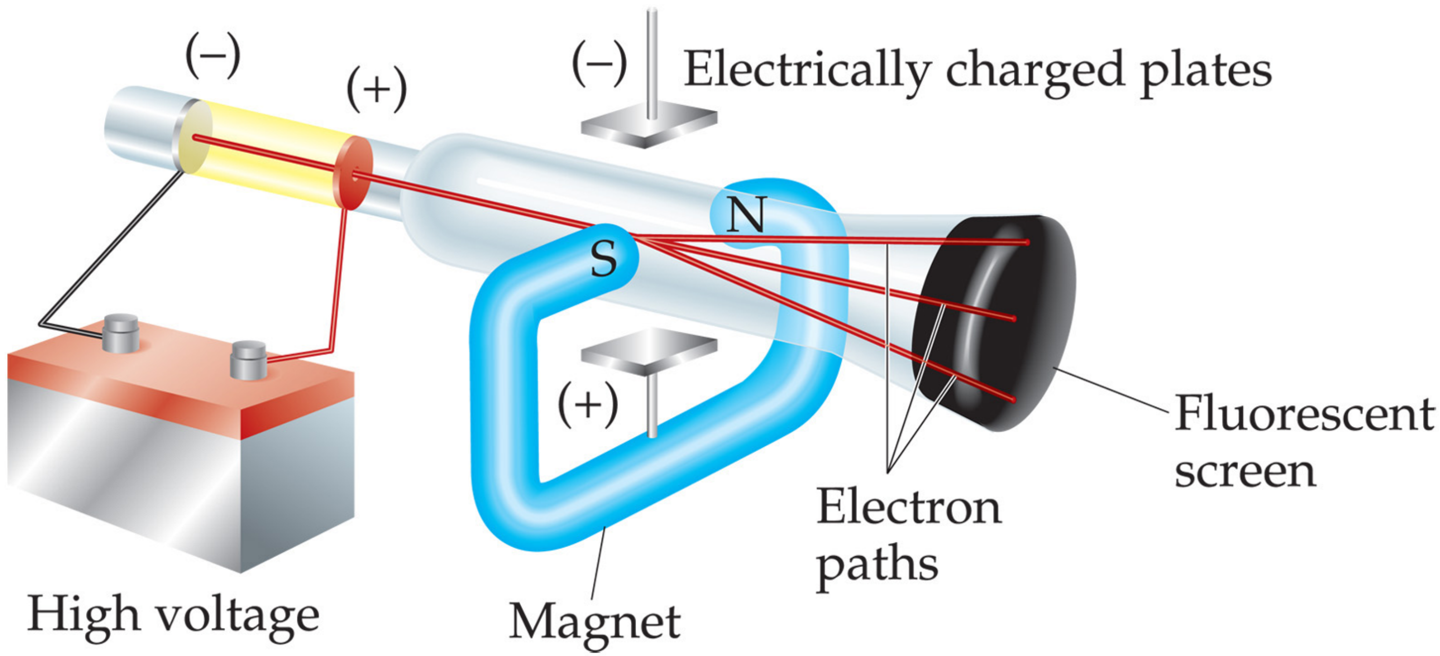
Atoms, Molecules, and Ions

The Electron



- Streams of negatively charged particles were found to emanate from cathode tubes.
- J. J. Thompson (1897).
- Maybe atoms weren't completely indivisible after all.

The Electron

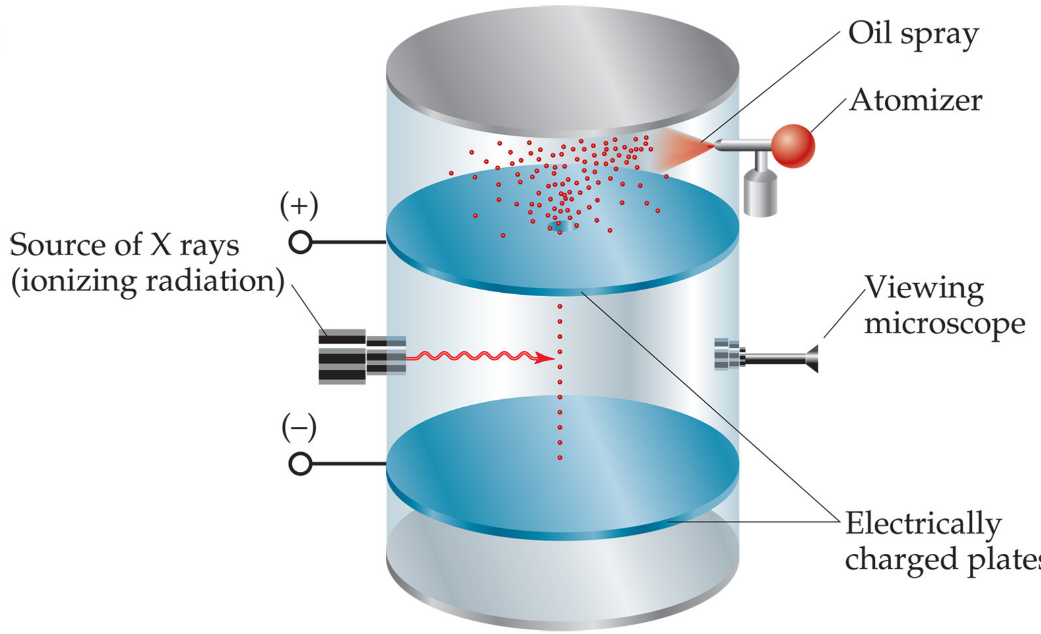


Thompson measured the charge/mass ratio of the electron to be 1.76×10^8 coulombs/g.

How? by manipulating the magnetic and electric fields and observing the change in the beam position on a fluorescent screen.

Millikan Oil Drop Experiment

measured charge of electron
Univ. Chicago (1909).



How?

Vary the electric field (E) until the drops stop.

Vary the charge (q) on the drop with more X-rays. Get a multiple of 1.6×10^{-19} Coulombs. The charge of 1 electron.

Eq = mg

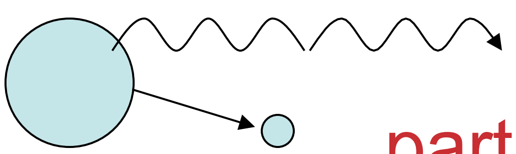
You set E, measure mass of drop (m) & know g. Find q.

Major result: q couldn't be any number. It was a multiple of 1.6×10^{-19} Coulombs

Atoms, Molecules, and Ions

Radioactivity:

- The spontaneous emission of radiation by an atom.
- First observed by Henri Becquerel.
- (1903 Nobel Prize with Pierre and Marie Curie)
- Also studied by Marie and Pierre Curie.



"rays" not particles

particles of some sort.

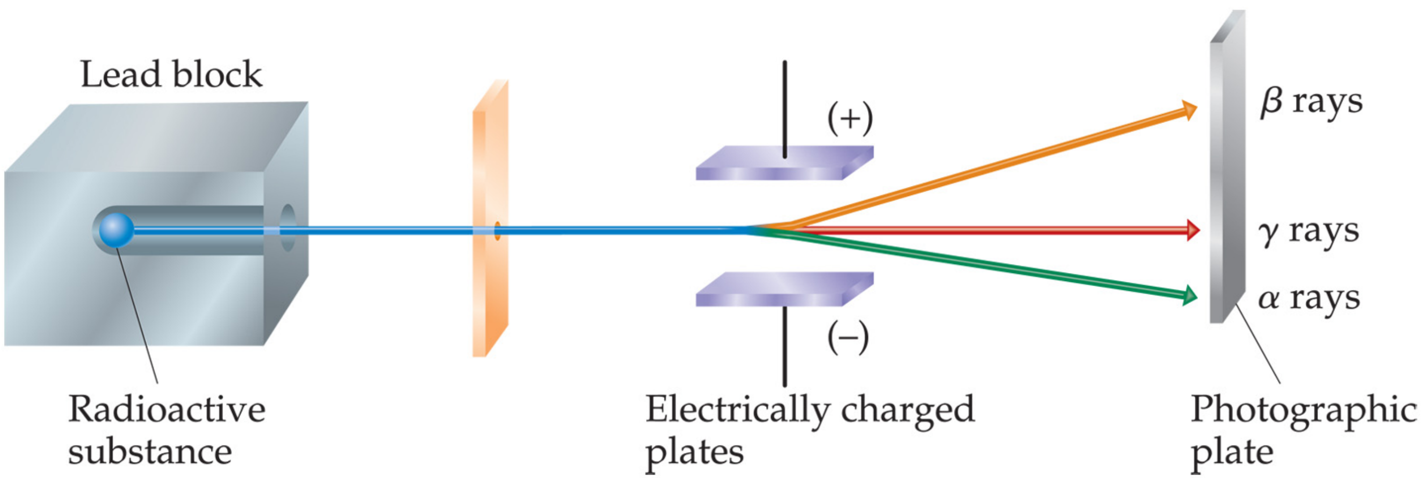
Stuff comes out of atoms, "subatomic particles"

Radioactivity

- Three types of radiation were discovered by Ernest Rutherford: (memorize the 3 types of particle)
  - α particles, attracted to negative electrode, so they have a positive charge, much more mass than negative stuff (turn out to be He nuclei)
  - β particles, attracted to positive electrode, so

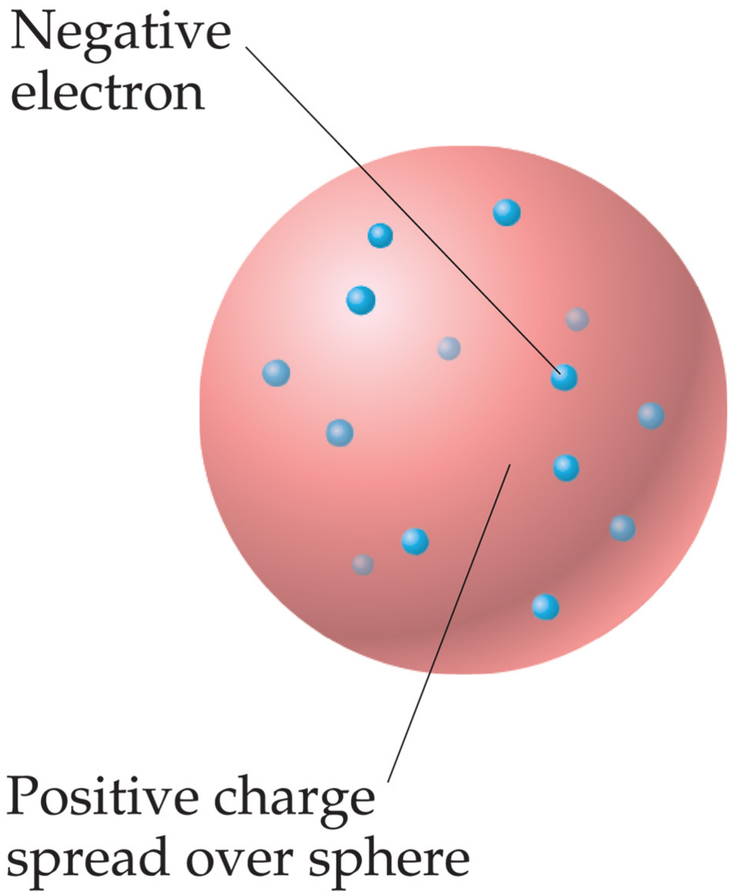
they have a negative charge, 1000s of times less massive (turn out to be electrons coming from nucleus).

  - γ rays, no charge, no mass, like light.



Atoms, Molecules, and Ions

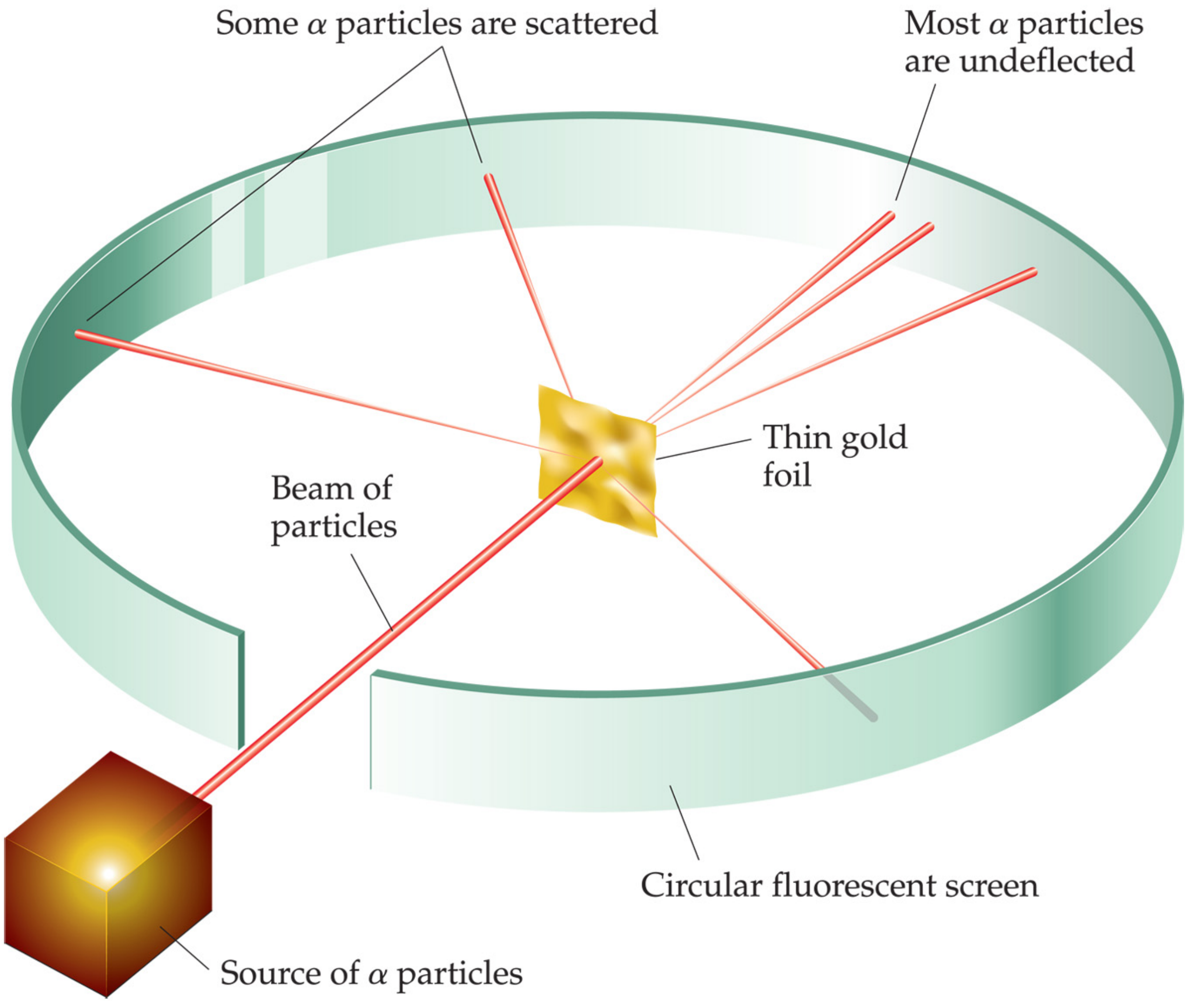
The Atom, circa 1900:



- "Plum pudding" model, put forward by Thompson.
- Positive sphere of matter with negative electrons imbedded in it.
- most of the volume = positive stuff because most of the mass is positive
- Expectation: density more or less uniform throughout.

Discovery of the Nucleus

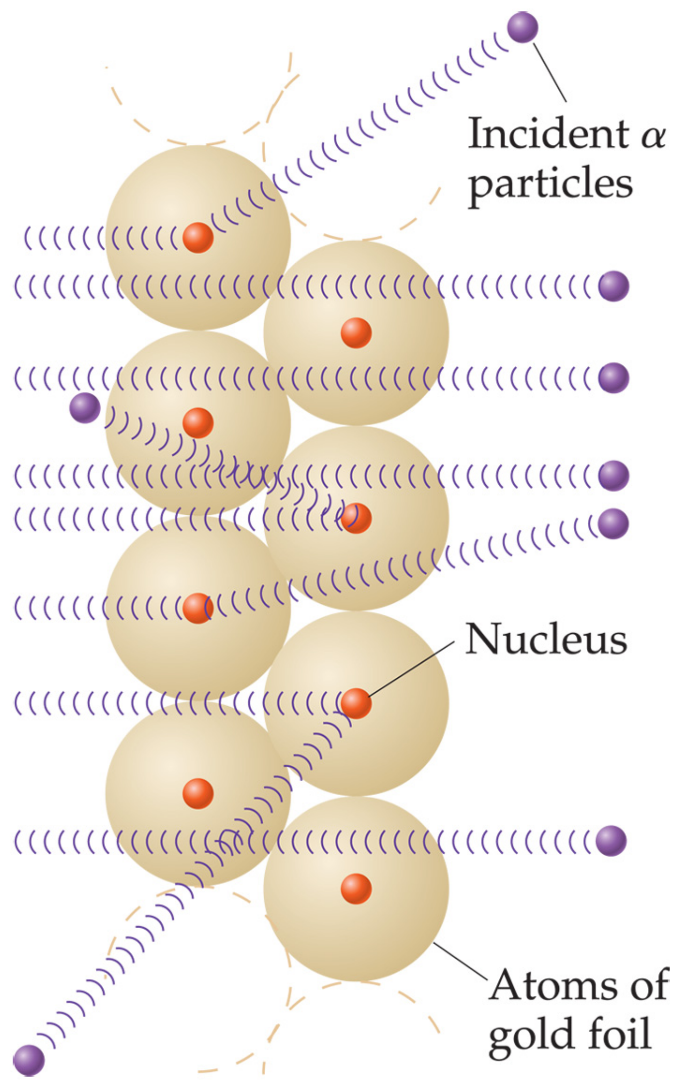
The Gold Foil Experiment



Rutherford and Marsden shot α particles at a thin sheet of gold foil and observed the pattern of scatter of the particles.

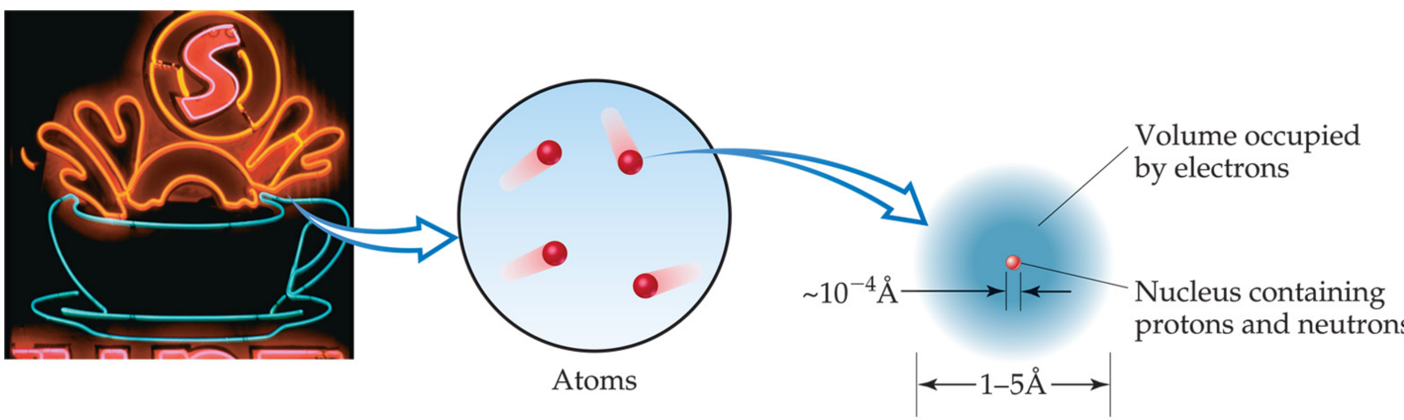
The Nuclear Atom

- **Virtually all the particles went straight through**
  - Most of the atom essentially empty
- **A few particles deflected,some straight back.**
  - A very small part of the atom is very dense, impenetrable.
  - The mass must be concentrated.
- The size of nucleus will be proportional to the # of highly scattered particles.



The Nuclear Atom

- Rutherford and Marsden postulated a very small, dense nucleus with the negative electrons around the outside of the atom.
- Most of the volume of the atom is empty space.



Other Subatomic Particles

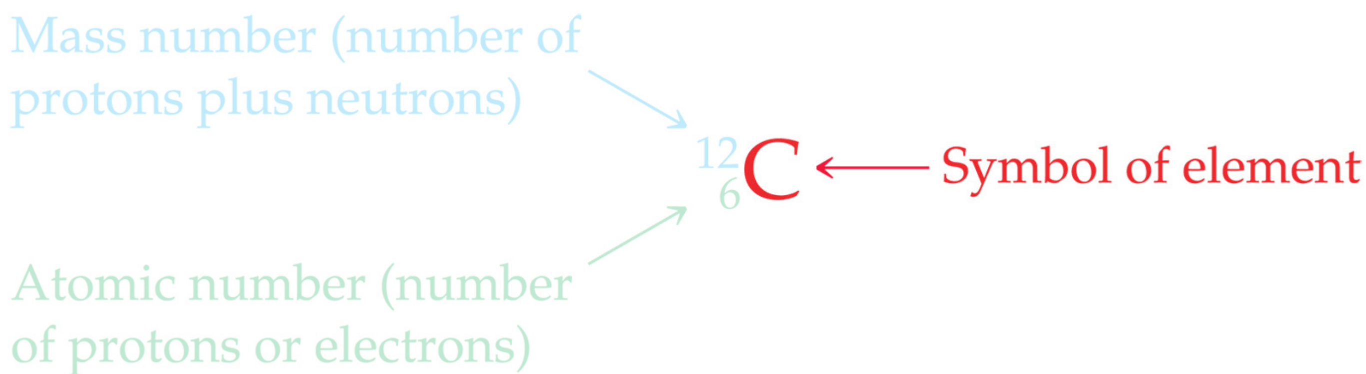
- Protons were discovered by Rutherford in 1919.
  - Have the positive charge in the atom.
- Neutrons were discovered by James Chadwick in 1932.
  - Have mass like proton, but no charge. Why was it harder to discover them?

Subatomic Particles

- Protons and electrons are the only particles that have a charge.
- Protons and neutrons have similar mass.
- The mass of an electron is so small we will often ignore it.

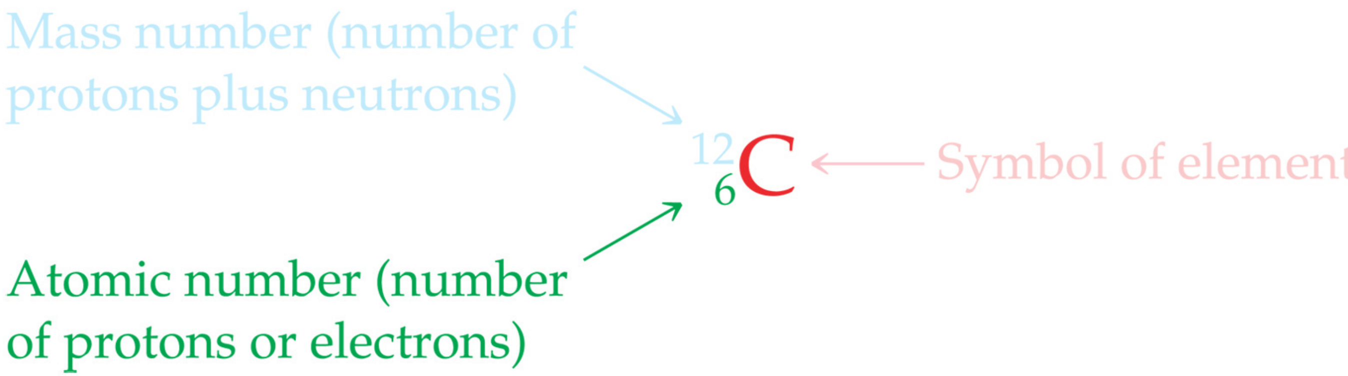
Particle	Charge	Mass (amu)
Proton	Positive (1+)	1.0073
Neutron	None (neutral)	1.0087
Electron	Negative (1−)	5.486×10^{-4}

Symbols of Elements: depicting the subatomic particles



Elements are symbolized by one or two letters.

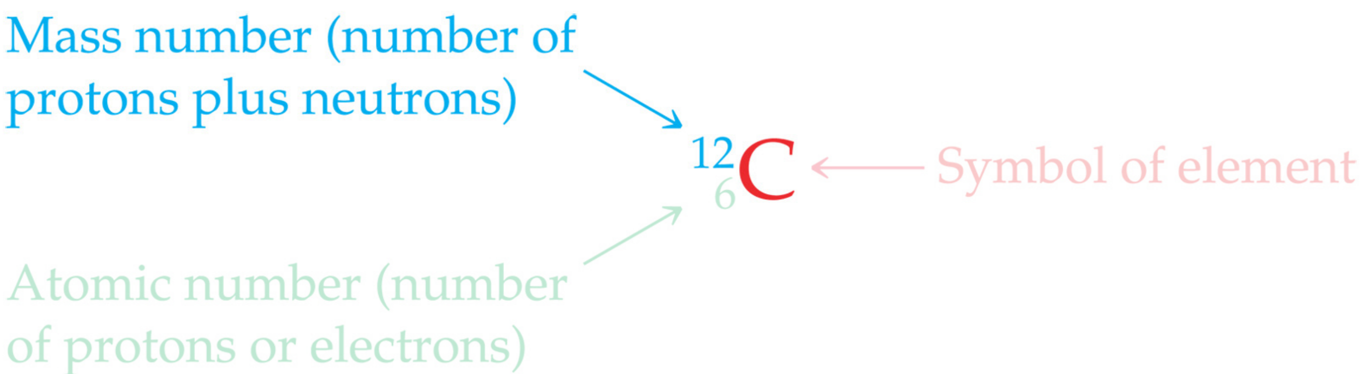
Atomic Number



All atoms of the same element have the same number of protons:

The atomic number (Z)

Atomic Mass



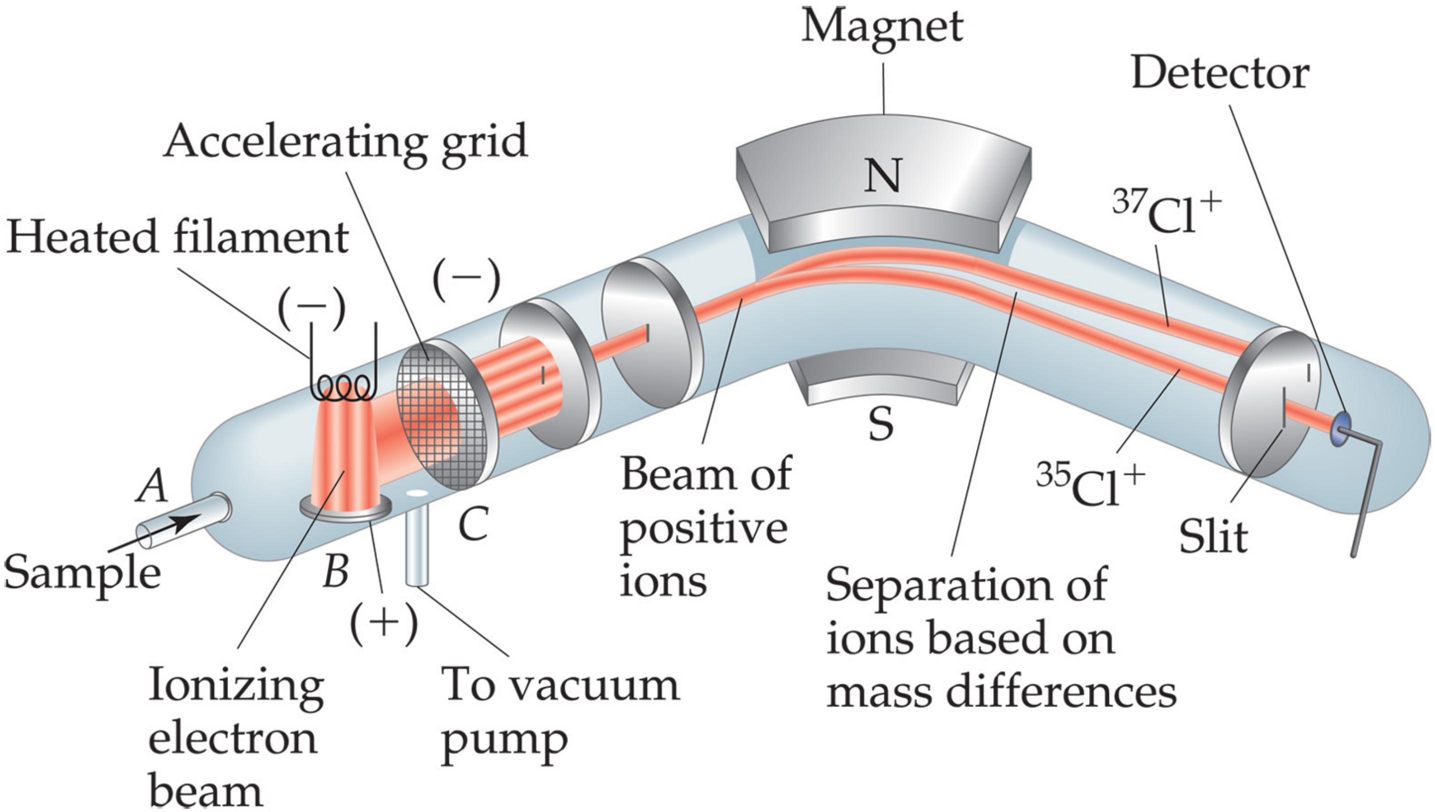
The mass of an atom in atomic mass units (amu) is approximately the total number of protons and neutrons in the atom.

Isotopes:

- Elements are defined by the number of protons.
- Isotopes are atoms of the same element with different masses.
- Isotopes have different numbers of neutrons.

	$^{11}_{6}C$	$^{12}_{6}C$	$^{13}_{6}C$	$^{14}_{6}C$
# Neutrons	5	6	7	8

Atomic Mass



Atomic and molecular masses can be measured with great accuracy with a mass spectrometer. Heavier ion turns less in the magnetic field (more momentum, because of higher mass (mv)) (magnetic moments of ions similar). Aston, 1919.

Average Mass

- Because in the real world all the elements exist as mixtures of isotopes.
- And we measure many many atoms at a time

"Natural abundance"

  - Average mass is calculated from the isotopes of an element weighted by their relative abundances.

Average mass, example

Isotope	abundance	Atomic mass
^{24}Mg	78.99%	23.98504 amu
^{25}Mg	10.00%	24.98584 amu
^{26}Mg	11.01%	25.98259 amu

Given the above data, what is the average molecular mass of magnesium (Mg)?

$$.7899(23.98504)+0.1000(24.98584)+0.1101(25.98259)=$$

$$18.95 + 2.499 + 2.861 = 24.31$$

Periodic Table:

	1A 1	2A 2	3B 3	4B 4	5B 5	6B 6	7B 7	8B 8	8B 9	8B 10	1B 11	2B 12	3A 13	4A 14	5A 15	6A 16	7A 17	8A 18
1	1 H																	2 He
2	3 Li	4 Be											5 B	6 C	7 N	8 O	9 F	10 Ne
3	11 Na	12 Mg											13 Al	14 Si	15 P	16 S	17 Cl	18 Ar
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5	37 Rb	38 Sr	39 Y	40 Zr	41 Nb	42 Mo	43 Tc	44 Ru	45 Rh	46 Pd	47 Ag	48 Cd	49 In	50 Sn	51 Sb	52 Te	53 I	54 Xe
6	55 Cs	56 Ba	71 Lu	72 Hf	73 Ta	74 W	75 Re	76 Os	77 Ir	78 Pt	79 Au	80 Hg	81 Tl	82 Pb	83 Bi	84 Po	85 At	86 Rn
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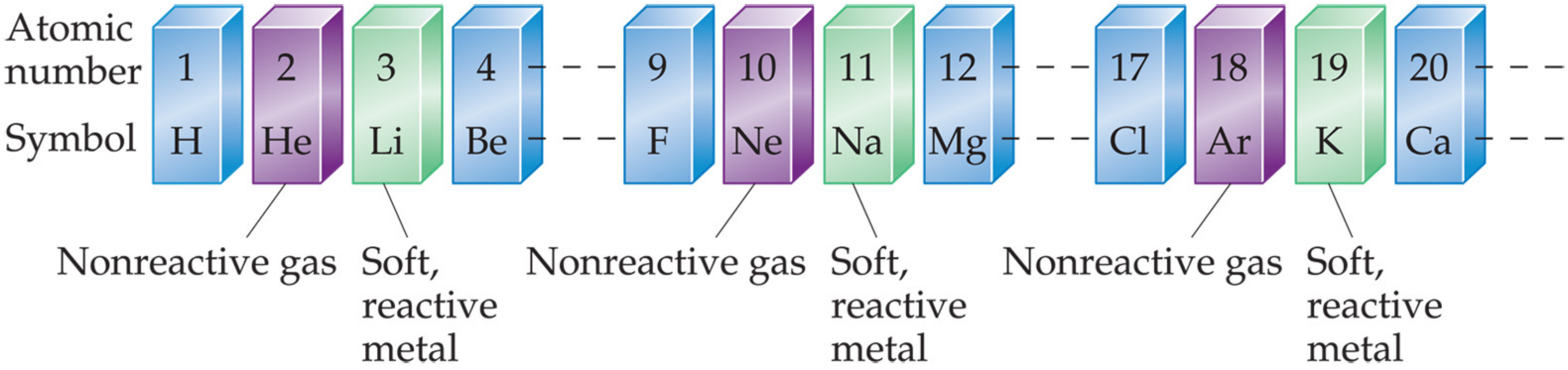
Metals

Metalloids

Nonmetals

- A systematic catalog of elements.
- Elements are arranged in order of atomic number.
- But why like this?

Periodicity



When one looks at the chemical properties of elements, one notices a repeating pattern of reactivities.

Periodic Table

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	1	2	3	4	5	6	7	8	9	10	11	12	13	14	15	16	17	18
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Metals

Metalloids

Nonmetals

- The rows on the periodic chart are periods.
- Columns are groups.
- Elements in the same group have similar chemical properties.
- Derived empirically, no theory to explain it.

Groups

Group	Name	Elements
1A	Alkali metals	Li, Na, K, Rb, Cs, Fr
2A	Alkaline earth metals	Be, Mg, Ca, Sr, Ba, Ra
6A	Chalcogens	O, S, Se, Te, Po
7A	Halogens	F, Cl, Br, I, At
8A	Noble gases (or rare gases)	He, Ne, Ar, Kr, Xe, Rn

These five groups are known by their names.

You gotta know these very well.

Periodic Table

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Metals

Metalloids

Nonmetals

Nonmetals are on the upper right-hand corner of the periodic table (with the exception of H).

Atoms, Molecules, and Ions

Periodic Table

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Metals

Metalloids

Nonmetals

Metalloids border the stair-step line (with the exception of Al and Po, which are both metals).

Atoms, Molecules, and Ions

Periodic Table

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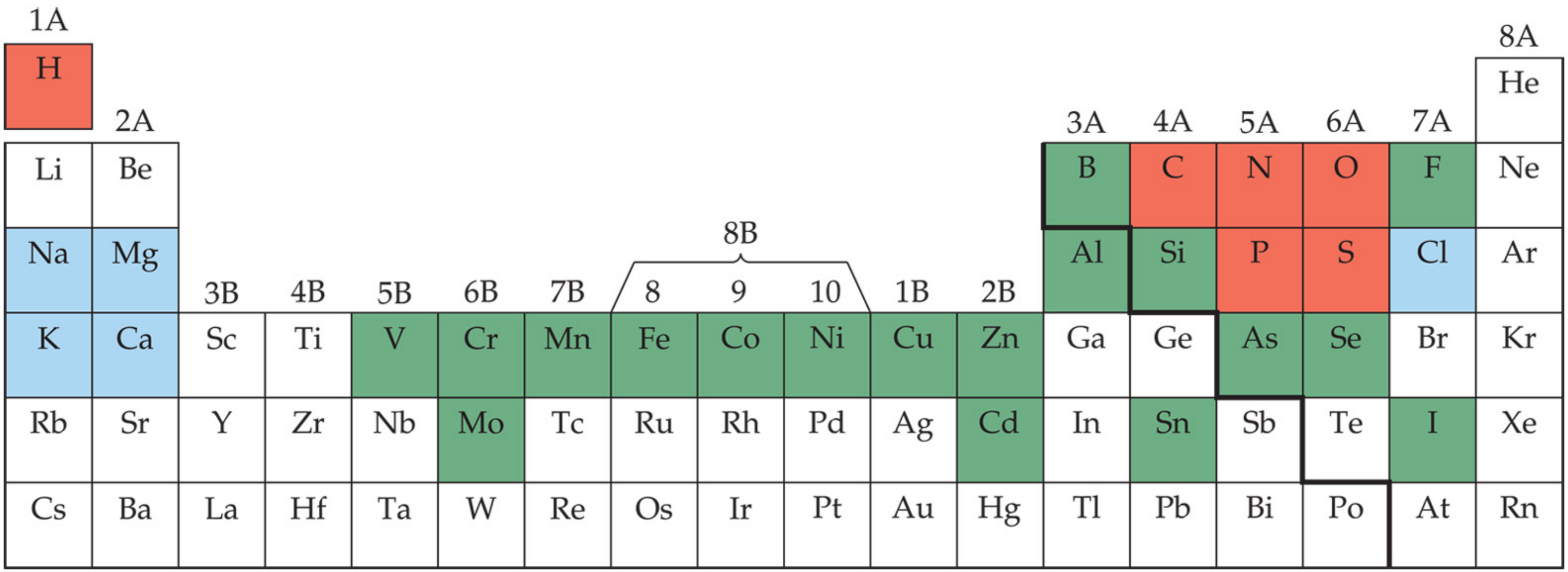
Metals

Metalloids

Nonmetals

Metals are on the left side of the chart.

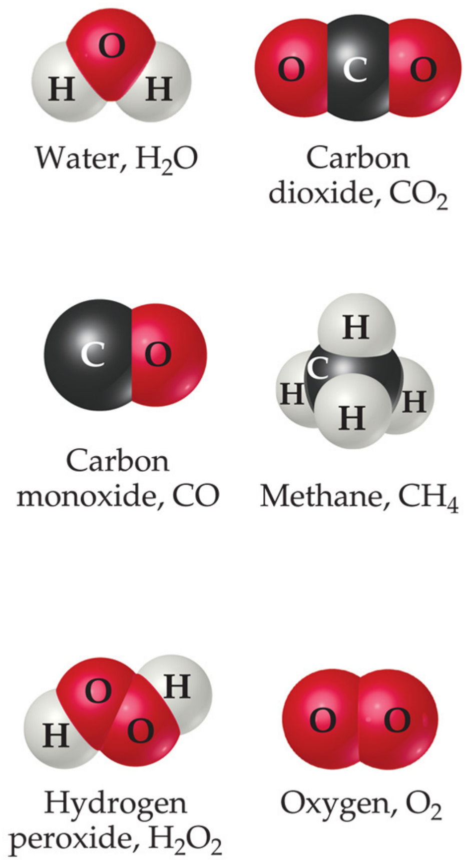
Elements of life



- Elements required for living organisms.
- Red, most abundant
- blue, next most abundant
- Green, trace amounts.

Atoms, Molecules, and Ions

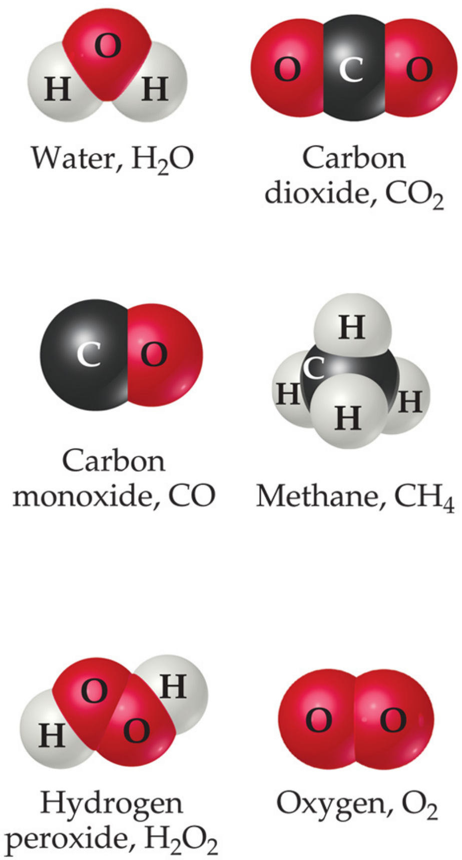
Chemical Formulas



The subscript to the right of the element

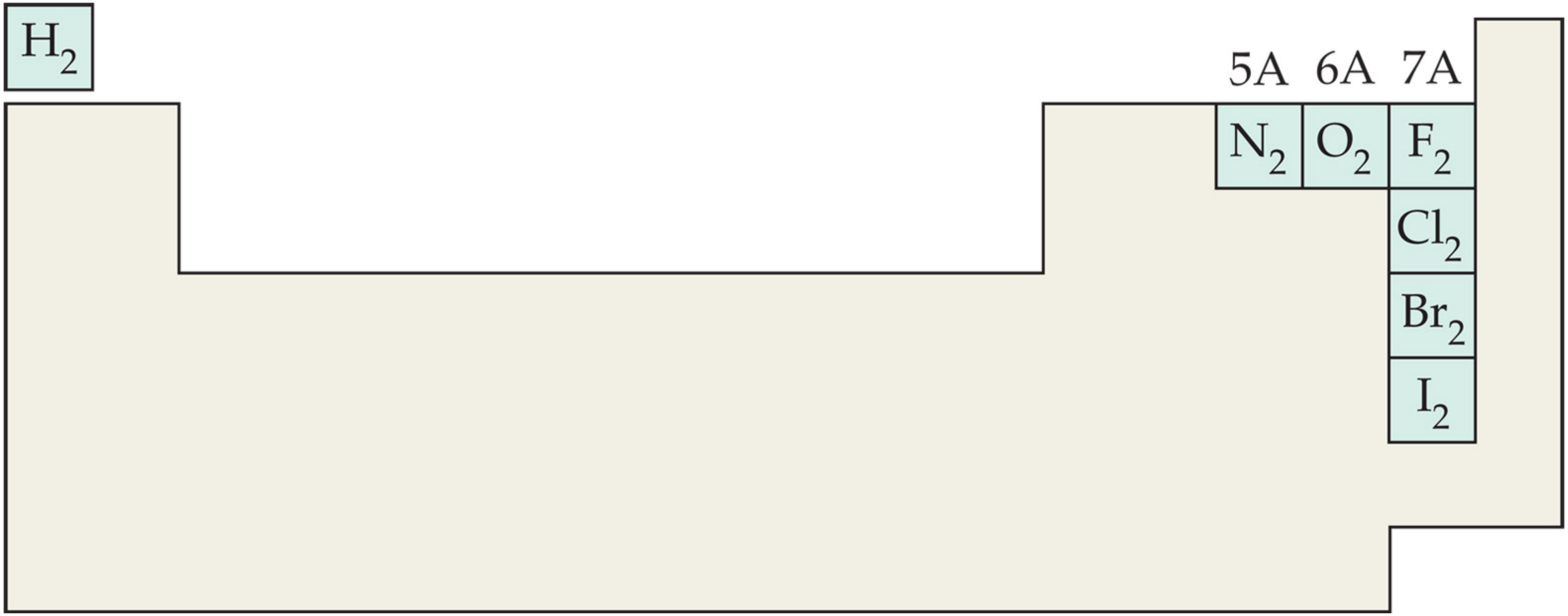
tells the number of atoms of that element in the compound.

Molecular Compounds



Molecular compounds are composed of **molecules** and almost always contain only nonmetals.

Diatomic Molecules



These seven elements occur naturally as molecules containing two atoms.

You should **know** these guys.

Types of Formulas

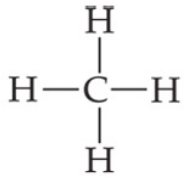
- Empirical formulas give the lowest whole-number ratio of atoms of each element in a compound.
- Molecular formulas give the exact number of atoms of each element in a compound.

Example: ethane:

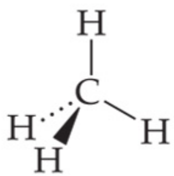
Empirical formula: CH_3

Molecular formula: C_2H_6

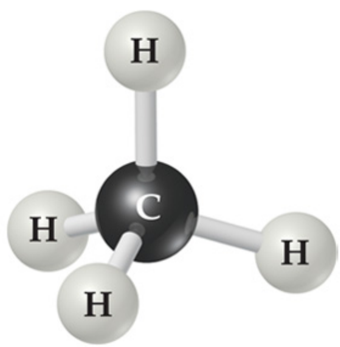
Types of Formulas



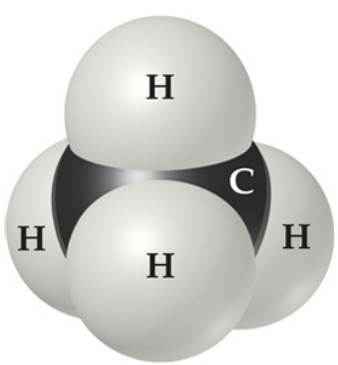
Structural formula



Perspective drawing



Ball-and-stick model

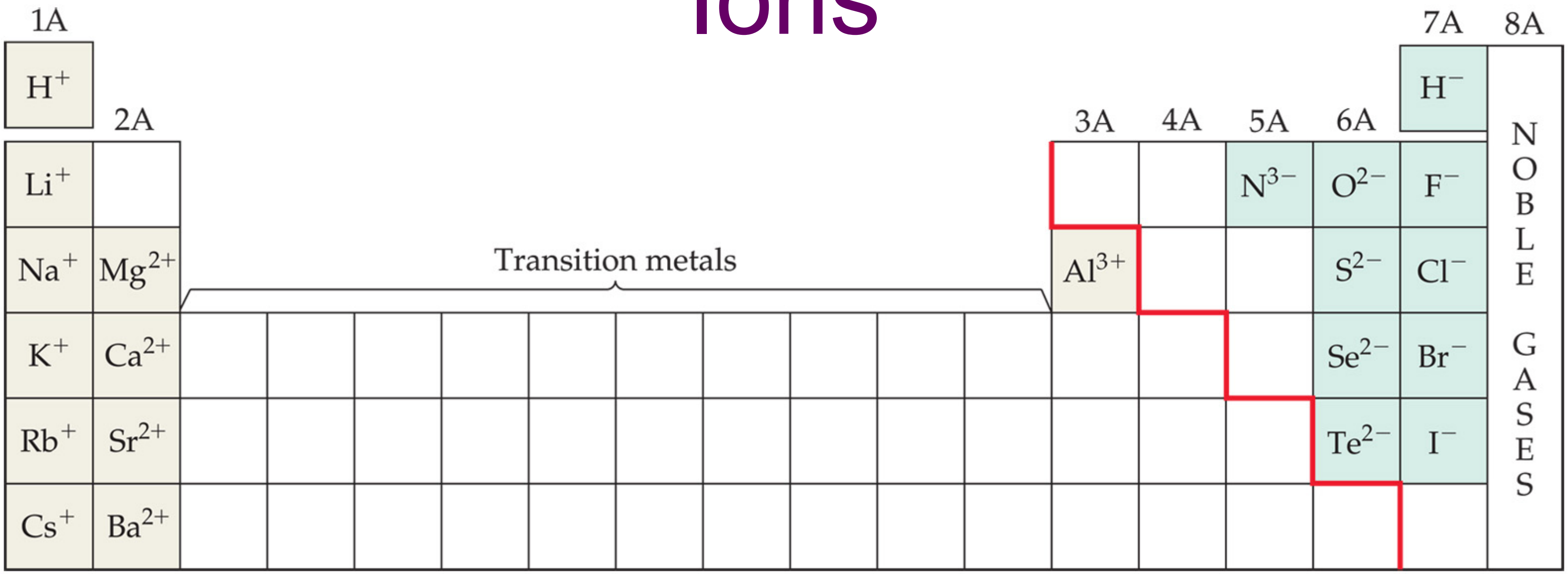


Space-filling model

- Structural formulas show the order in which atoms are bonded.

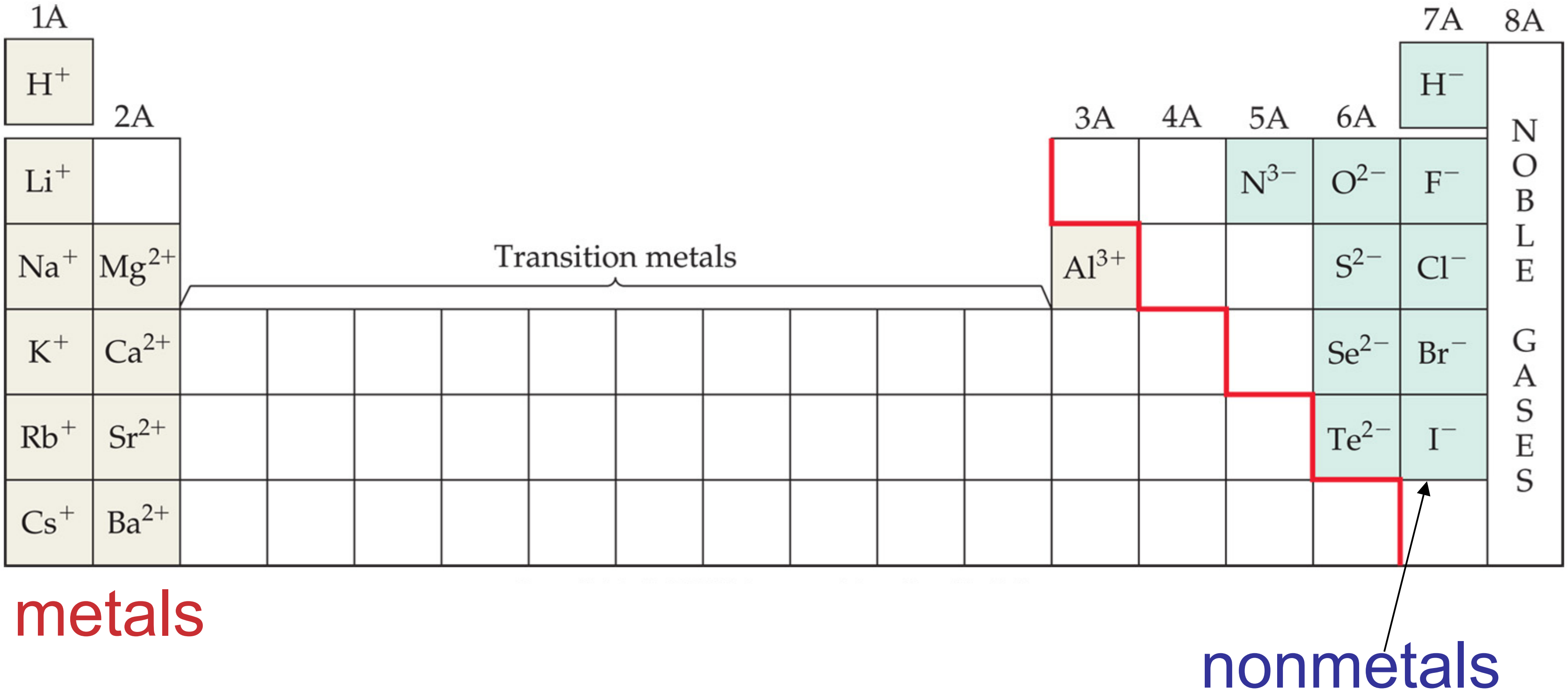
- Perspective drawings also show the three-dimensional array of atoms in a compound.

Ions



- When atoms lose or gain electrons, they become ions. Often they lose or gain electrons to have the same number of electrons as the nearest noble gas.
  - Cations are positive and are formed by elements on the left side of the periodic chart (**metals**).
  - Anions are negative and are formed by elements on the right side of the periodic chart (**nonmetals**).

Mono-atomic ions



- Metals usually become cations (+)
- Nonmetals usually become anions (-)

Ionic compounds

- A metal will give up electrons to a nonmetal forming a cation (+) (the metal), and an anion (-) (the nonmetal).

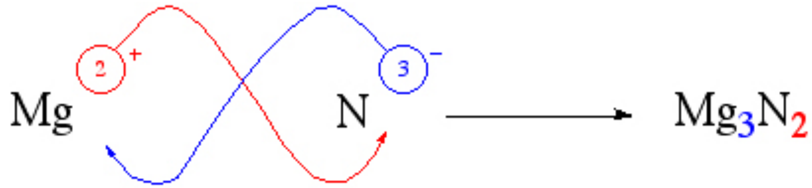
$Na + Cl \longrightarrow Na^{+} + Cl^{-} \longrightarrow NaCl$

$Mg + 2Cl \longrightarrow Mg^{2+} + 2Cl^{-} \longrightarrow MgCl_2$

Note, everybody gains or loses electrons to be like the nearest noble gas.

Compounds are always electrically neutral!!

Writing Formulas



- Because compounds are electrically neutral, one can determine the formula of a compound this way:
  - The charge on the cation becomes the subscript on the anion.
  - The charge on the anion becomes the subscript on the cation.
  - If these subscripts are not in the lowest whole-number ratio, divide them by the greatest common factor.

Mg^{2+} O^{2-} $\longrightarrow$ MgO Not Mg_2O_2

Common Cations

Charge	Formula	Name	Formula	Name
1+	* **H^{+}**	**Hydrogen ion**	* **NH_4^{+}**	**Ammonium ion**
	* Li^{+}	Lithium ion	* Cu^{+}	Copper(I) or cuprous ion
	* **Na^{+}**	**Sodium ion**		
	* **K^{+}**	**Potassium ion**		
	* Cs^{+}	Cesium ion		
	* **Ag^{+}**	**Silver ion**		
2+	* **Mg^{2+}**	**Magnesium ion**	Co^{2+}	Cobalt(II) or cobaltous ion
	* **Ca^{2+}**	**Calcium ion**	* **Cu^{2+}**	**Copper(II)** or cupric ion
	* Sr^{2+}	Strontium ion	* **Fe^{2+}**	**Iron(II)** or ferrous ion
	* Ba^{2+}	Barium ion	Mn^{2+}	Manganese(II) or manganous ion
	* **Zn^{2+}**	**Zinc ion**	Hg_2^{2+}	Mercury(I) or mercurous ion
	* Cd^{2+}	Cadmium ion	**Hg^{2+}**	**Mercury(II)** or mercuric ion
			* Ni^{2+}	Nickel(II) or nickelous ion
			* **Pb^{2+}**	**Lead(II)** or plumbous ion
			Sn^{2+}	Tin(II) or stannous ion
3+	* **Al^{3+}**	**Aluminum ion**	* Cr^{3+}	Chromium(III) or chromic ion
			* **Fe^{3+}**	**Iron(III)** or ferric ion

*The most common ions are in boldface.

*You should know these.

Common Anions

Charge	Formula	Name
1−	H^-	Hydride ion
	F^-	**Fluoride ion**
	Cl^-	**Chloride ion**
	Br^-	**Bromide ion**
	I^-	**Iodide ion**
	CN^-	Cyanide ion
	OH^-	**Hydroxide ion**
2−	**O^{2-}**	**Oxide ion**
	O_2^{2-}	Peroxide ion
	S^{2-}	**Sulfide ion**
3−	N^{3-}	Nitride ion

*The most common ions are in boldface.

Polyatomic anions

I_3^-	triiodide	HPO_4^{2-}	hydrogen phosphate
O_2^-	Superoxide	$H_2PO_4^-$	dihydrogen phosphate
OH^-	hydroxide	PO_4^{-3}	Phosphate
CN^-	cyanide	ClO^-	hypochlorite
SCN^-	thiocyanate	ClO_2^-	chlorite
NO_3^-	nitrate	ClO_3^-	chlorate
NO_2^-	nitrite	ClO_4^-	perchlorate
SO_3^{-2}	sulfite	MnO_4^-	Permanganate
HSO_3^-	bisulfite	CrO_4^{-2}	Chromate
SO_4^{-2}	sulfate	$Cr_2O_7^{-2}$	Dichromate
HSO_4^-	bisulfate		
HCO_3^-	bicarbonate		
CO_3^{-2}	carbonate		
$CH_3CO_2^-$	Acetate		

You must memorize **ALL** of these!

Atoms, Molecules, and Ions

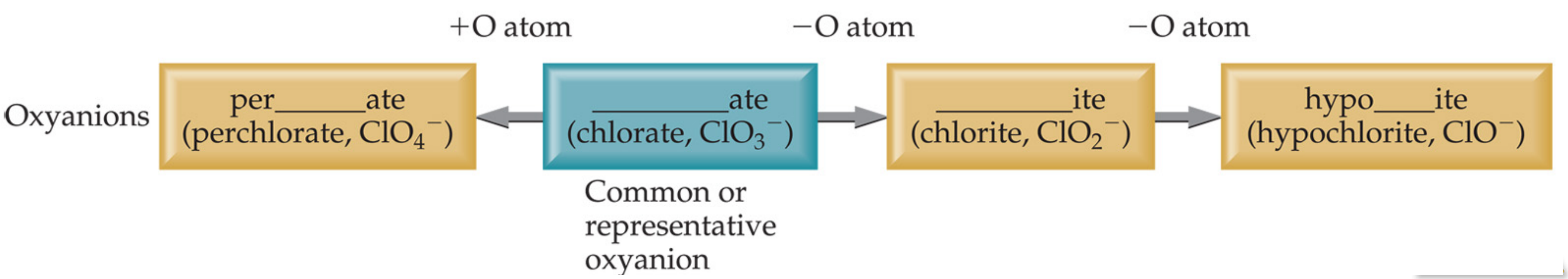
Patterns in Oxyanion Nomenclature

- When there are **only two** oxyanions involving the same element:
  - The one with fewer oxygens ends in -*ite*
    - NO_2^- : nitrite; SO_3^{2-} : sulfite
  - The one with more oxygens ends in -*ate*
    - NO_3^- : nitrate; SO_4^{2-} : sulfate

Patterns in Oxyanion Nomenclature

When there are more than two:

- The one with the fewest oxygens has the prefix *hypo*- and ends in -*ite*
  - ClO^- : hypochlorite
- The one with the second fewest oxygens ends in -*ite*
  - ClO_2^- : chlorite
- The one with the second most oxygens ends in *-ate*
  - ClO_3^- : chlorate
- The one with the most oxygens has the prefix *per*- and ends in -*ate*
  - ClO_4^- : perchlorate



Inorganic Nomenclature

- name of cation goes first.
- If anion is element, change ending to -*ide*;
- If anion is polyatomic ion, simply write the name of the polyatomic ion.
- If the cation can have more than one possible charge, write the charge as a Roman numeral in parentheses. (Fe(II), Fe(III))

Examples
naming inorganic compounds

- Write the name of the cation.
- If the anion is an element, change its ending to -*ide*; if the anion is a polyatomic ion, simply write the name of the polyatomic ion.
- If the cation can have more than one possible charge, write the charge as a Roman numeral in parentheses.

$NaCl$	sodium chloride
NH_4NO_3	ammonium nitrate
$Fe(SO_4)$	Iron(II) sulfate
KCN	potassium cyanide
$RbOH$	Rubidium hydroxide
$LiC_2H_3O_2$	lithium acetate
$NaClO_3$	sodium chlorate
$NaClO_4$	sodium perchlorate
K_2CrO_4	potassium chromate
NaH	Sodium hydride

Examples
naming inorganic compounds

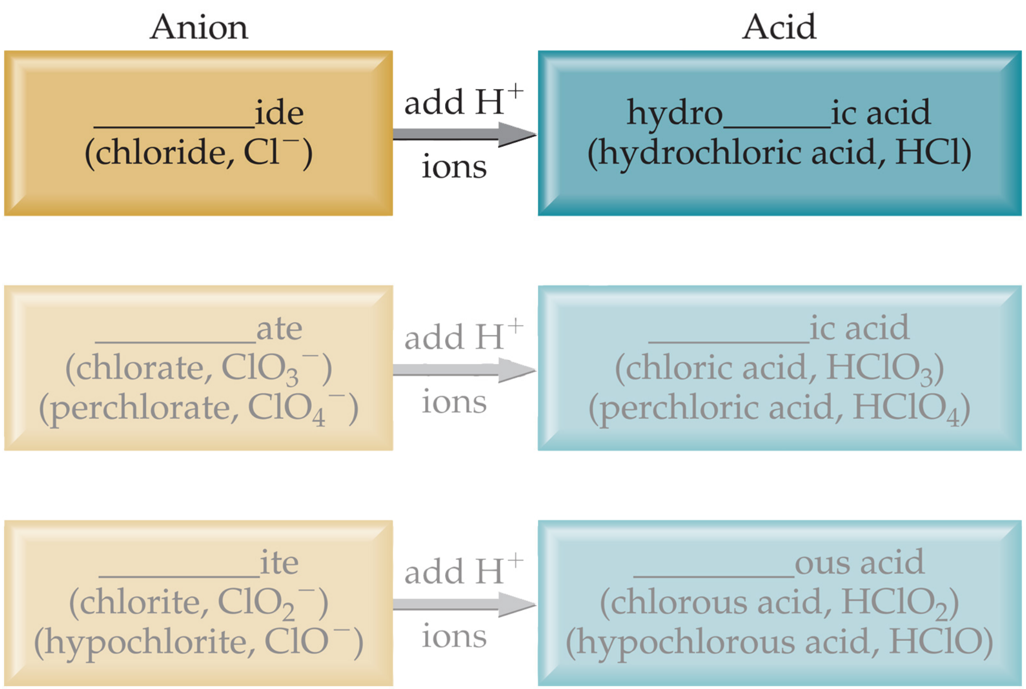
- Write the name of the cation.
- If the anion is an element, change its ending to -*ide*; if the anion is a polyatomic ion, simply write the name of the polyatomic ion.
- If the cation can have more than one possible charge, write the charge as a Roman numeral in parentheses.

potasium permanganate	$KMnO_4$
Calcium carbonate	$CaCO_3$
Calcium bicarbonate	$Ca(HCO_3)_2$
ammonium dichromate	$NH_4(Cr_2O_7)$
potassium phosphate	K_3PO_4
Lithium oxide	Li_2O (O^{2-} is the anion)
sodium peroxide	Na_2O_2 (O_2^{2-} is the anion)
Calcium sulfide	CaS

Hydrogen

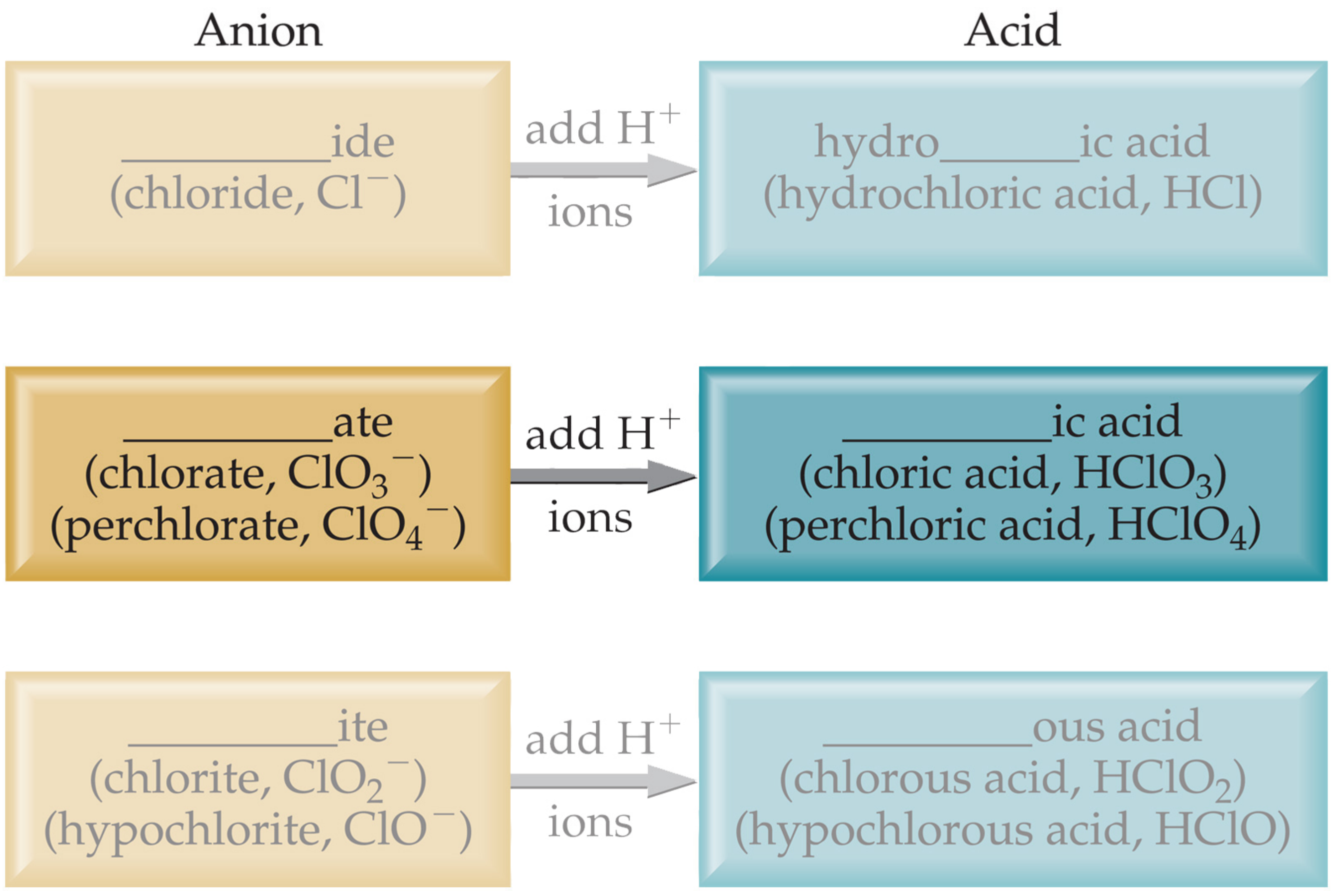
- H can be cation or anion
- H^- hydride
- H^+ (the cation of an inorganic compound) makes an acid, naming different.

Acid Nomenclature



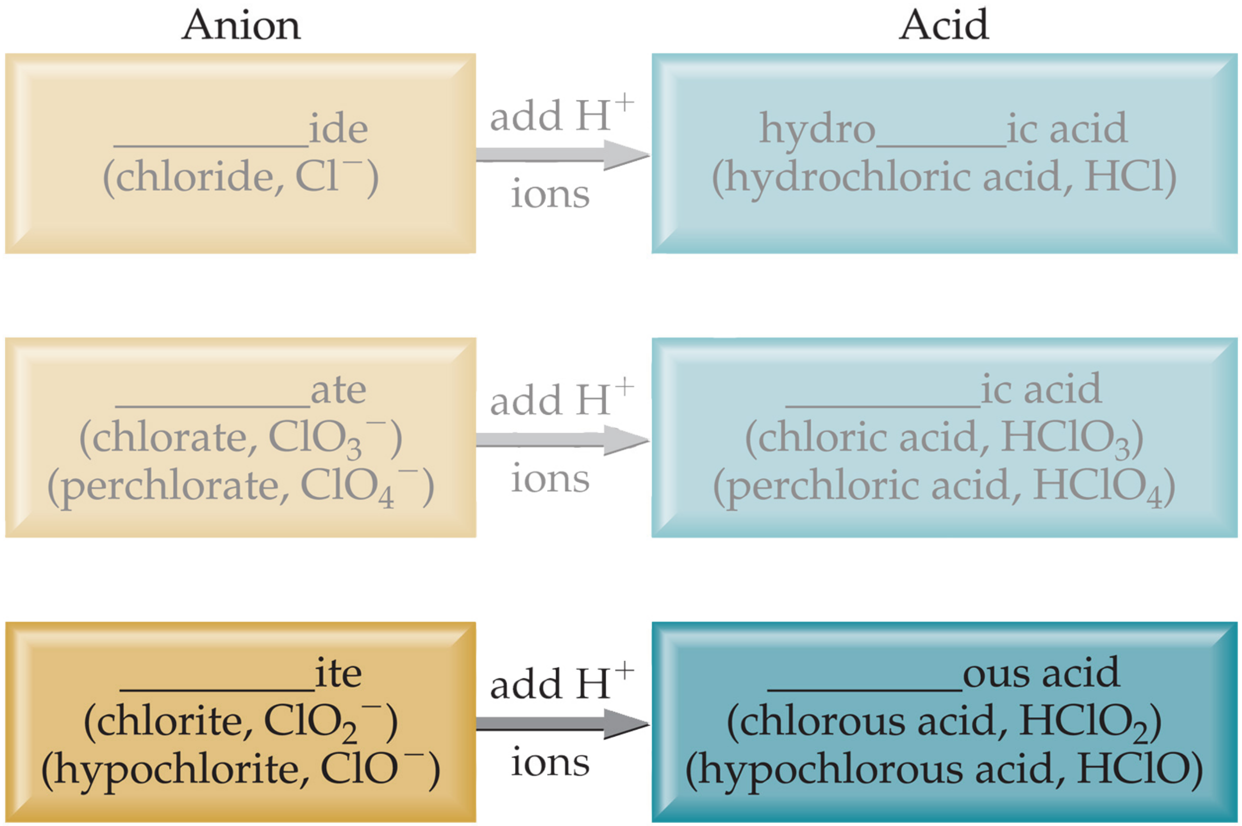
- If the anion in the acid ends in *-ide*, change the ending to *-ic acid* and add the prefix *hydro-* :
  - HCl: hydrochloric acid
  - HBr: hydrobromic acid
  - HI: hydroiodic acid

Acid Nomenclature



- If the anion in the acid ends in -*ate*, change the ending to -*ic acid:*
  - $HClO_3$: chloric acid
  - $HClO_4$: perchloric acid

Acid Nomenclature



- If the anion in the acid ends in -*ite*, change the ending to -*ous acid*:
  - HClO: hypochlorous acid
  - $HClO_2$: chlorous acid

Naming Binary Compounds (2 nonmetals)

Prefix	*Meaning*
Mono-	1
Di-	2
Tri-	3
Tetra-	4
Penta-	5
Hexa-	6
Hepta-	7
Octa-	8
Nona-	9
Deca-	10

- less electronegative atom (element closest to the lower lefthand corner of periodic table).
- A prefix is used to denote the number of atoms of each element in the compound (*mono*- is not used on the first element listed, however.)

Nomenclature of Binary Compounds (two nonmetals)

Prefix	*Meaning*
Mono-	1
Di-	2
Tri-	3
Tetra-	4
Penta-	5
Hexa-	6
Hepta-	7
Octa-	8
Nona-	9
Deca-	10

- The ending on the more electronegative element is changed to -*ide.*
  - CO_2: carbon dioxide
  - CCl_4: carbon tetrachloride

Nomenclature of Binary Compounds

Prefix	*Meaning*
Mono-	*1*
Di-	*2*
Tri-	*3*
Tetra-	*4*
Penta-	*5*
Hexa-	*6*
Hepta-	*7*
Octa-	*8*
Nona-	*9*
Deca-	*10*

If the prefix ends with *a* or *o* and the name of the element begins with a vowel, the two successive vowels are often merged into one:

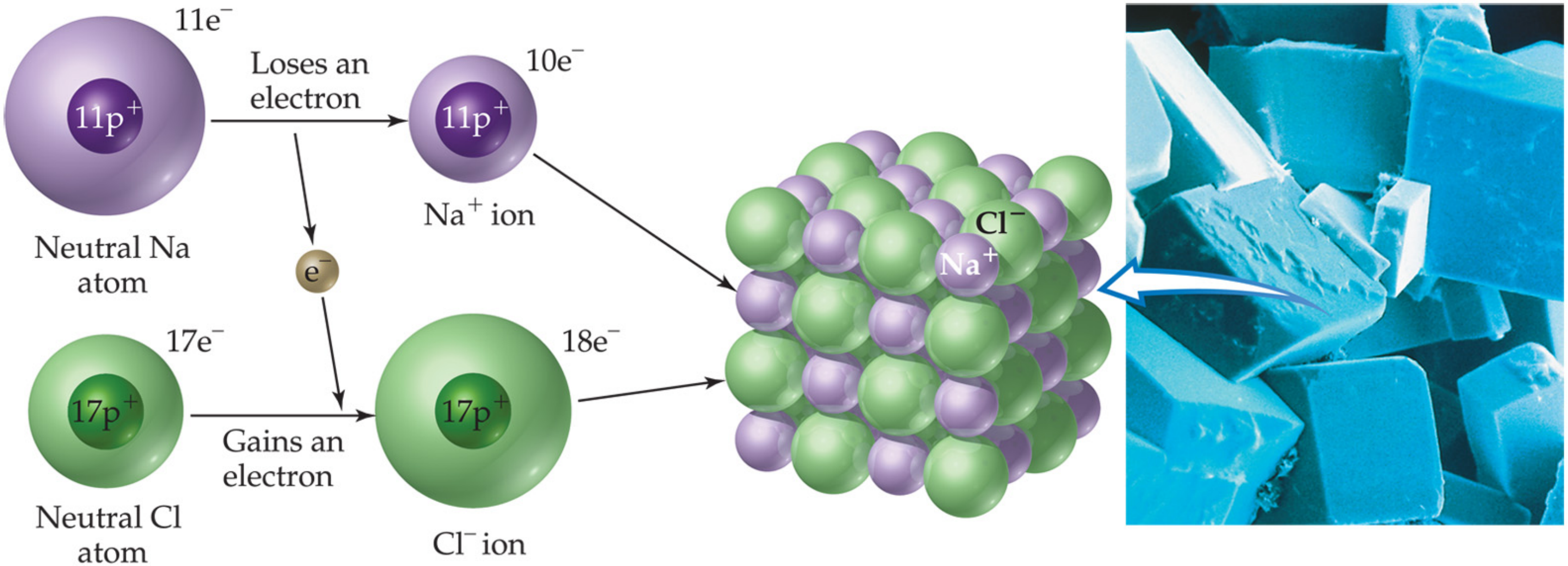
N_2O_5: dinitrogen pentoxide
not: dinitrogen pentaoxide

Nomenclature of binary compounds

- carbon dioxide — CO_2
- carbon tetrafluoride — CF_4
- nitrogen triiodide — NI_3
- oxygen difluoride — OF_2
- phosphorous pentachloride — PCl_5
- hydrogen sulfide — H_2S
- tetraphosphorous decoxide — P_4O_{10}

Ionic Bonds

Ionic compounds (such as NaCl) are generally formed between metals and nonmetals.



Barking Dog

$2HNO_3 + 2Cu \text{ ------->} \ NO + NO_2 + 2Cu^{2+} + 2H^+$

$3\ NO + CS_2 \text{ ->} \ 3/2\ N_2 + CO + SO_2 + 1/8\ S_8$

$4\ NO + CS_2 \text{ ->} \ 2\ N_2 + CO_2 + SO_2 + 1/8\ S_8$

Atoms, Molecules, and Ions