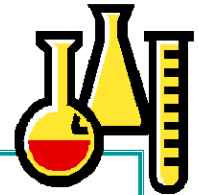

Chapter 2

Atoms, Molecules, and Ions

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Preview

- Reviewing the fundamental chemical laws.
- Understanding the modern view of atomic structure, molecules, and ions.
- Being familiar with main groups of elements in the periodic table.
- Naming simple compounds, ionic compounds, and getting the chemical formulas from the names.

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The Atomic Theory

In 1808, John Dalton presented his theory.

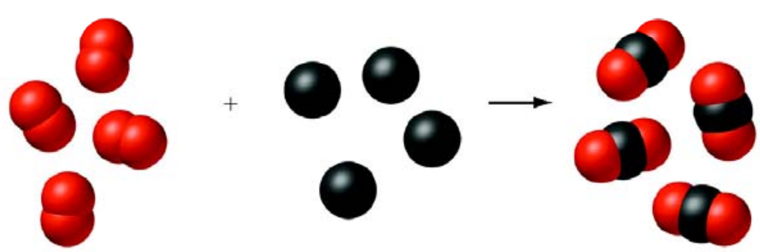
1. An element is made up from extremely small particles called ***atoms***. Atoms of a given element are identical, but are different for different elements.

   Dalton had no idea what an atom would look like!

2. Elements combine to form ***chemical compounds***.
3. A ***chemical reaction*** involves rearrangement of atoms; it doesn't create or destroy them.

Atoms » Elements » Molecules (Compounds)

The Atomic Theory



Combination of oxygen and carbon to form carbon dioxide

Chapter 2 Section 1

Law of Definite Proportion

- By Joseph Proust.

Different samples of a given compound always contain the same elements in the ***same mass ratio***.

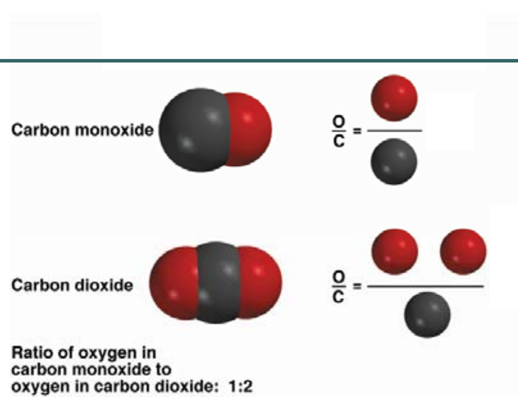
Sample	Mass of O (g)	Mass of C (g)	Ratio (g O : g C)
123 g carbon dioxide	89.4	33.6	2.66:1
50.5 g carbon dioxide	36.7	13.8	2.66:1
88.6 g carbon dioxide	64.4	24.2	2.66:1

Sample	Mass of O (g)	Mass of C (g)	Ratio (g O : g C)
16.3 g carbon monoxide	9.31	6.99	1.33:1
25.9 g carbon monoxide	14.8	11.1	1.33:1
88.4 g carbon monoxide	50.5	37.9	1.33:1

Chapter 2 Section 1

What Did Dalton Observe in CO Molecules?

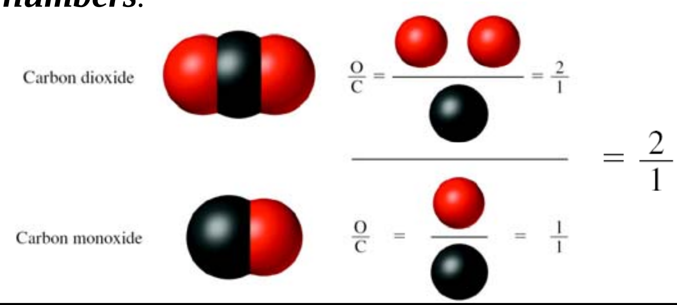
Mass of oxygen that combines with 1g of carbon		Ratio of mass of oxygen that combines with 1g of carbon
1.33g		$\frac{1}{1}$ or $\frac{2}{1}$ etc.
2.66g		$\frac{2}{1}$ or $\frac{4}{1}$ etc.



Law of Multiple Proportions

- By John Dalton.

  If two elements can combine to form more than one compound with each other, the masses of one element that combine with a fixed mass of the other element are in ***ratios of small whole numbers***.



Law of Multiple Proportions

- For several compounds of nitrogen (N) and oxygen (O), the following results were observed:

	Mass of Nitrogen that combines with 1g of Oxygen
Compound A	1.750 g
Compound B	0.8750 g
Compound C	0.4375 g

Chapter 2 Section 1

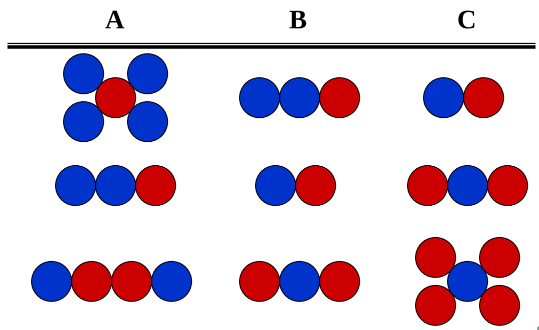
Law of Multiple Proportions

	Mass of **Nitrogen** that combines with 1g of **Oxygen**
Compound **A**	1.750 g
Compound **B**	0.8750 g
Compound **C**	0.4375 g

$$\frac{\text{Mass of N in A}}{\text{Mass of N in C}} = \frac{4}{1}$$

$$\frac{\text{Mass of N in B}}{\text{Mass of N in C}} = \frac{2}{1}$$

The mass ratios shown can be readily described on basis of the **ratios of number of atoms**.

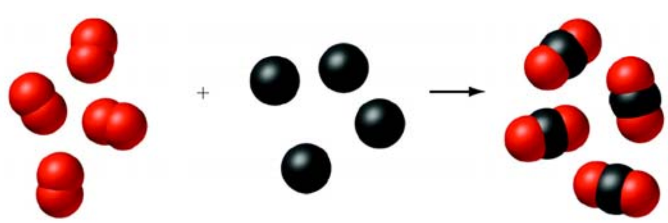


Chapter 2 Section 1

Law of Conservation of Mass

- Matter can be ***neither created nor destroyed***.

  Because matter is made up of atoms that are unchanged (masses and properties) in a chemical reaction, it follows that mass must be conserved as well.



Combination of oxygen and carbon to form carbon dioxide

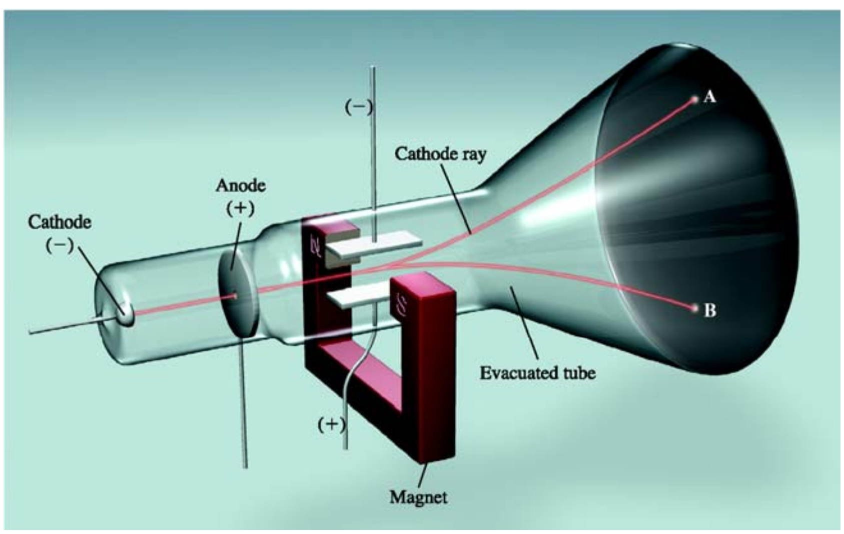
Chapter 2 Section 2

Early Imagination of the Atom

- What did Dalton think about the structure of an atom?
  - Extremely small.
  - Invisible.
  - Has a mass.
  - The smallest size ever of matter.
  - No internal structure.
- By mid 1800's it became evident that atoms **are** divisible - there is an internal structure to the atom. (subatomic particles)

Chapter 2 Section 2

Cathode Ray Experiment



Chapter 2 Section 2

Cathode Ray Experiment

- The Deflection of the ray by a magnet indicates that the ray is made up of negatively charged particles.

Thomson measured the charge-to-mass ratio as: $\frac{e}{m} = -1.76 \times 10^{8}\ \text{C/g}$

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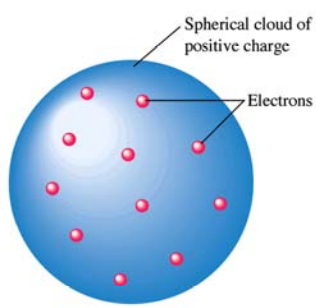
Chapter 2 Section 2

Cathode Ray Experiment

- Cathode ray experiments revealed important conclusions about the structure of the atom.
  - The ray is a stream of negatively charged particles (later on called electrons).
  - All atoms must contain electrons.
  - Since the atom is neutral overall, it must have a positively charged component.



Plum-pudding model suggested by Thomson

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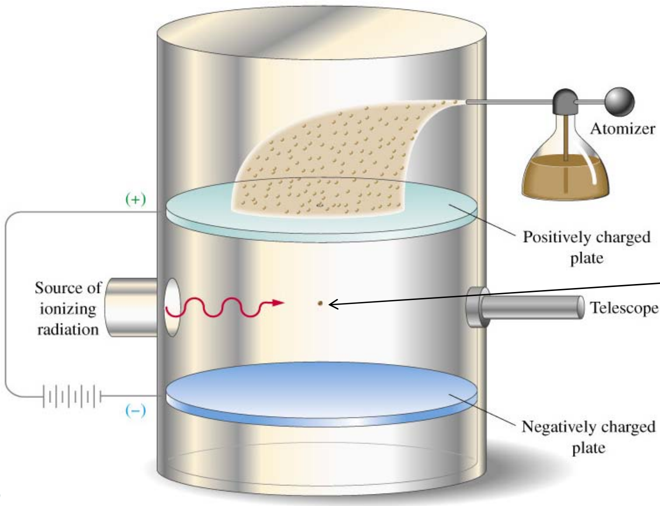
Chapter 2 Section 2

Mass of the Electron

- Millikan's Experiment (1917).



Oil droplets

Mass of the electron:

$$\frac{-1.6022 \times 10^{-19}\ \text{C}}{-1.76 \times 10^{8}\ \text{C/g}}$$

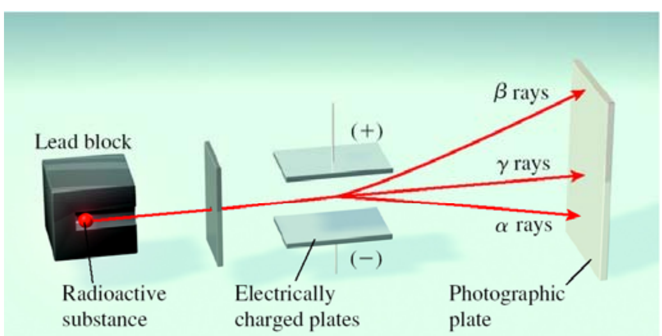
$$= 9.10 \times 10^{-28}\ \text{g}$$

Chapter 2 Section 2

Radioactivity

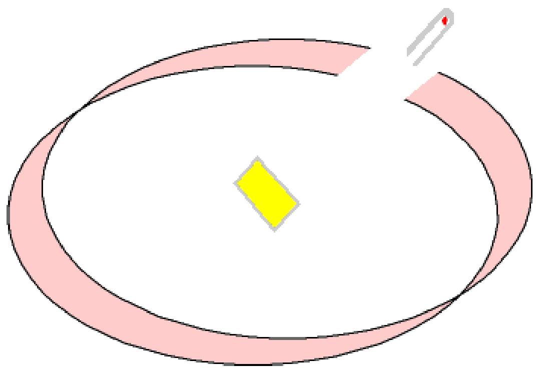
Types of spontaneous radioactive emission:

- α particles: have +ve charge and have mass that is 7300 time the mass of electron
- β particles: high-speed electrons.
- γ particles: high-energy light.



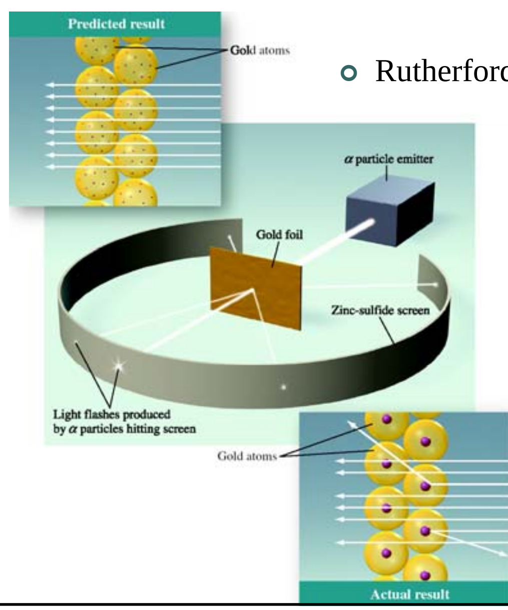
Rutherford Experiment

The Rutherford gold foil experiment used an alpha particle gun.



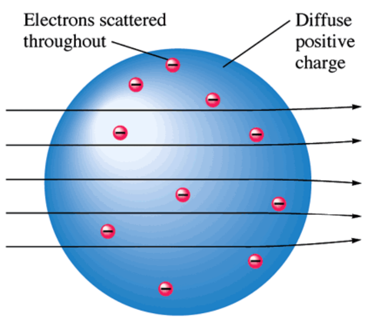
The Proton and the Nucleus

- Rutherford Experiment (1910)

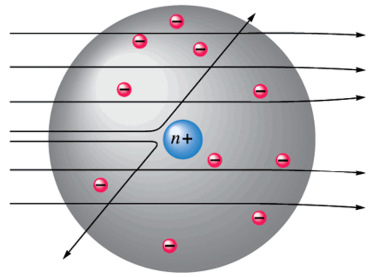


The Nuclear Atom

- Rutherford's Model vs. Thomson's Model.



Thomson's model
(The plum-pudding model)



Rutherford's model
(The nuclear atom)

The Nuclear Atom

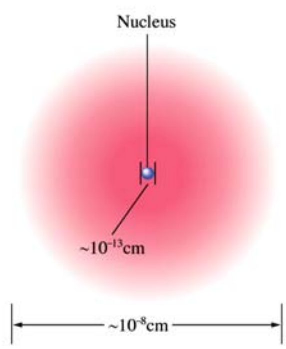
Main components of atoms:

Outside the nucleus:

- Electrons: are responsible for the chemistry of the atom.

Inside the nucleus:

- Protons: are positively charged particles whose charge is equal in magnitude to that for electrons.



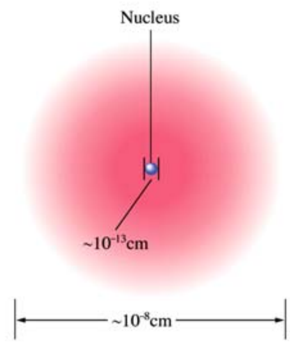
The simplest view of the atom

Chapter 2 Section 2

The Nuclear Atom

- Nucleus is very tiny in terms of size.
- Each proton carries exactly the opposite charge of an electron.
- Almost all the atomic mass is concentrated in it (very dense)!!
- The mass of the proton is 1.67×10^{-24} g.

  If a nucleus were to have the size of a pea, it would weigh 250,000,000,000 kg!



10^{-8} cm = 100 picometer

1 picometer (pm) = 1×10^{-12} m

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Chapter 2 Section 2

The Nuclear Atom

- Rutherford's model left one problem:
  - If H has a mass of 1, then He should have a mass of 2.
  - But its mass is 4!
- J. Chadwick (1932) discovered the ***neutrons***; massive but uncharged particles.

TABLE 2.1 Masses and Charges of Subatomic Particles

Particle	Mass (g)	Charge (C)	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 resulted in a small change to Millikan's original value.

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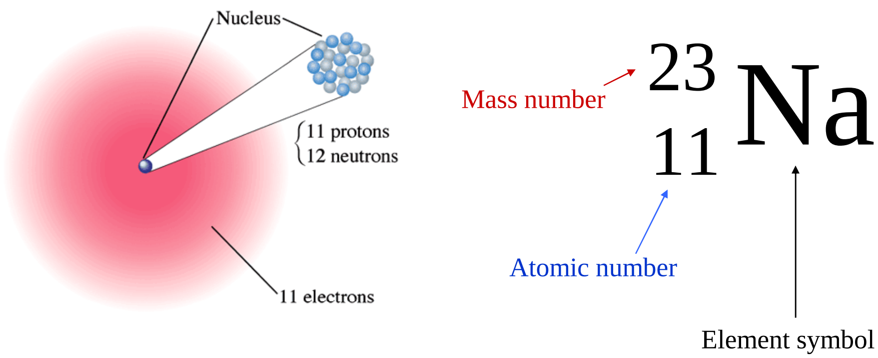
Chapter 2 Section 3

Atomic Number and Mass Number



Element symbol (Na) = Sodium (Note that it is neutral)

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

Atomic number (*Z*) = # of protons

For Na ion, the charge = # of protons – # of electrons

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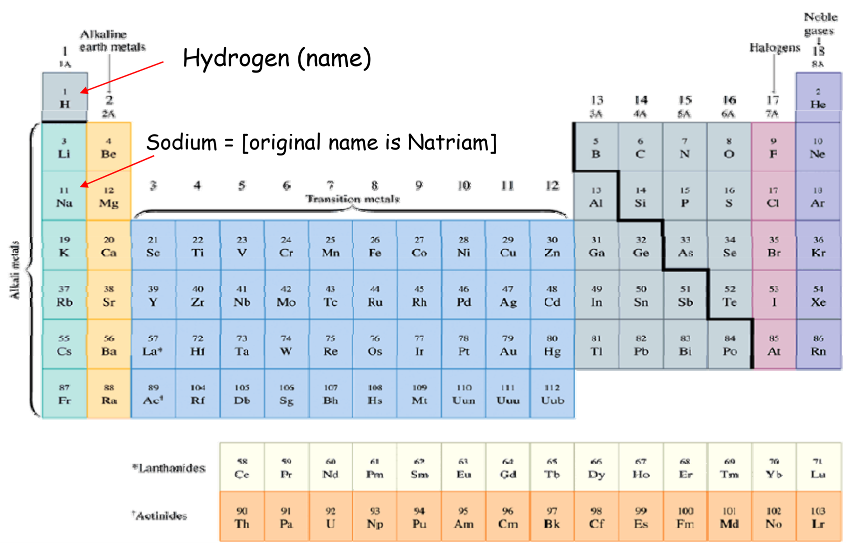
Chapter 2 Section 3

Atomic Symbols in the Periodic Table



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Chapter 2 Section 3

Exercise

What is the symbol for an ion with 63 protons, 60 electrons, and 88 neutrons?

Atomic number = # of protons = 63 => Eu

Atomic mass = # of protons + # neutrons = 63 + 88

Atomic charge = 63 – 60 = 3+

The symbol is $^{151}_{63}Eu^{3+}$

Another exercise: For $^{53}_{26}Fe^{2+}$

# of protons = 26

# of neutrons = 53 – 26

Net charge = 2+

# of electrons = 26 – 2 = 24

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Chapter 2 Section 3

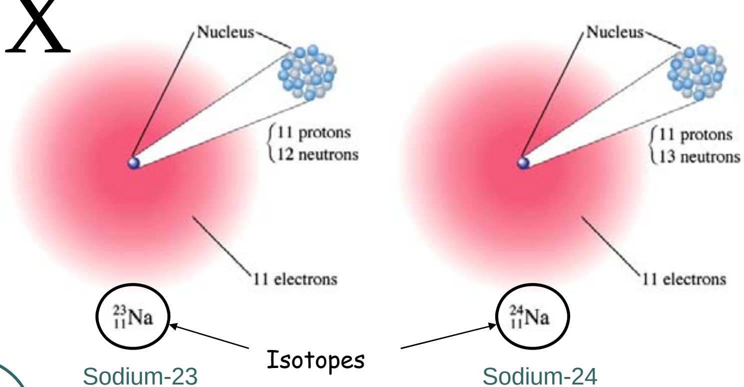
Isotopes

Mass number → A

Atomic number → Z

$^{A}_{Z}X$

Isotopes show almost identical properties



Ion (Cation)

$^{23}_{11}Na^{1+}$

# of protons = 11 # of neutrons = 12 # of electrons = 10

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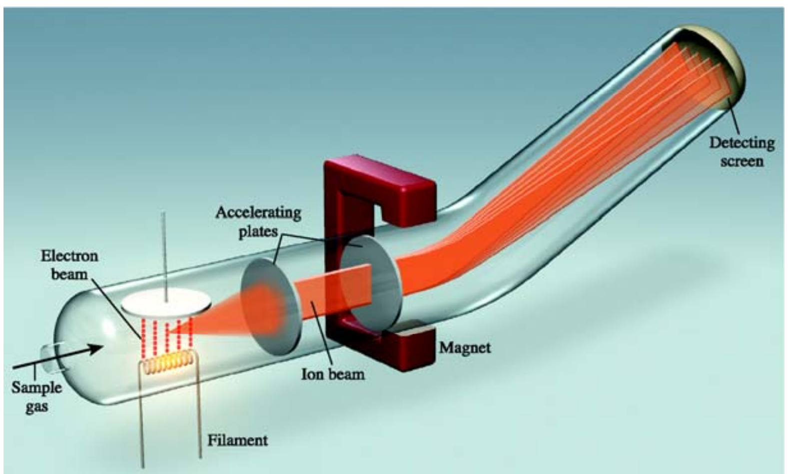
Isotopes

- Isotopes of Hydrogen
  - Hydrogen (protium) $^{1}_{1}H$
  - Deuterium $^{2}_{1}H$
  - Tritium $^{3}_{1}H$

- The chemical properties of an element are determined by the electrons and protons, not the neutrons. Thus, isotopes are chemically alike.

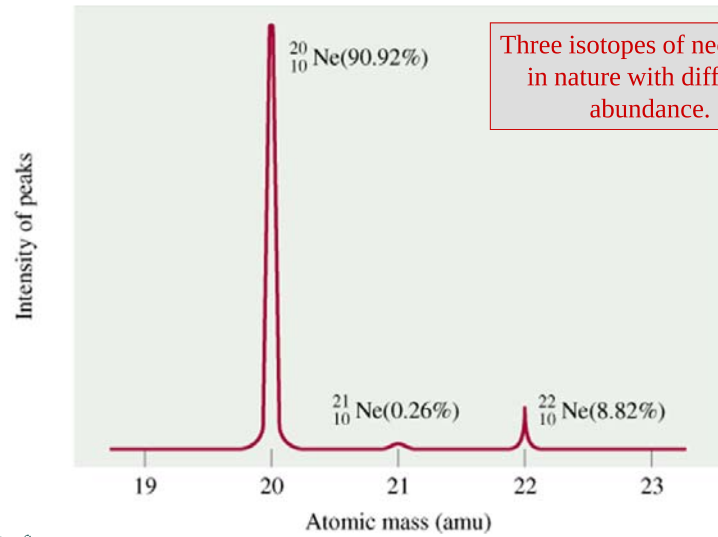
How are Atomic Masses Measured?



The Mass Spectrometer

Chapter 2 Section 3

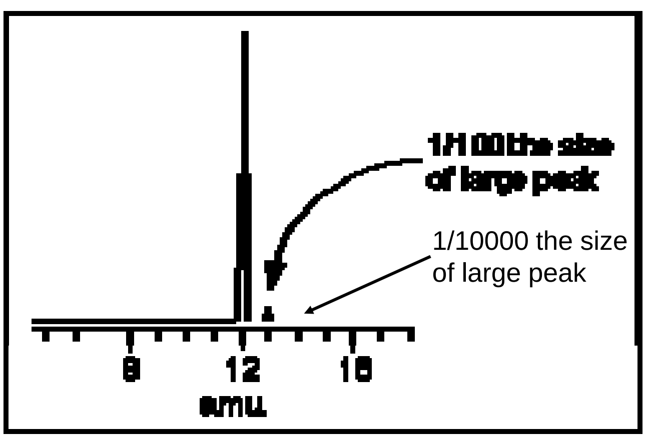
Mass Spectrum of Xenon



Three isotopes of neon exist in nature with different abundance.

Chapter 3 Section 3

Mass Spectrum of Carbon



Carbon exists in the form of three isotopes:

- ^{12}C (98.93%)
- ^{13}C (1.07%)
- ^{14}C (< 0.001%).

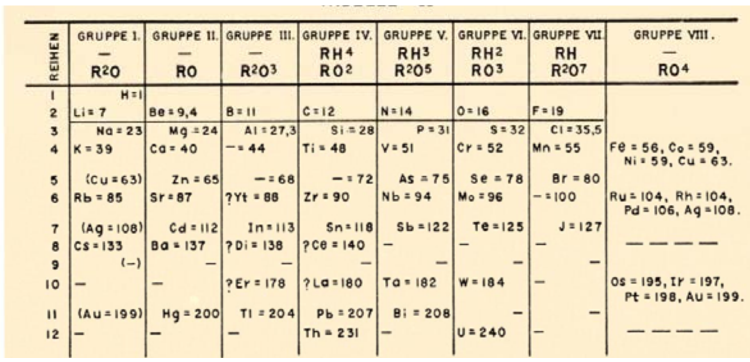
Three isotopes of carbon are present in nature

C-12 is the most abundant isotope of carbon.

Chapter 2 Section 4

The Periodic Table

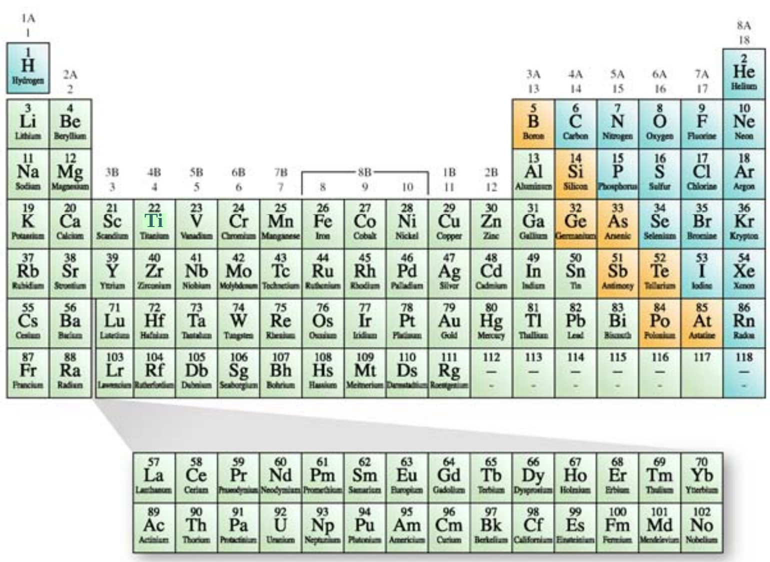
- Scientists noticed that chemical and physical properties of certain groups of elements are similar to one another.
- This led to the development of the ***periodic table***.



REIHEN	GRUPPE I. — R^2O	GRUPPE II. — RO	GRUPPE III. — R^2O^3	GRUPPE IV. RH^4 RO^2	GRUPPE V. RH^3 R^2O^5	GRUPPE VI. RH^2 RO^3	GRUPPE VII. RH R^2O^7	GRUPPE VIII. — RO^4
1	H = 1							
2	Li = 7	Be = 9,4	B = 11	C = 12	N = 14	O = 16	F = 19	
3	Na = 23	Mg = 24	Al = 27,3	Si = 28	P = 31	S = 32	Cl = 35,5	
4	K = 39	Ca = 40	— = 44	Ti = 48	V = 51	Cr = 52	Mn = 55	Fe = 56, Co = 59, Ni = 59, Cu = 63.
5	(Cu = 63)	Zn = 65	— = 68	— = 72	As = 75	Se = 78	Br = 80	
6	Rb = 85	Sr = 87	?Yt = 88	Zr = 90	Nb = 94	Mo = 96	— = 100	Ru = 104, Rh = 104, Pd = 106, Ag = 108.
7	(Ag = 108)	Cd = 112	In = 113	Sn = 118	Sb = 122	Te = 125	J = 127	
8	Cs = 133	Ba = 137	?Di = 138	?Ce = 140	—	—	—	— — — —
9	(—)	—	—	—	—	—	—	
10	—	—	?Er = 178	?La = 180	Ta = 182	W = 184	—	Os = 195, Ir = 197, Pt = 198, Au = 199.
11	(Au = 199)	Hg = 200	Tl = 204	Pb = 207	Bi = 208	—	—	
12	—	—	—	Th = 231	—	U = 240	—	— — — —

Chapter 2 Section 4

The Modern Periodic Table



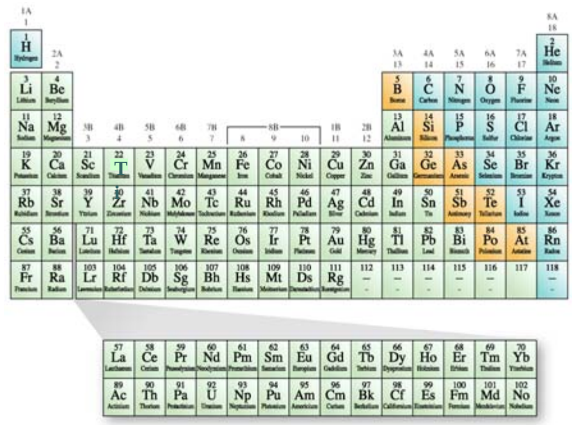
Chapter 2 Section 4

The Modern Periodic Table

- ***Periods*** : horizontal rows
- ***Families (Groups)*** : vertical columns

    Elements in the same family have similar chemical and physical properties

- Arranged in order of increasing atomic number



Chapter 2 Section 4

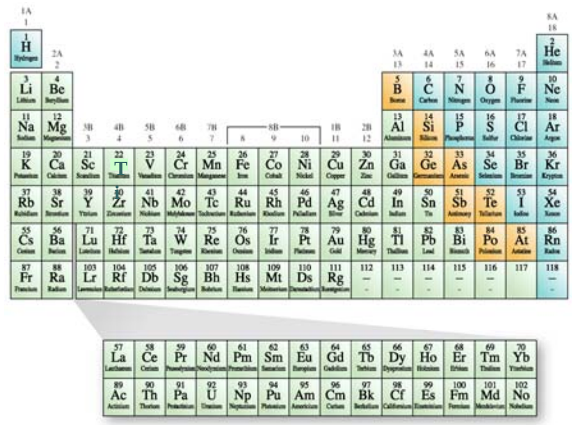
The Modern Periodic Table

- **Metals:** compose most of the periodic table. They have characteristic physical properties e.g.
    - High heat and electric conduction, malleability (hammered to sheet), ductility (pulled into wires).
    - Tend to lose electrons to form +ve ions. Fe^{2+} , Fe^{3+} , Na^{+} ,K^{+} , Ca^{2+} .
- **Nonmetals:** lack the physical properties of metals.
    - Tend to gain electrons to become –ve ions, like Cl^{-}, F^{-}, O^{2-}, S^{2-}.
    - Tend to bond with each other by forming (***covalent bonds***), such as Cl_2, HCl, N_2O, CO_2 etc.
    - react with metals to form salt (***ionic bonds***), such as NaCl, CaF_2, etc.
- **Metalloids:** have intermediate properties.
    - Examples are B, Si, Ge, etc.

The Modern Periodic Table

- The ***metallic properties*** increase as going from right to left across a period.
- The groups in the periodic table are given special names.

  • Alkali metals.
  • Alkaline earth metals.
  • Chalcogens
  • Halogens.
  • Nobel gasses.
  • Transition metals.



The Modern Periodic Table

-The groups in the periodic table are given special names.

• Alkali metals.
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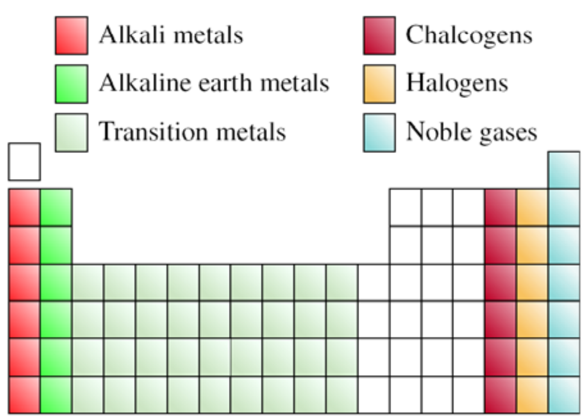


Chapter 2 Section 5

Atomic Masses

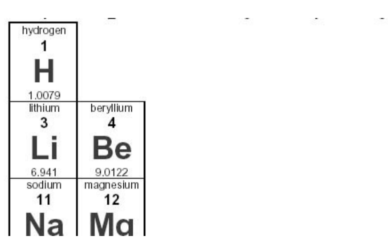
- Atomic mass is the mass of an atom expressed in *atomic mass unit* (amu).
- By definition used by modern systems, carbon-twelve ^{12}C is assigned a mass of *exactly* 12 amu.
- One ***atomic mass unit*** is defined as tha mass exactly equal to 1/12 the mass of one carbon-12 atom.

Chapter 2 Section 5

Atomic Masses

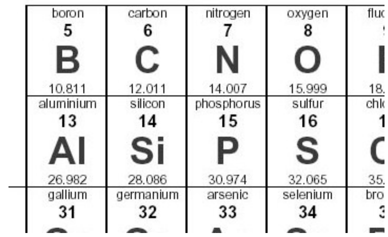
- Carbon-12 (12 amu) provides the standard for measuring the atomic mass of the rest of elements.
- Example: Hydrogen atom ^{1}H was found to be 8.3985% as massive as the C-12 atom. Can you find the atomic mass of a hydrogen atom in amu?

  Mass ^{1}H = 12 amu × 0.083985 = 1.0078 amu.



Average Atomic Masses

Why the carbon in the periodic table has a mass of 12.01 amu and not 12 amu??



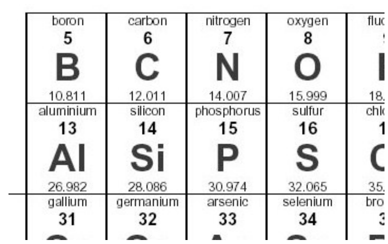
Carbon exists naturally as a mixture of three isotopes, ^{12}C, ^{13}C and ^{14}C and thus the atomic mass unit used for the carbon atom in the periodic table is the ***average value of the masses*** of those isotopes.

Average Atomic Masses

The ***average atomic mass*** (or just the ***atomic mass***) of the carbon atom = 98.89% of 12 amu (^{12}C) + 1.11% of 13.0034 amu (^{13}C)

= (0.9889)(12 amu) + (0.0111)(13.0034 amu)

= 12.011 amu



- That is applied for all the elements of the periodic table.

Chapter 2 Section 5

Average Atomic Masses

- Remember that, there is no a single carbon atom that has the mass of 12.010 amu. This is the ***average mass*** per carbon atom That is applied for all the elements of the periodic table.
- The mass of each element listed in the periodic table is an *average* value based on the isotopic composition of the naturally occurring element.

boron 5 B 10.811	carbon 6 C 12.011	nitrogen 7 N 14.007	oxygen 8 O 15.999	fluo
aluminium 13 Al 26.982	silicon 14 Si 28.086	phosphorus 15 P 30.974	sulfur 16 S 32.065	chl
gallium 31	germanium 32	arsenic 33	selenium 34	bro

Chapter 3 Section 5

Exercise

The element rhenium (Re) has two naturally occurring isotopes, ^{185}Re and ^{187}Re, with an average atomic mass of 186.207 amu. Rhenium is 62.60% ^{187}Re, and the atomic mass of ^{187}Re is 186.956 amu. Calculate the mass of ^{185}Re.

Atomic mass = (mass of isotope 1 x fractional abundance of 1) +
(mass of isotope 2 x fractional abundance of 2) +
(mass of isotope 3 x fractional abundance of 3) + ...

Exercise

The element rhenium (Re) has two naturally occurring isotopes, ^{185}Re and ^{187}Re, with an average atomic mass of 186.207 amu. Rhenium is 62.60% ^{187}Re, and the atomic mass of ^{187}Re is 186.956 amu. Calculate the mass of ^{185}Re.

$$\text{Atomic mass} = \sum_i m_i x_i$$

^{187}Re is 62.60% with a mass of 186.956 amu.

Mass of Re = 186.207 amu

= (186.956amu) (0.6260) + (?? amu) (0.3740)

Answer is

Mass of ^{185}Re = 184.9533 amu = 185.0 amu

Molecular Compounds & Ionic Compounds

- The force holds atoms together is called a ***chemical bond***.
- Some types of chemical bonds are
  - ***Covalent bonds***: Two atoms *"usually nonmetals"* can form a bond by sharing electrons to produce "***molecular compounds***".
  - ***Ionic bonds***: Two oppositely charged ions (a cation and an anion) *"a metal and a nonmetal"* can form a bond by attraction to produce "***ionic compounds***".

Chapter 2 Section 6

Molecular Compounds

- ***Molecule*** **:** combination of at least two atoms in a specific arrangement held together by *chemical bonds.*
  - May be an element or a compound.
  - H_2, hydrogen gas, is an element.
  - H_2O, water, is a compound.
  - They are also called "***binary compounds***".

N_2

CO

CH_4

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Chapter 2 Section 6

Molecular Compounds

- ***Diatomic Molecule:***
  - *Homonuclear* (2 of the same atoms)
    - Examples: H_2, N_2, O_2, F_2, Cl_2, Br_2, and I_2

  - *Heteronuclear* (2 different atoms)
    - Examples: CO and HCl

- ***Polyatomic Molecule:***
  - Contain more than 2 atoms
  - Most molecules
  - May contain more than one element
  - Examples: ozone, O_3; white phosphorus, P_4; water, H_2O, and methane (CH_4)

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Chapter 2 Section 6

Molecular Formulas

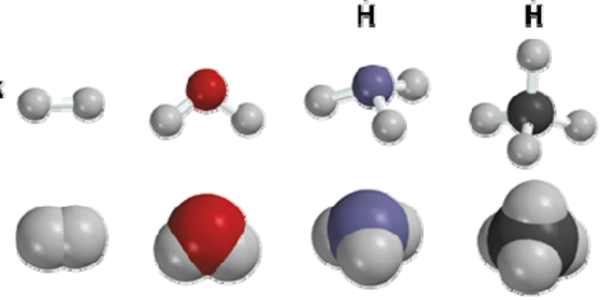
- ***Molecular formula***: shows exact number of atoms of each element in a molecule.
  - Subscripts indicate number of atoms of each element present in the formula.
  - Example: H_2O, NH_3, $C_{12}H_{22}O_{11}$ etc.
- ***Structural formula***: shows the general arrangement of atoms within the molecule.

H—O—H

Chapter 2 Section 6

Molecular Formulas

	Covalent-bonded Molecules			
	Hydrogen	Water	Ammonia	Methane
Molecular formula	H_2	H_2O	NH_3	CH_4
Structural formula	H—H	H—O—H	H—N—H | H	H | H—C—H | H
Ball-and-stick model				
Space-filling model				



Chapter 2 Section 6

Molecular Formulas

- ***Allotrope***: one of two or more distinct forms of an element.
  - oxygen, O_2 and ozone, O_3 (allotropic forms of oxygen)
  - diamond and graphite (allotropic forms of carbon)

Chapter 2 Section 6

Naming Molecular Compounds

TABLE 2.6 Prefixes Used to Indicate Number in Chemical Names

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

Only used for the second element

- It is also known as "***Nomenclature***".
- Binary molecular (or covalent) compounds are composed of two nonmetals:
  - Name the first element.
  - Name the second element changing ending to "*-ide*".
  - If the two elements form more than one type of binary molecular compounds then use *prefixes* to indicate number of atoms of each element.

Chapter 2 Section 6

Naming Molecular Compounds

TABLE 2.6 Prefixes Used to Indicate Number in Chemical Names

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

Only used for the second element

- HCl Hydrogen chloride
- SiC Silicon carbide
- NO Nitrogen monoxide
- N_2O Dinitrogen monoxide
- N_2O_5 Dinitrogen pentoxide
- SO_2 Sulfur dioxide
- PCl_3 Phosphorus trichloride

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Chapter 2 Section 6

Naming Molecular Compounds

- Name the following:
  - Cl_2O

    Dichlorine monoxide.
  - CBr_4

    Carbon tetrabromide.
  - ClO_2

    Chlorine dioxide.
  - SO

    Sulfur monoxide.

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Chapter 2 Section 6

Naming Molecular Compounds

- The names of molecular compounds containing hydrogen do not usually follow the systematic nomenclature guidelines.
  - B_2H_6 diborane
  - SiH_4 silane
  - NH_3 ammonia
  - PH_3 phosphine
  - H_2O water
  - H_2S hydrogen sulfide

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Chapter 2 Section 6

Naming Binary Acids

- Acids when are dissolved in water, they produce H^+ ions (protons) in the solutions.
  - Examples are: HCl, HBr.
- Binary acids:
  - Many have 2 names
    - **Pure substance:** HCl, *hydrogen chloride*.
    - **Aqueous solution:** when dissolved in water it is called *hydrochloric acid*.

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Naming Binary Acids

- In order to name binary acids:
  - Remove the "–*gen*" ending from hydrogen leaving "*hydro–*".
  - Change the "–*ide*" ending on the second element to "*–ic*".
  - Combine the two words and add the word "*acid*".
- Name the following:
  - HBr

    Hydrogen bromide ; Hydrobromic acid
  - H_2S

    Hydrogen sulfide ; Hydrosulfuric acid

Naming Organic Compounds

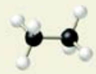
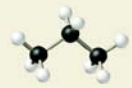
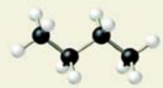
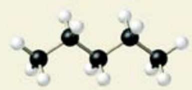
- ***Organic Compounds:*** contain carbon and hydrogen (sometimes with oxygen, nitrogen, sulfur and the halogens).
  - ***Hydrocarbons* :** contain only carbon and hydrogen.
  - ***Alkanes* :** simplest examples of hydrocarbons. Their names depend on the number of carbon atoms in the molecule.

- ***Inorganic Compounds:*** normally do not contain carbon.

Chapter 2 Section 6

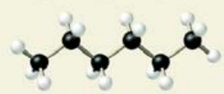
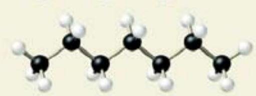
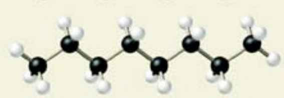
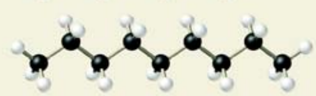
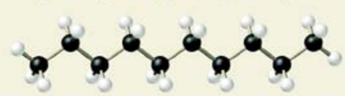
Alkanes

TABLE 2.5 Formulas, Names, and Models of Some Simple Alkanes

Formula	Name	Model
CH_4	Methane	
C_2H_6	Ethane	
C_3H_8	Propane	
C_4H_{10}	Butane	
C_5H_{12}	Pentane	

Chapter 2 Section 6

Alkanes

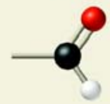
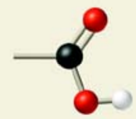
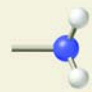
C_6H_{14}	Hexane	
C_7H_{16}	Heptane	
C_8H_{18}	Octane	
C_9H_{20}	Nonane	
$C_{10}H_{22}$	Decane	

Chapter 2 Section 6

Functional Groups

- Many derivatives of alkanes are derived by replacing a hydrogen with one or more ***functional groups.***
- Functional group determines chemical properties and is responsible for chemical reactions.

TABLE 2.6 Organic Functional Groups

Name	Functional Group	Model
Alcohol	−OH	
Aldehyde	−CHO	
Carboxylic acid	−COOH	
Amine	$-NH_2$	

Chapter 2 Section 6

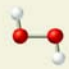
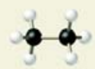
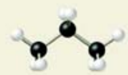
Empirical Formulas

- ***Empirical Formula:*** tells:
  - what elements are present in a molecule.
  - In what whole-number ratio they are combined.

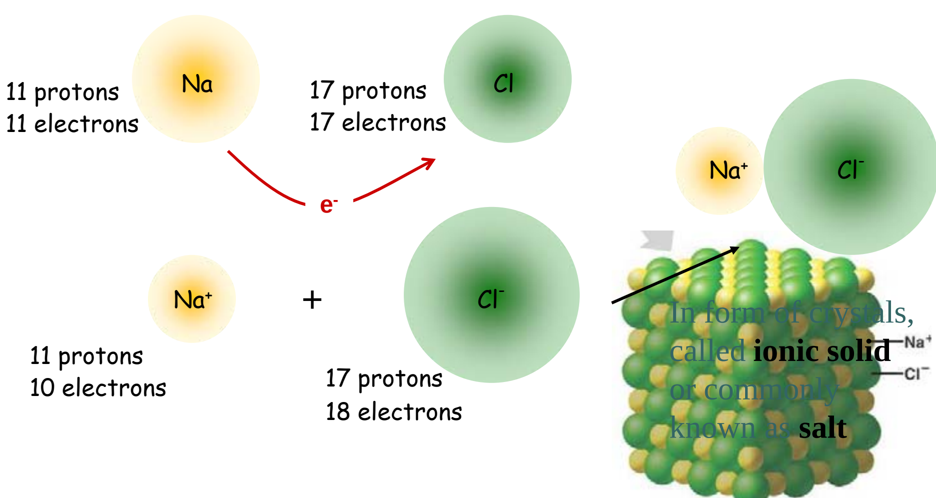
Molecular(true)	Empirical(simplest)
H_2O_2	HO
N_2H_4	NH_2
H_2O	H_2O

Molecular and Empirical Formulas

TABLE 2.7 Molecular and Empirical Formulas

Compound	Molecular Formula	Model	Empirical Formula	Model
Water	H_2O		H_2O	
Hydrogen peroxide	H_2O_2		HO	
Ethane	C_2H_6			
Propane	C_3H_8			
Acetylene	C_2H_2			
Benzene	C_6H_6			

Ionic Bonds and Ionic Compounds

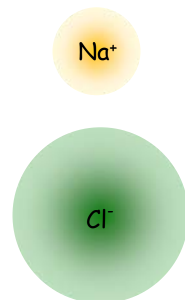


In form of crystals, called **ionic solid** or commonly known as **salt**

Chapter 2 Section 7

Ions and Ionic Compounds

- ***Ion***: an atom or *group* of atoms that has a net positive or negative charge.
- ***Monatomic ion*** **:** *one* atom with a positive or negative charge.
- ***Cation*** **:** ion with a net *positive* charge due to the loss of one or more electrons.
- ***Anion*** **:** ion with a net *negative* charge due the gain of one or more electrons.



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Chapter 2 Section 7

Common Monoatomic Ions

Cations:

- Lithium ion (Li^{+})
- Potassium ion (K^{+})
- Aluminum ion (Al^{3+})

Type I

- Iron (II) ion (Fe^{2+})
- Iron (III) ion (Fe^{3+})
- Lead (IV) (Pb^{4+})
- Lead (II) (Pb^{2+})

Type II

1A 1	2A 2	3B 3	4B 4	5B 5	6B 6	7B 7	8B 8	8B 9	8B 10	1B 11	2B 12	3A 13	4A 14	5A 15	6A 16	7A 17	8A 18
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+}										Hg_2^{2+} Hg^{2+}		Pb^{2+} Pb^{4+}				

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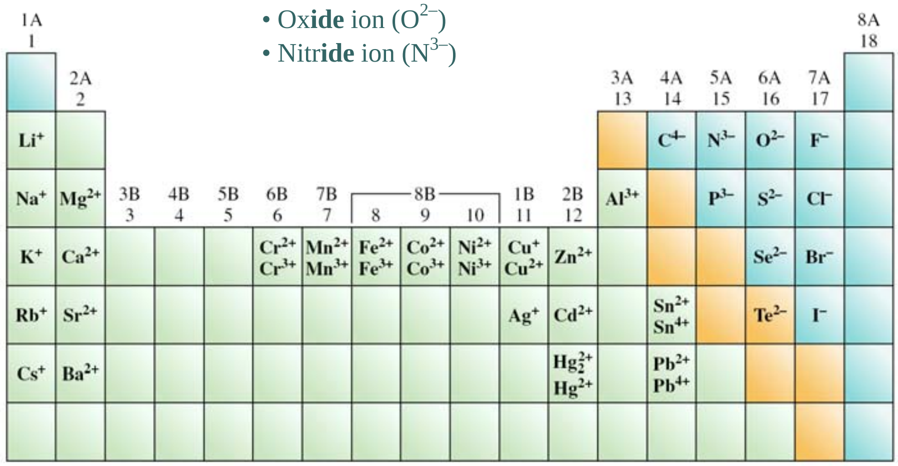
Common Monoatomic Ions

Anions:

- Fluor**ide** ion (F^-)
- Ox**ide** ion (O^{2-})
- Nitr**ide** ion (N^{3-})



Type I

Naming Binary Ionic Compounds

\+ -

It contains a +ve ion and a -ve ion.

TABLE 2.3 Common Monatomic Cations and Anions

Cation	Name	Anion	Name
H^+	Hydrogen	H^-	Hydride
Li^+	Lithium	F^-	Fluoride
Na^+	Sodium	Cl^-	Chloride
K^+	Potassium	Br^-	Bromide
Cs^+	Cesium	I^-	Iodide
Be^{2+}	Beryllium	O^{2-}	Oxide
Mg^{2+}	Magnesium	S^{2-}	Sulfide
Ca^{2+}	Calcium	N^{3-}	Nitride
Ba^{2+}	Barium	P^{3-}	Phosphide
Al^{3+}	Aluminum		
Ag^+	Silver		

1- Cations named first then anions.

2- Cation element has the same name without change.

3-Use –*ide* root to the anion name.

4-Double check the ionic charges to have the correct chemical formula.

5-You will need to practice this table.

6- You will need to be able to get names from formulas and vise versa.

Chapter 2 Section 7

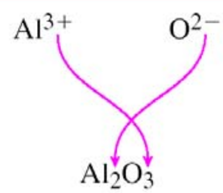
Type I

Naming Binary Ionic Compounds

TABLE 2.3 Common Monatomic Cations and Anions

Cation	Name	Anion	Name
H^+	Hydrogen	H^-	Hydride
Li^+	Lithium	F^-	Fluoride
Na^+	Sodium	Cl^-	Chloride
K^+	Potassium	Br^-	Bromide
Cs^+	Cesium	I^-	Iodide
Be^{2+}	Beryllium	O^{2-}	Oxide
Mg^{2+}	Magnesium	S^{2-}	Sulfide
Ca^{2+}	Calcium	N^{3-}	Nitride
Ba^{2+}	Barium	P^{3-}	Phosphide
Al^{3+}	Aluminum		
Ag^+	Silver		

Compound	Ions present	Name
NaCl	Na^+ , Cl^-	Sodium chloride
KI	K^+ , I^-	Potassium iodide
CaS	Ca^{2+} , S^{2-}	Calcium sulfide
MgO	Mg^{2+} , O^{2-}	Magnesium oxide
Al_2O_3	Al^{3+} , O^{2-}	Aluminum oxide



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Chapter 2 Section 7

Exercise

- Rb_2O
- Rubidium oxide.
- CaS
- Calcium sulfide.
- AlI_3
- Aluminum iodide.
- Strontium fluoride.
- SrF_2
- Aluminum selenide.
- Al_2Se_3
- Magnesium phosphide.
- Mg_3P_2

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Type II

Naming Binary Ionic Compounds

TABLE 2.4 Common Type II Cations

Ion	Systematic Name
Fe^{3+}	Iron(III)
Fe^{2+}	Iron(II)
Cu^{2+}	Copper(II)
Cu^{+}	Copper(I)
Co^{3+}	Cobalt(III)
Co^{2+}	Cobalt(II)
Sn^{4+}	Tin(IV)
Sn^{2+}	Tin(II)
Pb^{4+}	Lead(IV)
Pb^{2+}	Lead(II)
Hg^{2+}	Mercury(II)
Hg_2^{2+}*	Mercury(I)

- ***Only*** for metals that can form ***more than one type of cations***, the charge must be specified using Roman numerals
- Examples:
  - CuCl
    Copper(I) chloride.
  - $CuCl_2$
    Copper(II) chloride.
  - $CoCl_3$
    Cobalt(III) chloride.

Polyatomic Ions

- ***Polyatomic ions*** : ions that are a combination of two or more atoms.

TABLE 2.9 Common Polyatomic Ions

Name	Formula/Charge
Cations	
Ammonium	NH_4^+
Hydronium	H_3O^+
Mercury(I)	Hg_2^{2+}
Anions	
Azide	N_3^-
Carbonate	CO_3^{2-}
Chlorate	ClO_3^-
Chlorite	ClO_2^-
Chromate	CrO_4^{2-}
Cyanide	CN^-
Dichromate	$Cr_2O_7^{2-}$
Dihydrogen phosphate	$H_2PO_4^-$
Hydrogen carbonate or bicarbonate	HCO_3^-
Hydrogen phosphate	HPO_4^{2-}
Hydrogen sulfate or bisulfate	HSO_4^-
Hydroxide	OH^-
Hypochlorite	ClO^-
Nitrate	NO_3^-
Nitrite	NO_2^-
Oxalate	$C_2O_4^{2-}$
Perchlorate	ClO_4^-
Permanganate	MnO_4^-
Peroxide	O_2^{2-}
Phosphate	PO_4^{3-}
Phosphite	PO_3^{3-}
Sulfate	SO_4^{2-}
Sulfite	SO_3^{2-}
Thiocyanate	SCN^-

Chapter 2 Section 7

Polyatomic Ions

TABLE 2.5 Common Polyatomic Ions

Ion	Name	Ion	Name
Hg_2^{2+}	Mercury(I)	NCS^-	Thiocyanate
NH_4^+	Ammonium	CO_3^{2-}	Carbonate
NO_2^-	Nitrite	HCO_3^-	Hydrogen carbonate (bicarbonate is a widely used common name)
NO_3^-	Nitrate		
SO_3^{2-}	Sulfite		
SO_4^{2-}	Sulfate	ClO^-	Hypochlorite
HSO_4^-	Hydrogen sulfate (bisulfate is a widely used common name)	ClO_2^-	Chlorite
		ClO_3^-	Chlorate
		ClO_4^-	Perchlorate
OH^-	Hydroxide	$C_2H_3O_2^-$	Acetate
CN^-	Cyanide	MnO_4^-	Permanganate
PO_4^{3-}	Phosphate	$Cr_2O_7^{2-}$	Dichromate
HPO_4^{2-}	Hydrogen phosphate	CrO_4^{2-}	Chromate
$H_2PO_4^-$	Dihydrogen phosphate	O_2^{2-}	Peroxide
		$C_2O_4^{2-}$	Oxalate

oxoanions

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Chapter 2 Section 7

Exercise

- $BaSO_3$
  - Barium sulfite
- $K_2Cr_2O_7$
  - Potassium dichromate
- $CuMnO_4$
  - Copper(I) permanganate
- $NaNO_2$
  - Sodium nitrite
- Chromium(III) hydroxide
  - $Cr(OH)_3$
- Magnesium cyanide
  - $Mg(CN)_2$
- Lead(IV) carbonate
  - $Pb(CO_3)_2$
- Ammonium hypochlorite
  - NH_4ClO

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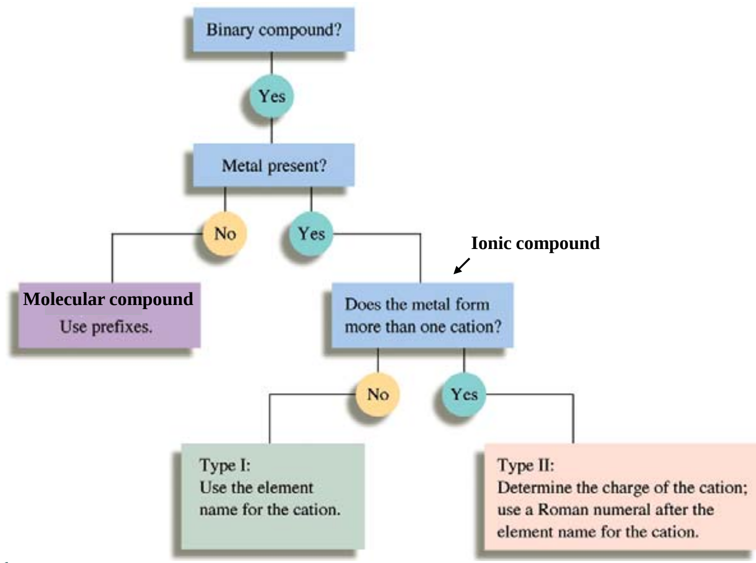
Chapter 2 Section 7

Naming Binary Compounds



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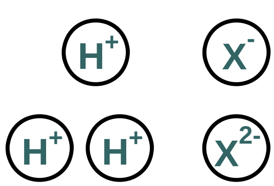
Chapter 2 Section 7

Naming Acids & Oxoacids

- ***Oxoacids*** **:** when ionized in water, they give H^+ ions (protons) and the corresponding polyatomic oxoanions in the solutions.
- Examples: HNO_3, H_2SO_3, and $HC_2H_3O_2$.
- When writing formulas, add the number of H^+ ions necessary to balance the corresponding oxoanion's negative charge.



where X is an oxoanion

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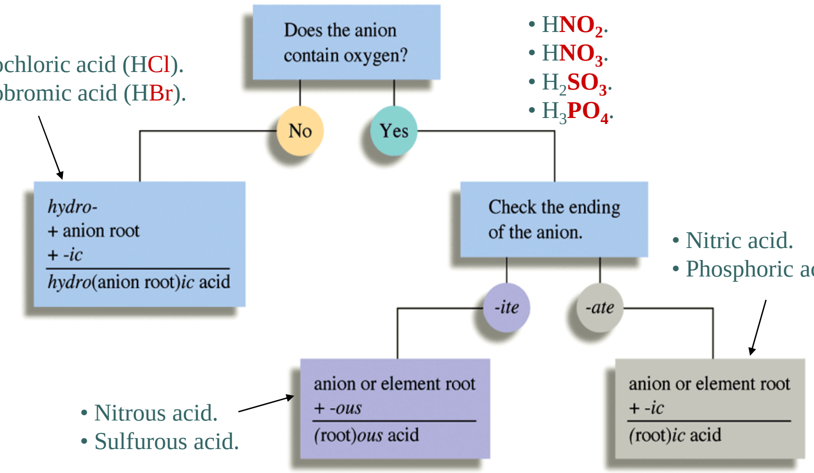
Chapter 2 Section 7

Naming Acids & Oxoacids

- Hydrochloric acid (HCl).
- Hydrobromic acid (HBr).

- HNO_2.
- HNO_3.
- H_2SO_3.
- H_3PO_4.



- Nitric acid.
- Phosphoric acid.

- Nitrous acid.
- Sulfurous acid.

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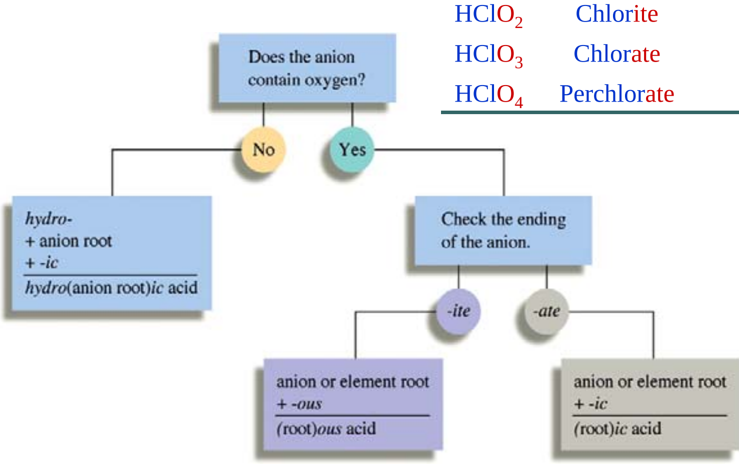
Chapter 2 Section 7

Naming Acids & Oxoacids

Acid	Anion	Name
HClO	Hypochlorite	
$HClO_2$	Chlorite	
$HClO_3$	Chlorate	
$HClO_4$	Perchlorate	



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Chapter 2 Section 7

Hydrates

- ***Hydrates*** : compounds that have a specific number of water molecules within their solid structure
  - Hydrated compounds may be heated to remove the water forming an ***anhydrous*** compound
- Name the compound and add the word hydrate. Indicate the number of water molecules with a prefix on hydrate.
  - Example: $CuSO_4 \cdot 5\ H_2O$
  - Copper (II) sulfate pentahydrate

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Chapter 2 Section 7

Exercise

The formulas and common names for several substances are given below. Give the systematic names for these substances.

a.	sugar of lead	$Pb(C_2H_3O_2)_2$	Lead(II) acetate
b.	blue vitrol	$CuSO_4$	Copper(II) sulfate
c.	quicklime	CaO	Calcium oxide
d.	Epsom salts	$MgSO_4$	Magnesium sulfate
e.	milk of magnesia	$Mg(OH)_2$	Magnesium hydroxide
f.	gypsum	$CaSO_4$	Calcium sulfate
g.	laughing gas	N_2O	Dinitrogen monoxide

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