Chapter 2: Atoms, Molecules, and Ions (Ch2 Chang, Chs 0 and 2 Jespersen)

Atoms are the basic building blocks of matter.

They are the smallest particles of an element that retain the chemical identity of the element, or, the basic unit of an element that can enter into chemical combination.

In 1808, Dalton introduced the term "*atoms*", and his Atomic Theory can be summarized by these 4 points:

1) Elements are composed of extremely small particles, called atoms.

2) All atoms of a given element are *identical*. The atoms of one element are different from atoms of a different element. (He did not say how).

3) Compounds are composed of atoms of *more than one element*. (Supported the Law of Definite Proportions, and predicted the Law of Multiple Proportions).

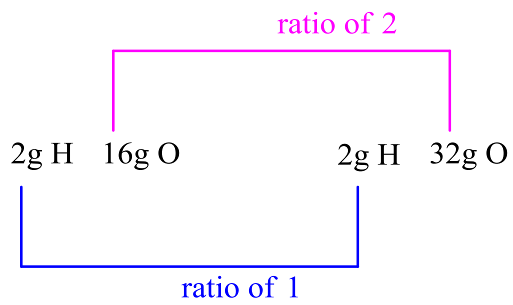
4) A chemical reaction only involves the separation, combination or rearrangement of *atoms*, but NOT their creation or destruction. (Supported the conservation of mass – that matter cannot be created or destroyed).

Proust's Law of Definite Proportions (1799) – different samples of the same compound always contain its constituent elements in the same proportion by mass.

(Any sample of CO_2 gas, from any source, would have the same ratio by mass of Carbon to Oxygen).

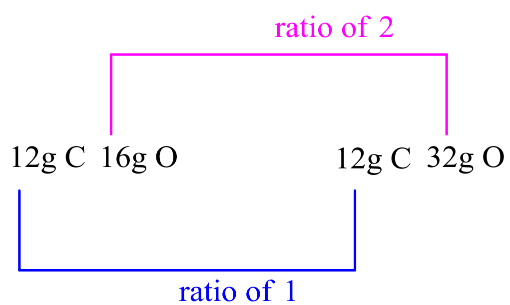
Dalton's Law of Multiple Proportions (1808) - if two elements A and B combine to form more than one compound, the masses of B that can combine with a given mass of A are in a ratio of small whole numbers.

For example: **H_2O** **H_2O_2**



H=1, O = 16 g/mol

CO **CO_2**



C=12, O = 16 g/mol

Notice we are **not** saying that the ratio will *come out* to be whole number (e.g. can get 5/3 for PCl_3 and PCl_5), but it must *involve* whole numbers.

Alternatively, Chemical formulas do not have fractions.

The Structure of Atoms

Dalton claimed atoms are indivisible and indestructible, but starting around 1850, people started to identify **sub-atomic** particles.

Nowadays we know atoms are composed of subatomic particles (electrons, protons and neutrons).

Faraday in 1834 passed electricity through an aqueous solution, and brought about chemical changes (which implies that chemicals are *related* to electricity).

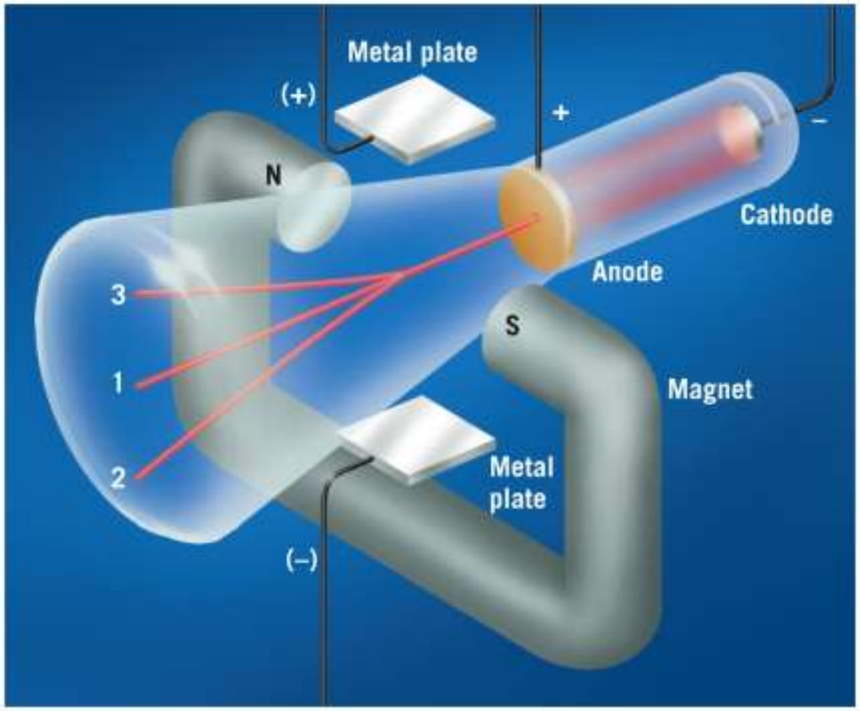
There are two types of electrical charge, positive (+) (+ve) and negative (-) (-ve).

Law of Electrostatic Attraction (1784): like charges repel one another, unlike (opposite) charges attract.

Thomson is credited with discovering the electron in 1897.

By passing current through a gas at low pressure (in a gas discharge tube), he generated species (*cathode rays*) that had very low mass (much less than atoms) and were negatively charged.

Cathode ray Tube and Electrons



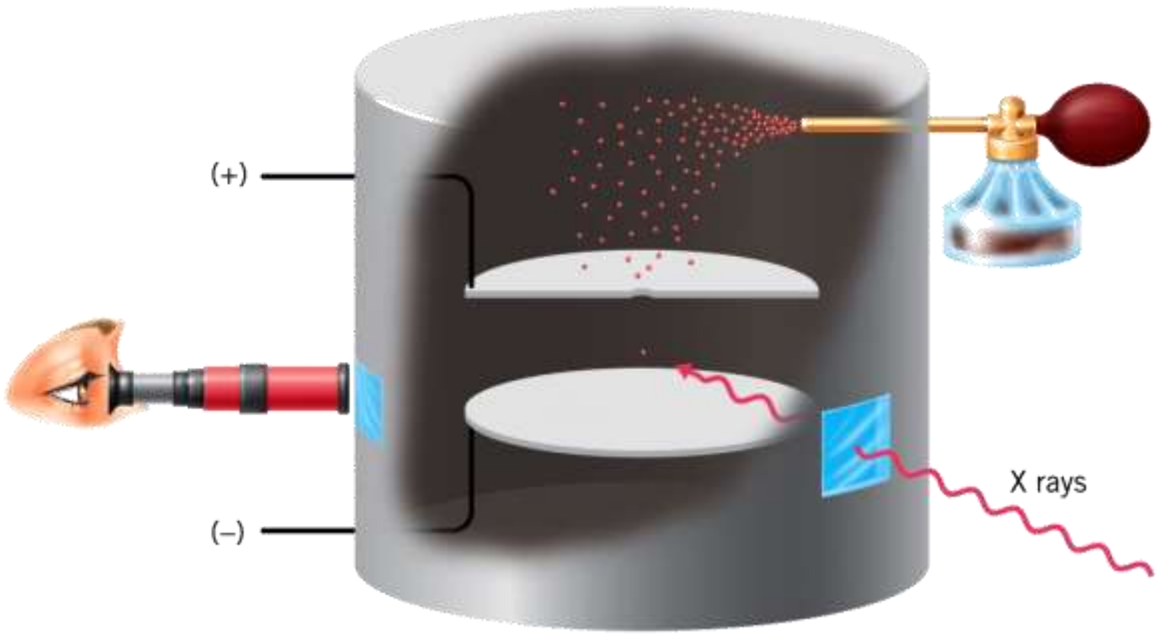
The cathode rays (cathode ray particles) were deflected by electric and magnetic fields.

Thomson obtained values of "charge to mass" for these species as1.76 x10^8 coulombs per gram.

They were negatively charged.

The same particles were produced using any (different) gas, implying they are *fundamental particles*.

The Millikan oil-drop experiment in 1909 established the charge of an electron as 1.6×10^{-19} C.



Combining the results of "charge per mass" and "charge", we can determine the mass of these species (electrons),

$$\text{Mass} = \frac{1.60 \times 10^{-19}\ \text{C}}{1.76 \times 10^{8}\ \text{C/g}} = 9.10 \times 10^{-28}\ \text{g}$$

Emission of electrons can also be described as a form of *radiation*.

Radiation - the emission and transmission of energy through space in the form of waves.

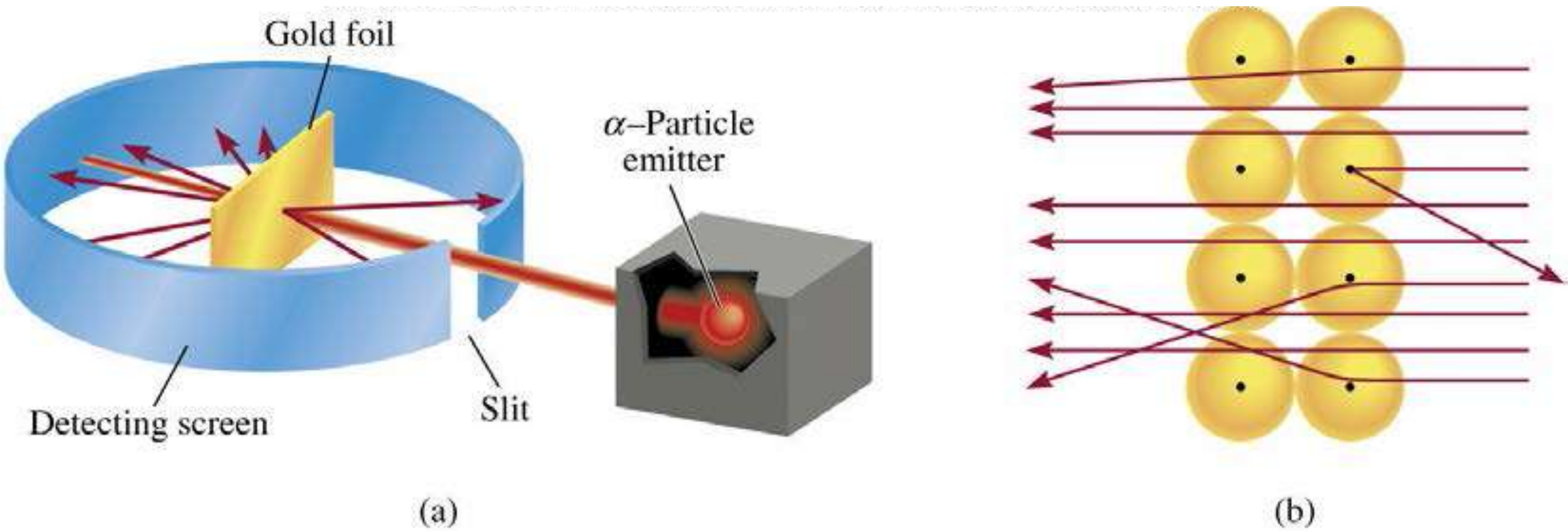
Radioactivity - spontaneous emission of particles and/or radiation.

Name	Symbols	Charge	mass (g/particle)
alpha particles	${}^{4}_{2}He$, ${}^{4}_{2}\alpha$	2+	6.65×10^{-24}
beta particle	${}^{0}_{-1}e$, ${}^{0}_{-1}\beta$	1-	9.11×10^{-28}
gamma ray	${}^{0}_{0}\gamma$, γ	0	0

The alpha particles played a key role in further probing the atomic structure.

The Structure of Atoms: Protons and Neutrons

The gold foil experiments by Rutherford between 1908 and 1913 established the modern picture of an atom.



Almost all the mass of an atom is in a small region at the center of the atom (*nucleus*).

They were able to determine the mass and charge of nuclei via the *deflection*.

But the number of protons they predicted (by charge) only accounted for about half of the mass of the nucleus.

This implied other non-charged particles (neutrons).

In 1932 Chadwick is credited with discovering the Neutron.

When discussing the mass of atoms we will use the **atomic mass unit** (**amu**).

1 amu is 1.66054 x10^{-24} g.

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

Most of the mass of an atom is in the nucleus (protons and neutrons).

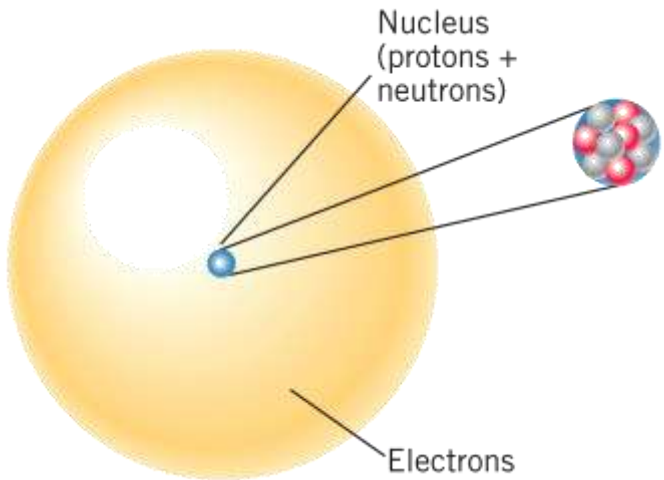
The size of atoms is small!

Atomic diameters are on the order of 1x10^{-10} m to 5x10^{-10} m, which also can be expressed as 100 - 500 pm.

In chemistry, a convenient (but non- SI) unit to express atomic diameters is the **angstrom** (**Å**).

1 Å = 10^{-10} m.

So atoms are around 1 - 5 Å in diameter.



Isotopes, Atomic Numbers and Mass Numbers

All atoms of an element have the same number of protons in the nucleus.

It is the number of protons that determines the type of atom (Elements).

In an atom, the number of electrons equals the number of protons (*electronically neutral*).

We use this general formula to describe an element, X:

$${}^{A}_{Z}X$$

Where: Z = Atomic Number (number of protons in the nucleus). This subscript is often omitted as the element (atomic symbol) defines it.

A = Mass Number (this *superscript* is the total number of protons and neutrons).

E.g. Carbon $\quad {}^{12}_{6}C \quad\quad {}^{12}C$

Atoms of the *same element* that differ in the number of neutrons (and therefore mass) are called isotopes.

$^{12}_{6}C$	^{12}C	"Carbon 12"	6 Protons, 6 Neutrons
$^{14}_{6}C$	^{14}C	"Carbon 14"	6 Protons, 8 Neutrons

$^{107}_{47}Ag$ (Silver with 60 Neutrons)

$^{109}_{47}Ag$ (Silver with 62 Neutrons)

An atom of a specific isotope is called a nuclide.

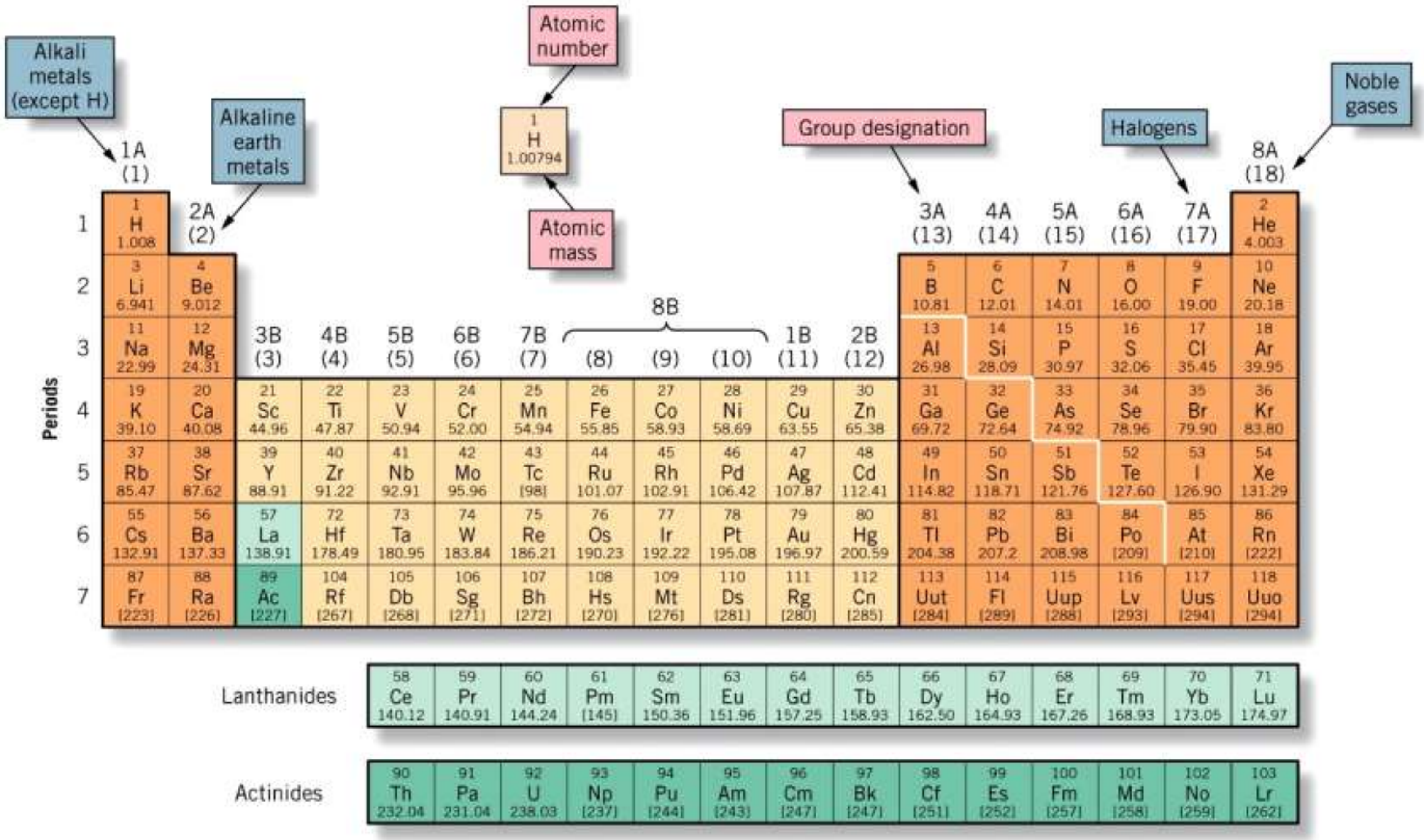
Differences in *elements* (and *isotopes*) are due to the *differences* in the *number* of subatomic particles.

(Chemists explain things at the atomic or molecular level).

Now we are ready to address the Elements…

Elements and the Periodic Table

Horizontal Rows are called Periods; vertical Columns are called Groups.



It is called a *Periodic* Table since it summarizes periodic (*repeating*) physical and/or chemical properties of elements.

Mendeleev (Russian) and Meyer (German) in 1869 are credited with this type of description.

- Noted repeating properties
- Arranged by increasing atomic mass

Rows are called periods. Often refer to elements as first, second, third, … row/period elements.

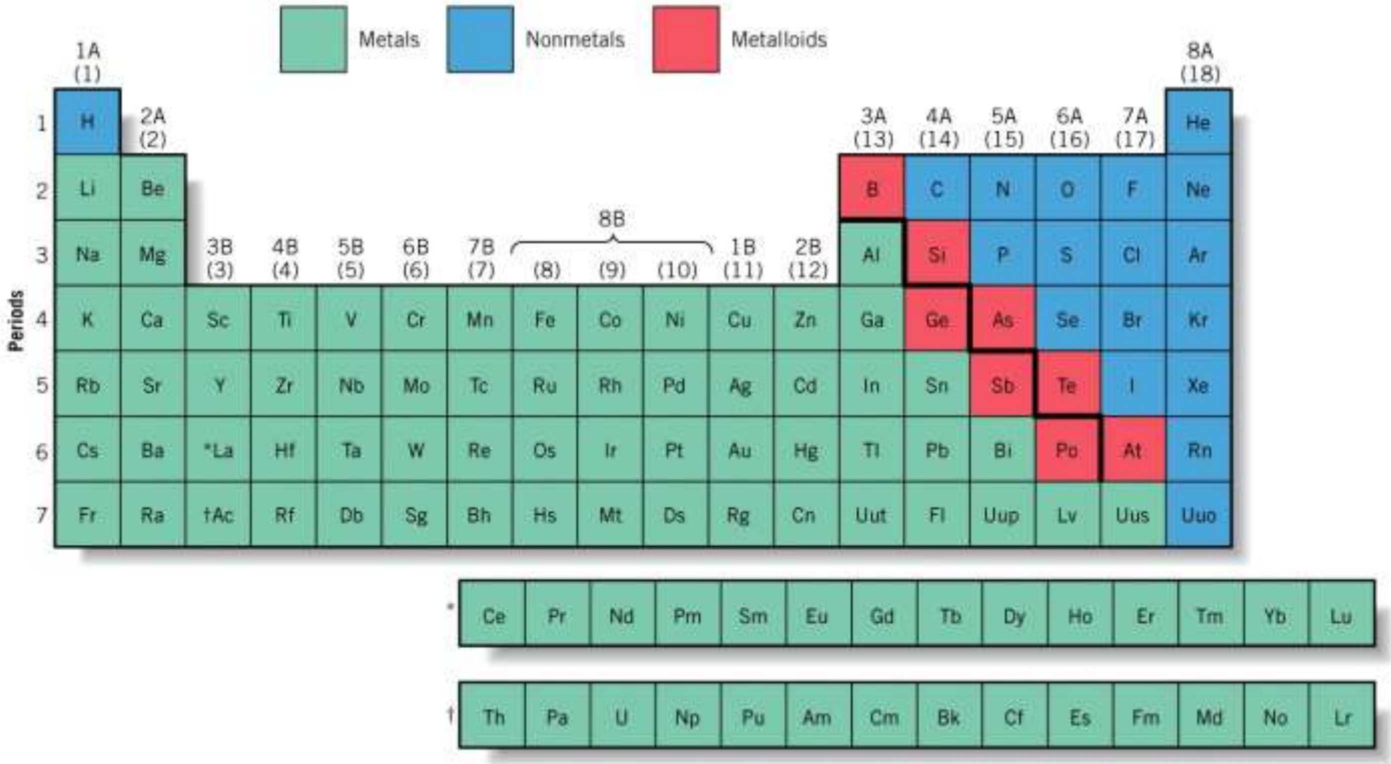
Columns are called groups or families.

- Identified by numbers
- 1 – 18 standard international (1A – 8A longer columns and 1B – 8B shorter columns).

Elements in the longer columns (the A groups) are called Representative Elements, or Main Group Elements.

Transition metals are located in the shorter (B) columns; Inner transition metal groups are the *Lanthanides* and *Actinides*.

Metals, Non-Metals and Metalloids



The elements can also be categorized into **three** other classifications (and their *crude* definitions):

Metal (a *good* conductor of heat and electricity);

Non-metal (usually a *poor* conductor of heat and electricity);

Metalloid (intermediate properties between those of a metal and non-metal).

Molecules, Ions and Chemical Bonds

A molecule is an assembly of *two* or *more atoms* tightly bound together.

Molecules and Chemical Formulas

Chemical formulas tell us the composition substances (*elements*).

The subscripts in the formula tell us the *number* of that type of atom present in the molecule.

E.g. O_2 **two** oxygen atoms

O_3 three oxygen atoms

H_2O two hydrogen atoms and one oxygen atom

Molecules containing two atoms are called diatomic. Elements that **occur** as diatomic molecules include N_2, O_2, H_2, F_2, Cl_2, Br_2 and I_2.

When we speak of these elements we are referring to the diatomic form listed above.

Be careful, a Bromine *atom* (Br) is different to *elemental* Bromine (Br_2).

Molecular compounds are compounds that are composed of *molecules*.

Most molecular substances that we will encounter in this course contain only nonmetals.

Molecular and Empirical Formulas

There is a difference!

Molecular formulas indicate the *actual* number and types of atoms in a molecule.

Empirical formulas give only the *relative* number of atoms of each type.

The subscripts are always the smallest whole number ratio.

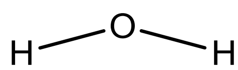
E.g.	**MF**	**EF**
	H_2O_2	HO
	C_2H_4	CH_2
	C_6H_{12}	CH_2
	H_2O	H_2O

(What is the point of EF's ?

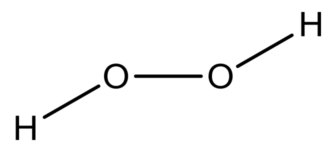
Often *experimental* results will provide the EF. Additional information is often needed to determine the actual MF).

Picturing Molecules

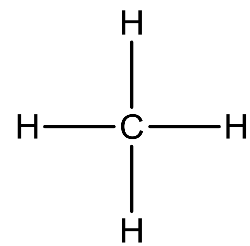
You will encounter many different ways to *indicate* which atoms are attached to which within a molecule.



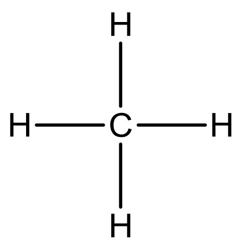
Water



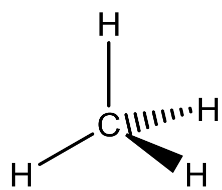
Hydrogen Peroxide



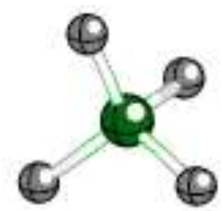
Methane



Lewis Structure
(Structural Drawing)



Sticks and Wedges
(Perspective Drawing)



Ball and Stick

Space Filling

Addition or *removal* of electrons from a neutral atom results in the formation of a charged particle called an Ion.

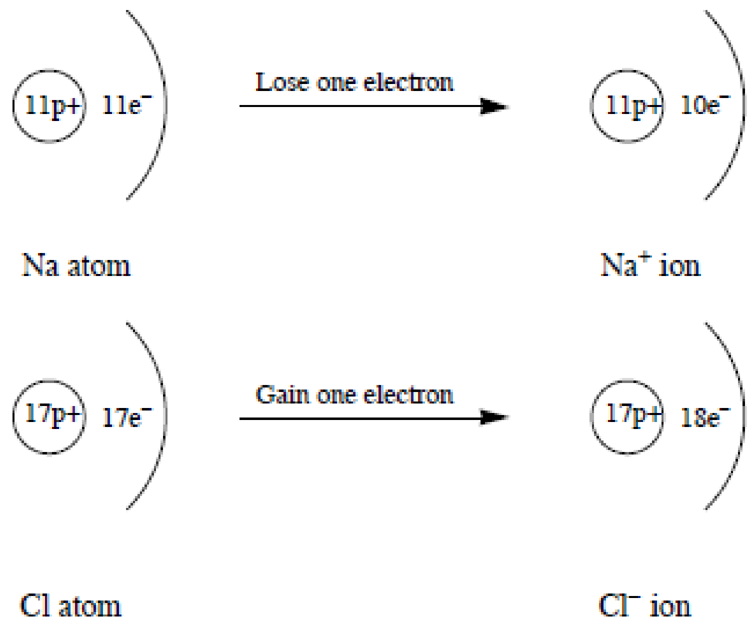
An ion with positive charge is called a cation.

A negatively charged ion is called an anion.

The net charge is represented by a *superscript*.

Superscripts +, 2+, and 3+ mean a net charge resulting from the *loss* of one, two, or three electrons.

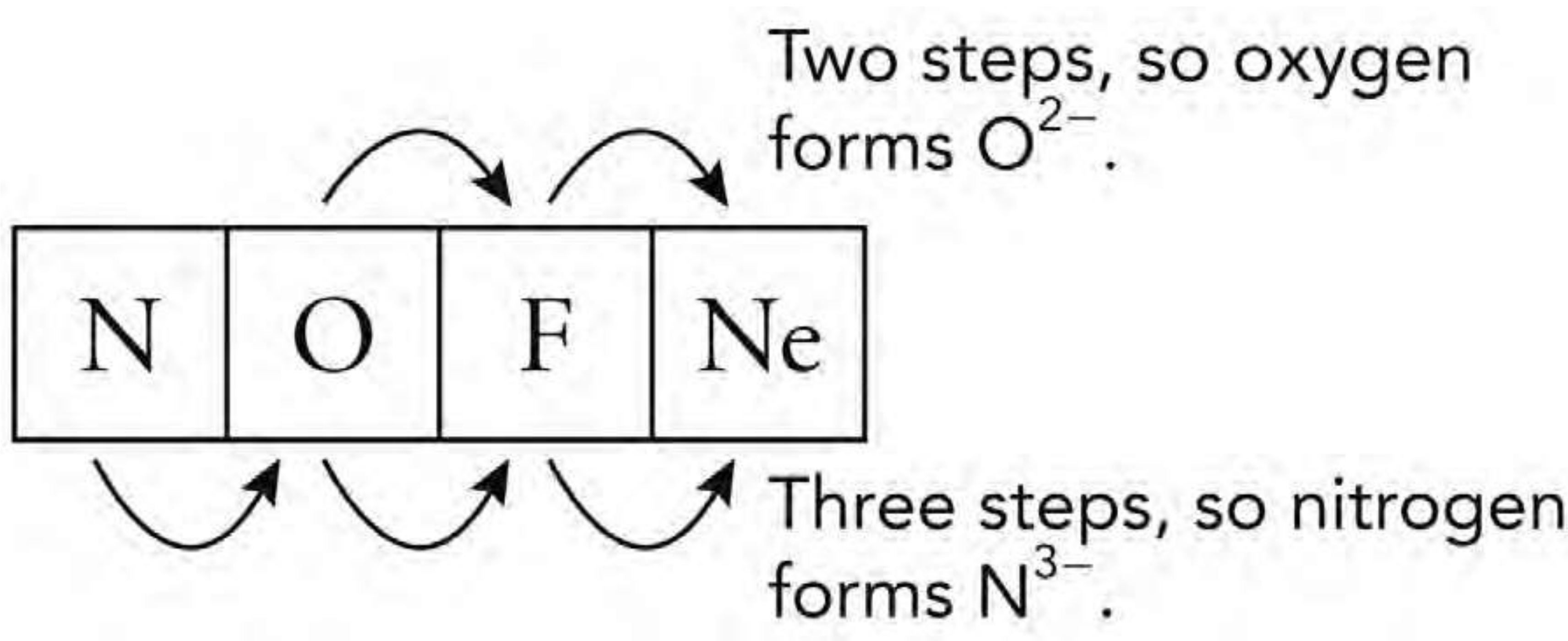
Superscripts −, 2−, and 3− mean a net charge resulting from the *gain* of one, two, or three electrons.



In general, metal atoms lose electrons (→ cations), whereas nonmetals tend to gain electrons (→ anions).

Predicting Ionic Charges

Many atoms *gain* or *lose* electrons so as to end up with as many electrons as the *closest* noble gas.



The Periodic Table helps you *predict* ion formation:

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^+													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+}				

Group 1A atoms form 1+ ions

Group 2A atoms form 2+ ions

Group 7A atoms form 1– ions

Group 6A atoms form 2– ions

Naming Chemical Compounds

There are four common types of Inorganic Compounds:

- ionic
- molecular
- acids and bases
- hydrates.

Ionic compounds are generally combinations of metals with nonmetals.

Molecular compounds are generally composed of nonmetals only.

Then there are also Organic Compounds – compounds that contain *chains o*f connected Carbon atoms.

Inorganic compounds are any compounds that are **not** Organic.

Ionic Compounds

Ionic compounds contain positively charged ions and negatively charged ions (= salts).

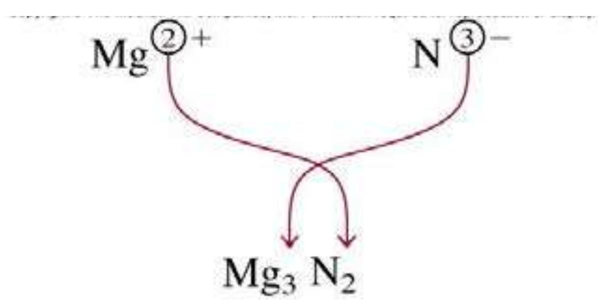
Generally, cations are *metal* ions and anions are *nonmetal* ions.

Only empirical formulas can be written for most ionic compounds.

These are given such that the **total** positive charge *equals* the **total** negative charge.

NaCl	Na^+	Cl^-	+1 and −1 make neutral
$BaCl_2$	Ba^{2+}	Cl^- (Two of them)	+2 and −1 and −1 (= −2) make neutral
Mg_3N_2	Mg^{2+} (Three of them)	N^{3-} (Two of them)	+6 and −6 make neutral

Sometimes this "*cross over*" technique can provide a shortcut to the formula:



Names and Formulas of Ionic Compounds

They are traditionally written with the cation first, then the anion.

Part 1. **Positive Ions (Cations)**

a. Cations formed from metal atoms have the *same name* as the metal.

b. If a metal can form cations of *differing charges*, the positive charge is given by a Roman numeral in parentheses following the name of the metal (called the Stock system – published in 1919):

Fe^{2+} Iron (II) (olden days – *Ferrous*)

Fe^{3+} Iron (III) (olden days – *Ferric*)

c. Cations formed from nonmetals atoms have names that end in **-ium**:

NH_4^+ Ammonium ion

Part 2. **Negative Ions (Anions)**

a. Monatomic (*one-atom*) anions have names formed by dropping the ending of the name of the element and adding the ending **-ide**.

H^-	Hydride	C^{4-}	Carbide	N^{3-}	Nitride	O^{2-}	Oxide	F^-	Fluoride
		Si^{4-}	Silicide	P^{3-}	Phosphide	S^{2-}	Sulfide	Cl^-	Chloride
				As^{3-}	Arsenide	Se^{2-}	Selenide	Br^-	Bromide
						Te^{2-}	Telluride	I^-	Iodide

b. Polyatomic (*many-atom*) anions containing oxygen have names ending in **-ate** or **-ite**.

(-*ates* have **more O** than –*ites*)

E.g.	NO_2^-	Nitrite	NO_3^-	Nitrate
	SO_3^{2-}	Sulfite	SO_4^{2-}	Sulfate
	PO_4^{3-}	Phosphate		
	CO_3^{2-}	Carbonate		

c. Anions derived by adding H^+ to an *oxyanion* are named by adding as a prefix the word **hydrogen** or **dihydrogen**, as appropriate.

E.g.	HCO_3^-	Hydrogencarbonate		
	HSO_4^-	Hydrogensulfate		
	HPO_4^{2-}	Hydrogenphospate	$H_2PO_4^-$	Dihydrogenphosphate

An oxyanion (*oxoanion*) is an ion with the generic formula $\mathbf{A_xO_y^{z-}}$ such as CO_3^{2-}, or SO_4^{2-}.

Polyatomic cations and anions you should know:

Ion	Name (Alternate name in parentheses)	Ion	Name (Alternate name in parentheses)
NH_4^+	Ammonium ion	CO_3^{2-}	Carbonate ion
Hg_2^{2+}	Mercury(I) ion	HCO_3^-	Hydrogen carbonate ion (bicarbonate ion)[b]
H_3O^+	Hydronium ion[a]	SO_3^{2-}	Sulfite ion
OH^-	Hydroxide ion	HSO_3^-	Hydrogen sulfite ion (bisulfite ion)[b]
CN^-	Cyanide ion	SO_4^{2-}	Sulfate ion
NO_2^-	Nitrite ion	HSO_4^-	Hydrogen sulfate ion (bisulfate ion)[b]
NO_3^-	Nitrate ion	SCN^-	Thiocyanate ion
ClO^- or OCl^-	Hypochlorite ion	$S_2O_3^{2-}$	Thiosulfate ion
ClO_2^-	Chlorite ion	CrO_4^{2-}	Chromate ion
ClO_3^-	Chlorate ion	$Cr_2O_7^{2-}$	Dichromate ion
ClO_4^-	Perchlorate ion	PO_4^{3-}	Phosphate ion
MnO_4^-	Permanganate ion	HPO_4^{2-}	Monohydrogen phosphate ion
$C_2H_3O_2^-$	Acetate ion	$H_2PO_4^-$	Dihydrogen phosphate ion
$C_2O_4^{2-}$	Oxalate ion		

[a]You will only encounter this ion in aqueous solutions.

[b]You will often see and hear the alternate names for these ions.

Part 3: **Put them together**

Ionic compounds are the cation name *followed* by the anion name.

If there is more than one *polyatomic* ion present in the compound, the ion formula is placed in parentheses followed by a subscript indicating the number of ions present.

E.g.	**$BaBr_2$**	barium bromide
	Al**$(NO_3)_3$**	aluminum nitrate
	$CuCO_3$	copper (II) carbonate (or *cupric* carbonate)
	$NaHCO_3$	sodium hydrogencarbonate (or sodium *bicarbonate*)

Molecular Compounds

Binary Molecular Compounds (*Binary* meaning *two* different elements)

Rules for Naming:

1. The name of the element farthest to the left in the periodic table is written first.
2. If both elements are in the *same* group in the periodic table, the lower one is named first.
3. The name of the *second* element is given an **-ide** ending.
4. Greek prefixes (*mono*, *di*, *tri*, *tetra*, *penta*, etc.) are used to indicate the number of atoms of each element.

(The prefix *mono*- is never used with the first element but is with the second. When the prefix ends in **a** or **o**, and the name of the second element begins with a vowel (such as oxide) the **a** or **o** is often dropped).

E.g.	NF_3	nitrogen trifluoride
	P_2O_5	diphosphorous pentoxide
	CO	carbon monoxide
	SF_6	sulfur hexafluoride
	S_2F_{10}	disulfur decafluoride

Acids and Bases

Names and Formulas of Acids

Acid **-** a substance that yields *hydrogen ions* (H^+) when dissolved in water.

1. Acids Based on Anions Whose Name End in **-ide**.

Anions whose names end in **-ide** have associated acids that have the prefix **hydro-** and an **-ic** ending.

E.g.

$$\underset{\substack{\text{hydrogen chloride} \\ \text{(molecular)}}}{\text{HCl (g)}} \xrightarrow{H_2O} \underset{\text{hydrochloric acid}}{\text{HCl (aq)}}$$

Anion	**Corresponding Acid**	**Anion**	**Corresponding Acid**
Cl^- (chloride)	HCl (hydrochloric acid)	F^- (fluoride)	HF (hydrofluoric acid)
S^{2-} (sulfide)	H_2S (hydrosulfuric acid)	CN^- (cyanide)	HCN (hydrocyanic acid)

2. Acids Based on Anions Whose Names End in **-ate** or **-ite**.

Anions whose names end in **–ate** have associated acids with an **-ic** ending,

whereas anions whose names end in **-ite** have acids with an **-ous** ending.

Anion	**Acid**
Hypochlorite, ClO^{-}	Hypochlorous acid, $HClO$
Chlorite, ClO_2^{-}	Chlorous acid, $HClO_2$
Nitrite, NO_2^{-}	Nitrous acid, HNO_2
Chlorate, ClO_3^{-}	Chloric acid, $HClO_3$
Nitrate, NO_3^{-}	Nitric acid, HNO_3
Perchlorate, ClO_4^{-}	Perchloric acid, $HClO_4$

(*Hypo* comes from the Greek word for "under", *per* is the Latin for "all over", relating to the element's ability to combine to oxygens).

If all of the hydrogens are **not** removed, you must indicate the number of hydrogens remaining.

E.g.	H_3PO_4	phosphoric acid	(remember –*ate* → –*ic acid*)
	NaH_2PO_4	sodium dihydrogen phosphate	
	Na_2HPO_4	sodium hydrogen phosphate	
	Na_3PO_4	sodium phosphate	

Bases

Bases - substances that yield hydroxide ions (OH^-) when dissolved in water.

E.g.	Sodium hydroxide	$NaOH$	(which is actually Na^+ and OH^-)
	Potassium hydroxide	KOH	
	Barium hydroxide	$Ba(OH)_2$	
	Ammonia	NH_3	

Since $NH_3(g) + H_2O(l) \rightarrow NH_4^+(aq) + OH^-(aq)$

Hydrates

Hydrated ionic compounds (i.e. hydrates) have *a specific number of water molecules* in their chemical formulas.

In the solid, these water molecules (also called "*waters of hydration*") are part of the structure of the compound.

To name the Hydrates:

1. The ionic compound (*without* the waters of hydration) is named first by using the rules for naming ionic compounds.

2. Greek prefixes are attached to the word "*hydrate*" to indicate the number of water molecules per formula unit for the compound.

3. When the chemical formula for a hydrated ionic compound is written, the formula for the ionic compound is separated from the waters of hydration by a centered "dot".

$BaCl_2 \cdot 2H_2O$	barium chloride dihydrate
$MgSO_4 \cdot 7H_2O$	magnesium sulfate heptahydrate

Organic compounds

These are molecules notable for their *chains* of consecutive Carbon atoms.

There are very, very many, but the simplest are the Hydrocarbons (only contain *Carbon* and *Hydrogen*).

Alkanes: C_nH_{2n+2} (**single** bonds between carbons)

Alkenes: C_nH_{2n} (single and **double** bonds between carbons)

Alkynes: C_nH_{2n-2} (single and **triple** bonds between carbons)

Simple Alkanes – They are named according to the *number* of contained Carbon atoms.

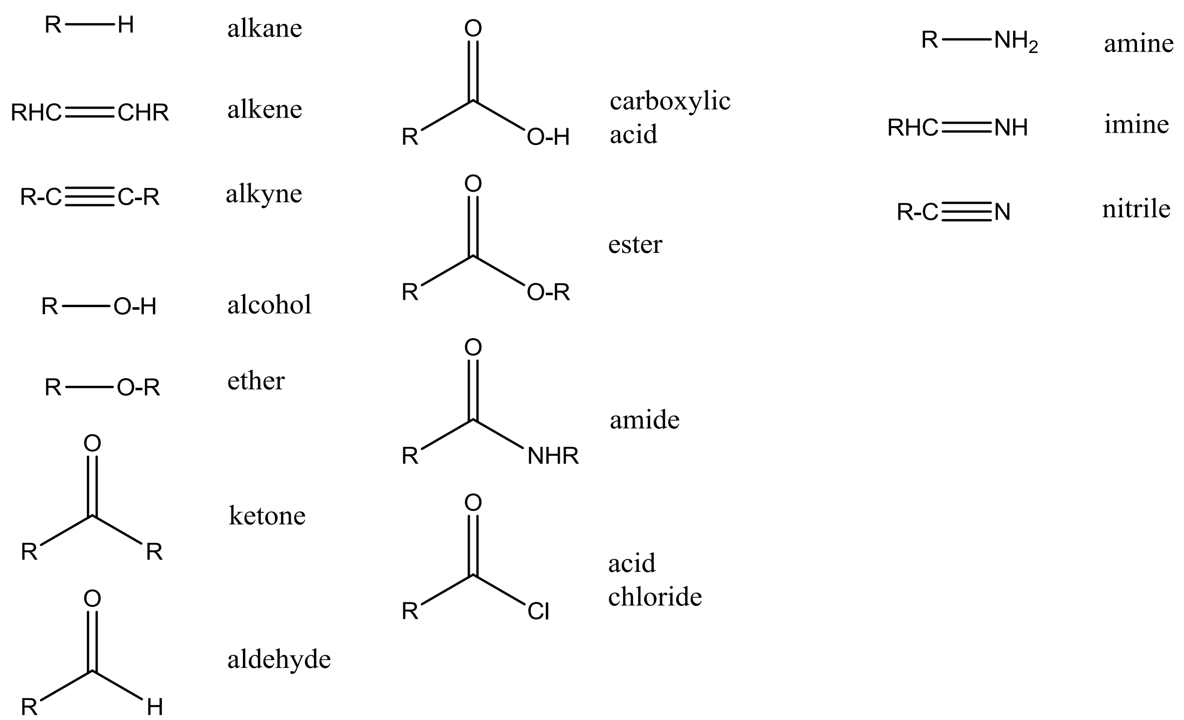
Number of carbons	Name	formula
1	Methane	CH_4
2	Ethane	CH_3CH_3
3	Propane	$CH_3CH_2CH_3$
4	Butane	$CH_3CH_2CH_2CH_3$
5	Pentane	$CH_3(CH_2)_3CH_3$
6	Hexane	$CH_3(CH_2)_4CH_3$
7	Heptane	$CH_3(CH_2)_5CH_3$
8	Octane	$CH_3(CH_2)_6CH_3$
9	Nonane	$CH_3(CH_2)_7CH_3$
10	Decane	$CH_3(CH_2)_8CH_3$

Functional Groups

All organic molecules belong to certain classes or families, as determined by their functionality (*reactive parts*).

These are unique bonding permutations found in different organic molecules (*structural motifs*).

These are some of the most common functional groups:



R = alkyl group, CH_3-, CH_3CH_2-, etc.

Naming Summary flowchart