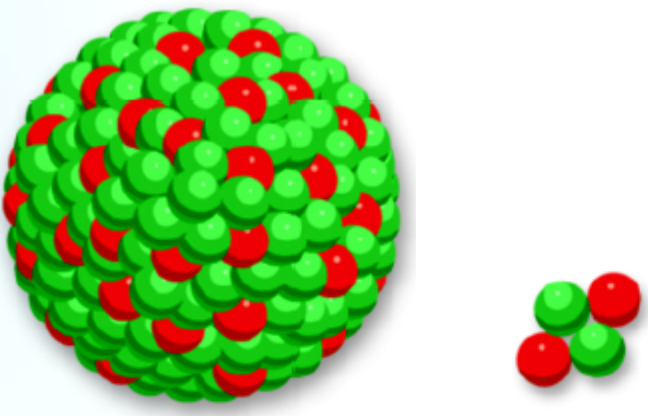
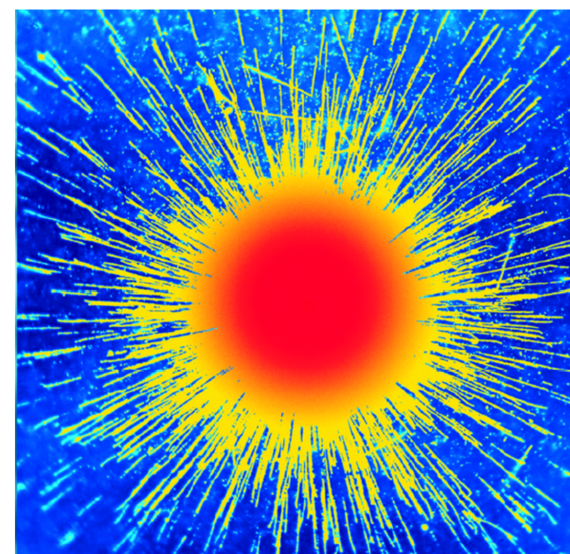


Atoms, Molecules and Ions

Chapter 2



Dalton's Atomic Theory (1808)

1. Elements are composed of extremely small particles called ***atoms***.

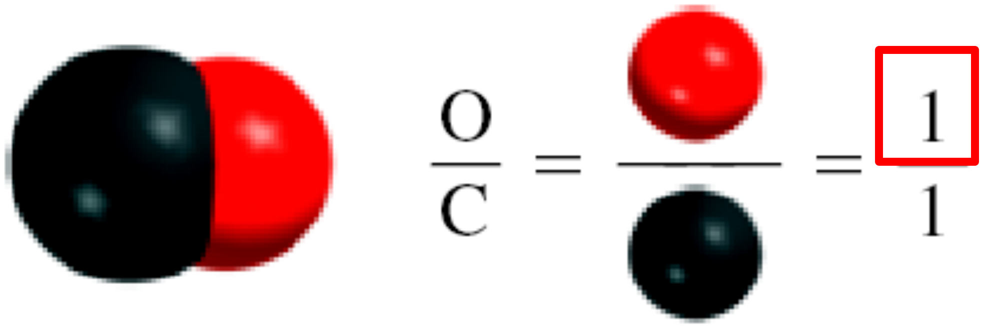
2. All ***atoms*** of a given element are identical, having the same size, mass and chemical properties. The atoms of one element are different from the atoms of all other elements.

3. ***Compounds*** are composed of atoms of more than one element. In any compound, the ratio of the numbers of atoms of any two of the elements present is either an integer or a simple fraction.

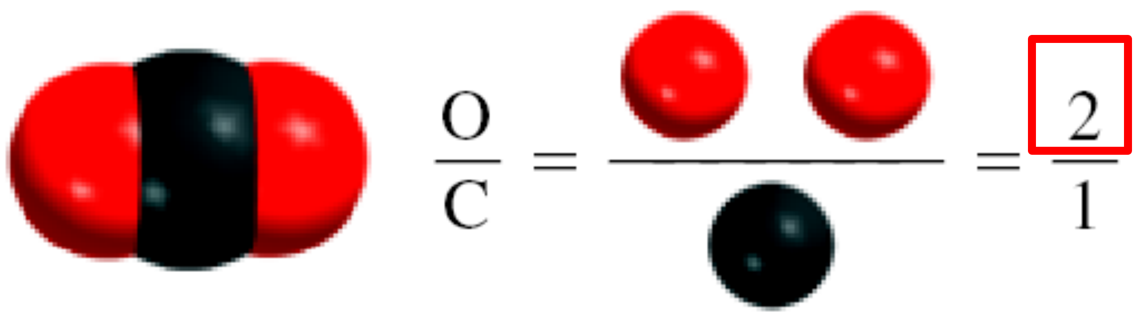
4. A ***chemical reaction*** involves only the separation, combination, or rearrangement of atoms; it does not result in their creation or destruction.

Dalton's Atomic Theory

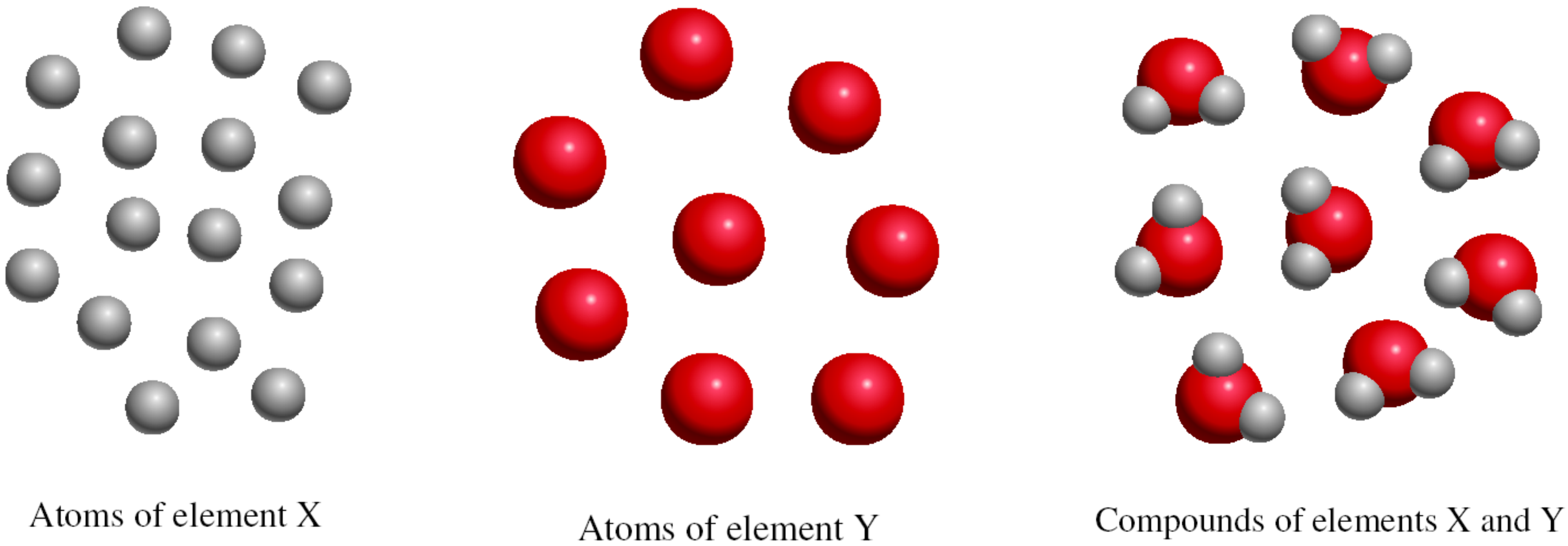
Carbon monoxide



Carbon dioxide



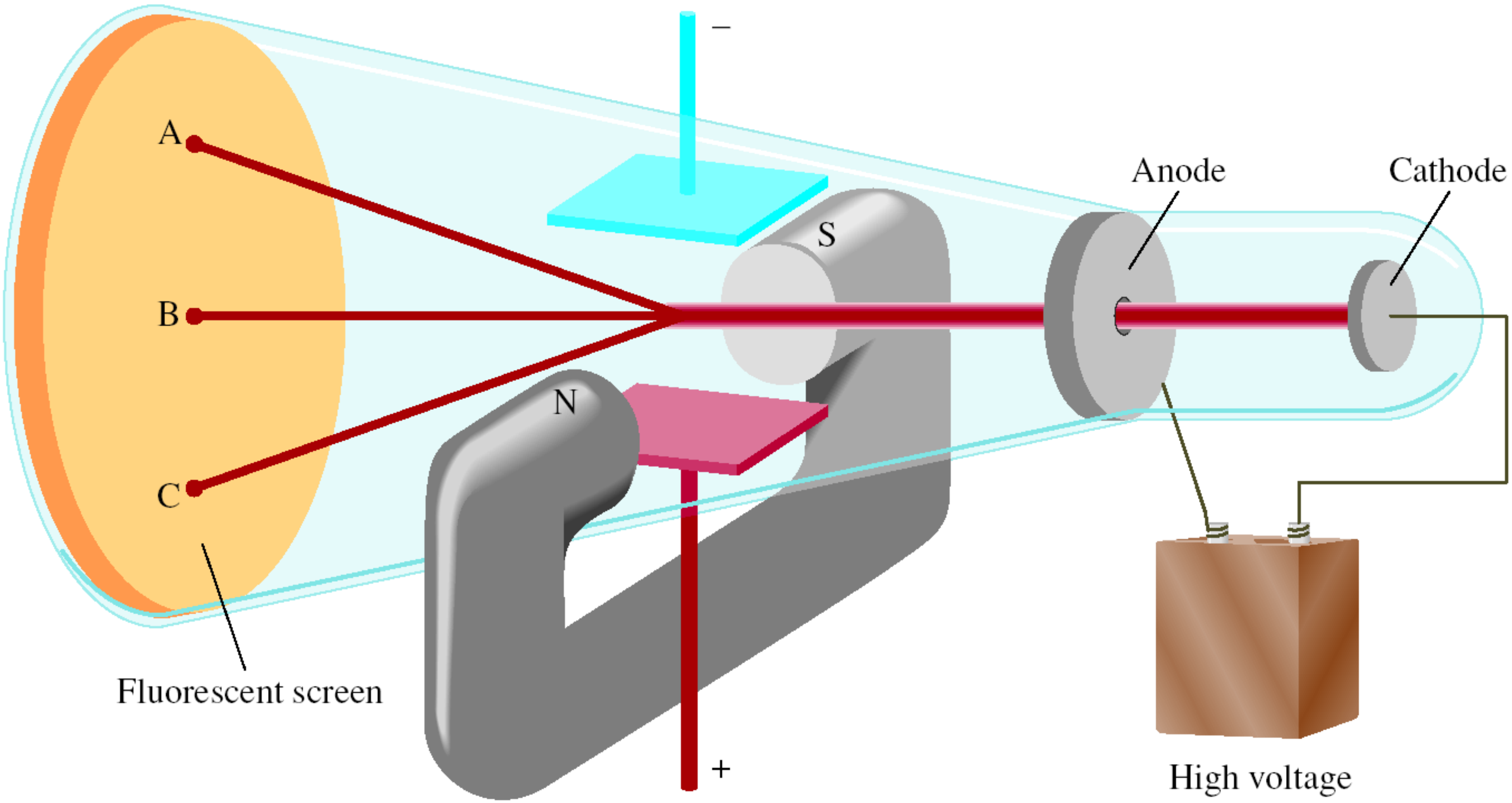
Law of Multiple Proportions



$$16\,X \quad + \quad 8\,Y \longrightarrow 8\,X_2Y$$

Law of Conservation of Mass

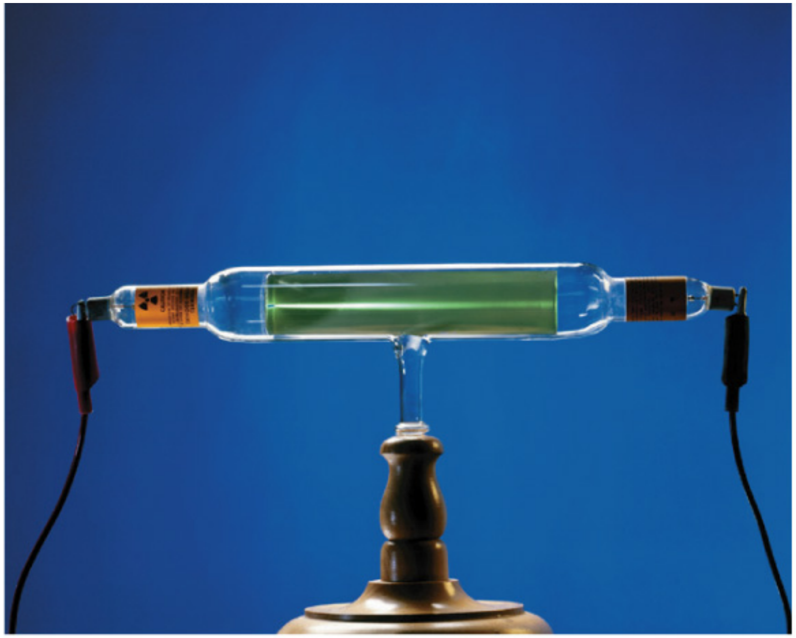
Cathode Ray Tube

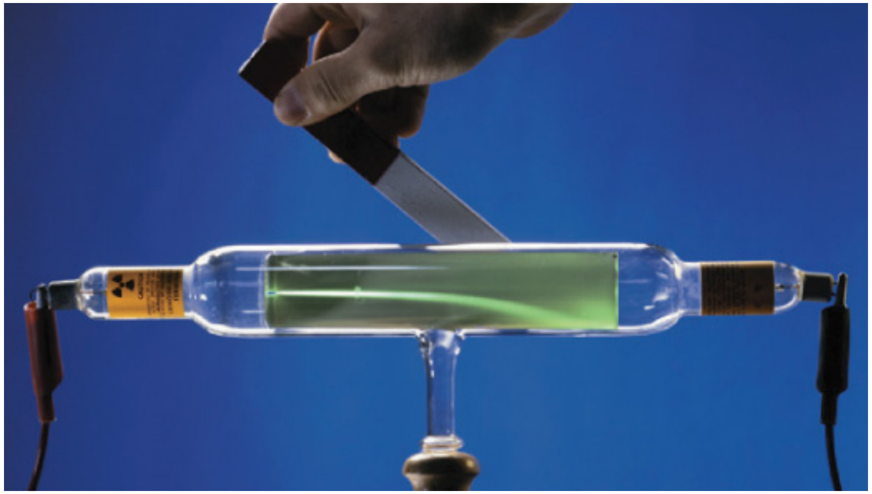


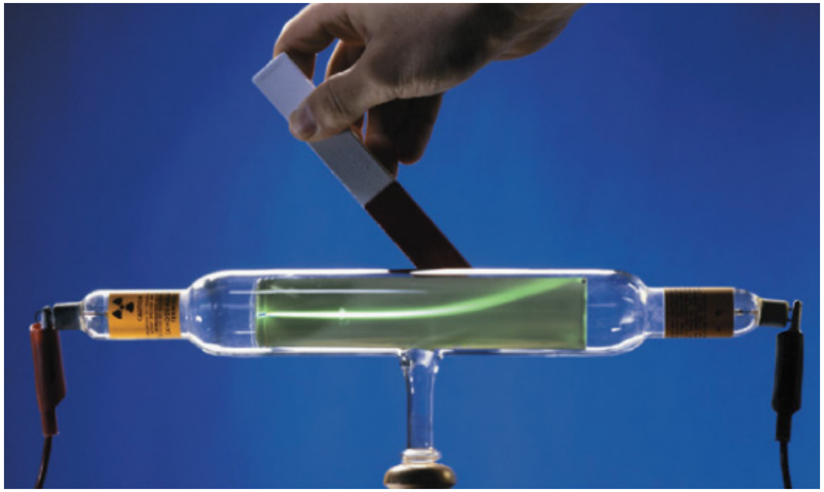
J.J. Thomson, measured mass/charge of e-

(1906 Nobel Prize in Physics)

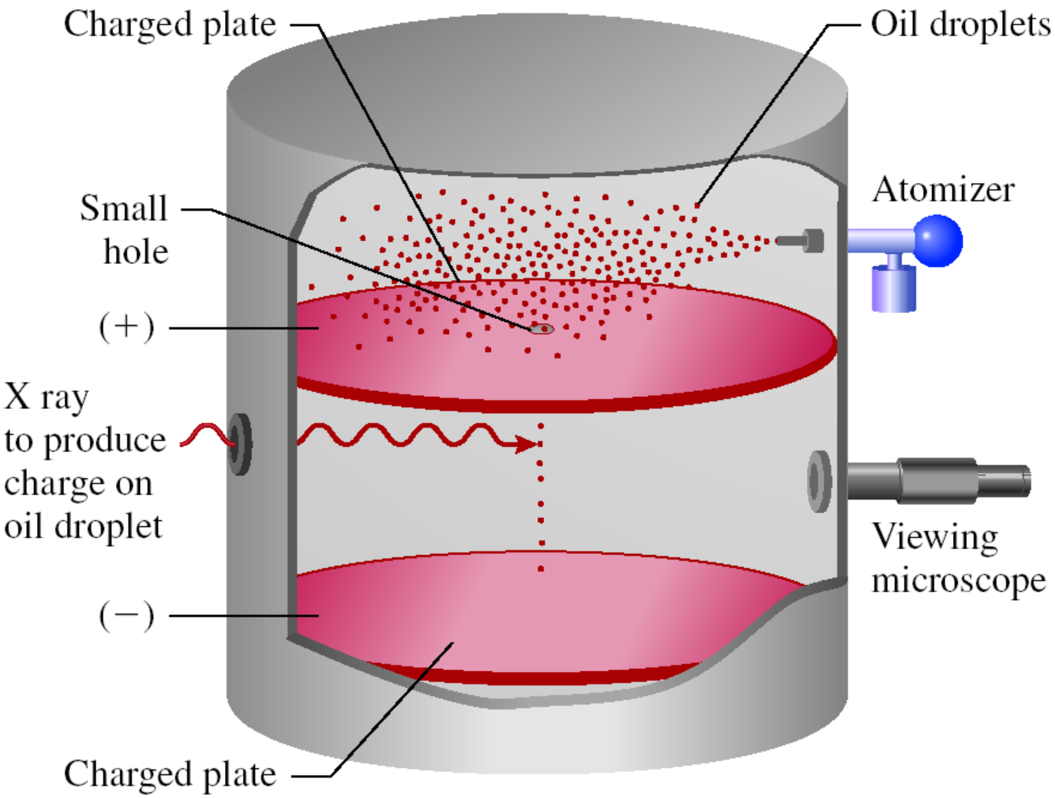
Cathode Ray Tube







Millikan's Experiment



Measured mass of e^-

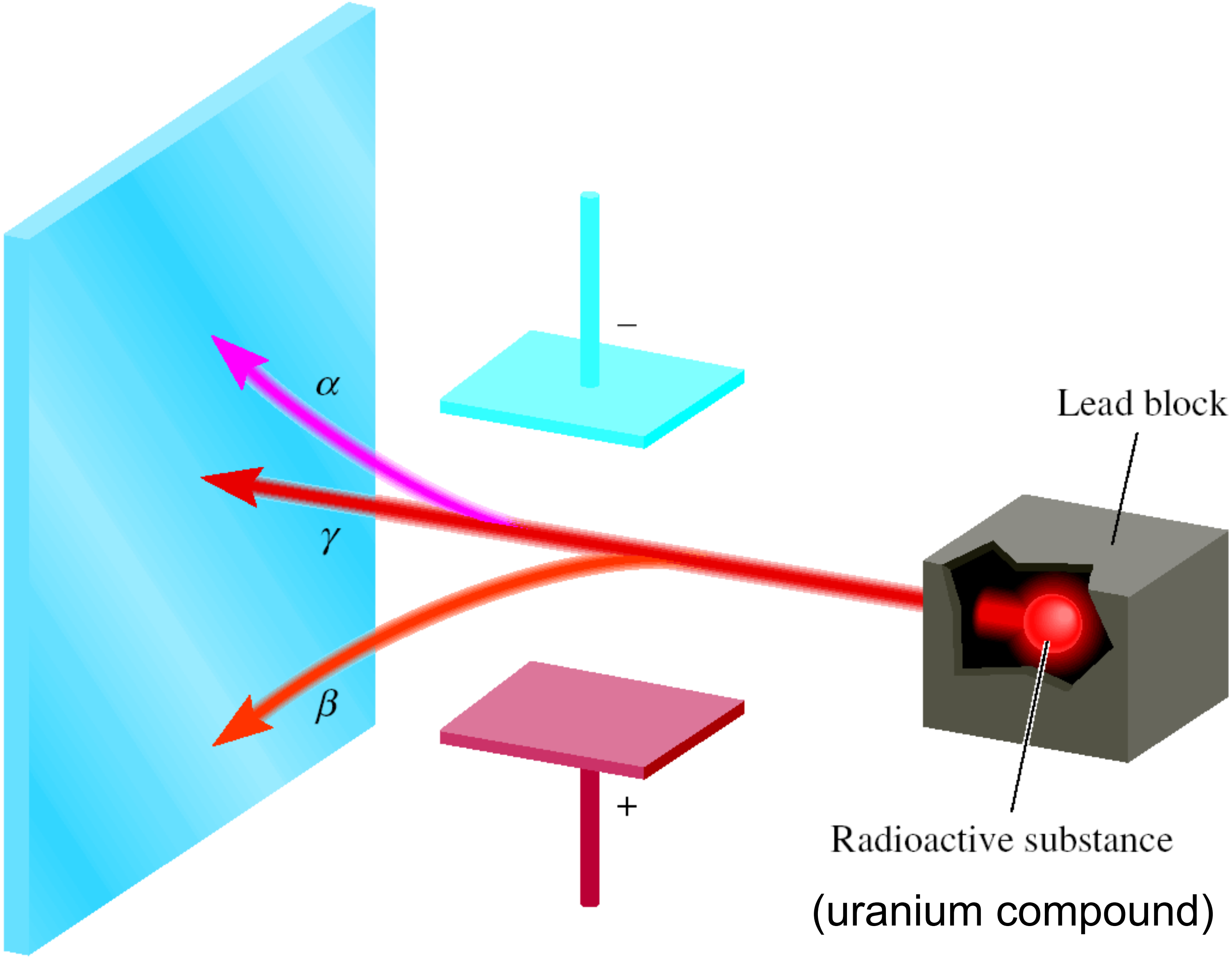
(1923 Nobel Prize in Physics)

e^- charge = -1.60×10^{-19} C

Thomson's charge/mass of e^- = -1.76×10^{8} C/g

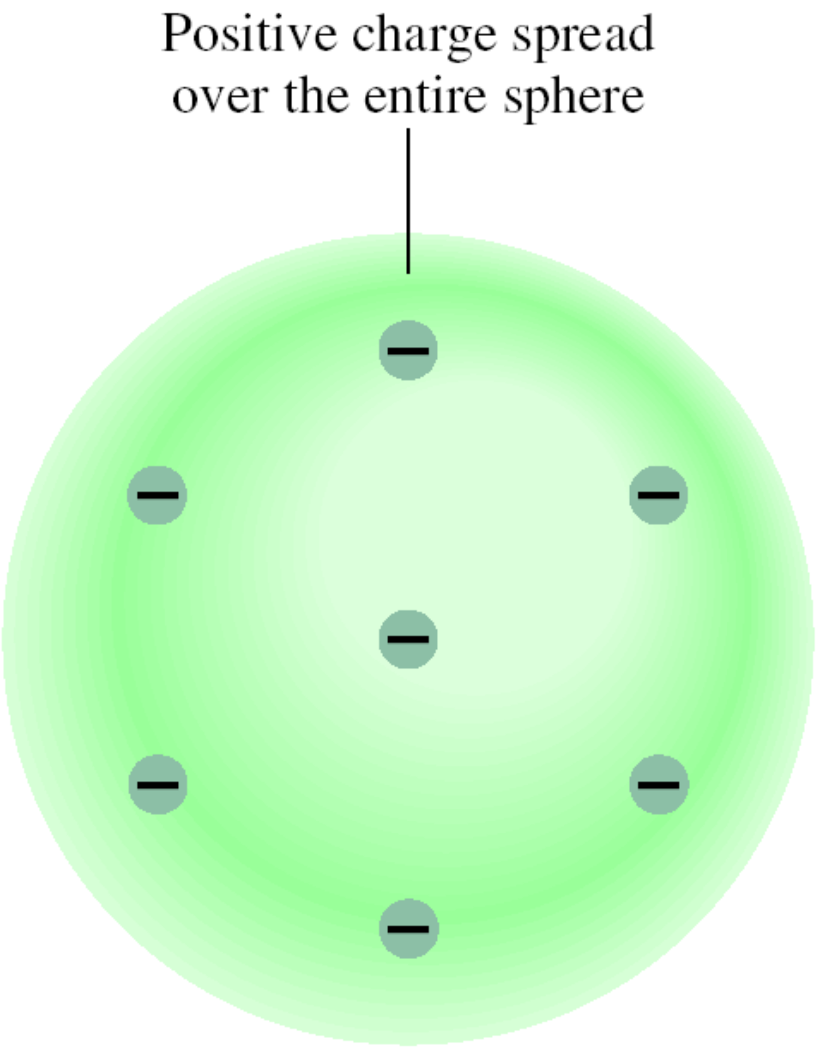
e^- mass = 9.10×10^{-28} g

Types of Radioactivity



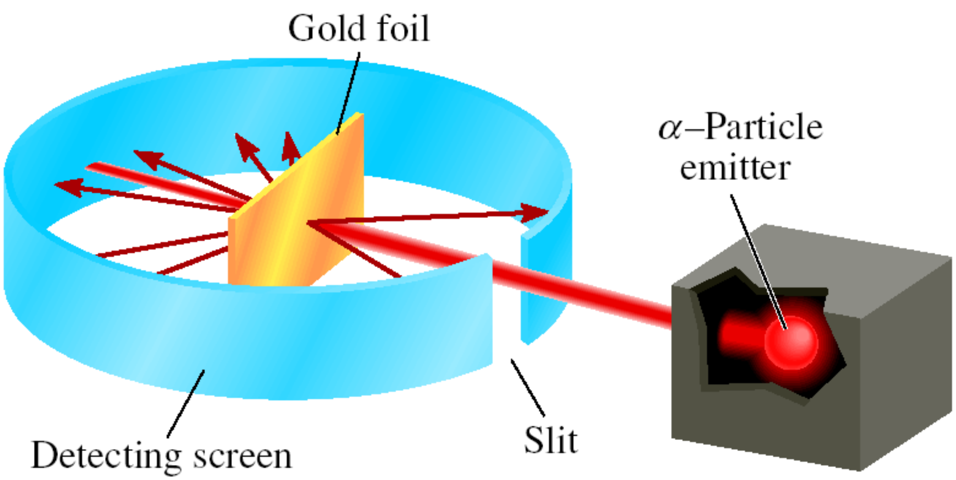
(uranium compound)

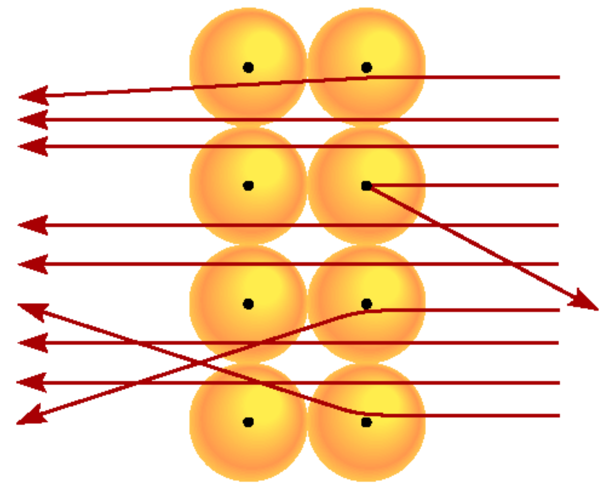
Thomson's Model



Rutherford's Experiment

(1908 Nobel Prize in Chemistry)

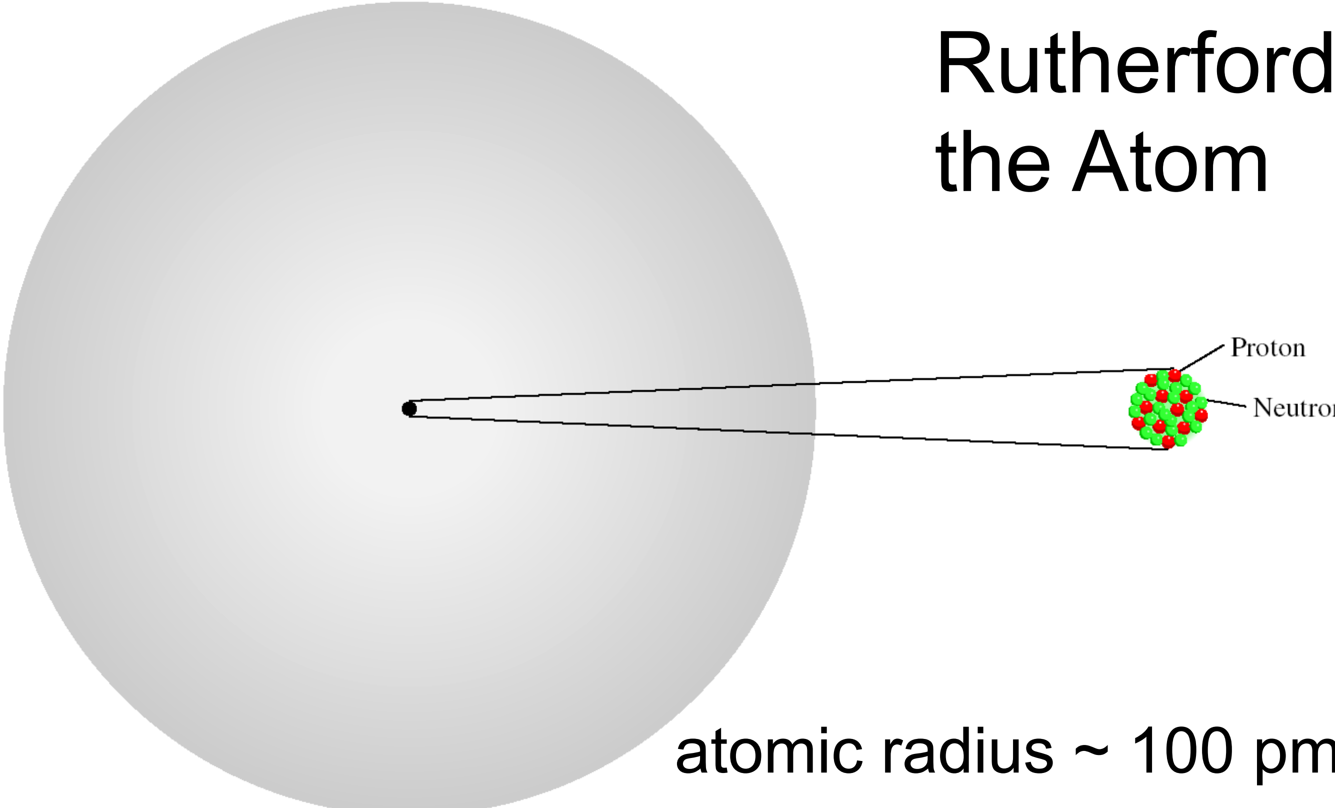




α particle velocity ~ 1.4 x 10^7 m/s
(~5% speed of light)

1. atoms positive charge is concentrated in the nucleus
2. proton (p) has opposite (+) charge of electron (-)
3. mass of p is 1840 x mass of e^- (1.67 x 10^{-24} g)

Rutherford's Model of the Atom



atomic radius ~ 100 pm = 1×10^{-10} m

nuclear radius ~ 5×10^{-3} pm = 5×10^{-15} m

"If the atom is the Houston Astrodome, then the nucleus is a marble on the 50-yard line."

Chadwick's Experiment (1932)
(1935 Noble Prize in Physics)

H atoms - 1 p; He atoms - 2 p

mass He/mass H should = 2

measured mass He/mass H = 4

$$\alpha + {}^{9}Be \longrightarrow {}^{1}n + {}^{12}C + \text{energy}$$

neutron (n) is neutral (charge = 0)

n mass ~ p mass = 1.67×10^{-24} g

TABLE 2.1 Mass and Charge of Subatomic Particles

		Charge	
Particle	Mass (g)	Coulomb	Charge Unit
Electron*	9.10938×10^{-28}	-1.6022×10^{-19}	-1
Proton	1.67262×10^{-24}	$+1.6022 \times 10^{-19}$	$+1$
Neutron	1.67493×10^{-24}	0	0

*More refined measurements have given us a more accurate value of an electron's mass than Millikan's.

mass p ≈ mass n ≈ 1840 x mass e^-

Atomic number, Mass number and Isotopes

Atomic number (Z) = number of protons in nucleus

Mass number (A) = number of protons + number of neutrons

= atomic number (Z) + number of neutrons

Isotopes are atoms of the same element (X) with different numbers of neutrons in their nuclei

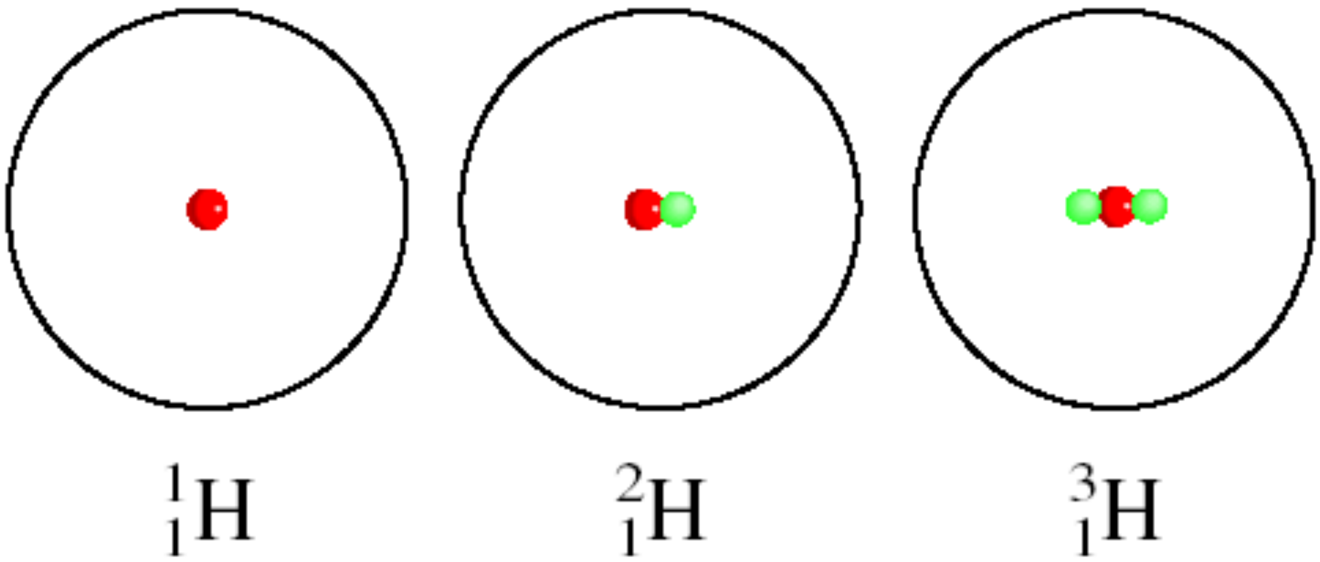
Mass Number → $^{A}_{Z}X$ ← Element Symbol

Atomic Number →

$^{1}_{1}H$ $^{2}_{1}H$ (D) $^{3}_{1}H$ (T)

$^{235}_{92}U$ $^{238}_{92}U$

The Isotopes of Hydrogen



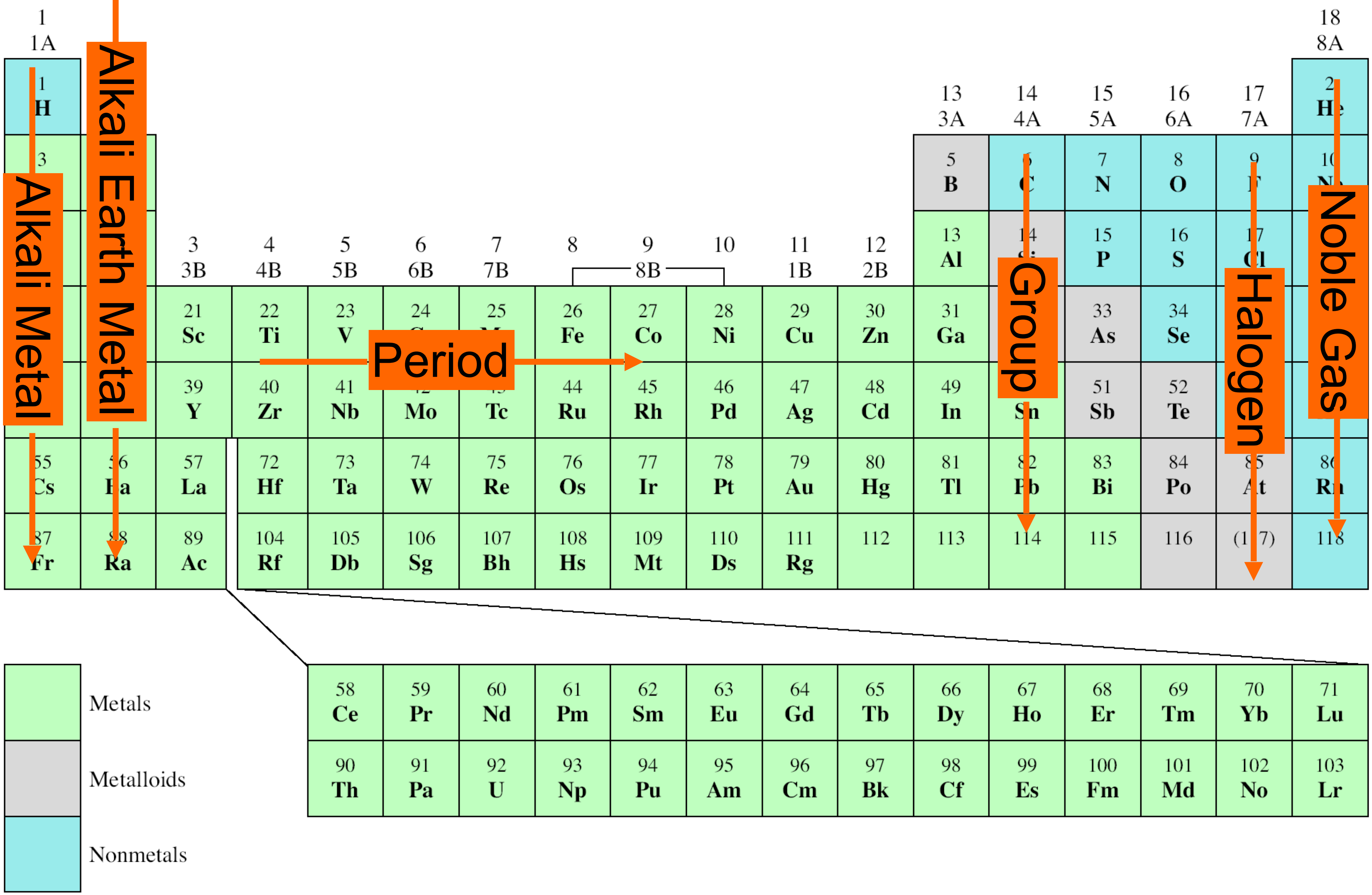
How many protons, neutrons, and electrons are in $^{14}_{6}C$?

6 protons, 8 (14 - 6) neutrons, 6 electrons

How many protons, neutrons, and electrons are in $^{11}_{6}C$?

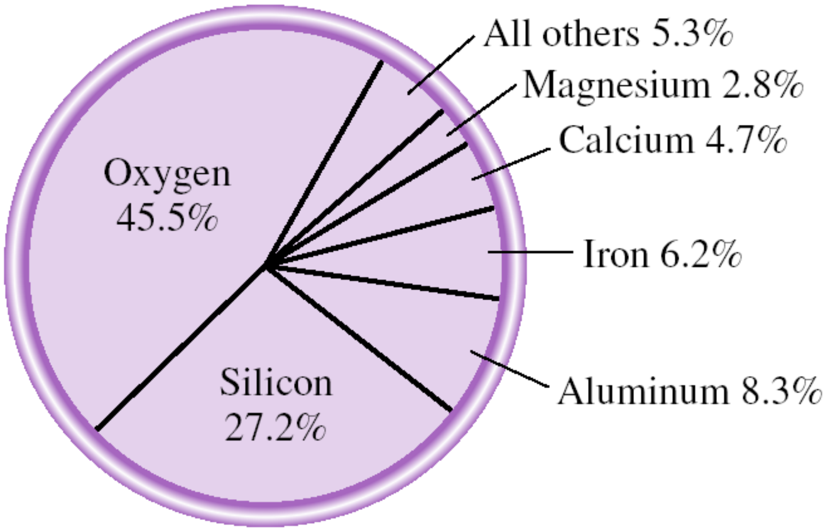
6 protons, 5 (11 - 6) neutrons, 6 electrons

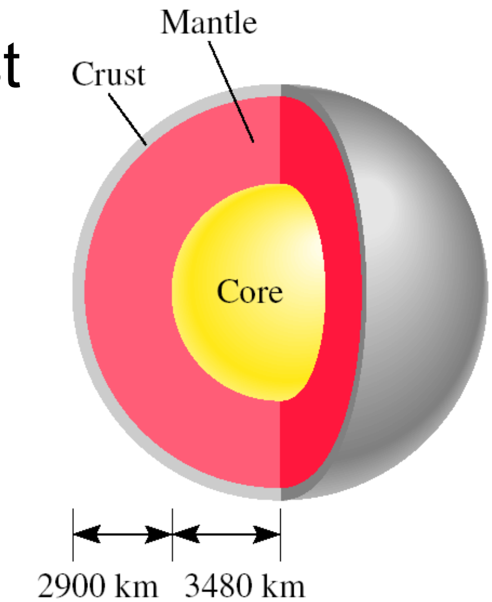
The Modern Periodic Table



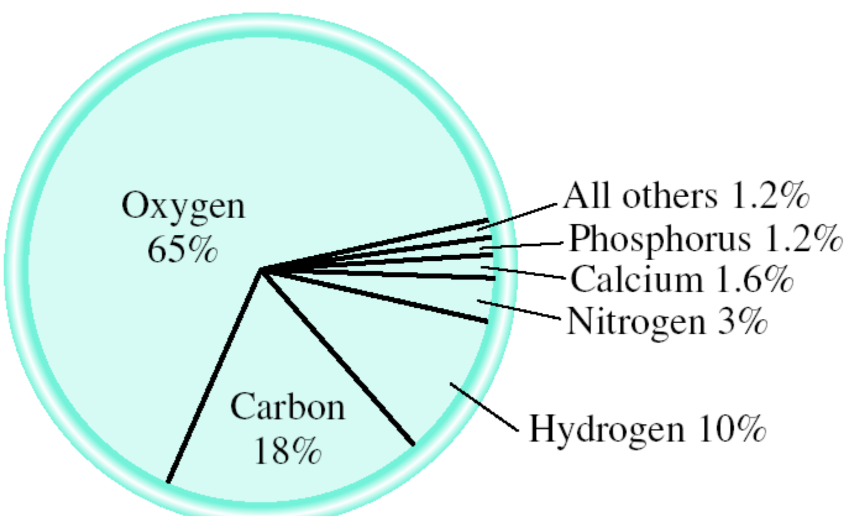
Chemistry In Action

Natural abundance of elements in Earth's crust

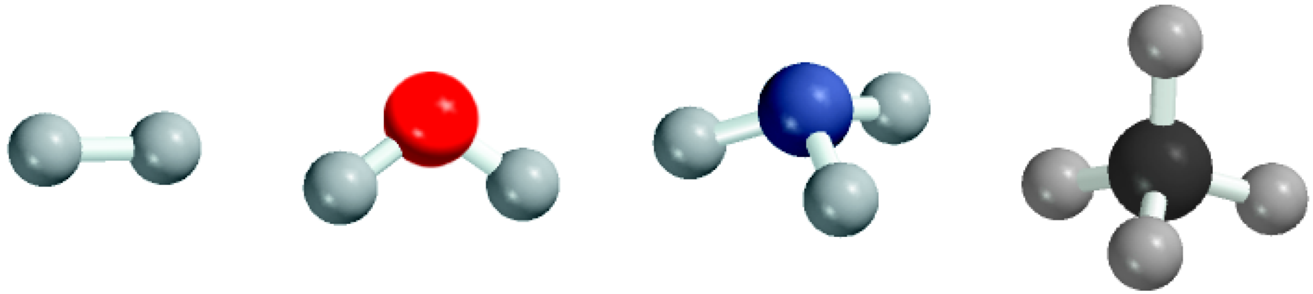




Natural abundance of elements in human body



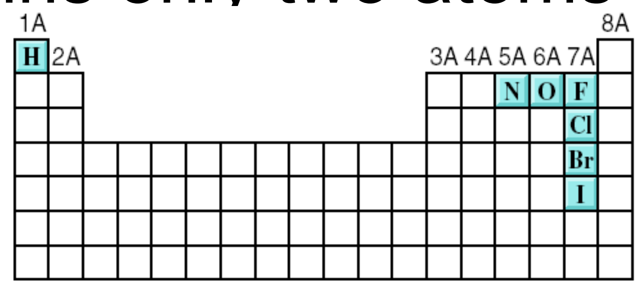
A ***molecule*** is an aggregate of two or more atoms in a definite arrangement held together by chemical forces



H_2 H_2O NH_3 CH_4

A ***diatomic molecule*** contains only two atoms

H_2, N_2, O_2, Br_2, HCl, CO



diatomic elements

A ***polyatomic molecule*** contains more than two atoms

O_3, H_2O, NH_3, CH_4

An ***ion*** is an atom, or group of atoms, that has a net positive or negative charge.

cation – ion with a positive charge

If a neutral atom **loses** one or more electrons it becomes a cation.

Na

11 protons
11 electrons

Na^+

11 protons
10 electrons

anion – ion with a negative charge

If a neutral atom **gains** one or more electrons it becomes an anion.

Cl

17 protons
17 electrons

Cl^-

17 protons
18 electrons

A ***monatomic ion*** contains only one atom

Na^{+}, Cl^{-}, Ca^{2+}, O^{2-}, Al^{3+}, N^{3-}

A ***polyatomic ion*** contains more than one atom

OH^{-}, CN^{-}, NH_4^{+}, NO_3^{-}

Common Ions Shown on the Periodic Table

1 1A	2 2A	3 3B	4 4B	5 5B	6 6B	7 7B	8	9 8B	10	11 1B	12 2B	13 3A	14 4A	15 5A	16 6A	17 7A	18 8A
Li^{+}													C^{4-}	N^{3-}	O^{2-}	F^{-}	
Na^{+}	Mg^{2+}											Al^{3+}		P^{3-}	S^{2-}	Cl^{-}	
K^{+}	Ca^{2+}				Cr^{2+} Cr^{3+}	Mn^{2+} Mn^{3+}	Fe^{2+} Fe^{3+}	Co^{2+} Co^{3+}	Ni^{2+} Ni^{3+}	Cu^{+} Cu^{2+}	Zn^{2+}				Se^{2-}	Br^{-}	
Rb^{+}	Sr^{2+}									Ag^{+}	Cd^{2+}		Sn^{2+} Sn^{4+}		Te^{2-}	I^{-}	
Cs^{+}	Ba^{2+}									Au^{+} Au^{3+}	Hg_2^{2+} Hg^{2+}		Pb^{2+} Pb^{4+}				

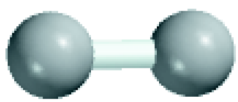
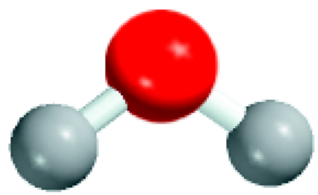
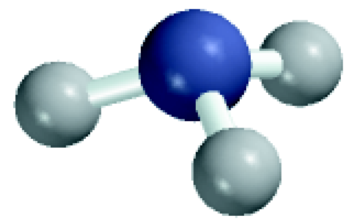
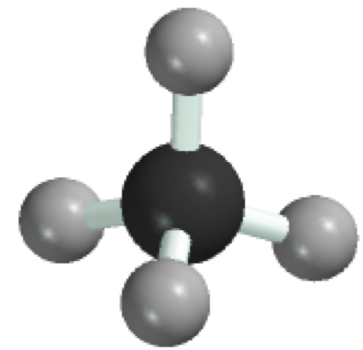
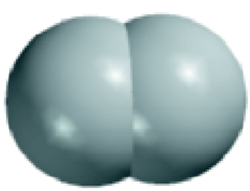
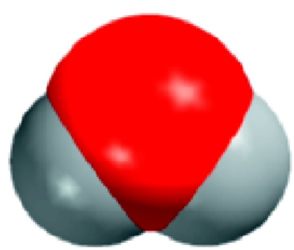
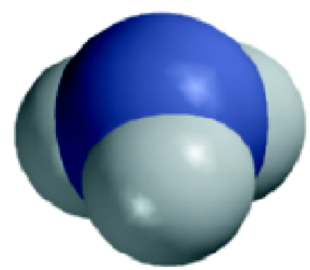
How many protons and electrons are in ${}^{27}_{13}Al^{3+}$?

13 protons, 10 (13 – 3) electrons

How many protons and electrons are in ${}^{78}_{34}Se^{2-}$?

34 protons, 36 (34 + 2) electrons

Formulas and Models

	Hydrogen	Water	Ammonia	Methane
Molecular formula	H_2	H_2O	NH_3	CH_4
Structural formula	$H-H$	$H-O-H$	$\begin{array}{c} H-N-H \\ \vert \\ H \end{array}$	$\begin{array}{c} H \\ \vert \\ H-C-H \\ \vert \\ H \end{array}$
Ball-and-stick model				
Space-filling model				

A ***molecular formula*** shows the exact number of atoms of each element in the smallest unit of a substance

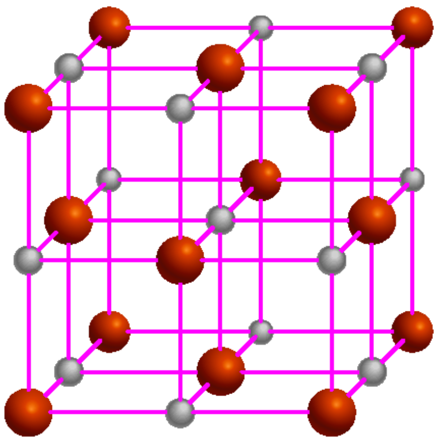
An ***empirical formula*** shows the simplest whole-number ratio of the atoms in a substance

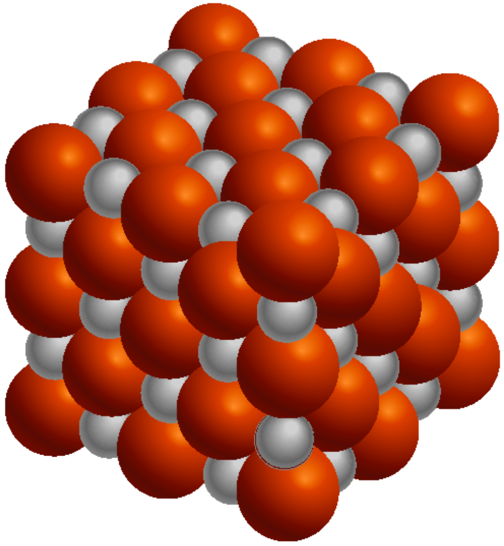
molecular	**empirical**
H_2O	H_2O
$C_6H_{12}O_6$	CH_2O
O_3	O
N_2H_4	NH_2

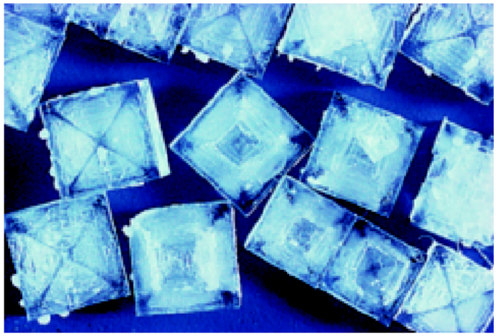
ionic compounds consist of a combination of cations and an anions

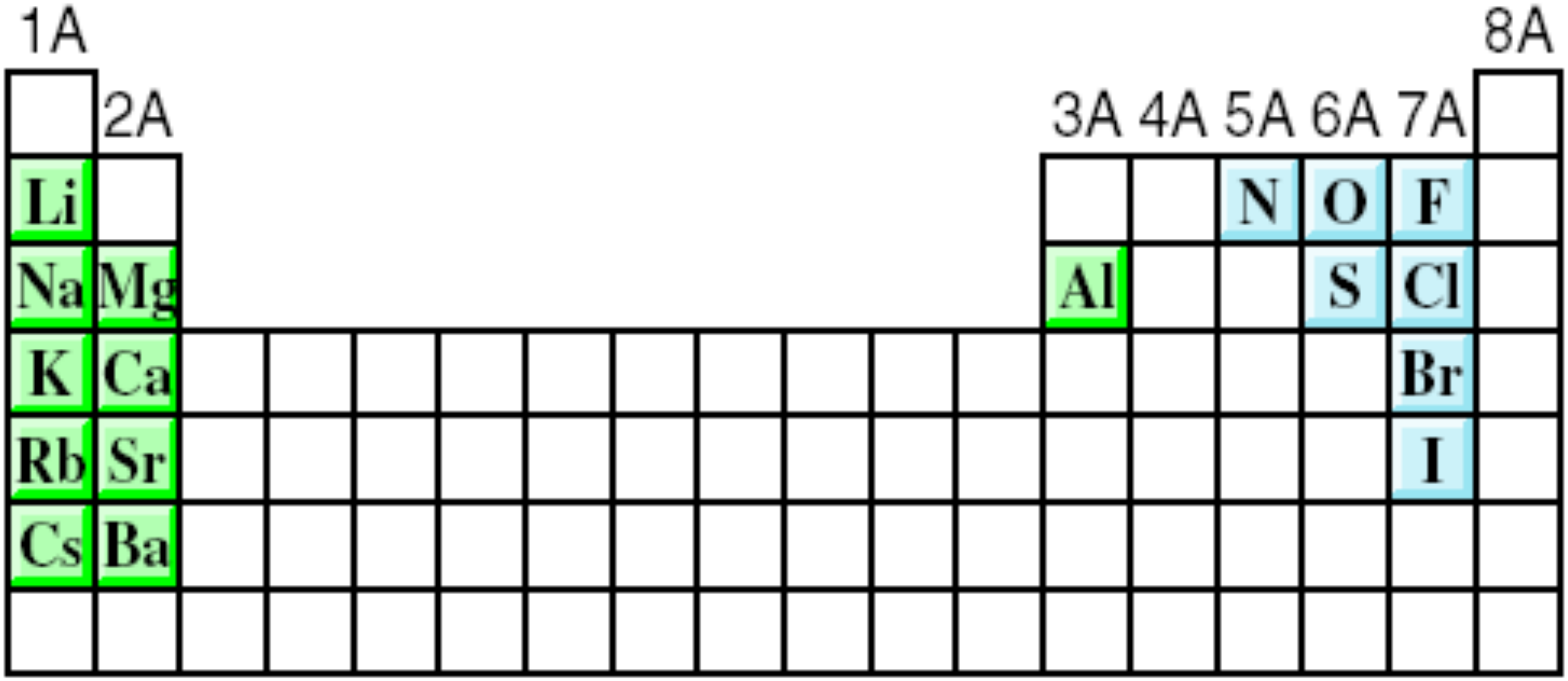
- The formula is usually the same as the empirical formula
- The sum of the charges on the cation(s) and anion(s) in each formula unit must equal zero

The ionic compound NaCl









The most reactive **metals** (green) and the most reactive **nonmetals** (blue) combine to form ionic compounds.

Formula of Ionic Compounds

$2 \times +3 = +6$ $3 \times -2 = -6$

Al_2O_3

Al^{3+} O^{2-}

$1 \times +2 = +2$ $2 \times -1 = -2$

$CaBr_2$

Ca^{2+} Br^{-}

$1 \times +2 = +2$ $1 \times -2 = -2$

Na_2CO_3

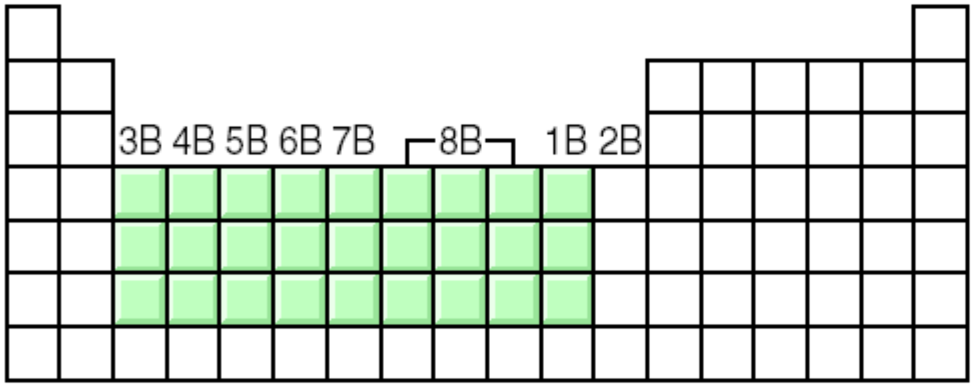
Na^{+} CO_3^{2-}

Chemical Nomenclature

- **Ionic Compounds**
  - Often a metal + nonmetal
  - Anion (nonmetal), add “ide” to element name

$BaCl_2$	barium chloride
K_2O	potassium oxide
$Mg(OH)_2$	magnesium hydroxide
KNO_3	potassium nitrate

- Transition metal ionic compounds
  - indicate charge on metal with Roman numerals



$FeCl_2$	2 Cl^- -2 so Fe is +2	iron(II) chloride
$FeCl_3$	3 Cl^- -3 so Fe is +3	iron(III) chloride
Cr_2S_3	3 S^{-2} -6 so Cr is +3 (6/2)	chromium(III) sulfide

TABLE 2.2 The "-ide" Nomenclature of Some Common Monatomic Anions According to Their Positions in the Periodic Table

Group 4A	Group 5A	Group 6A	Group 7A
C carbide (C^{4-})*	N nitride (N^{3-})	O oxide (O^{2-})	F fluoride (F^-)
Si silicide (Si^{4-})	P phosphide (P^{3-})	S sulfide (S^{2-})	Cl chloride (Cl^-)
		Se selenide (Se^{2-})	Br bromide (Br^-)
		Te telluride (Te^{2-})	I iodide (I^-)

*The word "carbide" is also used for the anion C_2^{2-}.

TABLE 2.3 Names and Formulas of Some Common Inorganic Cations and Anions

Cation	Anion
aluminum (Al^{3+})	bromide (Br^-)
ammonium (NH_4^+)	carbonate (CO_3^{2-})
barium (Ba^{2+})	chlorate (ClO_3^-)
cadmium (Cd^{2+})	chloride (Cl^-)
calcium (Ca^{2+})	chromate (CrO_4^{2-})
cesium (Cs^+)	cyanide (CN^-)
chromium(III) or chromic (Cr^{3+})	dichromate ($Cr_2O_7^{2-}$)
cobalt(II) or cobaltous (Co^{2+})	dihydrogen phosphate ($H_2PO_4^-$)
copper(I) or cuprous (Cu^+)	fluoride (F^-)
copper(II) or cupric (Cu^{2+})	hydride (H^-)
hydrogen (H^+)	hydrogen carbonate or bicarbonate (HCO_3^-)
iron(II) or ferrous (Fe^{2+})	hydrogen phosphate (HPO_4^{2-})
iron(III) or ferric (Fe^{3+})	hydrogen sulfate or bisulfate (HSO_4^-)
lead(II) or plumbous (Pb^{2+})	hydroxide (OH^-)
lithium (Li^+)	iodide (I^-)
magnesium (Mg^{2+})	nitrate (NO_3^-)
manganese(II) or manganous (Mn^{2+})	nitride (N^{3-})
mercury(I) or mercurous (Hg_2^{2+})*	nitrite (NO_2^-)
mercury(II) or mercuric (Hg^{2+})	oxide (O^{2-})
potassium (K^+)	permanganate (MnO_4^-)
rubidium (Rb^+)	peroxide (O_2^{2-})
silver (Ag^+)	phosphate (PO_4^{3-})
sodium (Na^+)	sulfate (SO_4^{2-})
strontium (Sr^{2+})	sulfide (S^{2-})
tin(II) or stannous (Sn^{2+})	sulfite (SO_3^{2-})
zinc (Zn^{2+})	thiocyanate (SCN^-)

*Mercury(I) exists as a pair as shown.

- **Molecular compounds**
  - Nonmetals or nonmetals + metalloids
  - Common names
    - H_2O, NH_3, CH_4,
  - Element furthest to the left in a period and closest to the bottom of a group on periodic table is placed first in formula
  - If more than one compound can be formed from the same elements, use prefixes to indicate number of each kind of atom
  - Last element name ends in *ide*

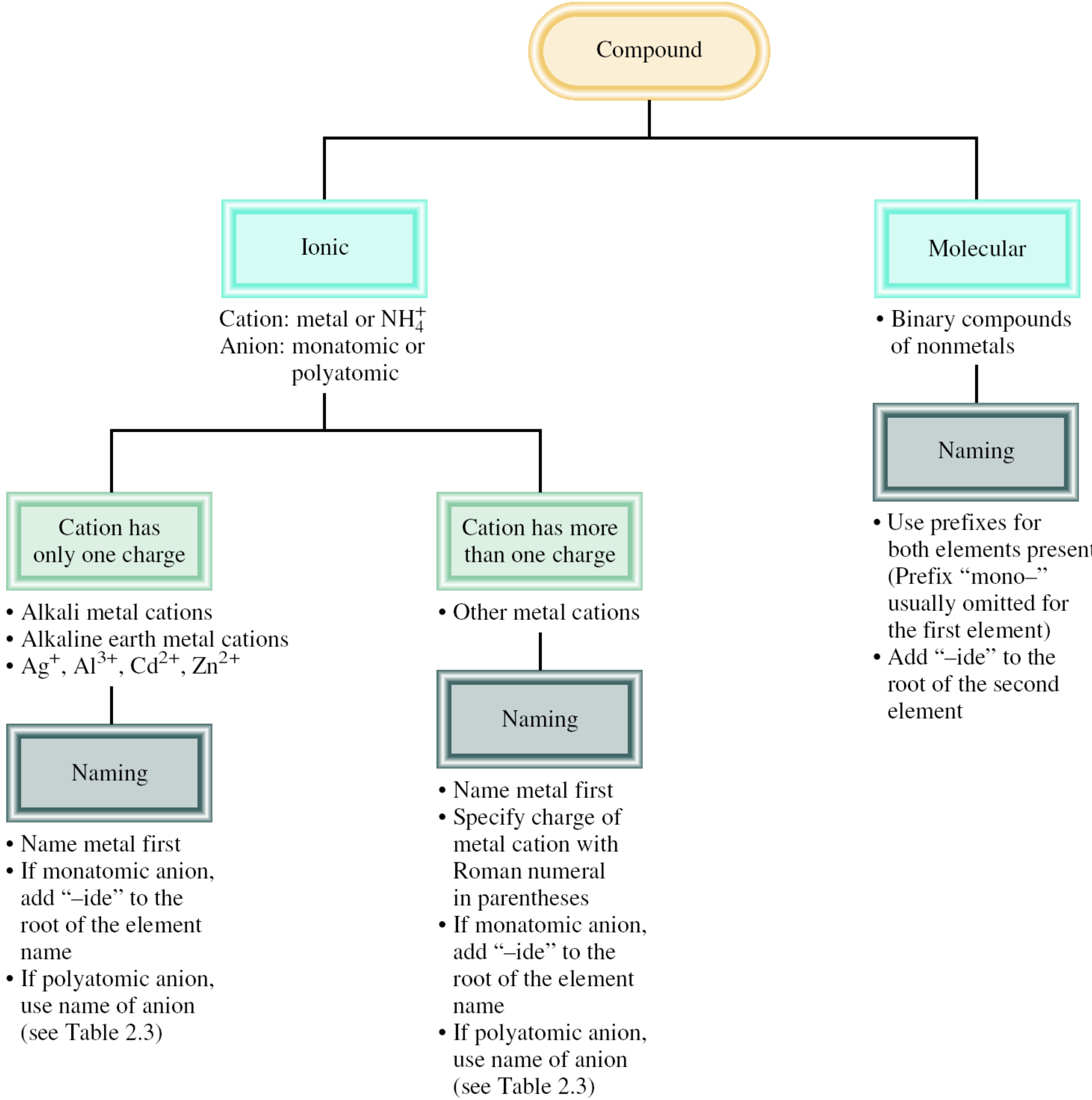
TABLE 2.4

Greek Prefixes Used in Naming Molecular Compounds

Prefix	Meaning
mono-	1
di-	2
tri-	3
tetra-	4
penta-	5
hexa-	6
hepta-	7
octa-	8
nona-	9
deca-	10

Molecular Compounds

HI	hydrogen iodide
NF_3	nitrogen trifluoride
SO_2	sulfur dioxide
N_2Cl_4	dinitrogen tetrachloride
NO_2	nitrogen dioxide
N_2O	dinitrogen monoxide


Compound
Ionic
Cation: metal or NH_4^+
Anion: monatomic or polyatomic
Cation has only one charge
• Alkali metal cations
• Alkaline earth metal cations
• Ag^+, Al^{3+}, Cd^{2+}, Zn^{2+}
Naming
• Name metal first
• If monatomic anion, add "–ide" to the root of the element name
• If polyatomic anion, use name of anion (see Table 2.3)
Cation has more than one charge
• Other metal cations
Naming
• Name metal first
• Specify charge of metal cation with Roman numeral in parentheses
• If monatomic anion, add "–ide" to the root of the element name
• If polyatomic anion, use name of anion (see Table 2.3)
Molecular
• Binary compounds of nonmetals
Naming
• Use prefixes for both elements present (Prefix "mono–" usually omitted for the first element)
• Add "–ide" to the root of the second element

An ***acid*** can be defined as a substance that yields hydrogen ions (H^+) when dissolved in water.

For example: HCl gas and HCl in water

- Pure substance, hydrogen chloride

- Dissolved in water (H_3O^+ and Cl^-), hydrochloric acid

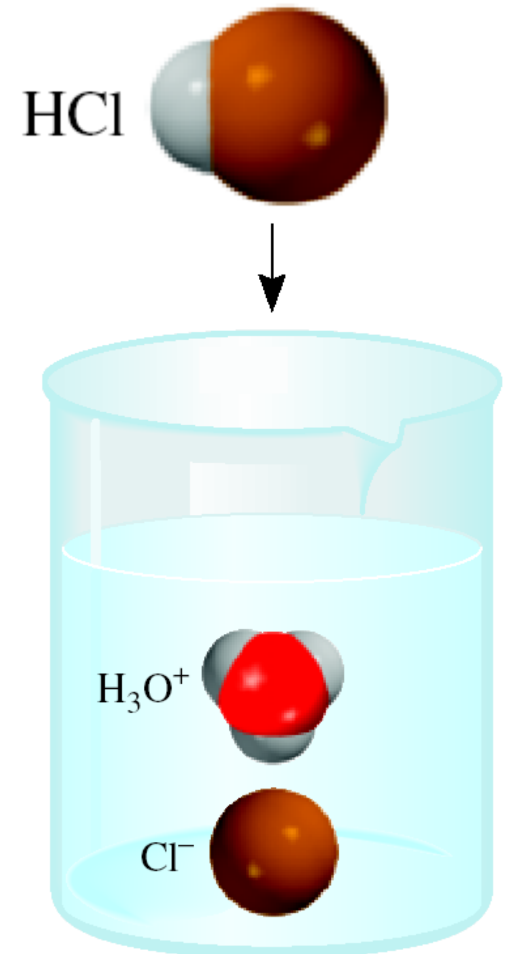
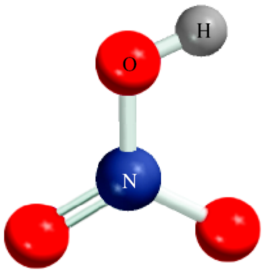


TABLE 2.5 Some Simple Acids

Anion	Corresponding Acid
F^- (fluoride)	HF (hydrofluoric acid)
Cl^- (chloride)	HCl (hydrochloric acid)
Br^- (bromide)	HBr (hydrobromic acid)
I^- (iodide)	HI (hydroiodic acid)
CN^- (cyanide)	HCN (hydrocyanic acid)
S^{2-} (sulfide)	H_2S (hydrosulfuric acid)

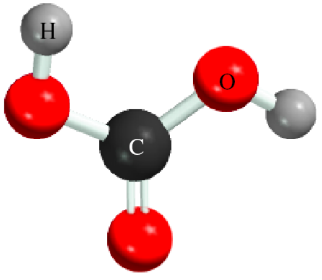
An ***oxoacid*** is an acid that contains hydrogen, oxygen, and another element.

HNO_3 nitric acid



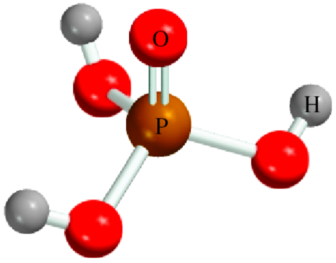
HNO_3

H_2CO_3 carbonic acid



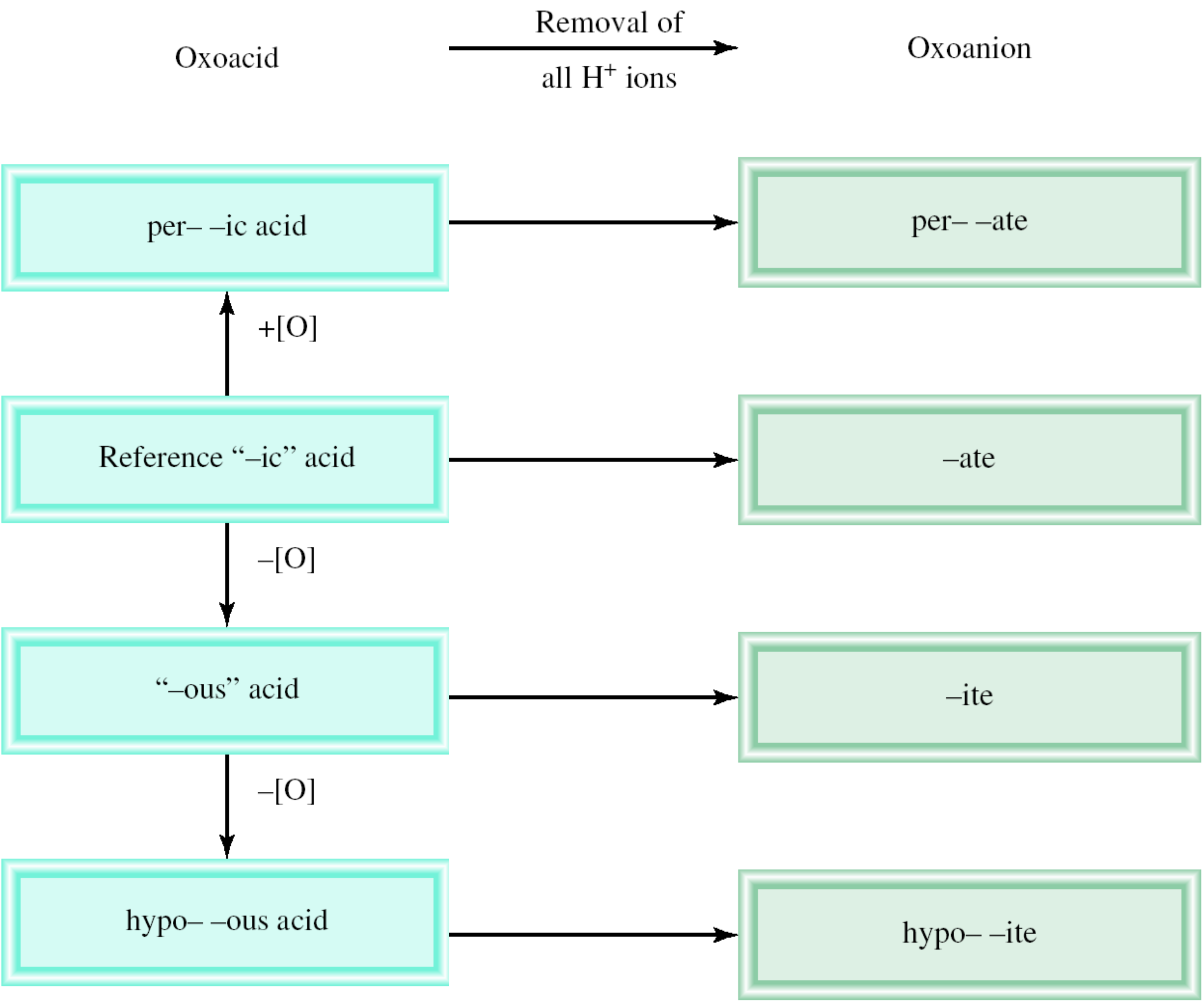
H_2CO_3

H_3PO_4 phosphoric acid



H_3PO_4

Naming Oxoacids and Oxoanions



The rules for naming ***oxoanions,*** *anions of oxoacids,* are as follows:

1. When all the H ions are removed from the "-ic" acid, the anion's name ends with "-ate."
2. When all the H ions are removed from the "-ous" acid, the anion's name ends with "-ite."
3. The names of anions in which one or more but not all the hydrogen ions have been removed must indicate the number of H ions present.

For example:

– $H_2PO_4^-$ dihydrogen phosphate
– HPO_4^{2-} hydrogen phosphate
– PO_4^{3-} phosphate

TABLE 2.6 Names of Oxoacids and Oxoanions That Contain Chlorine

Acid	Anion
$HClO_4$ (perchloric acid)	ClO_4^- (perchlorate)
$HClO_3$ (chloric acid)	ClO_3^- (chlorate)
$HClO_2$ (chlorous acid)	ClO_2^- (chlorite)
$HClO$ (hypochlorous acid)	ClO^- (hypochlorite)

A ***base*** can be defined as a substance that yields hydroxide ions (OH^-) when dissolved in water.

$NaOH$	sodium hydroxide
KOH	potassium hydroxide
$Ba(OH)_2$	barium hydroxide

Hydrates are compounds that have a specific number of water molecules attached to them.

$BaCl_2 \cdot 2H_2O$	barium chloride dihydrate
$LiCl \cdot H_2O$	lithium chloride monohydrate
$MgSO_4 \cdot 7H_2O$	magnesium sulfate heptahydrate
$Sr(NO_3)_2 \cdot 4H_2O$	strontium nitrate tetrahydrate

$CuSO_4 \cdot 5H_2O$ →

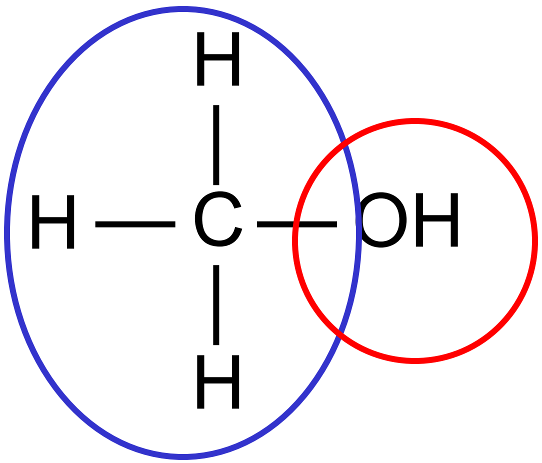
← $CuSO_4$

TABLE 2.7 Common and Systematic Names of Some Compounds

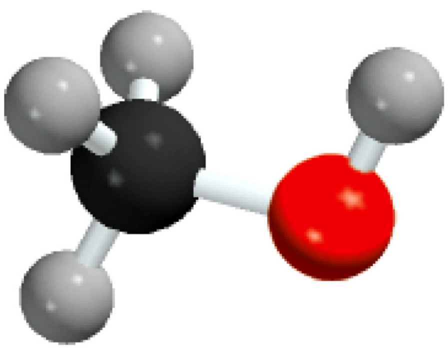
Formula	Common Name	Systematic Name
H_2O	Water	Dihydrogen monoxide
NH_3	Ammonia	Trihydrogen nitride
CO_2	Dry ice	Solid carbon dioxide
$NaCl$	Table salt	Sodium chloride
N_2O	Laughing gas	Dinitrogen monoxide
$CaCO_3$	Marble, chalk, limestone	Calcium carbonate
CaO	Quicklime	Calcium oxide
$Ca(OH)_2$	Slaked lime	Calcium hydroxide
$NaHCO_3$	Baking soda	Sodium hydrogen carbonate
$Na_2CO_3 \cdot 10H_2O$	Washing soda	Sodium carbonate decahydrate
$MgSO_4 \cdot 7H_2O$	Epsom salt	Magnesium sulfate heptahydrate
$Mg(OH)_2$	Milk of magnesia	Magnesium hydroxide
$CaSO_4 \cdot 2H_2O$	Gypsum	Calcium sulfate dihydrate

Organic chemistry is the branch of chemistry that deals with carbon compounds

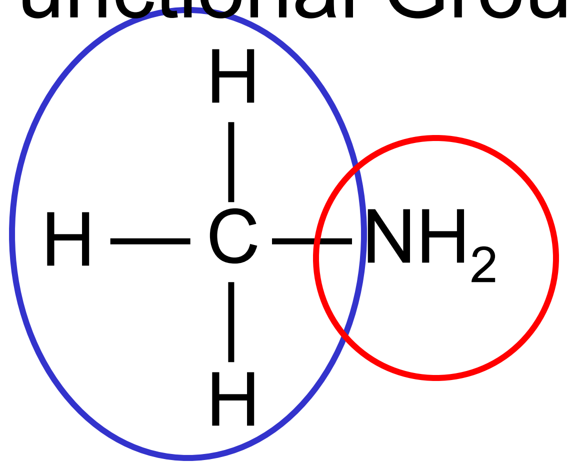
Functional Groups



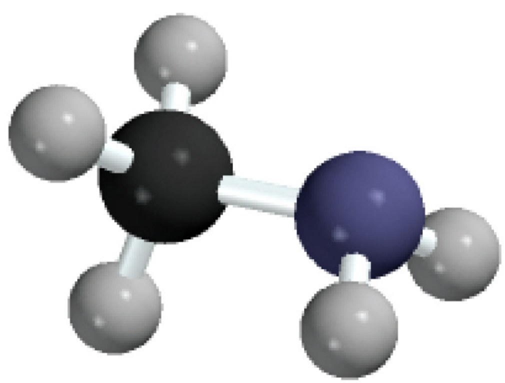
methanol



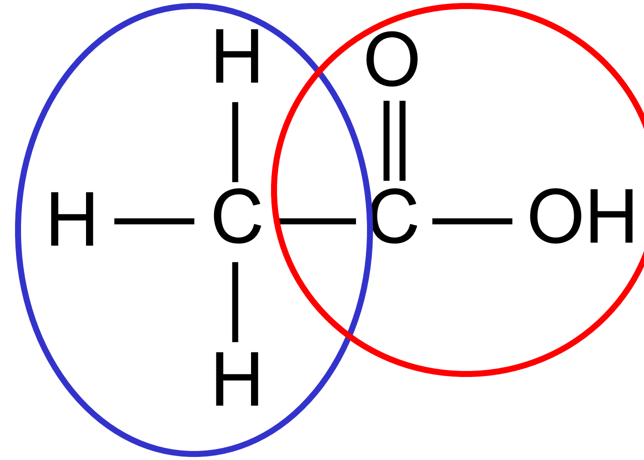
CH_3OH



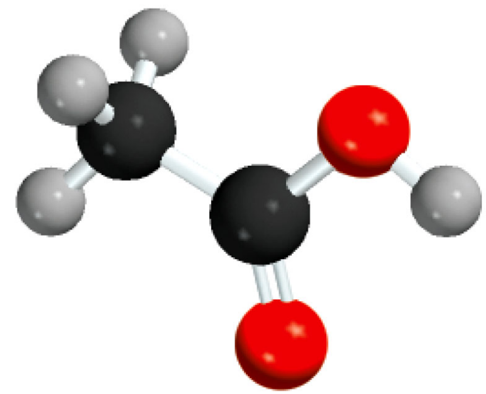
methylamine



CH_3NH_2

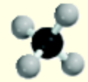
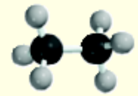
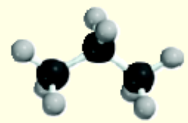
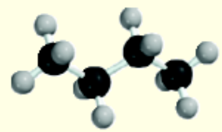
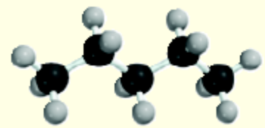
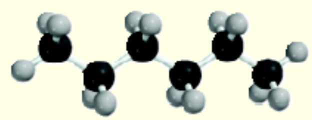
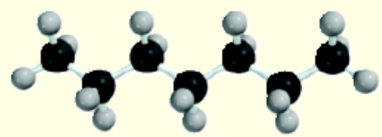
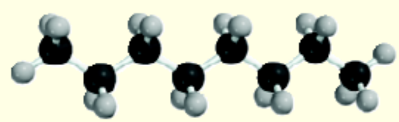
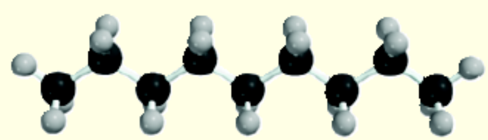


acetic acid



CH_3COOH

TABLE 2.8 The First Ten Straight-Chain Alkanes

Name	Formula	Molecular Model
Methane	CH_4	
Ethane	C_2H_6	
Propane	C_3H_8	
Butane	C_4H_{10}	
Pentane	C_5H_{12}	
Hexane	C_6H_{14}	
Heptane	C_7H_{16}	
Octane	C_8H_{18}	
Nonane	C_9H_{20}	
Decane	$C_{10}H_{22}$