Introduction and Review:

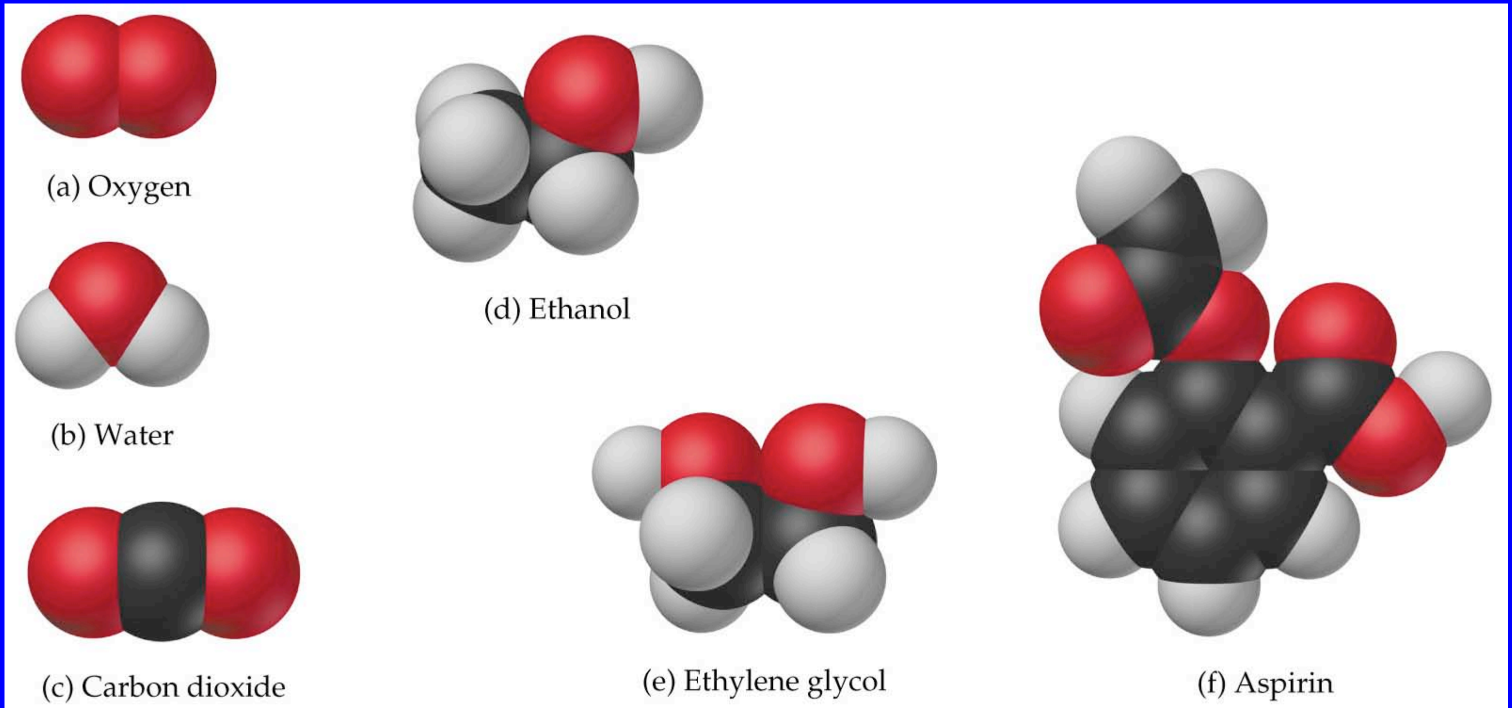
Chapter 1: Matter – Its Properties and Measurement (Review on your own)

Chapter 2: Atoms and Atomic Theory

Chapter 3: Chemical Compounds

What in the world isn't Chemistry

- Matter is the physical material of the universe.
- On the microscopic level, matter consists of atoms and molecules.
- As we know, molecules may consist of the same type of atoms or different types of atoms.



We encounter chemistry in our everyday lives

Lavoisier, 1743-1794

Dalton (1766 – 1844)

Conservation of Matter - A chemical equation must be balanced for mass.

Atomic Theory - A chemical equation must have the same number of atoms of the same kind on both sides.

Dalton's Atomic Theory

- ✓ All matter is made of small, indestructible particles called "atoms".
- ✓ All atoms of a given element are identical and atoms differ for different elements.
- ✓ Compounds are formed when atoms of more than one element combine
- ✓ Atoms of different elements combine in simple, whole number ratios (e.g. 1:1 as in AB; or 1:2 as in AB_2)

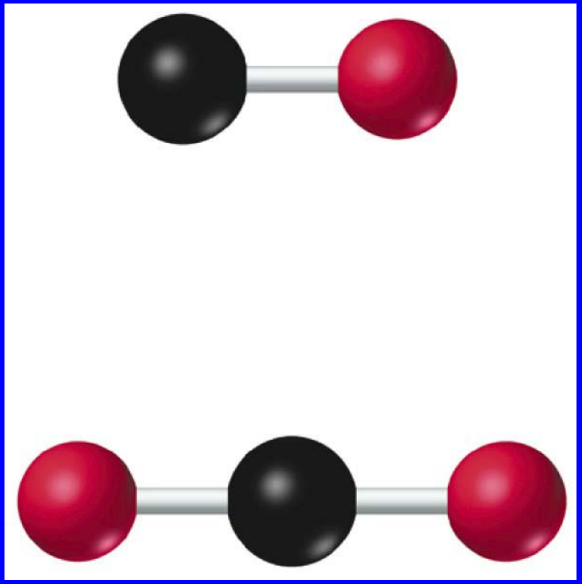
Consequences of Dalton's Theory

Combinations of elements are in ratios of small whole numbers (molecules)

Carbon monoxide (CO): 1.33 g of oxygen combines with 1.0 g of carbon or O/C = 1.33

Carbon dioxide (CO_2): the ratio of O/C is 2.667

2.667/1.33 = 2.0 which is the ratio of oxygen in CO_2/oxygen in CO



What are Atoms?

Atoms are the basic unit of an element that can enter into chemical combination.

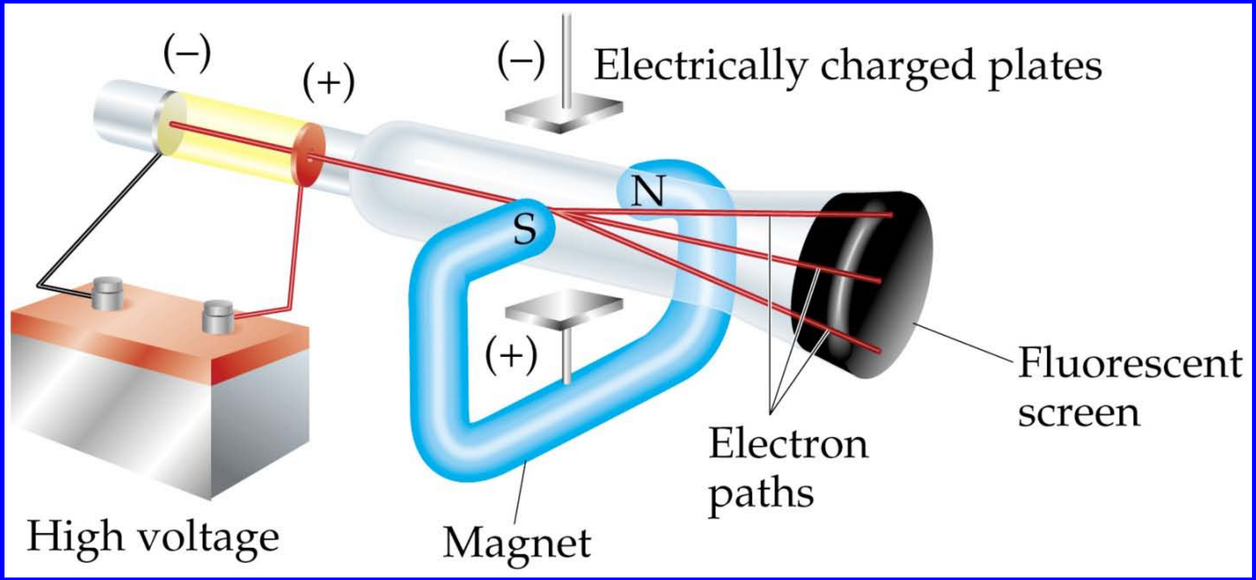
Dalton – Atoms are small and indivisible

– not quite!

Atomic Structure

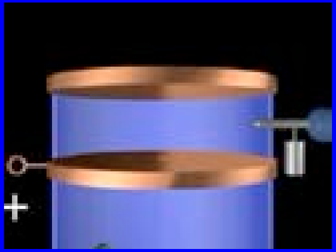
J.J. Thompson's Experiment (1898 – 1903)

Z/m of the electron = 1.76×10^{8} C/g



Milliken's Oil Drop Experiment (1909)

charge of the electron = -1.6022×10^{-19} C

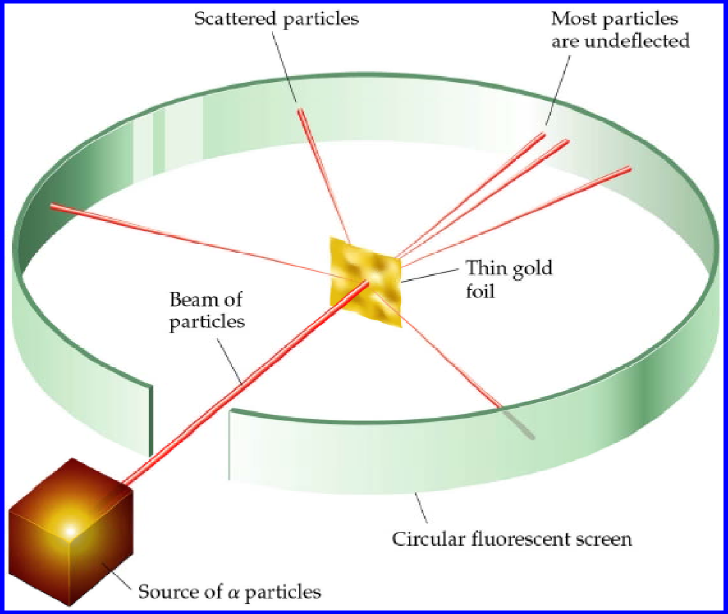


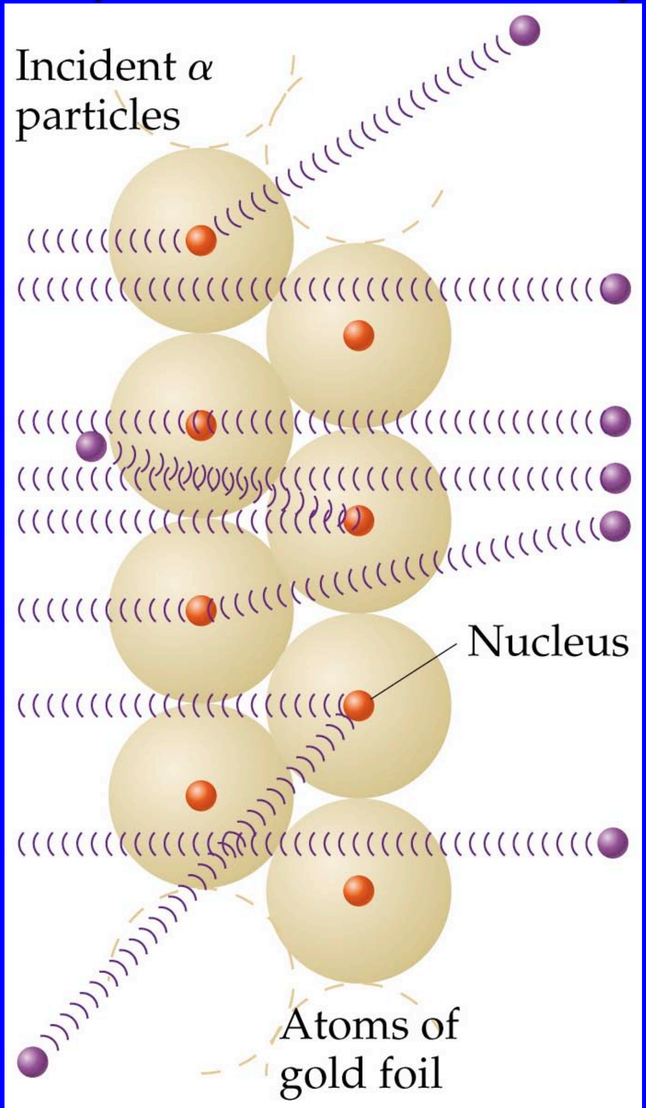
Therefore: mass electron = (e^- charge)/(e^- charge-to-mass ratio)

= 9.10×10^{-28} g (9.10939×10^{-28} g)

Atomic Structure

Nuclear Model of the Atom (Rutherford: 1910)

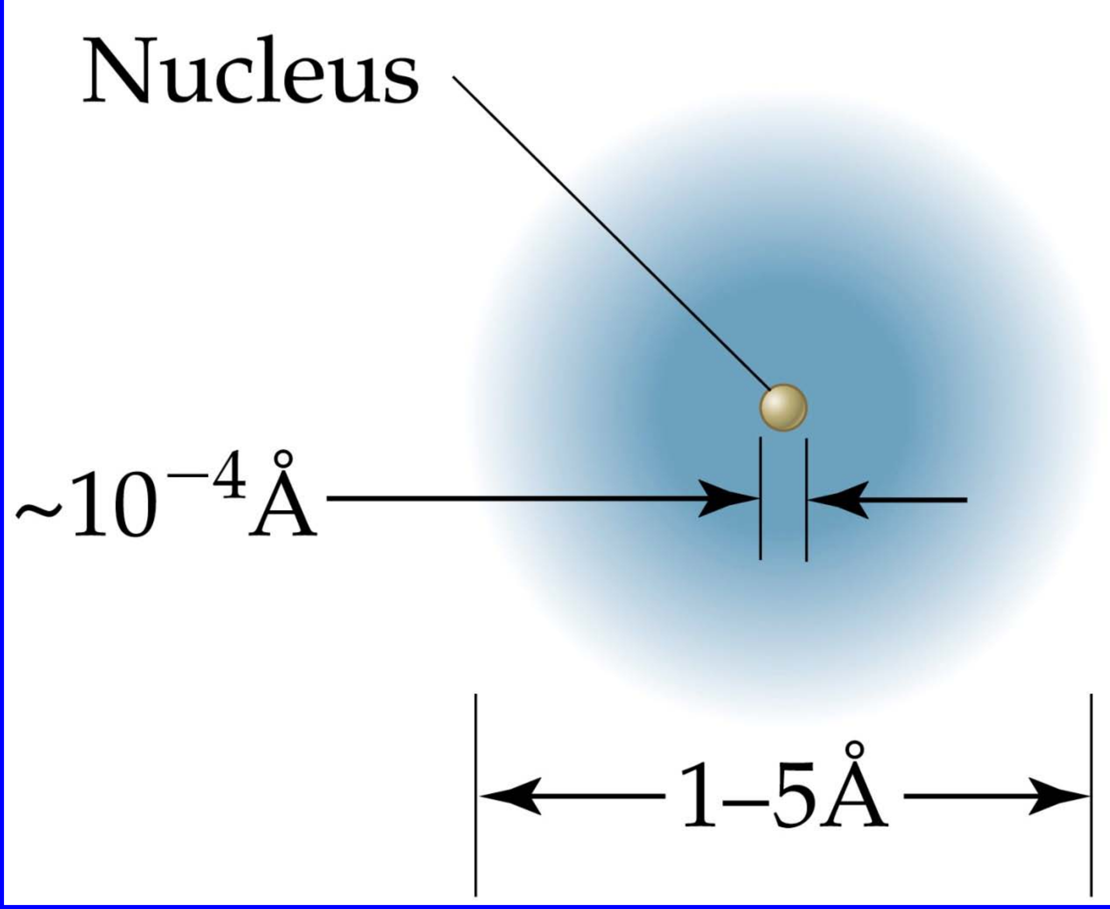




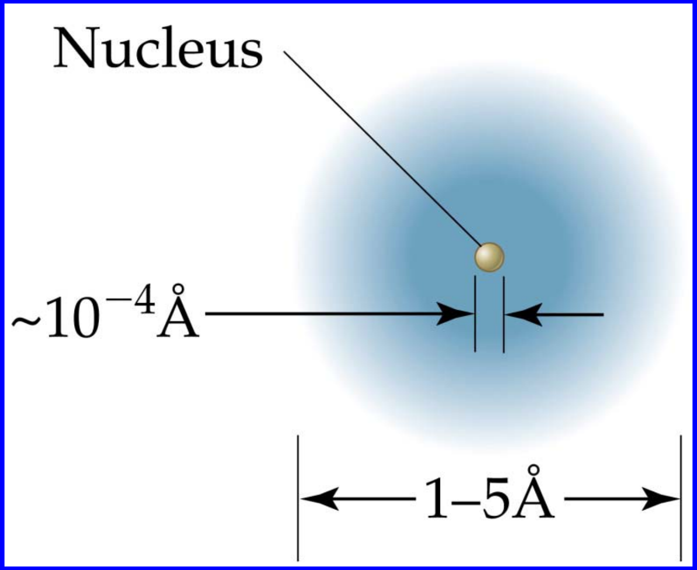
The Modern View of Atomic Structure

<u>Nucleus</u>

- nucleus contains protons and neutrons
- nuclear radius is 20,000 times smaller than the atomic radius!
  (5 cm *vs* 1 km)



The Modern View of Atomic Structure



Particle	mass	charge
Proton	1.67×10^{-24} g	1.60×10^{-19} C
Neutron	1.67×10^{-24} g	No charge
Electron	9.11×10^{-28} g	-1.60×10^{-19} C

Mass of atom is mostly from nucleus, volume of atom is mostly from electron cloud.

Scale of Atoms

Useful units:

- ☆ **1 u (atomic mass unit) = 1.66054 x 10^{-24} kg**
- ☆ **1 pm (picometer) = 1 x 10^{-12} m**
- ☆ **1 Å (Angstrom) = 1 x 10^{-10} m = 100 pm = 1 x 10^{-8} cm**

❖ The heaviest atom has a mass of only 4.5 x 10^{-22} g and a diameter of only 4 x 10^{-10} m.

❖ Typical C-C bond length 1.54 Å (1.5 x 10^{-10} m)

Isotopes, Atomic Numbers, and Mass Numbers

- Atomic number Z = number of protons

  determines the number of electrons thus the element identity
- Atomic Mass - Σ of protons and neutrons
- By convention, for element X, we write ${}_{Z}^{A}X$.
- Isotopes have the same Z but different A (variable number of neutrons)
- No two elements have the same value of Z.

Isotopes

$^{1}_{1}H$ Hydrogen

$^{2}_{1}H$ Deuterium

$^{3}_{1}H$ Tritium

How many electrons does each isotope have?

In nature, elements occur as a mixture of isotopes

Atomic Weights

- ^{1}H weighs 1.6735×10^{-24} g and ^{16}O 2.6560×10^{-23} g.
- We define: mass of ^{12}C = 12 u. Where 12 u is an ***exact*** number (with an infinite number of zeros past the decimal). The mass of every other element in atomic mass units is determined by measuring its mass relative to the mass of ^{12}C.
- Using atomic mass units:

$$1\ u = 1.66054 \times 10^{-24}\ g$$

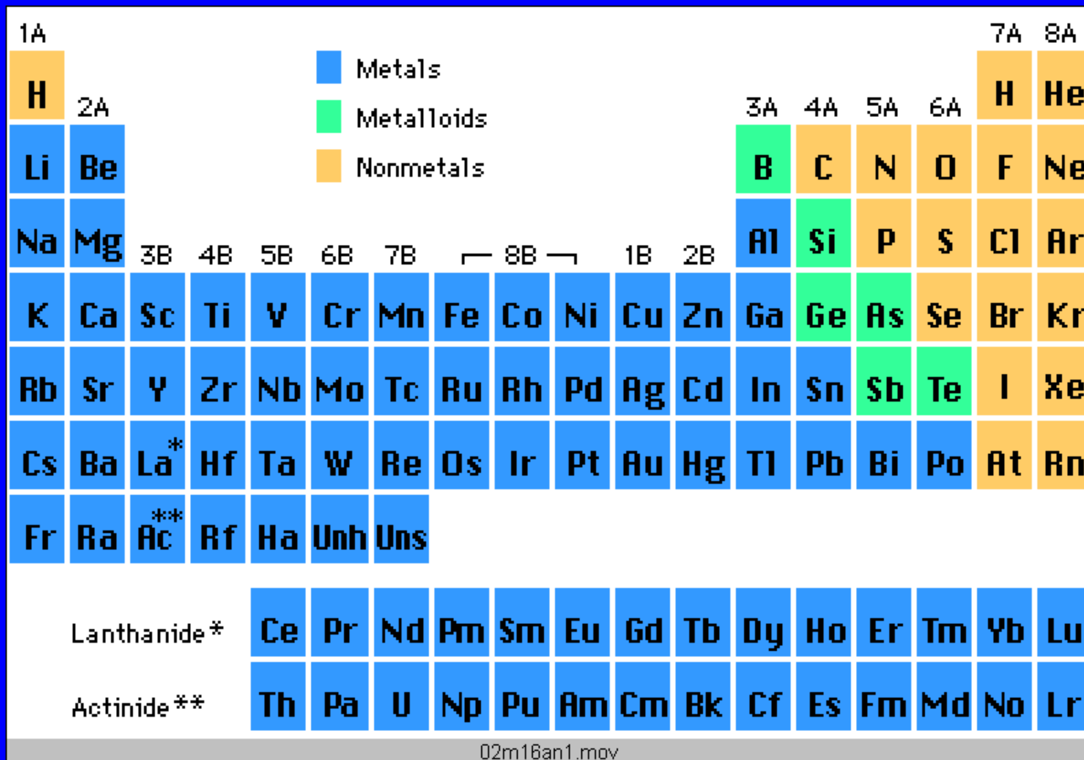
$$1\ g = 6.02214 \times 10^{23}\ u$$

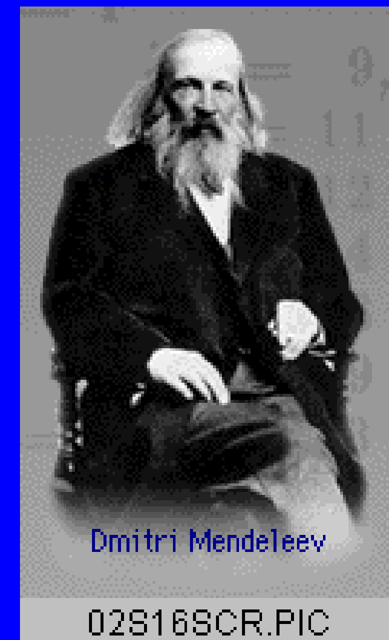
Remember that we must include the isotopes when calculating atomic weights. In nature C: 98.892 % ^{12}C + 1.108 % ^{13}C.

- Average mass of C:
  (0.98892)(12) + (0.0108)(13.00335) = 12.011 amu.

- What is the mass of a single C atom??

The Periodic Table





(1834 – 1907)

- Based on the reactions and properties of the elements and their compounds
- Reflects the underlying electronic structure of the elements

The Periodic Table

- Columns in the periodic table are called *groups* (numbered from 1A to 8A or 1 to 18).
- Rows in the periodic table are called *periods*.
- Metals are located on the left hand side of the periodic table (most of the elements are metals).
- Non-metals are located in the top right hand side of the periodic table.
- Elements with properties similar to both metals and non-metals are called metalloids and are located at the interface between the metals and non-metals.

The Periodic Table

Alkali Metals

Alkaline Earths

Halogens

Main Group

Noble Gases

Transition Metals

1 1A	2 2A	3 3B	4 4B	5 5B	6 6B	7 7B	8 8B	9 8B	10 8B	11 1B	12 2B	13 3A	14 4A	15 5A	16 6A	17 7A	18 8A
1 H 1.00794																	2 He 4.00260
3 Li 6.941	4 Be 9.01218											5 B 10.811	6 C 12.011	7 N 14.0067	8 O 15.9994	9 F 18.9984	10 Ne 20.1797
11 Na 22.9898	12 Mg 24.3050											13 Al 26.9815	14 Si 28.0855	15 P 30.9738	16 S 32.066	17 Cl 35.4527	18 Ar 39.948
19 K 39.0983	20 Ca 40.078	21 Sc 44.9559	22 Ti 47.88	23 V 50.9415	24 Cr 51.9961	25 Mn 54.9381	26 Fe 55.847	27 Co 58.9332	28 Ni 58.693	29 Cu 63.546	30 Zn 65.39	31 Ga 69.723	32 Ge 72.61	33 As 74.9216	34 Se 78.96	35 Br 79.904	36 Kr 83.80
37 Rb 85.4678	38 Sr 87.62	39 Y 88.9059	40 Zr 91.224	41 Nb 92.9064	42 Mo 95.94	43 Tc (98)	44 Ru 101.07	45 Rh 102.906	46 Pd 106.42	47 Ag 107.868	48 Cd 112.411	49 In 114.818	50 Sn 118.710	51 Sb 121.757	52 Te 127.60	53 I 126.904	54 Xe 131.29
55 Cs 132.905	56 Ba 137.327	57 *La 138.906	72 Hf 178.49	73 Ta 180.948	74 W 183.84	75 Re 186.207	76 Os 190.23	77 Ir 192.22	78 Pt 195.08	79 Au 196.967	80 Hg 200.59	81 Tl 204.383	82 Pb 207.2	83 Bi 208.980	84 Po (209)	85 At (210)	86 Rn (222)
87 Fr (223)	88 Ra 226.025	89 †Ac 227.028	104 Rf (261)	105 Db (262)	106 Sg (263)	107 Bh (262)	108 Hs (265)	109 Mt (266)	110 (269)	111 (272)	112 (272)		114 (287)		116 (289)		118 (293)

*Lanthanide series	58 Ce 140.115	59 Pr 140.908	60 Nd 144.24	61 Pm (145)	62 Sm 150.36	63 Eu 151.965	64 Gd 157.25	65 Tb 158.925	66 Dy 162.50	67 Ho 164.930	68 Er 167.26	69 Tm 168.934	70 Yb 173.04	71 Lu 174.967
†Actinide series	90 Th 232.038	91 Pa 231.036	92 U 238.029	93 Np 237.048	94 Pu (244)	95 Am (243)	96 Cm (247)	97 Bk (247)	98 Cf (251)	99 Es (252)	100 Fm (257)	101 Md (258)	102 No (259)	103 Lr (260)

Main Group

Lanthanides and Actinides

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
4	19 K	20 Ca															21 Sc	22 Ti	23 V	24 Cr	25 Mn	26 Fe	27 Co	28 Ni	29 Cu	30 Zn	31 Ga	32 Ge	33 As	34 Se	35 Br	36 Kr
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	57 La	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	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
7	87 Fr	88 Ra	89 Ac	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	104 Rf	105 Db	106 Sg	107 Bh	108 Hs	109 Mt	110 Ds	111	112		114				

- Arranged in order of increasing mass, the periodic table looks like this.
- Because rows 6 and 8 are quite long, 14 of these elements are separated from the rest.
- They are called the lanthanides and the actinides. (WHY?)

Counting Atoms

- Recall that we must account for all of the atoms in a chemical reaction.
- Physically counting atoms is impossible.
- We must be able to relate measured mass to numbers of atoms.
  - using atoms by the gram

Avogadro's number


Amedeo Avogadro (1776-1856)

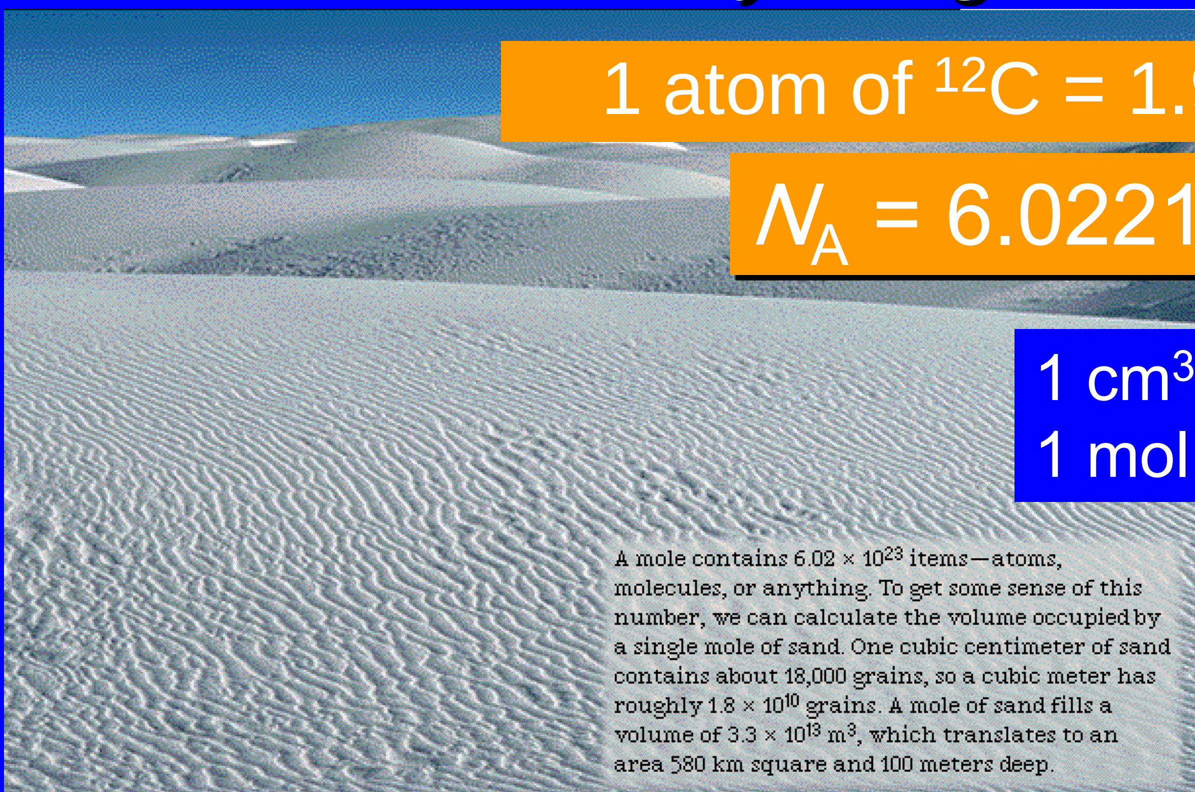
The mole is an amount of substance that contains the same number of elementary entities as there are carbon-12 atoms in *exactly* 12 g of carbon-12.

1 atom of ^{12}C = 1.992646 x 10^{-23}g

N_A = 6.02214199 x 10^{23} mol^{-1}

1 cm^3 sand = 18,000 grains

1 mol sand = 580 km x 100m



Molar Mass

- The molar mass, M, is the mass of one mole of a substance.

M (g/mol ^{12}C) = A (g/atom ^{12}C) x N_A (atoms ^{12}C /mol ^{12}C)

Elements occur as a mixture of their isotopes. The molar mass of an element is the mass of 1 mole of the naturally occurring *mixture of isotopes*.

Thus: 1 mole of O = 15.9994 g

1 mole of H = 1.0079 g

1 mole of C = 12.011 g

"molecular weight and atomic weights"
numerical values of the molar mass
not actually weights!!

Molar Mass

- Most elements are found on the earth as a mixture of isotopes in a constant ratio.
- A single atom of each isotope has a unique mass.
- Since we cannot deal with single atoms we think of a mole of atoms. How much does a mole of atoms of an isotope weigh?
- Let's look at ^{13}C….it has an atomic mass of 2.15928×10^{-23} g, what is the mass of one mole of ^{13}C?
- Is this the mass for carbon on the periodic table?

Example 2-9

Combining Several Factors in a Calculation—Molar Mass, the Avogadro Constant, Percent Abundance.

Potassium-40 is one of the few naturally occurring radioactive isotopes of elements of low atomic number. Its percent natural abundance among K isotopes is 0.012%. How many ^{40}K atoms do you ingest by drinking one cup of whole milk containing 371 mg of K?

Want atoms of ^{40}K, need atoms of K,

Want atoms of K, need moles of K,

Want moles of K, need mass and M(K).

Convert strategy to plan

Convert mass of K(mg K) into moles of K (mol K)

m_K(mg) x (1g/1000mg) ➔ m_K (g) x 1/M_K (mol/g) ➔ n_K(mol)

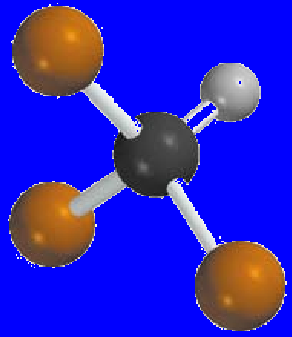
n_K = (371 mg K) x (10^{-3} g/mg) x (1 mol K) / (39.10 g K) = 9.49 x 10^{-3} mol K

Convert moles of K into atoms of ^{40}K

n_K(mol) x N_A ➔ atoms K x 0.012% ➔ atoms ^{40}K

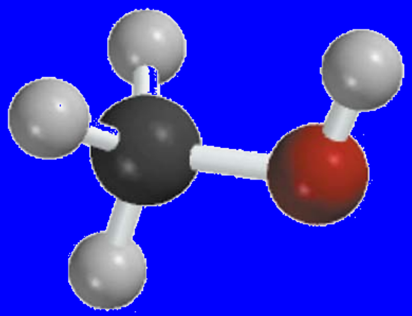
atoms ^{40}K = (9.49 x 10^{-3} mol K) x (6.022 x 10^{23} atoms K/mol K) x (1.2 x 10^{-4} ^{40}K/K) = 6.9 x 10^{17} ^{40}K atoms

Chapter 3 Molecules



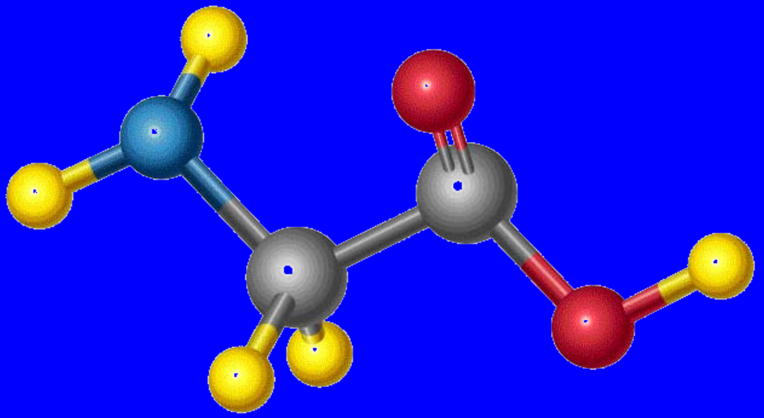
Only the noble gases exist in nature as single atoms

The atoms of all other elements combine to form **MOLECULES**



e.g. H_2 a **DIATOMIC** molecule

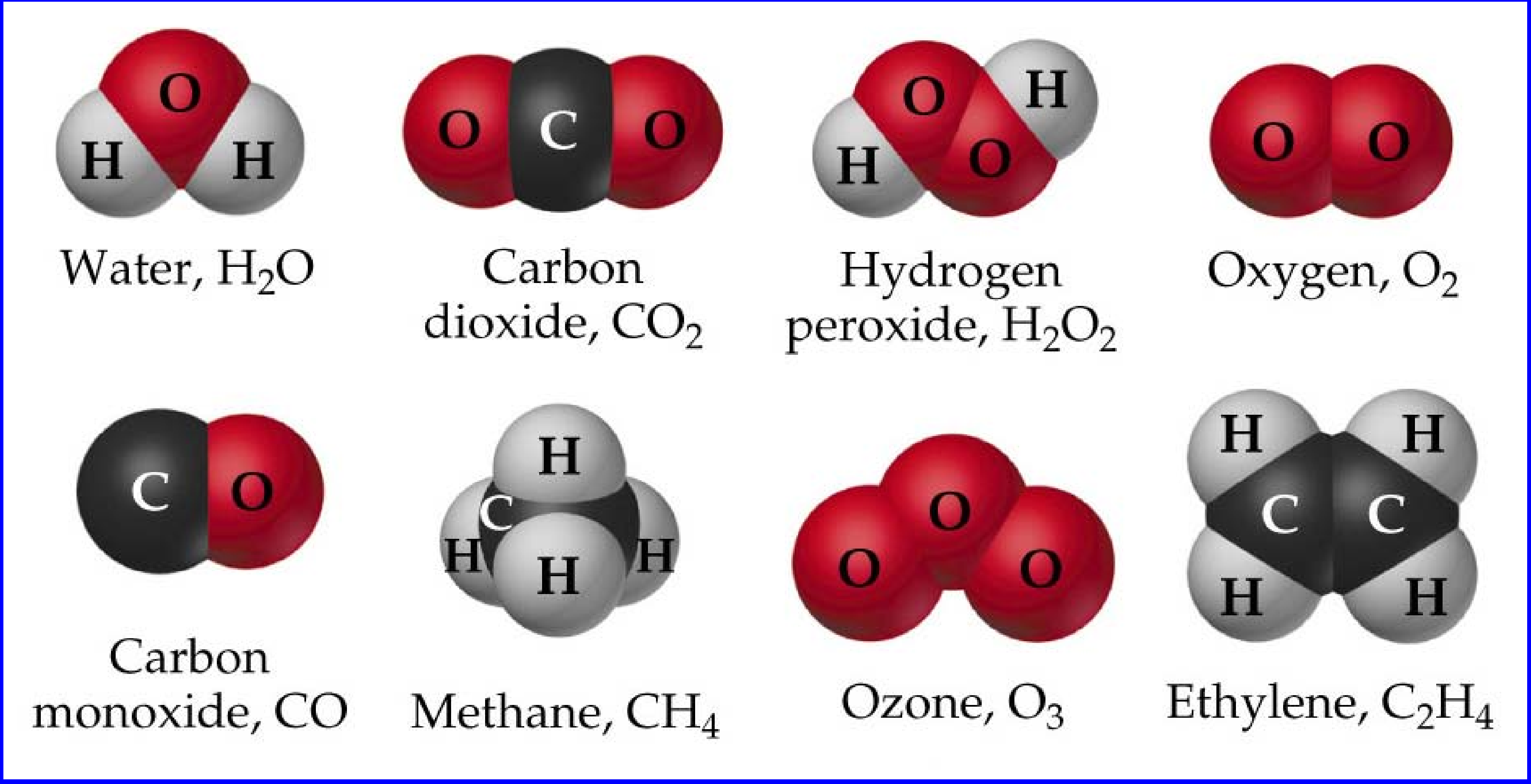
H_2O a **POLYATOMIC** molecule



In a neutral atom or molecule # protons = # electrons

Molecules and Molecular Compounds

Molecular Structural and Empirical Formulas



Molecular Mass/Molecular Weight

Molecular Mass

molecular mass = sum of atomic mass of each atom in the molecule (in atomic mass units = u)

Determine the molecular mass of H_2O:

$$2 \text{ atoms H} \times \frac{1.008 \text{ u}}{1 \text{ H atom}} = 2.016 \text{ u}$$

$$1 \text{ atom O} \times \frac{15.999 \text{ u}}{1 \text{ O atom}} = 15.999 \text{ u}$$

molecular mass

1 molecule H_2O = 2.016 u + 15.999 u = 18.015 u

Molar Mass of Molecules

mass in grams of 1 mole of the molecule

What is the molar mass of H_2O?

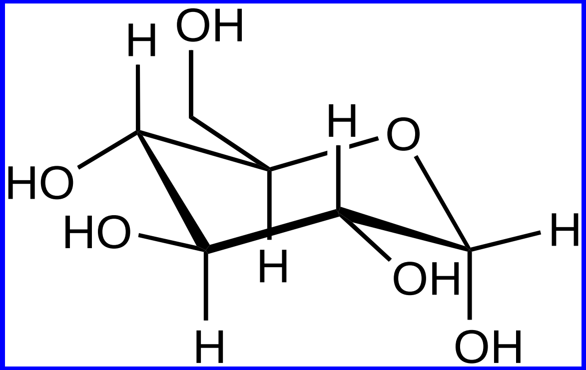
molar mass

1 mole H_2O = 2.016 g for two moles of H
+ 15.999 g for one mole of O = 18.015 g/mol

Molecular Mass: mass of one molecule 18.015 u

Molar Mass: mass of one mole of molecules 18.015 g/mol

Molecular mass



Glucose

Molecular formula $C_6H_{12}O_6$
Empirical formula CH_2O

Molecular Mass: Use the naturally occurring mixture of isotopes,

$$6 \times 12.01 + 12 \times 1.01 + 6 \times 16.00 = 180.18$$

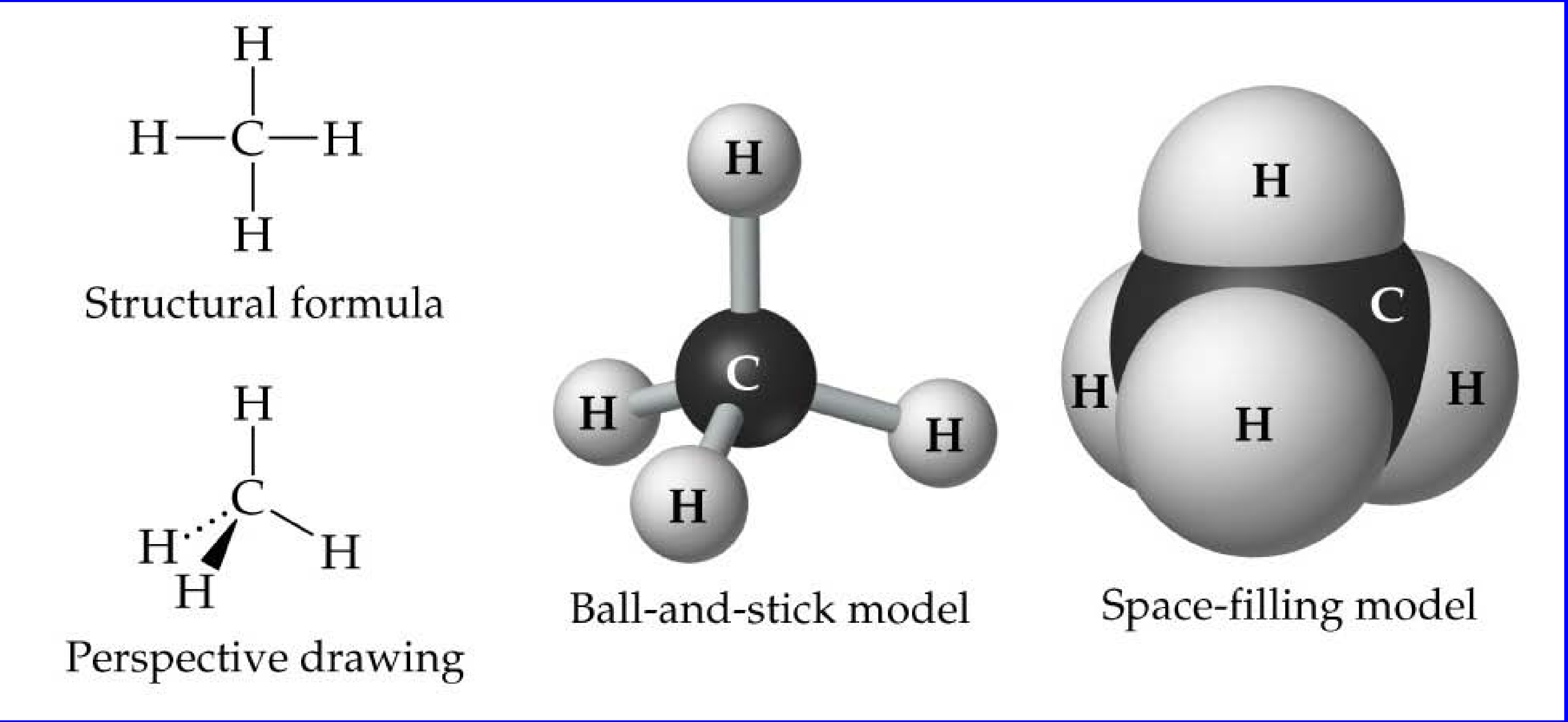
Exact Mass: Use the most abundant isotopes,

$$6 \times 12.000000 + 12 \times 1.007825 + 6 \times 15.994915$$
$$= 180.06339$$

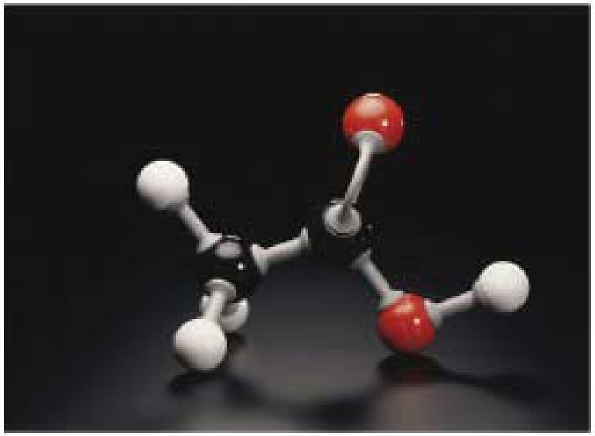
Molecules and Molecular Compounds

Picturing Molecules

- Molecules occupy three dimensional space.



Molecular Compounds

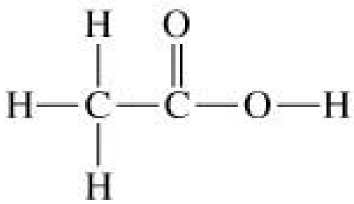


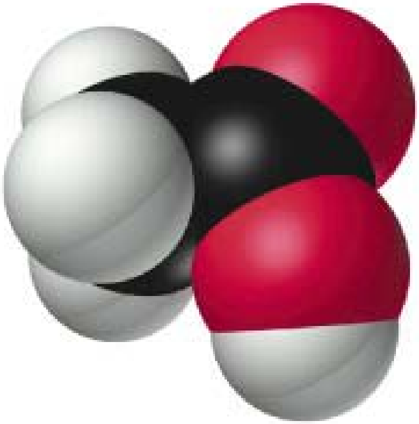
Molecular model:
("ball and stick")

Empirical formula: CH_2O

Molecular formula: $C_2H_4O_2$

Structural formula:

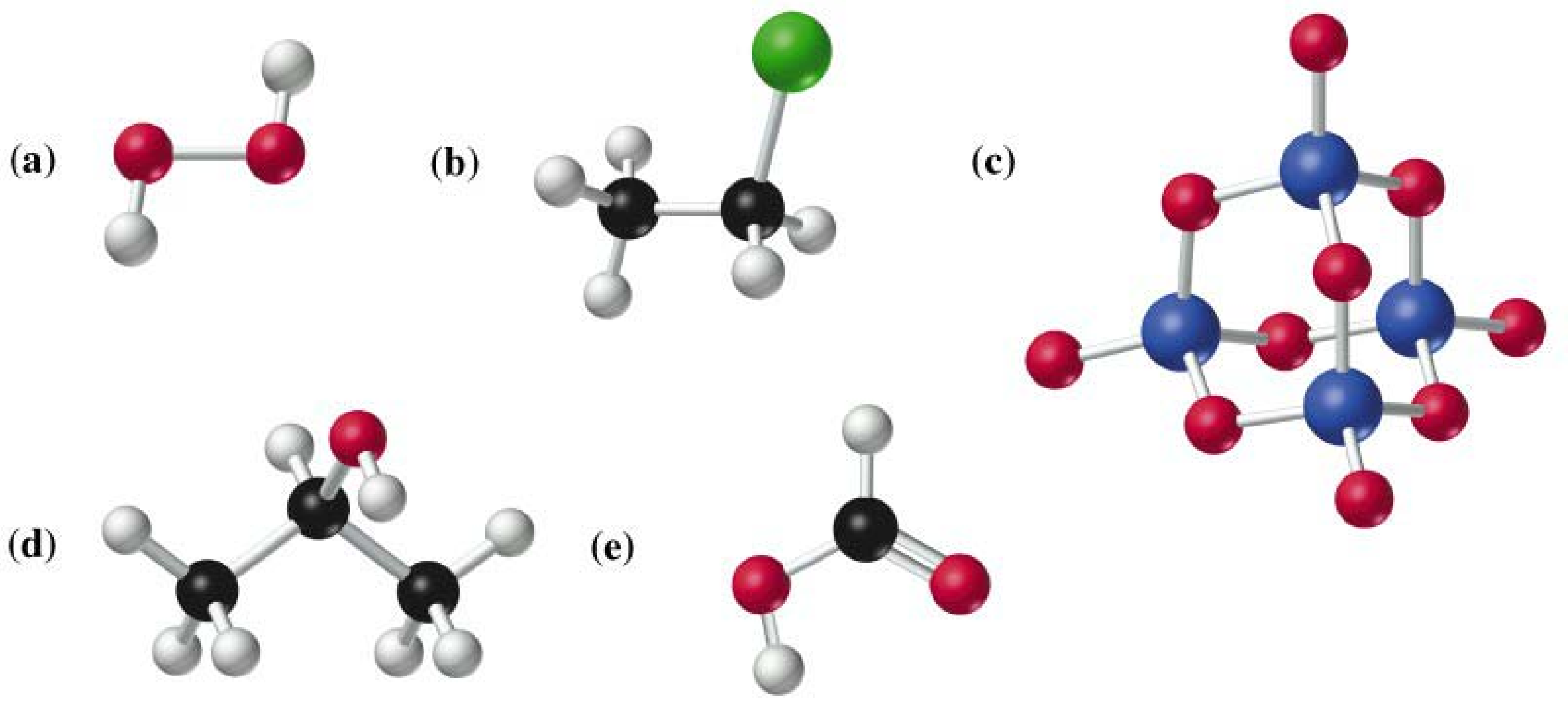




Molecular model:
("space-filling")

Some Molecules

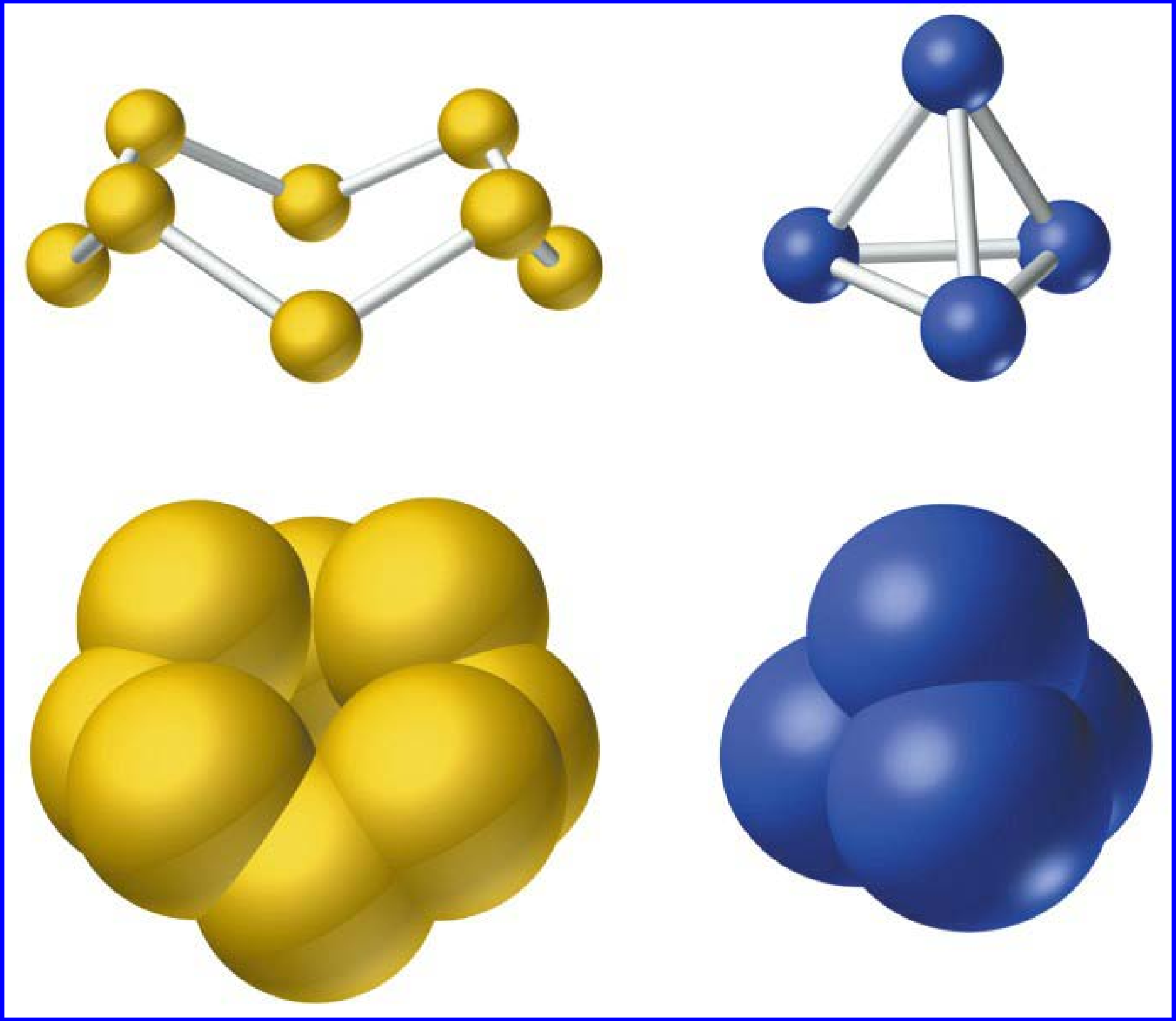
H_2O_2 CH_3CH_2Cl P_4O_{10}



$CH_3CH(OH)CH_3$ HCO_2H

Inorganic Molecules

S_8



P_4

Ions

Atoms are electronically neutral → they do not have a charge.
But an atom can lose or gain electrons (it cannot lose or gain protons…**WHY NOT?**)
An *ION* – is a particle with an unequal number of protons and electrons.

When an atom or molecule loses electrons, it becomes positively charged.
- For example, when Na loses an electron it becomes Na^{+}.

- Positively charged ions are called *cations*.

- When an atom or molecule gains electrons, it becomes negatively charged.
    - For example when Cl gains an electron it becomes Cl^{-}.
- Negatively charged ions are called *anions*.
- An atom or molecule can lose more than one electron.

Ionic compounds

- Atoms of almost all elements can gain *or* lose electrons to form charged species called *ions.*
- Compounds composed of ions are known as *ionic compounds.*

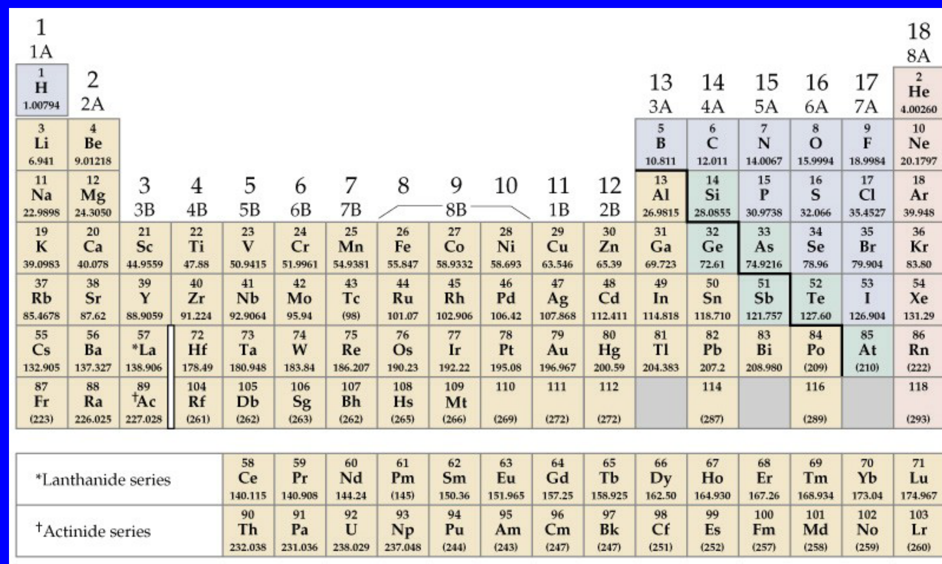
Whereas many metals form monatomic cations, only six nonmetallic elements commonly form anions

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

Oxidation States

Metals tend to lose electrons.

$Na \rightarrow Na^{+} + e^{-}$



Non-metals tend to gain electrons.

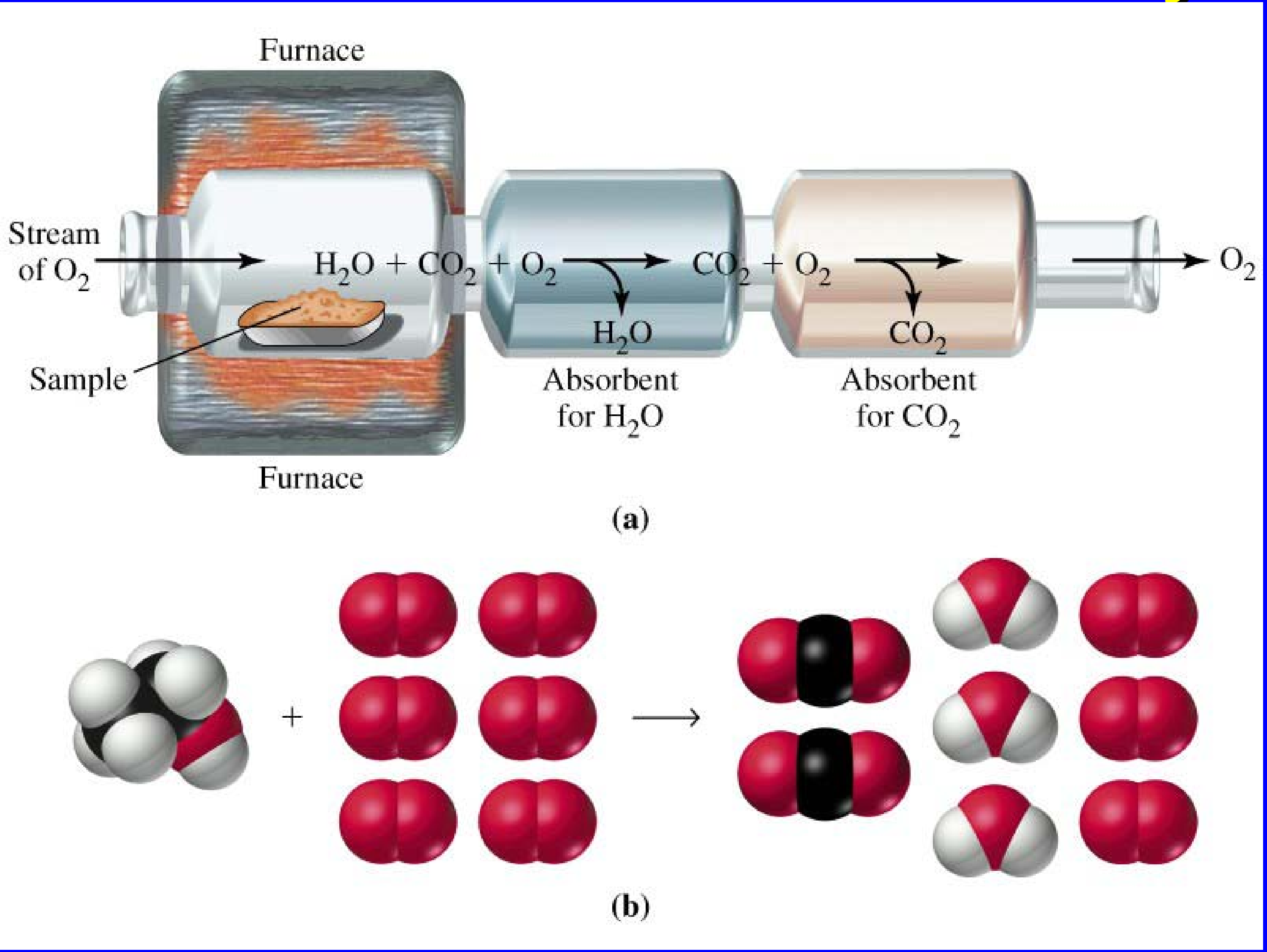
$Cl + e^{-} \rightarrow Cl^{-}$

Reducing agents

Oxidizing agents

We use the Oxidation State to keep track of the number of electrons that have been gained or lost by an element.

Determination of Empirical Formula Combustion Analysis



An Example of Determining the Empirical and Molecular Formulas of a Compound from Combustion Analysis Data.

Complete combustion of a 1.505 g sample of an unknown compound consisting of C, H and S yields 3.149 g CO_2, 0.645 g H_2O and 1.146 g of SO_2. What is the empirical formula for the unknown?

Step 1: Convert the mass of each product into moles of C, H and S.

3.149 g CO_2 x 1mol/44.010g CO_2 x 1mol C/1mole CO_2
= 0.07155 mol C

0.645 g H_2O x 1mol/18.02g H_2O x 2mol H/mol H_2O
= 0.0716 mol H

1.146 g of SO_2 x 1 mol/64.06 SO_2 x 1mol S/mol SO_2
= 0.01789 mol S

Example

Step 1: Convert the mass of each product into moles of C, H and S.

0.07155 mol C, 0.0716 mol H, 0.01789 mol S

Step 2: Write a tentative formula

$C_{0.07155}\ H_{0.0716}\ S_{0.01789} = C_4H_4S$

By the way:

0.07155 mol C = 0.859 g
0.0716 mol H = 0.0716 g
0.01789 mol S = 0.574 g
1.505 g

Example 2

Determining the Empirical and Molecular Formulas of a Compound from Its Mass Percent Composition.

When an unknown compound is decomposed into its constituent elements, it is found to contain 71.65% Cl, 24.27% C, and 4.07% H by mass. What is the empirical formula for the unknown?

Step 1: Determine the mass of each element in a 100g sample.

71.65 g Cl, 24.27 g C and 4.07 g H in the sample.

Example 2

Step 2: Convert masses to amounts in moles (molar mass).

$$71.65 \text{ g Cl} \times \frac{1 \text{ mol Cl}}{35.45 \text{ g Cl}} = 2.021 \text{ mol Cl}$$

$$24.27 \text{ g C} \times \frac{1 \text{ mol C}}{12.01 \text{ g C}} = 2.021 \text{ mol C}$$

$$4.07 \text{ g H} \times \frac{1 \text{ mol H}}{1.008 \text{ g H}} = 4.04 \text{ mol H}$$

Step 3: Write a tentative formula. $Cl_{2.021}C_{2.021}H_{4.04}$

Step 4: Convert to small whole numbers. $Cl_1C_1H_2$ or $ClCH_2$

Possible molecular formulas:
$Cl_2C_2H_4$ or $Cl_3C_3H_6$ or $Cl_4C_4H_8$ etc.

Example 2

Step 4: Convert to small whole numbers. $Cl_1C_1H_2$ or $ClCH_2$

Possible molecular formulas:
$Cl_2C_2H_4$ or $Cl_3C_3H_6$ or $Cl_4C_4H_8$ etc.

We experimentally determine the molar mass to be 98.96 g/mol. Determine the molecular formula of the unknown.

Empirical formula mass is 49.48 g/mol
This is half of the actual molar mass so the correct formula is

$Cl_2C_2H_4$

Naming Compounds

Binary Compounds of Metal/Nonmetal – (Ionic Compounds)

The metal element appears first and retains its elemental name.

The nonmetal element begins with a root derived from its element name and ends with the suffix –ide.

NaCl	sodium chloride
$MgBr_2$	magnesium bromide
Al_2O_3	aluminum oxide

Naming Compounds

Nonmetallic Binary Compounds – (Covalent Compounds)

The element that appears first retains its elemental name. The second element begins with a root derived from its element name and ends with the suffix –ide.

SO_2 sulfur dioxide

When there is more than one atom of a given element in the formula, the name of the element usually contains a prefix to specify the number of atoms present.

SO_3 sulfur trioxide
CO_2 carbon dioxide

Naming Compounds

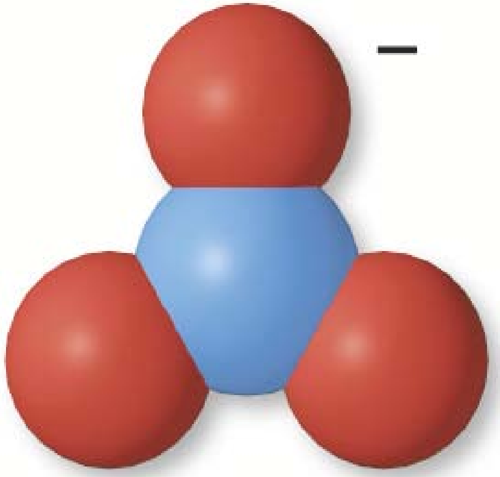
Number Prefixes for Chemical Names

Number	Prefix*	Example	Name
1	Mon(o)-	CO	Carbon monoxide
2	Di-	SiO_2	Silicon dioxide
3	Tri-	NI_3	Nitrogen triiodide
4	Tetr(a)-	CCl_4	Carbon tetrachloride
5	Pent(a)-	PCl_5	Phosphorus pentachloride
6	Hex(a)-	SF_6	Sulfur hexafluoride
7	Hept(a)-	IF_7	Iodine heptafluoride

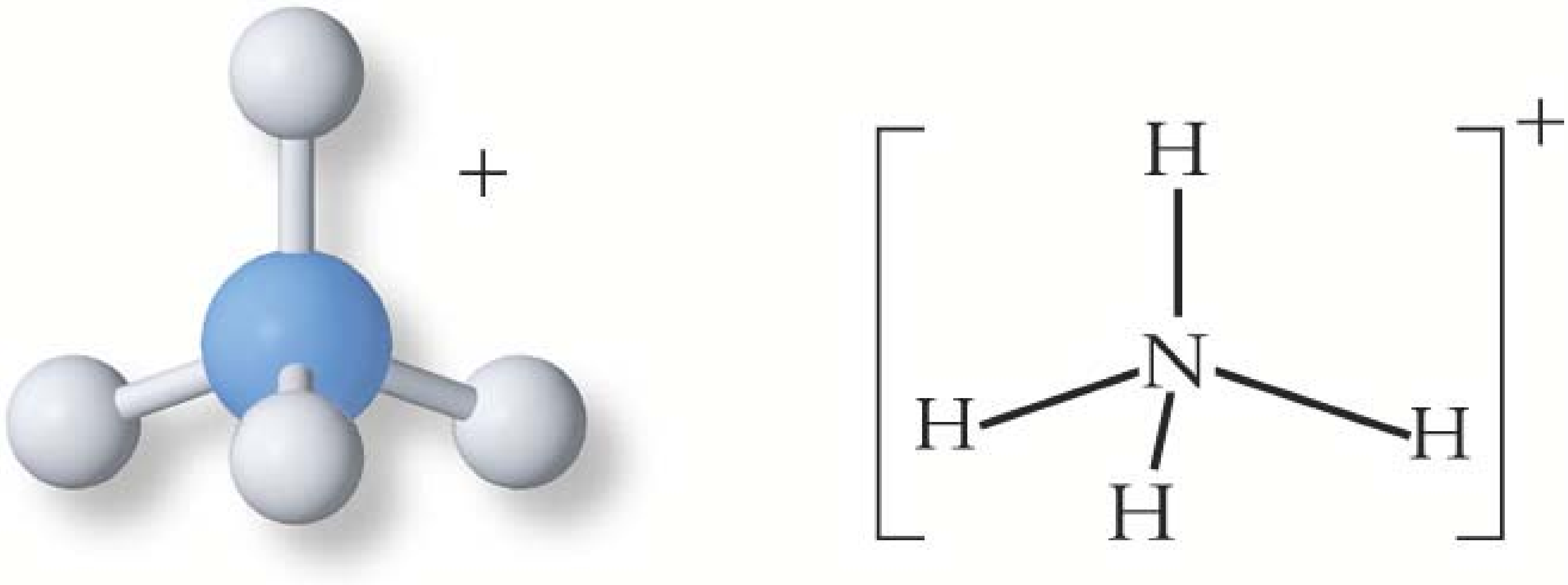
*If the numerical prefix ends with the letter "o" or "a" and the name of the element begins with a vowel, drop the last letter of the prefix.

Naming Compounds
Polyatomic ions

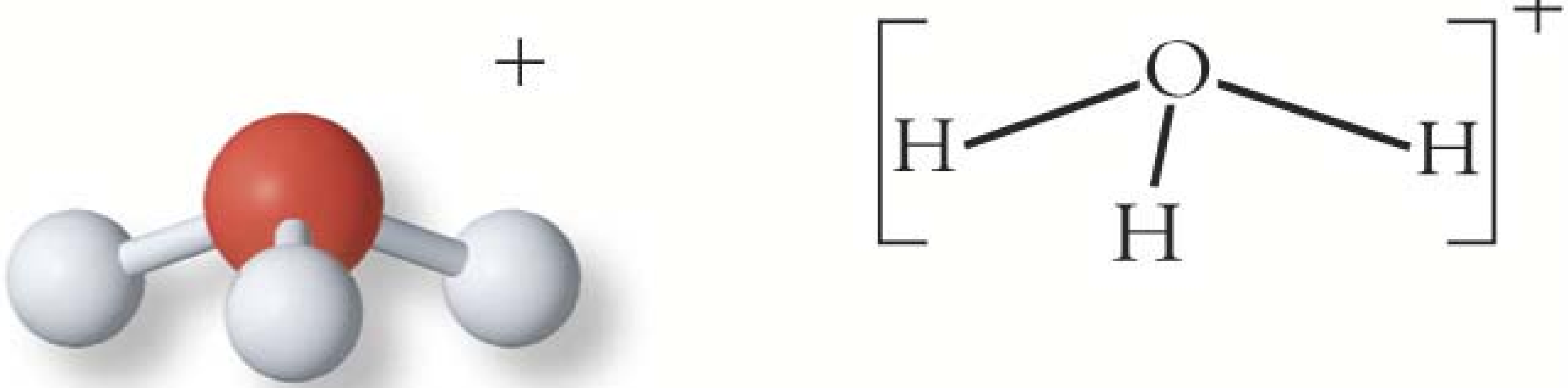
- A molecule with a charge is termed a polyatomic ion
- These ions can form ionic compounds with other ions.
- Therefore, some compounds can have both ionic and covalent bonds
- Example: $NaNO_3$



Nitrate anion
NO_3^-



Ammonium ion



Hydronium ion

Naming Compounds

Table 3.5 Common Polyatomic Ions

Formula	Name	Formula	Name
Cations		**Oxoanions**	
NH_4^+	Ammonium	SO_4^{2-}	Sulfate
H_3O^+	Hydronium	SO_3^{2-}	Sulfite
Hg_2^{2+}	Mercury(I)	NO_3^-	Nitrate
		NO_2^-	Nitrite
Diatomic Anions		PO_4^{3-}	Phosphate
OH^-	Hydroxide	MnO_4^-	Permanganate
CN^-	Cyanide	CrO_4^{2-}	Chromate
Anions with Carbon		$Cr_2O_7^{2-}$	Dichromate
CO_3^{2-}	Carbonate	ClO_4^-	Perchlorate
$CH_3CO_2^-$	Acetate	ClO_3^-	Chlorate
$C_2O_4^{2-}$	Oxalate	ClO_2^-	Chlorite
		ClO^-	Hypochlorite

Cations of Variable Charge

- When species involved form ions with more than one possible charge, it is necessary to state the charges of the compound.
- For example, copper forms compounds with Cu^{+} and Cu^{2+}.
- To indicate the charge in the name, a Roman numeral is added in parentheses.

  CuO = copper (II) oxide

  Cu_2O = copper (I) oxide
- This convention is ONLY used if multiple possible charges exist for a cation.