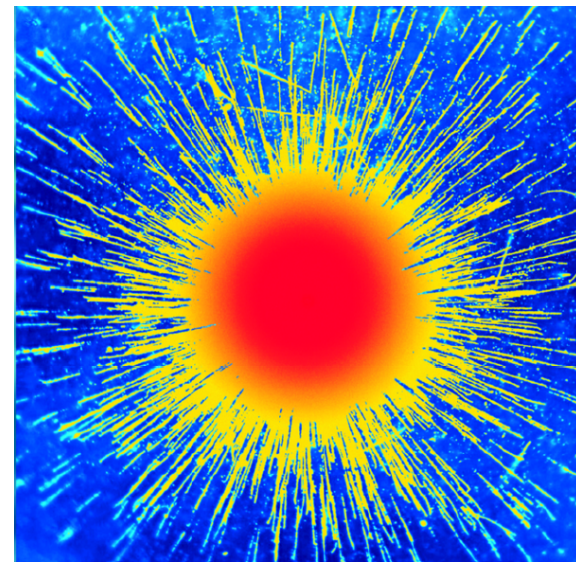


Atoms, Molecules and Ions

Chapter 2

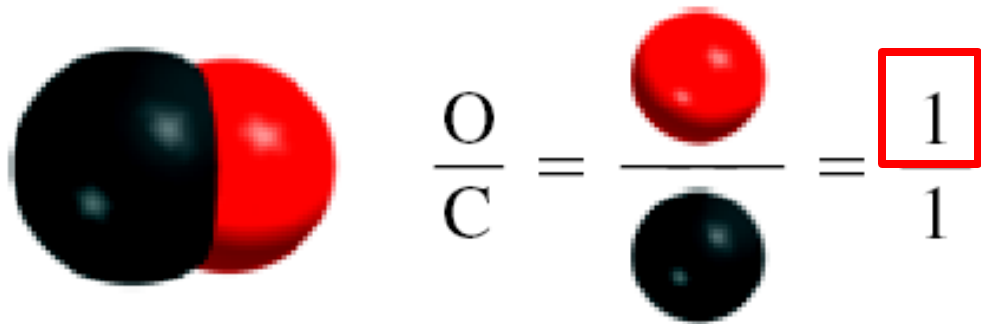


Dalton's Atomic Theory (1808)

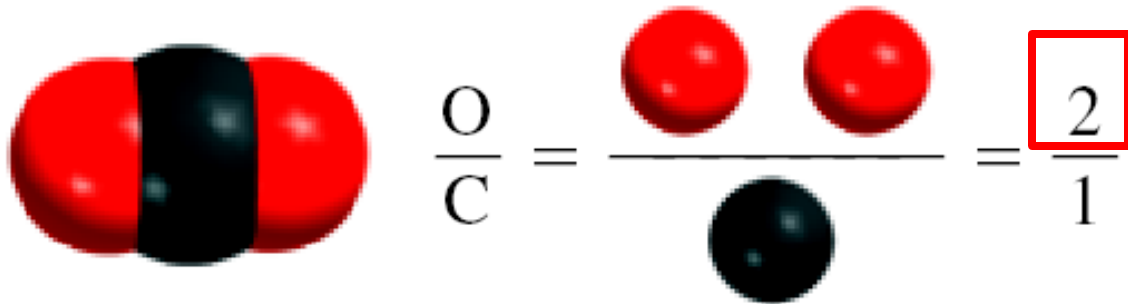
1. Elements are composed of extremely small particles called **atoms**.
2. All **atoms** of a given element are identical, having the same size, mass and chemical properties. The atoms of one element are different from the atoms of all other elements.
3. **Compounds** are composed of atoms of more than 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

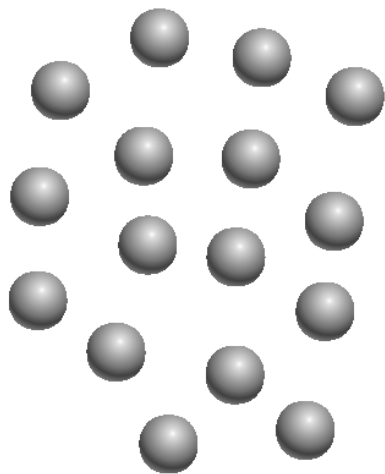
Carbon monoxide



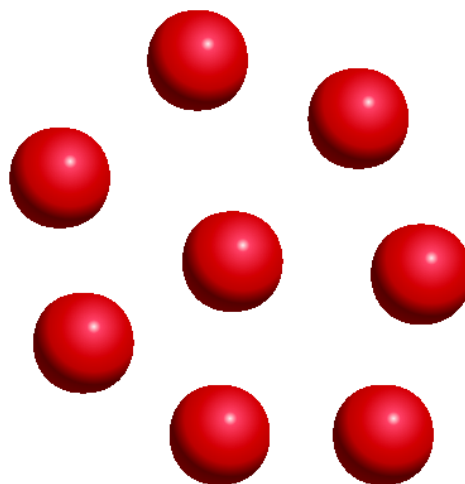
Carbon dioxide



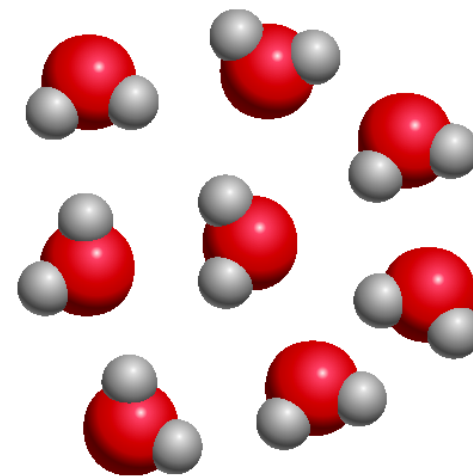
Law of Multiple Proportions



Atoms of element X



Atoms of element Y

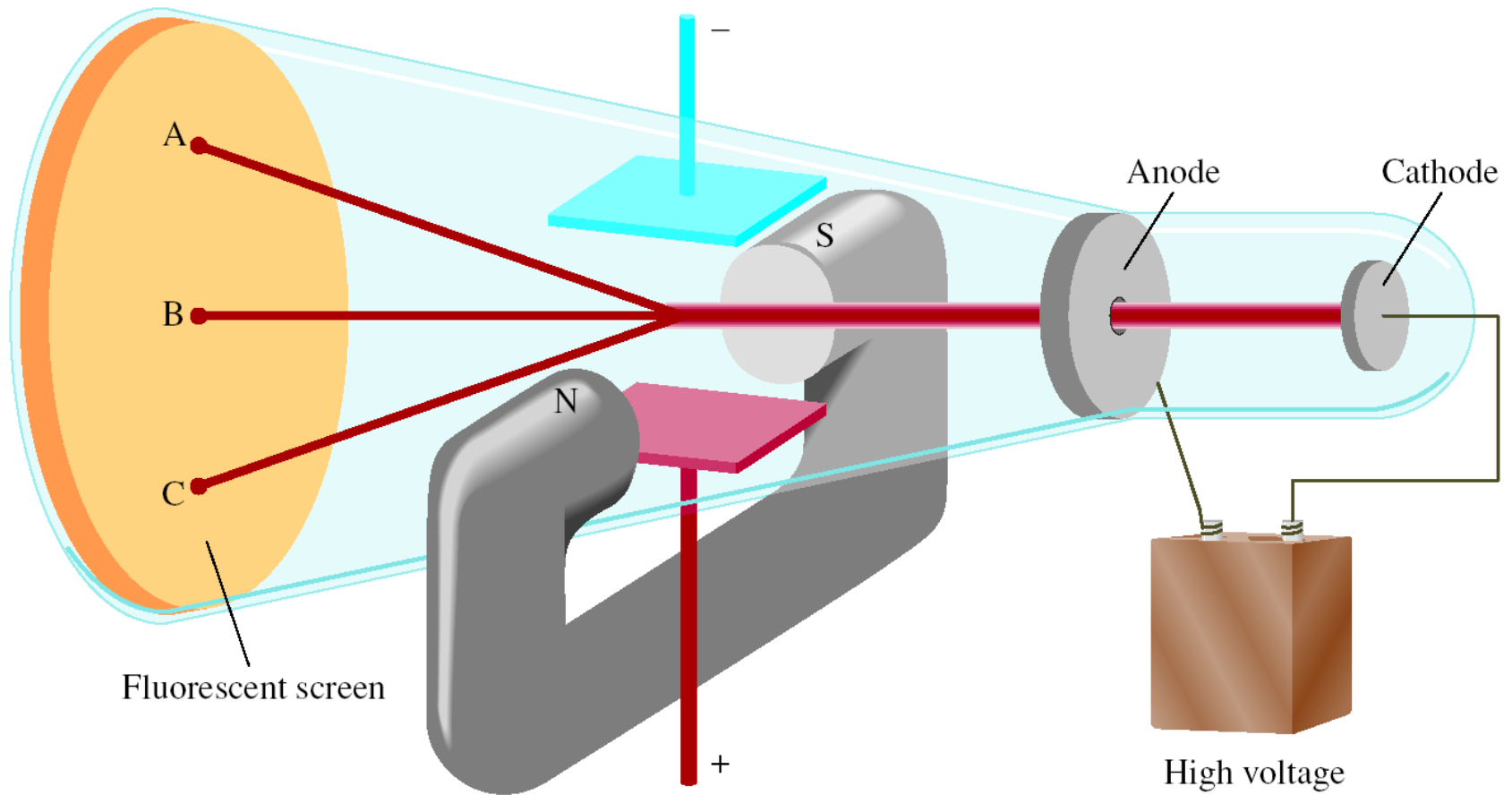


Compounds of elements X and Y



Law of Conservation of Mass

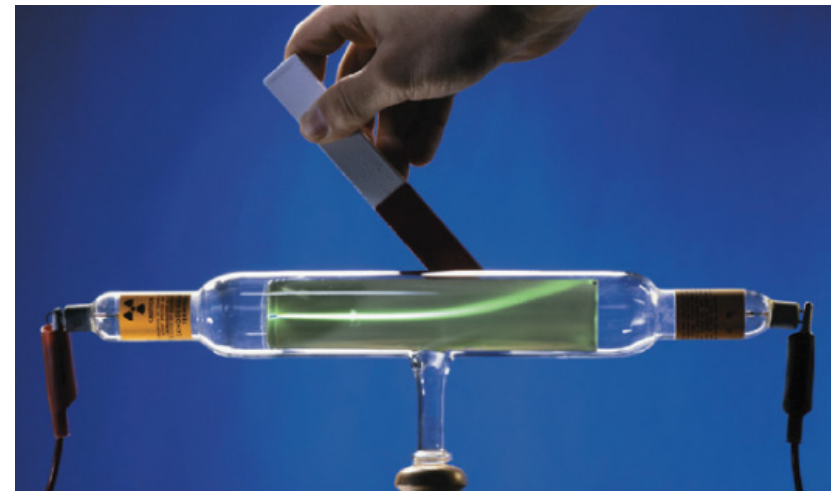
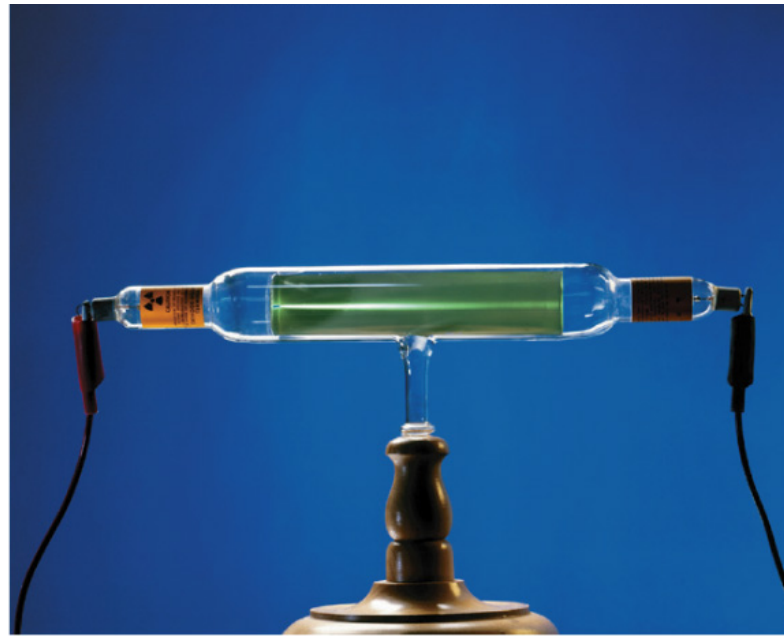
Cathode Ray Tube



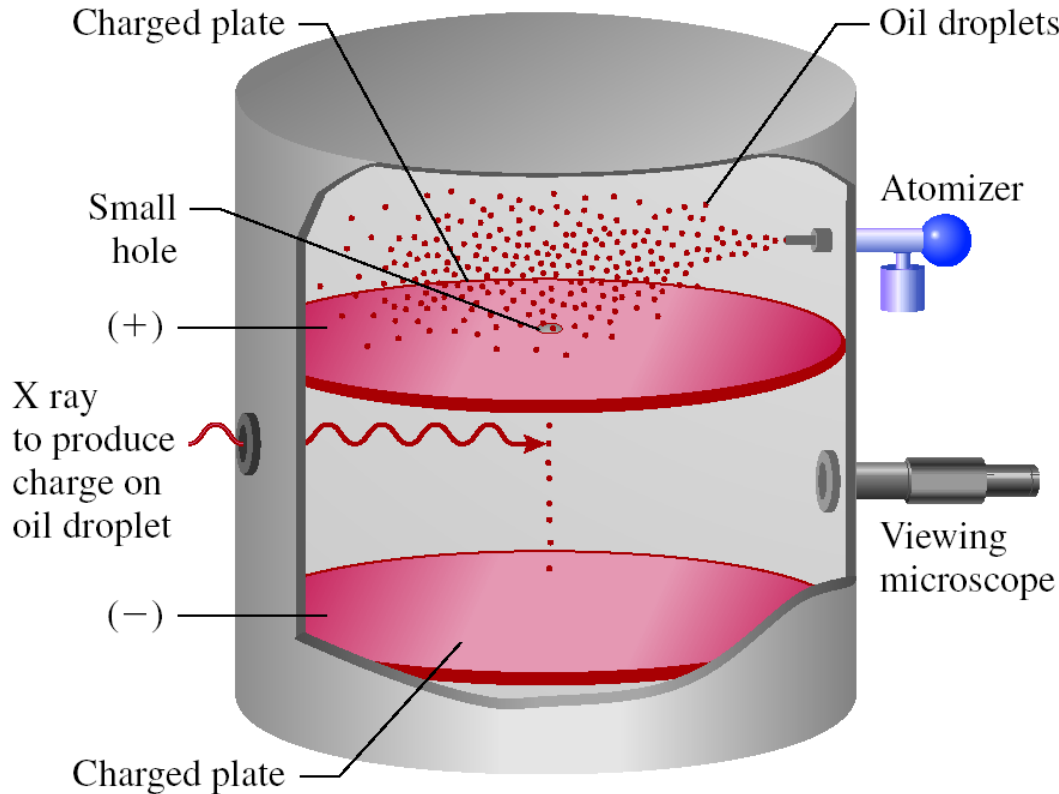
J.J. Thomson, **measured mass/charge of e^-**

(1906 Nobel Prize in Physics) 5

Cathode Ray Tube



Millikan's Experiment



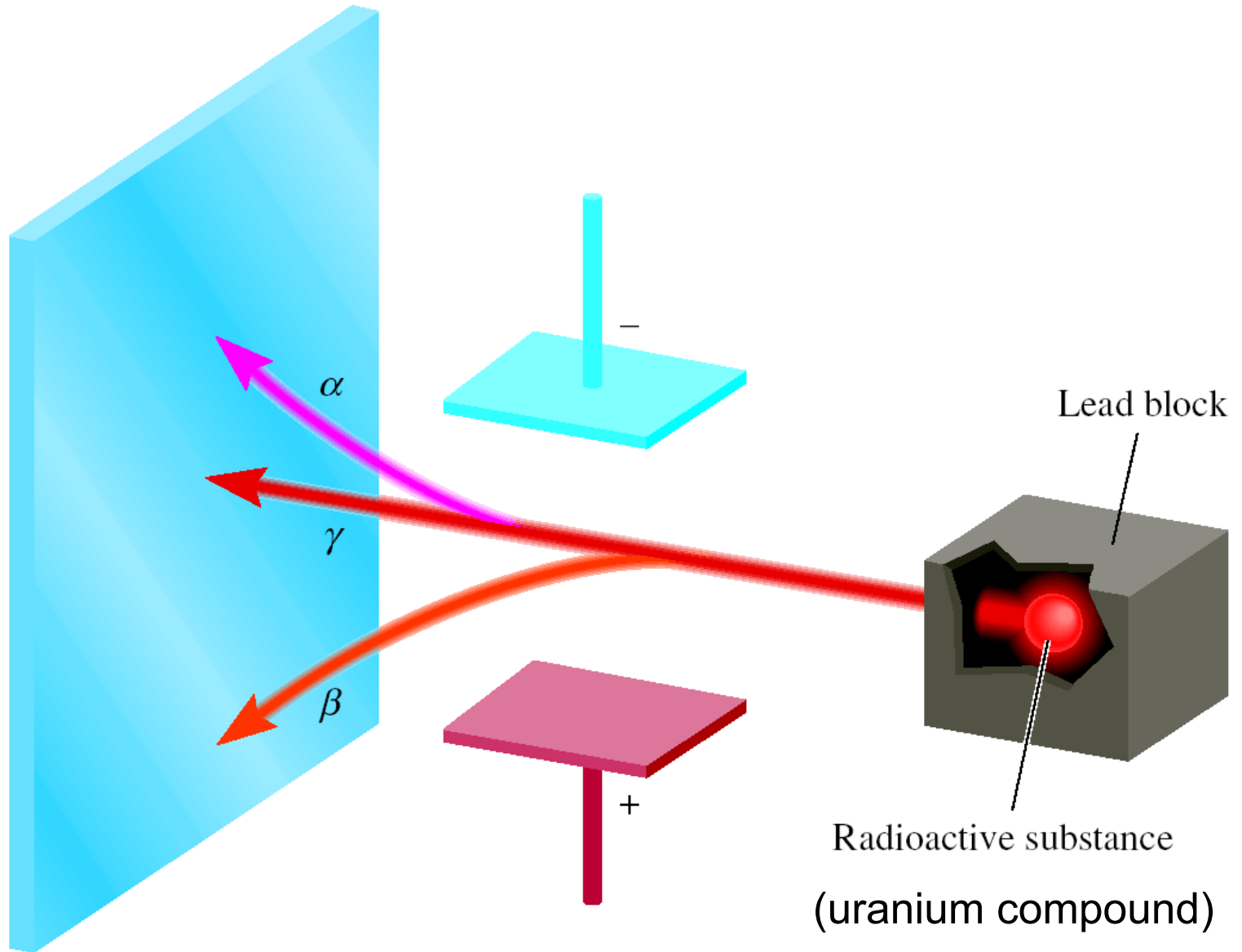
Measured mass of e^-
(1923 Nobel Prize in Physics)

$$e^- \text{ charge} = -1.60 \times 10^{-19} \text{ C}$$

$$\text{Thomson's charge/mass of } e^- = -1.76 \times 10^8 \text{ C/g}$$

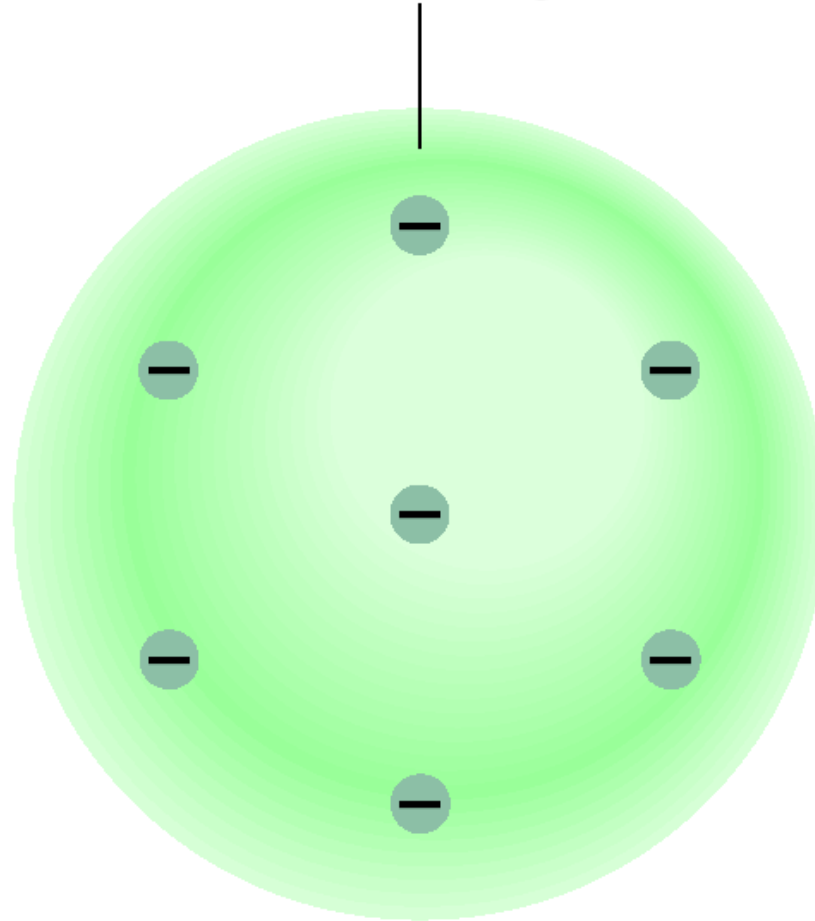
$$e^- \text{ mass} = 9.10 \times 10^{-28} \text{ g}$$

Types of Radioactivity



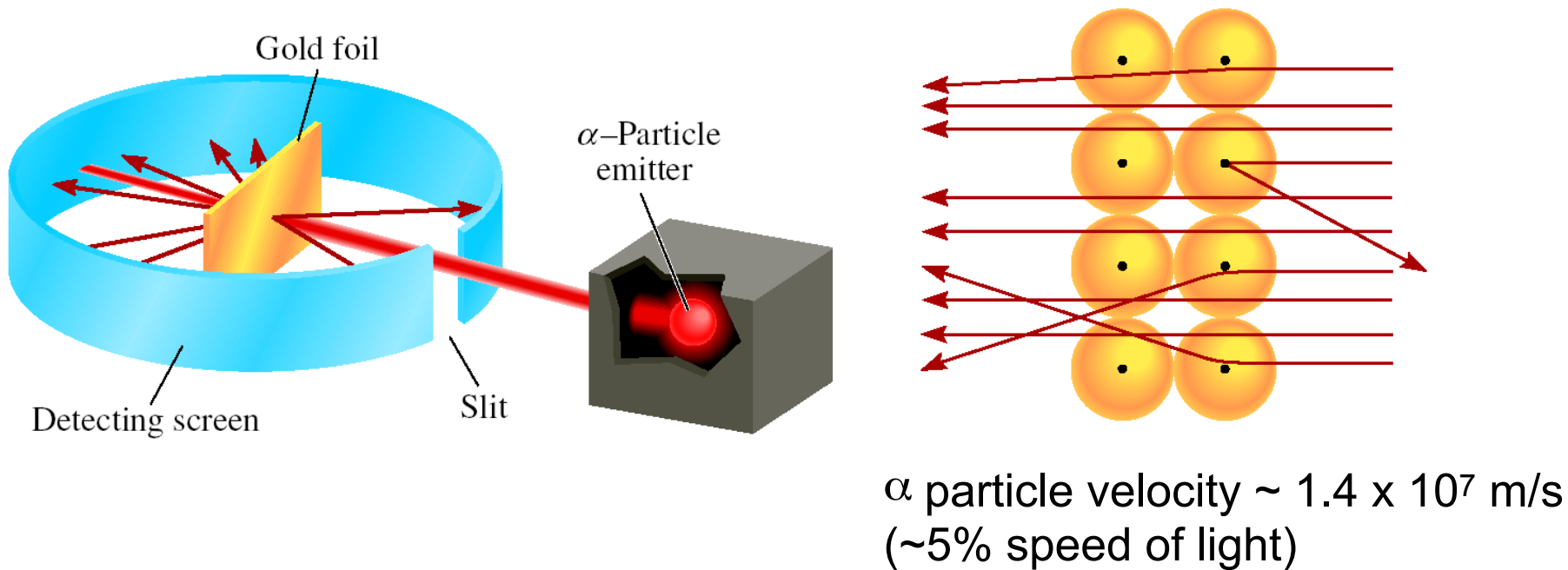
Thomson's Model

Positive charge spread
over the entire sphere



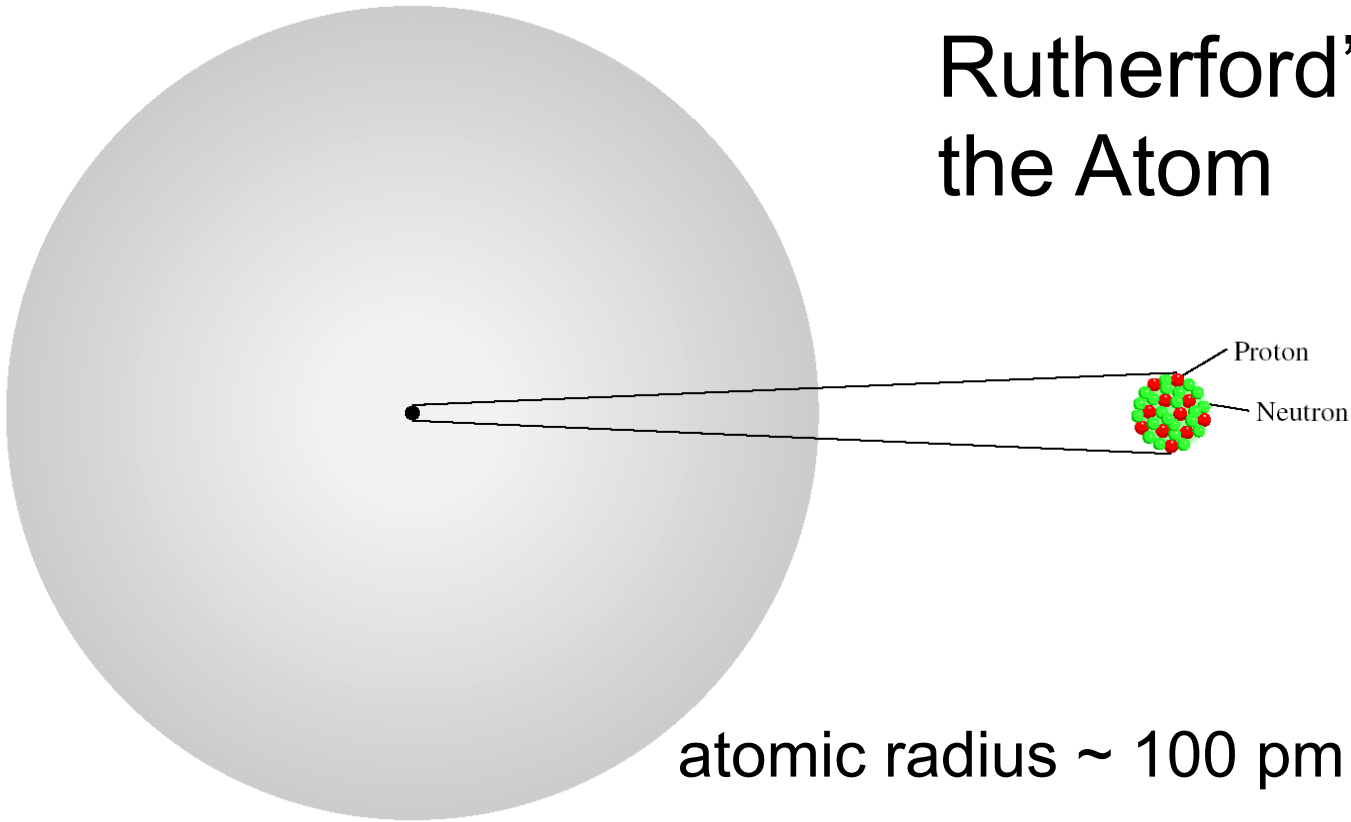
Rutherford's Experiment

(1908 Nobel Prize in Chemistry)



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⁻ (1.67×10^{-24} g)

Rutherford's Model of the Atom



atomic radius $\sim 100 \text{ pm} = 1 \times 10^{-10} \text{ m}$

nuclear radius $\sim 5 \times 10^{-3} \text{ pm} = 5 \times 10^{-15} \text{ m}$



“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



neutron (n) is neutral (charge = 0)

n mass \sim p mass = 1.67×10^{-24} g

TABLE 2.1 Mass and Charge of Subatomic Particles

Particle	Mass (g)	Charge	
		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 \approx mass n \approx 1840 x mass e⁻

Atomic number, Mass number and Isotopes

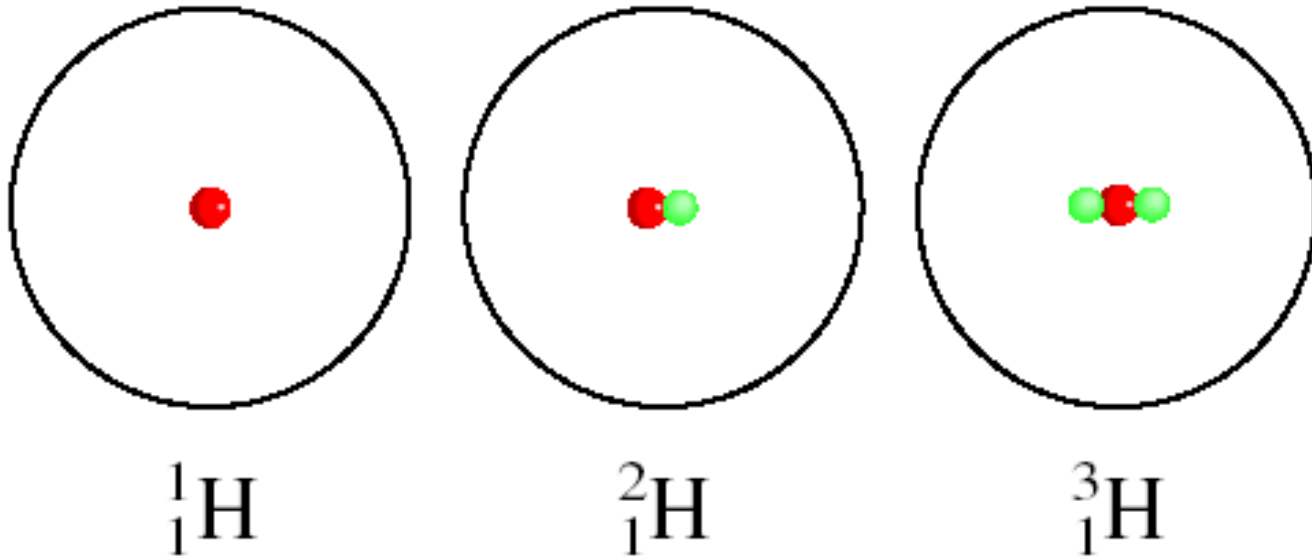
Atomic number (Z) = number of protons in nucleus

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



The Isotopes of Hydrogen



How many protons, neutrons, and electrons are in ${}^{14}_6\text{C}$?

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

How many protons, neutrons, and electrons are in ${}^{11}_6\text{C}$?

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

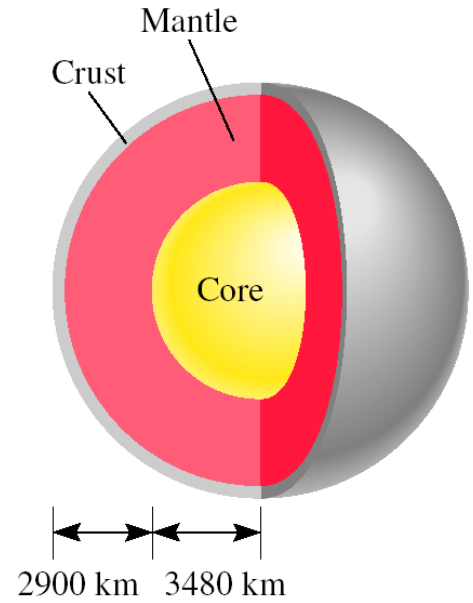
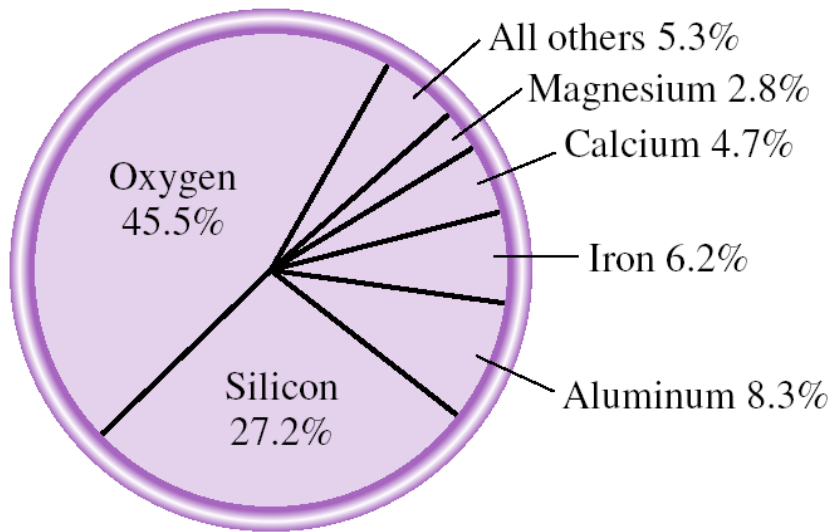
The Modern Periodic Table

1 1A														13 3A	14 4A	15 5A	16 6A	17 7A	18 8A
1 H														5 B	6 C	7 N	8 O	9 F	10 Ne
3														13 Al	14 Si	15 P	16 S	17 Cl	18 Ar
	3 3B	4 4B	5 5B	6 6B	7 7B	8 8B			11 1B	12 2B									
	21 Sc	22 Ti	23 V	24 Cr	25 Mn	26 Fe	27 Co	28 Ni	29 Cu	30 Zn	31 Ga			33 As	34 Se				
	39 Y	40 Zr	41 Nb	42 Mo	43 Tc	44 Ru	45 Rh	46 Pd	47 Ag	48 Cd	49 In			51 Sb	52 Te				
55 Cs	56 Ba	57 La	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		
87 Fr	88 Ra	89 Ac	104 Rf	105 Db	106 Sg	107 Bh	108 Hs	109 Mt	110 Ds	111 Rg	112	113	114	115	116	(117)	118		

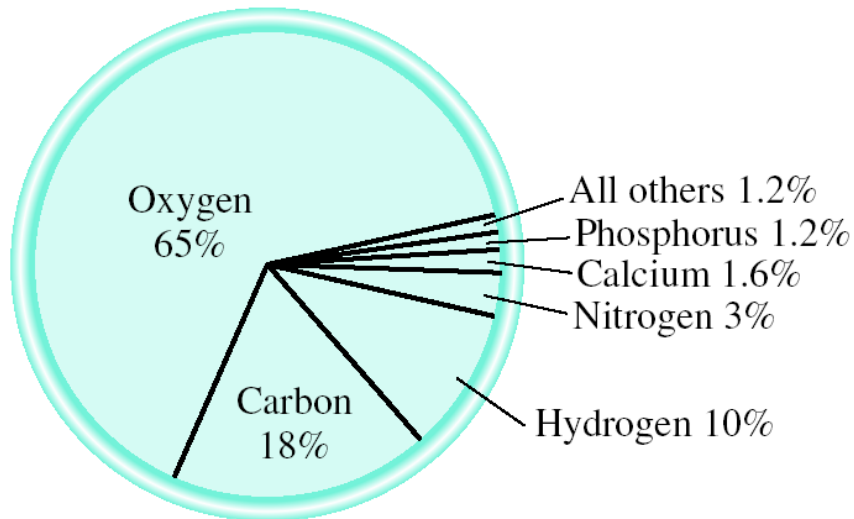
Metals	58 Ce	59 Pr	60 Nd	61 Pm	62 Sm	63 Eu	64 Gd	65 Tb	66 Dy	67 Ho	68 Er	69 Tm	70 Yb	71 Lu
Metalloids	90 Th	91 Pa	92 U	93 Np	94 Pu	95 Am	96 Cm	97 Bk	98 Cf	99 Es	100 Fm	101 Md	102 No	103 Lr
Nonmetals														

Chemistry In Action

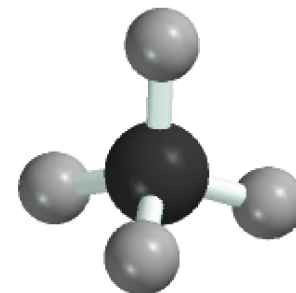
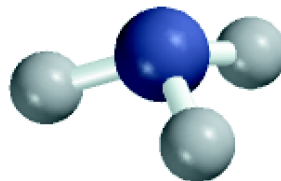
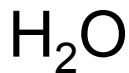
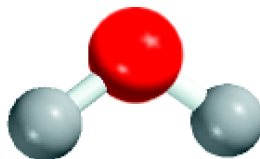
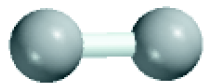
Natural abundance of elements in Earth's crust



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



A **diatomic molecule** contains only two atoms

H_2 , N_2 , O_2 , Br_2 , HCl , CO

1A	2A	3A	4A	5A	6A	7A	8A
H				N	O	F	
						Cl	
						Br	
						I	

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.



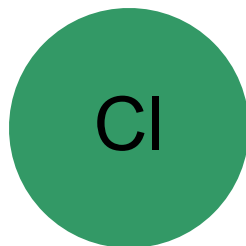
11 protons
11 electrons



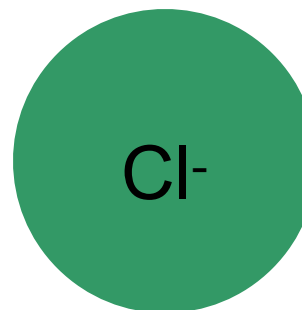
11 protons
10 electrons

anion – ion with a negative charge

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

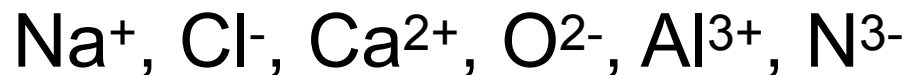


17 protons
17 electrons



17 protons
18 electrons

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



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



Common Ions Shown on the Periodic Table

1 1A	2 2A												13 3A	14 4A	15 5A	16 6A	17 7A	18 8A
Li ⁺														C ⁴⁻	N ³⁻	O ²⁻	F ⁻	
Na ⁺	Mg ²⁺	3 3B	4 4B	5 5B	6 6B	7 7B	8 8B		10	11 1B	12 2B		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⁺					

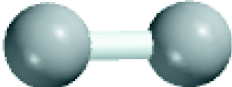
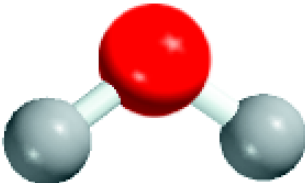
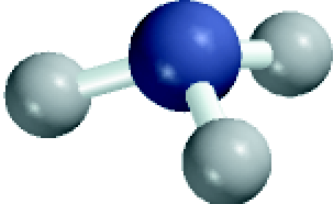
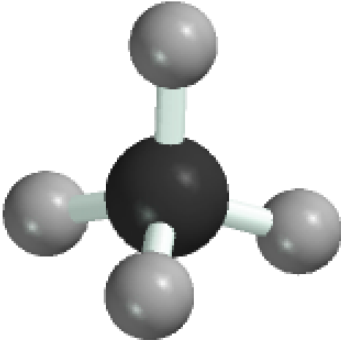
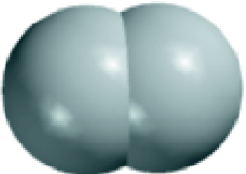

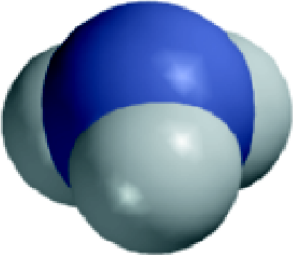
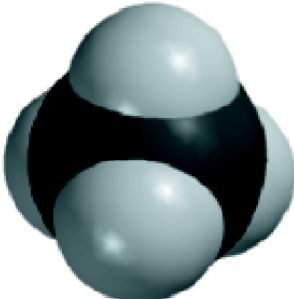
How many protons and electrons are in ${}_{13}^{27}\text{Al}^{3+}$?

13 protons, 10 (13 – 3) electrons

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

34 protons, 36 (34 + 2) electrons

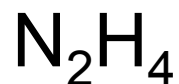
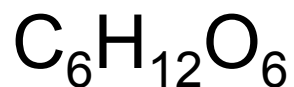
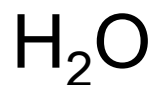
Formulas and Models

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

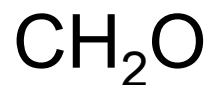
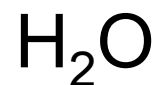
A ***molecular formula*** shows the exact number of atoms of each element in the smallest unit of a substance

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

molecular



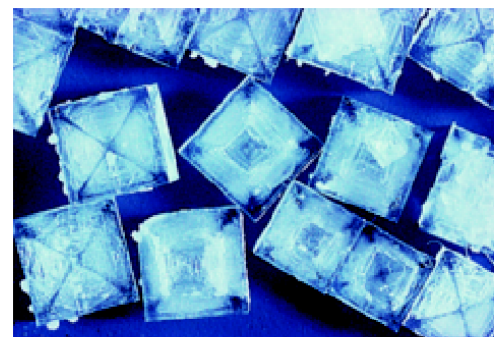
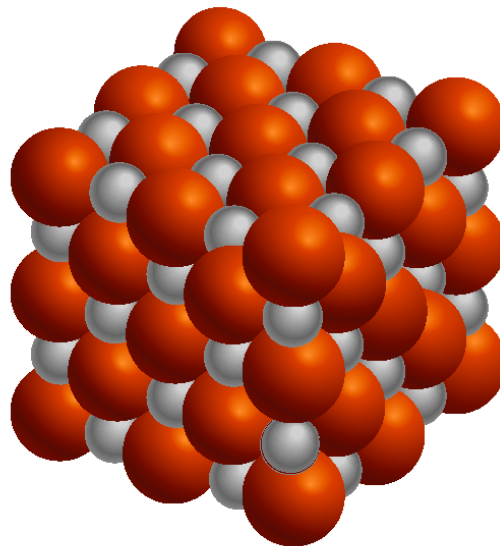
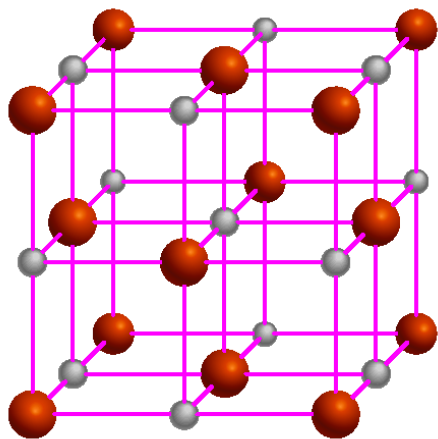
empirical



ionic compounds consist of a combination of cations and 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

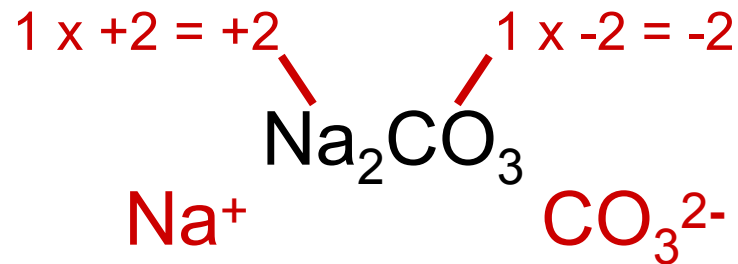
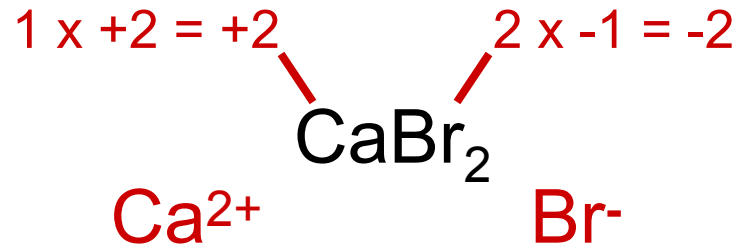
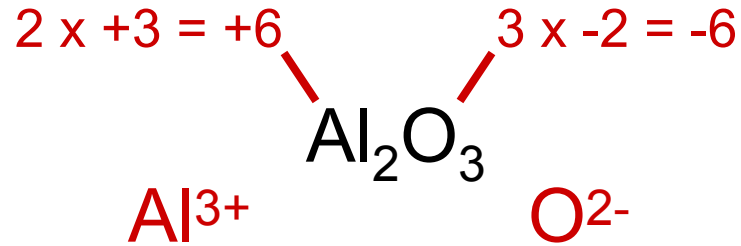
The ionic compound NaCl



1A	2A											3A	4A	5A	6A	7A	8A
Li														N	O	F	
Na	Mg											Al			S	Cl	
K	Ca															Br	
Rb	Sr															I	
Cs	Ba																

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

Formula of Ionic Compounds



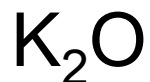
Chemical Nomenclature

- **Ionic Compounds**

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



barium chloride



potassium oxide

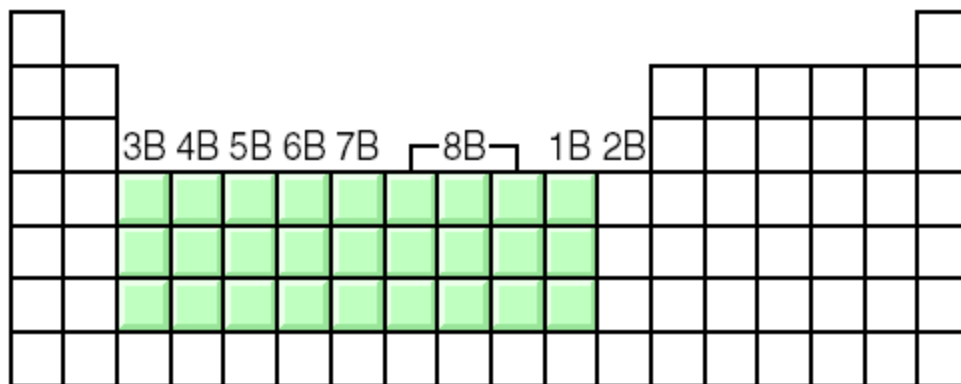


magnesium hydroxide



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 ($\text{Cr}_2\text{O}_7^{2-}$)
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

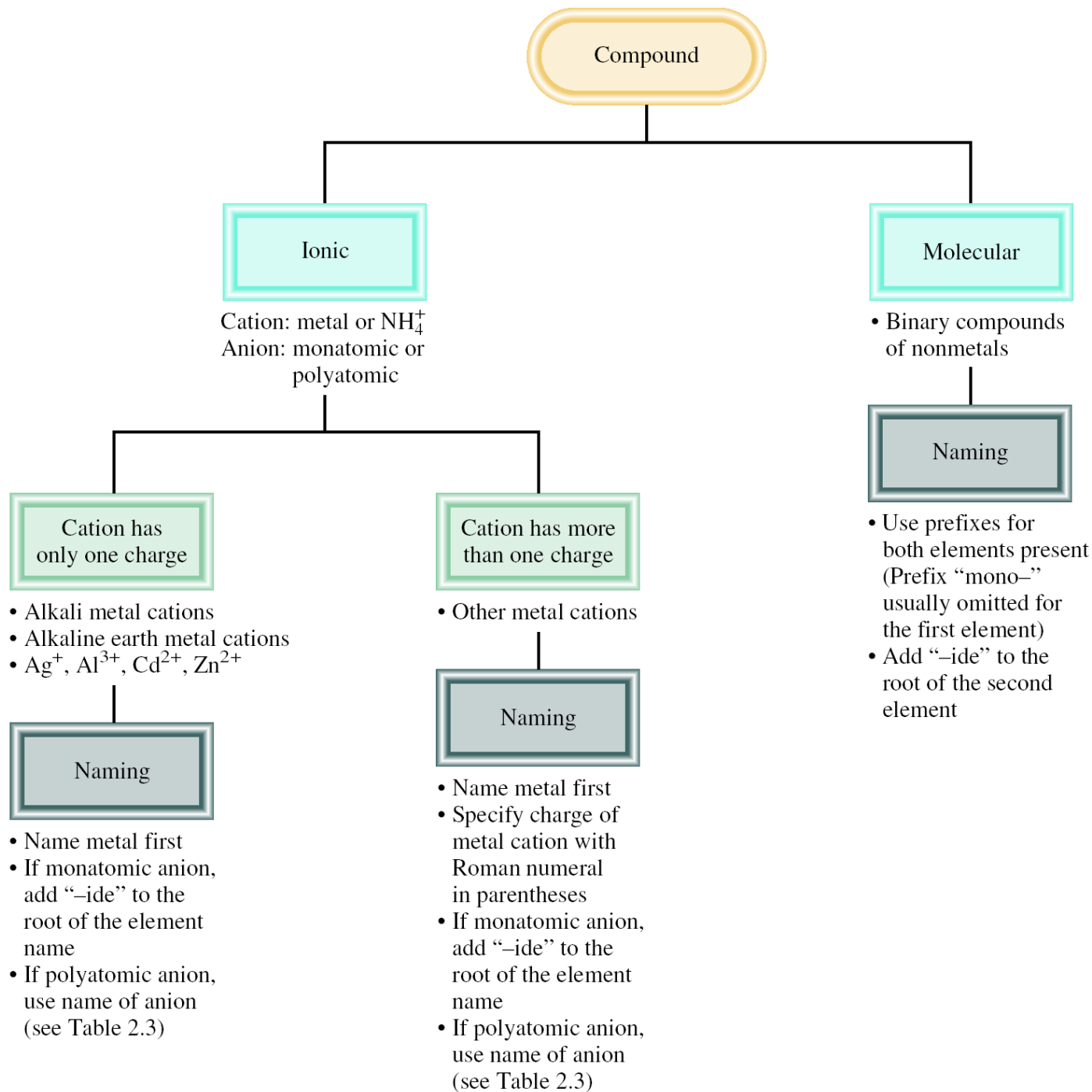
NF₃ nitrogen trifluoride

SO₂ sulfur dioxide

N₂Cl₄ dinitrogen tetrachloride

NO₂ nitrogen dioxide

N₂O dinitrogen monoxide



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

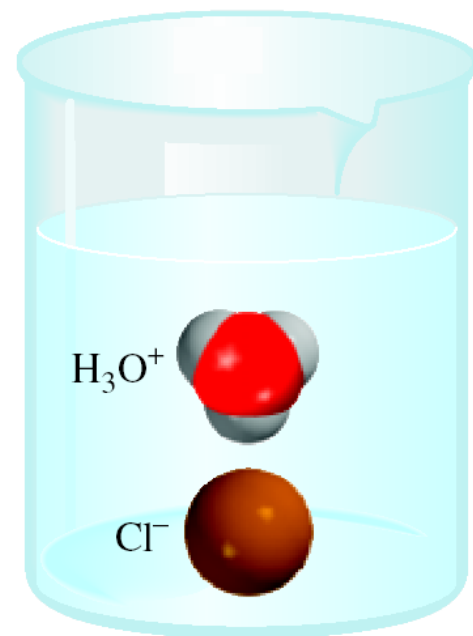


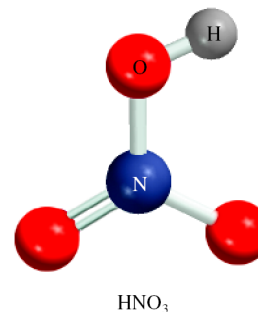
TABLE 2.5 Some Simple Acids

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)

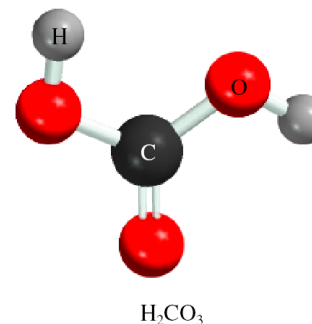
An **oxoacid** is an acid that contains hydrogen, oxygen, and another element.



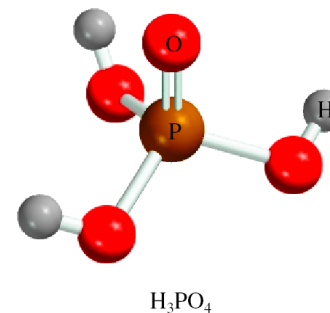
nitric acid



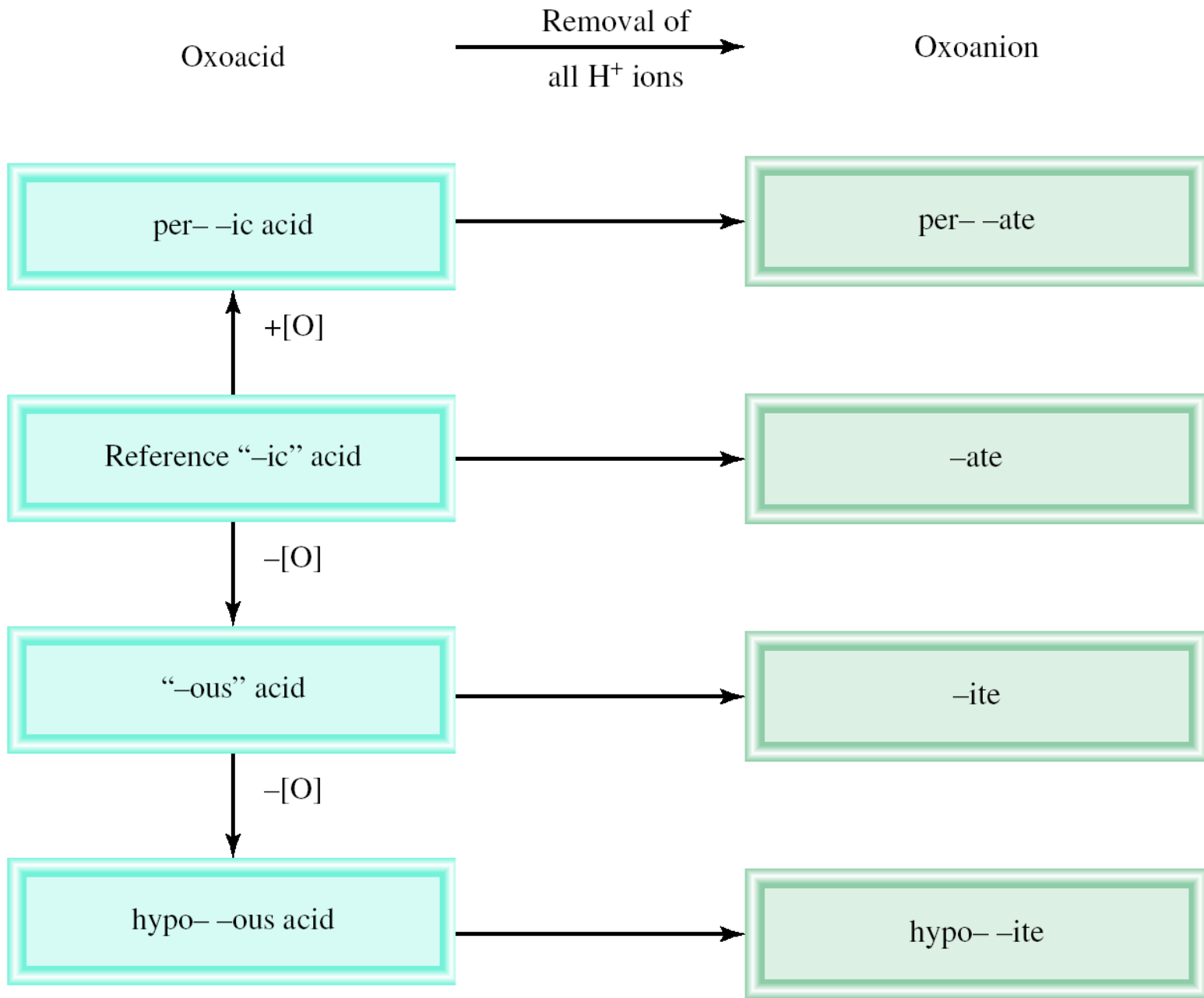
carbonic acid



phosphoric acid



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)

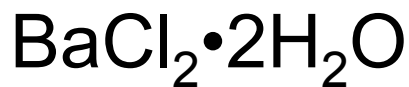
A **base** can be defined as a substance that yields hydroxide ions (OH^-) when dissolved in water.

NaOH sodium hydroxide

KOH potassium hydroxide

$\text{Ba}(\text{OH})_2$ barium hydroxide

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



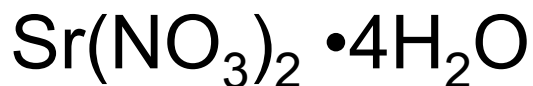
barium chloride dihydrate



lithium chloride monohydrate



magnesium sulfate heptahydrate



strontium nitrate tetrahydrate

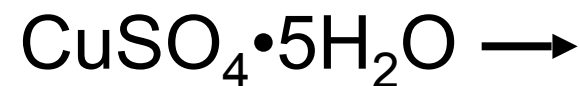
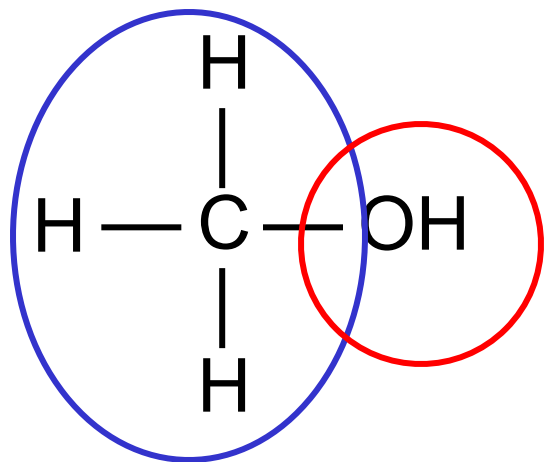


TABLE 2.7**Common and Systematic Names of Some Compounds**

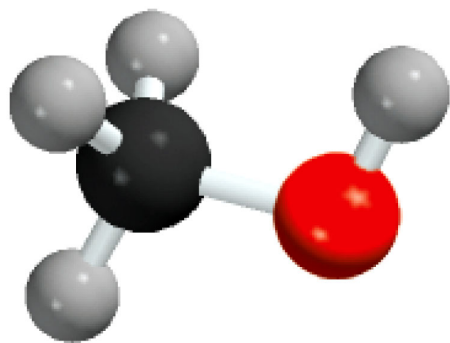
Formula	Common Name	Systematic Name
H_2O	Water	Dihydrogen monoxide
NH_3	Ammonia	Trihydrogen nitride
CO_2	Dry ice	Solid carbon dioxide
NaCl	Table salt	Sodium chloride
N_2O	Laughing gas	Dinitrogen monoxide
CaCO_3	Marble, chalk, limestone	Calcium carbonate
CaO	Quicklime	Calcium oxide
$\text{Ca}(\text{OH})_2$	Slaked lime	Calcium hydroxide
NaHCO_3	Baking soda	Sodium hydrogen carbonate
$\text{Na}_2\text{CO}_3 \cdot 10\text{H}_2\text{O}$	Washing soda	Sodium carbonate decahydrate
$\text{MgSO}_4 \cdot 7\text{H}_2\text{O}$	Epsom salt	Magnesium sulfate heptahydrate
$\text{Mg}(\text{OH})_2$	Milk of magnesia	Magnesium hydroxide
$\text{CaSO}_4 \cdot 2\text{H}_2\text{O}$	Gypsum	Calcium sulfate dihydrate

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

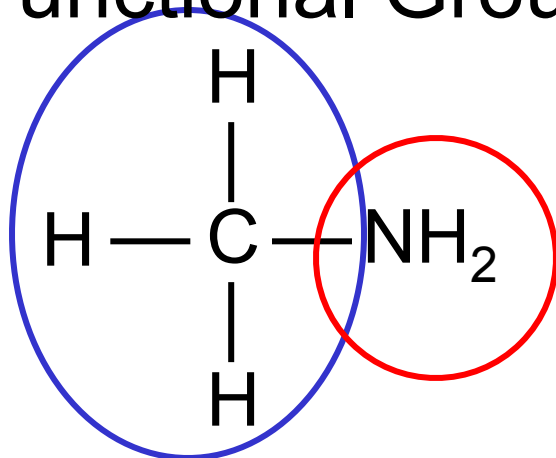
Functional Groups



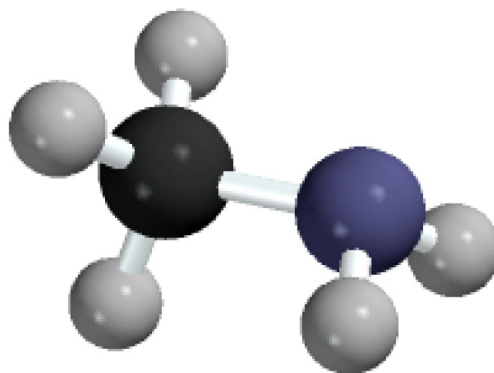
methanol



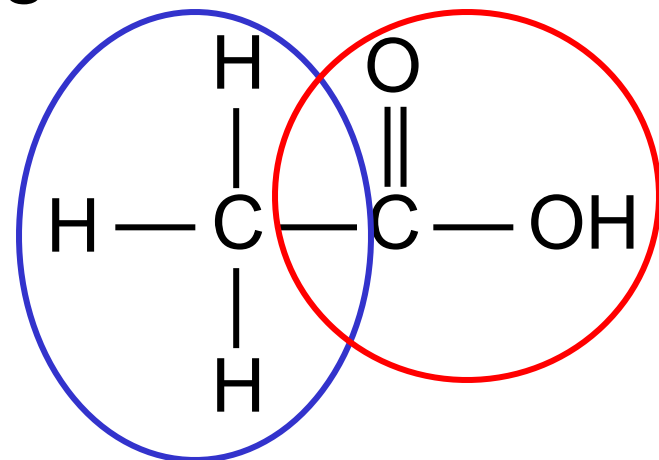
CH_3OH



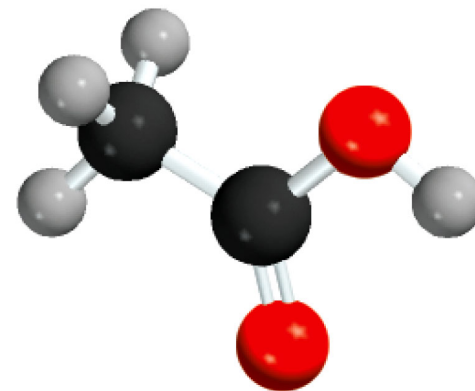
methylamine



CH_3NH_2




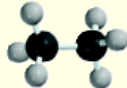
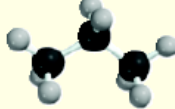
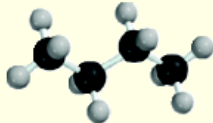
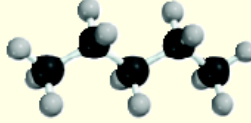
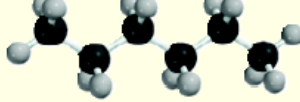
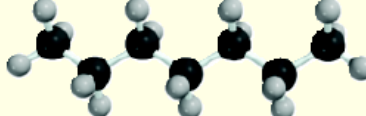
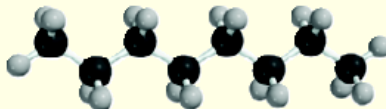
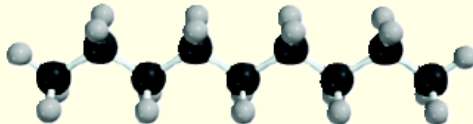
acetic acid



CH_3COOH

TABLE 2.8

The First Ten Straight-Chain Alkanes

Name	Formula	Molecular Model
Methane	CH ₄	
Ethane	C ₂ H ₆	
Propane	C ₃ H ₈	
Butane	C ₄ H ₁₀	
Pentane	C ₅ H ₁₂	
Hexane	C ₆ H ₁₄	
Heptane	C ₇ H ₁₆	
Octane	C ₈ H ₁₈	
Nonane	C ₉ H ₂₀	
Decane	C ₁₀ H ₂₂	