Chapter 2

Atoms, Molecules, and Ions

2.1 The Atomic Theory

- **5th century B.C.** - Greek philosopher Democritus proposed that all matter consists of very small, indivisible particles, which he named *atomos* (meaning uncuttable or indivisible)
  - No support was given to this theory by contemporaries Plato or Aristotle
- **1808** - English scientist and school teacher John Dalton formulated an accurate definition of the indivisible building blocks of matter that we call atoms

Lavoisier's Laws

- Law of Conservation of Mass: In a chemical reaction, the total mass of the products = the total mass of the reactants.

(Very close to true: E = mc2, so when energy is released or absorbed there are extremely tiny changes in mass)

- ***Law of definite proportions*** *states that* different samples of a given compound always contain the same elements in the same mass *ratio.*
  - Examples

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

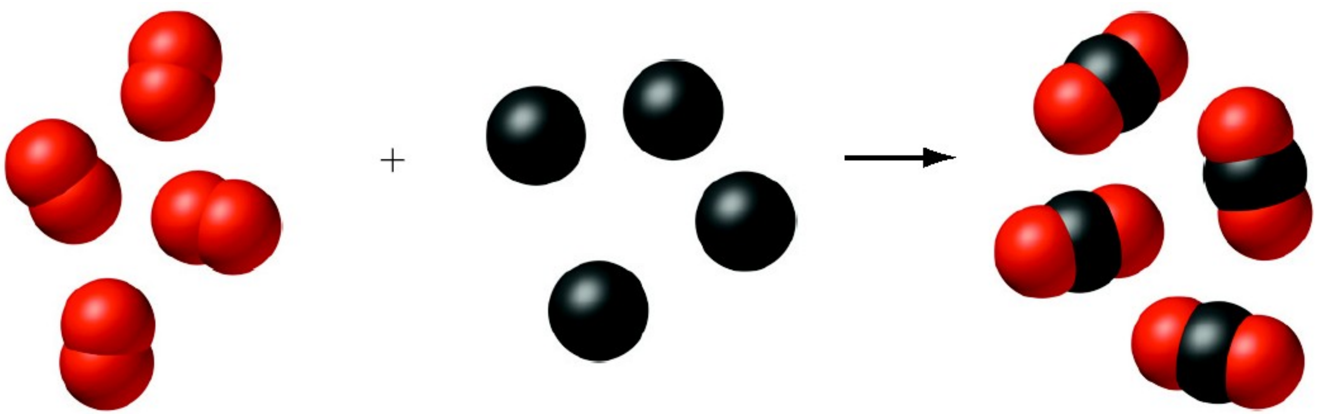
- ***Law of multiple proportions*** states that if two elements can combine to form more than one compound with each other, the masses of one element that combine with a fixed mass of the other element are in ratios of small whole numbers.
  - Example

$$\frac{\text{ratio of O to C in carbon dioxide}}{\text{ratio of O to C in carbon monoxide}} = \frac{2.66}{1.33} = 2{:}1$$

- Dalton’s atomic theory
  - Elements are composed of extremely small particles called atoms.
  - 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.

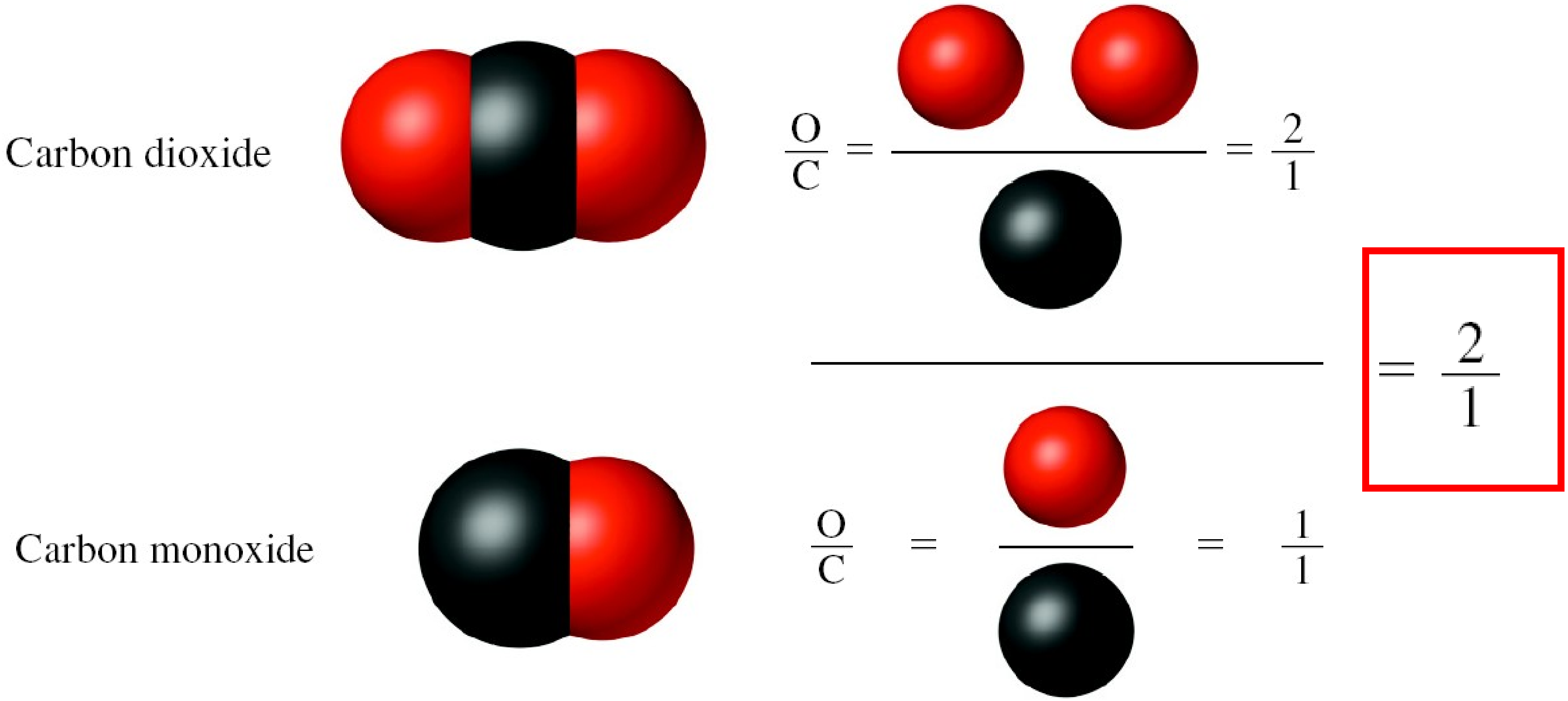
*Dalton wasn’t quite right on this point—more on “isotopes” later

- Compounds are composed of atoms of more than one element. In any given compound, the same types of atoms are always present in the same relative numbers.
- A chemical reaction rearranges atoms in chemical compounds; it does not create or destroy them.



Combination of oxygen and carbon to form carbon dioxide

Illustration of the Law of Multiple Proportions

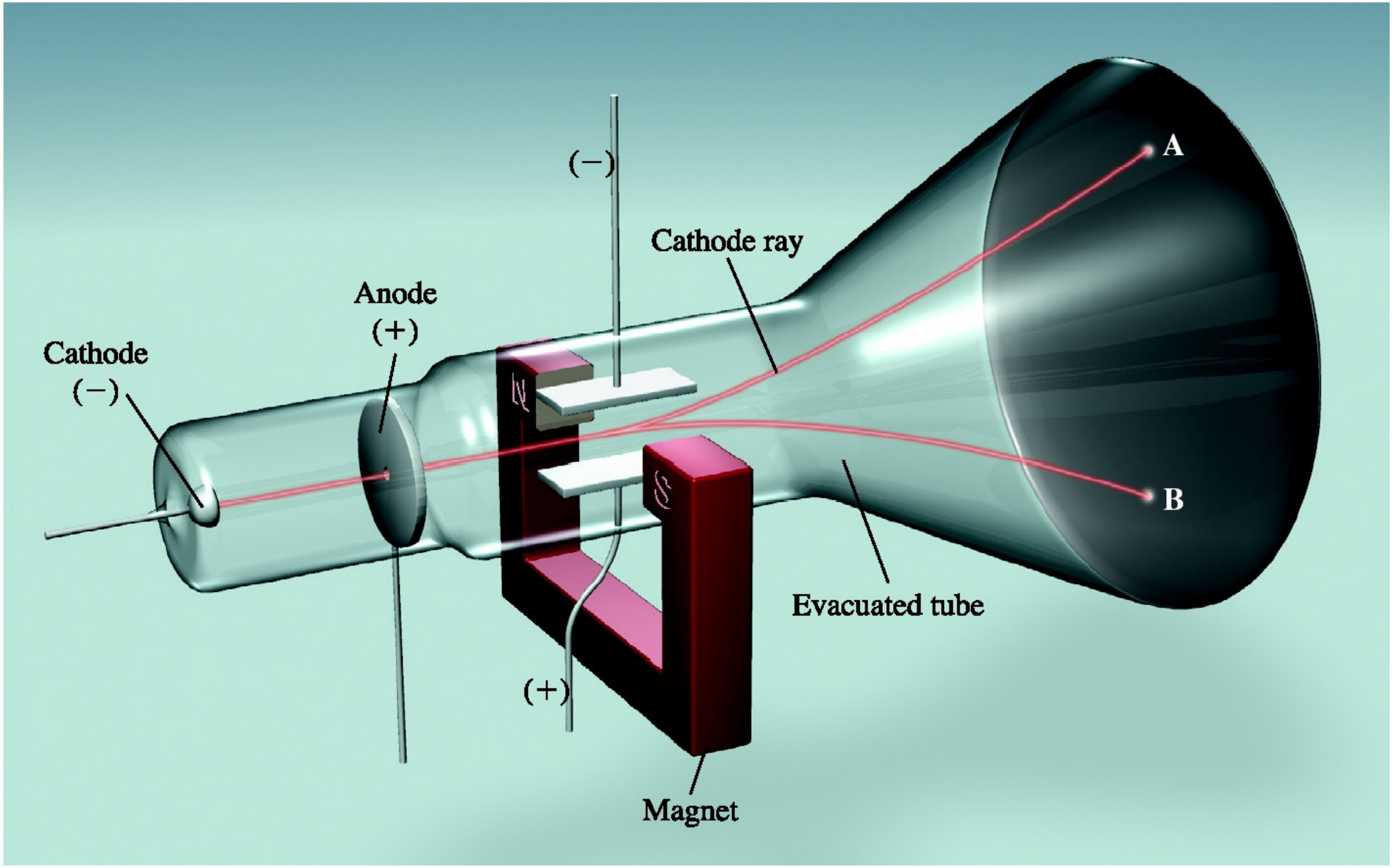


2.2 The Structure of the Atom

- Atoms are the basic unit of an element that can enter into a chemical reaction
- By mid 1800's it became evident that atoms **are** divisible - there is an internal structure to the atom. (subatomic particles)

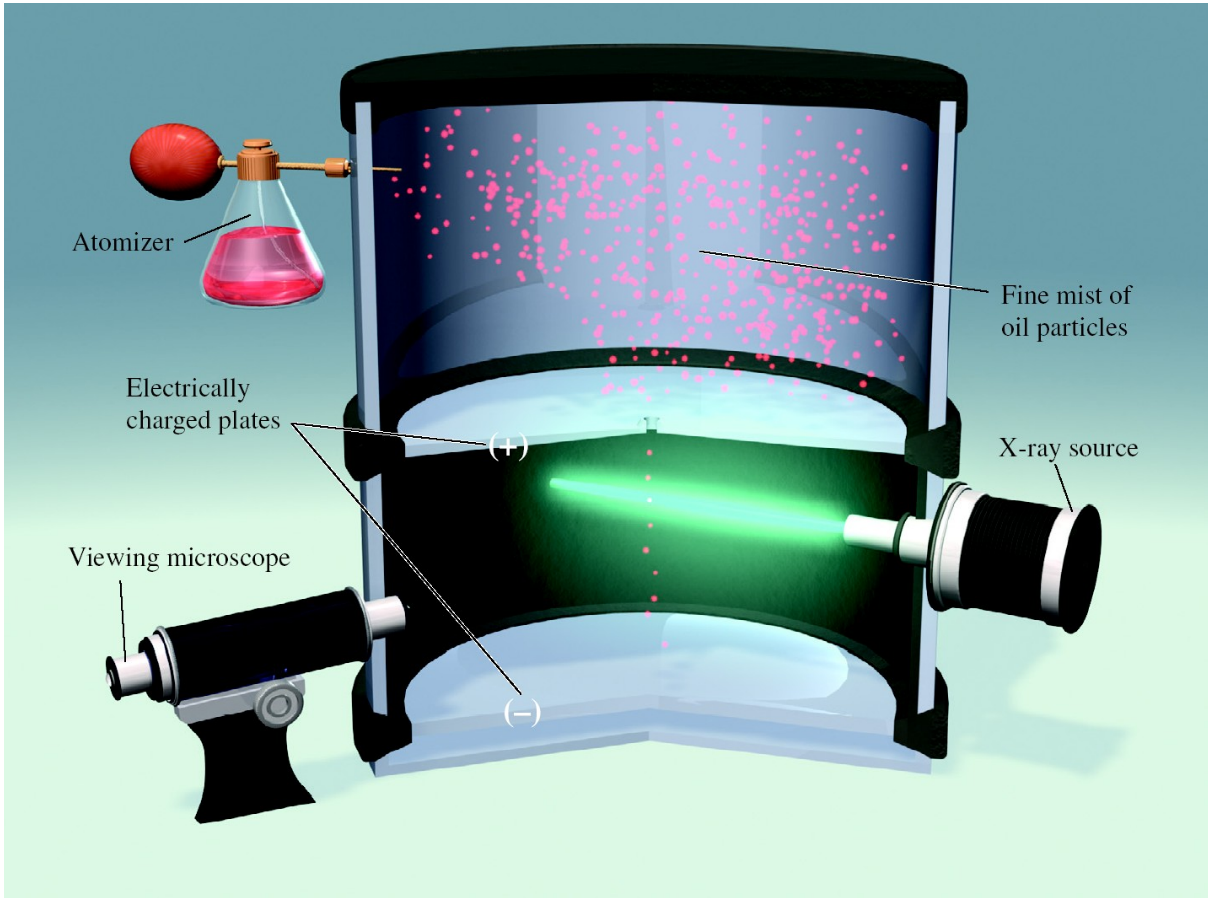
- Discovery of the electron
  - J.J. Thomson
    - discovered the ratio of the electric charge to the mass of an individual electron using the cathode ray tube
    - (-1.76×10^8 C/g); C = *coulomb*

Cathode Ray Experiment



Effect of a Magnetic Field on a Cathode Ray

- Millikan
  - Determined the charge on an electron
    - -1.66022×10^{-19} C



– The mass of the electron could be derived from

- Millikan's charge

  -1.66022×10^{-19} C

- Thomson's charge to mass ratio

  -1.76×10^{8} C/g

$$\text{mass of an electron} = \frac{\text{charge}}{\text{charge/mass}}$$

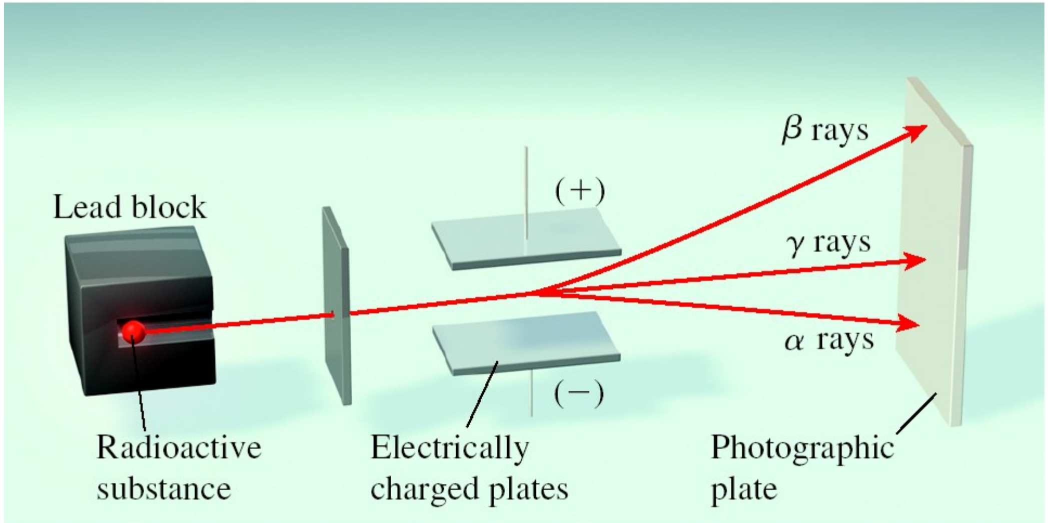
$$\frac{-1.6022 \times 10^{-19}\ \text{C}}{-1.76 \times 10^{8}\ \text{C/g}} = 9.10 \times 10^{-28}\ \text{g}$$

mass of electron

- Radioactivity – historical prespective
  - 1895 -Wilhelm Röentgen - German physicist
    - Noticed that cathode rays caused glass and metal to emit another type of ray
    - He named the rays “X rays” because of their mysterious nature
      - Caused fluorescence
      - Were not deflected by a magnet

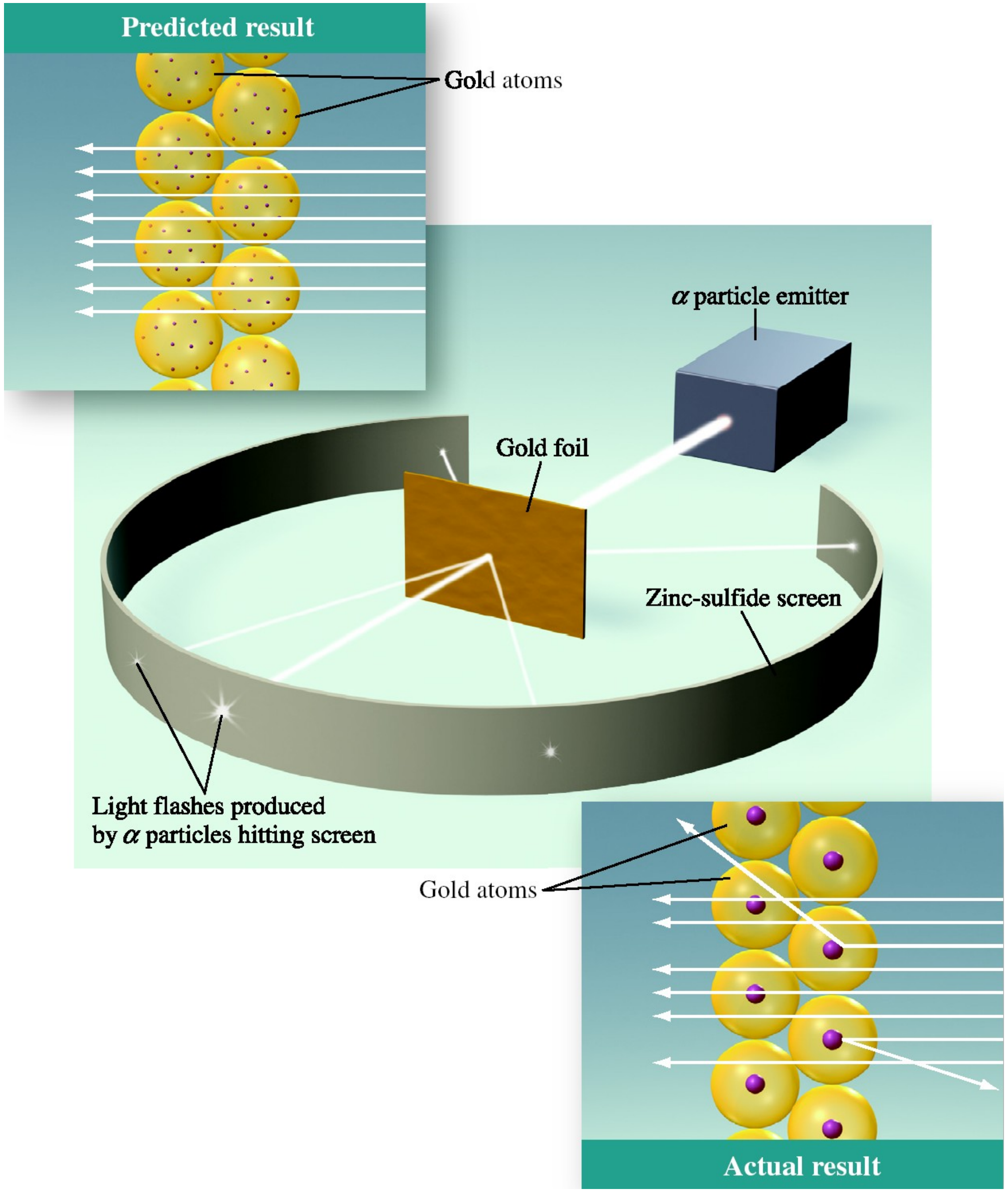
- Antoine Becquerel - French physicist
  - Accidentally discovered that uranium darkened photographic film
- Marie Curie (a student of Becquerel) suggested name “radioactivity”
  - Rays were highly energetic and not deflected by a magnet
  - However, rays arose spontaneously unlike the rays discovered by Röentgen

- Radioactive, describes a substance that spontaneously emits radiation
- Type of radioactivity
  - Alpha **(α)** - positively charged particles
  - Beta ***(β)*** - electrons
  - Gamma ***(γ)*** - no charge and are unaffected by external electric or magnetic fields.



- The Nuclear Atom
  - 1900’s - atoms consisted of positively charged matter with electrons scattered throughout (plum-pudding model by Thomson)
  - 1910 - Rutherford performed “gold-foil” experiment.
    - Proposed a new model of the atom
    - Nucleus contained
      - Positive charges (later called protons)
      - Most of the mass of the atom

Rutherford's Experiment



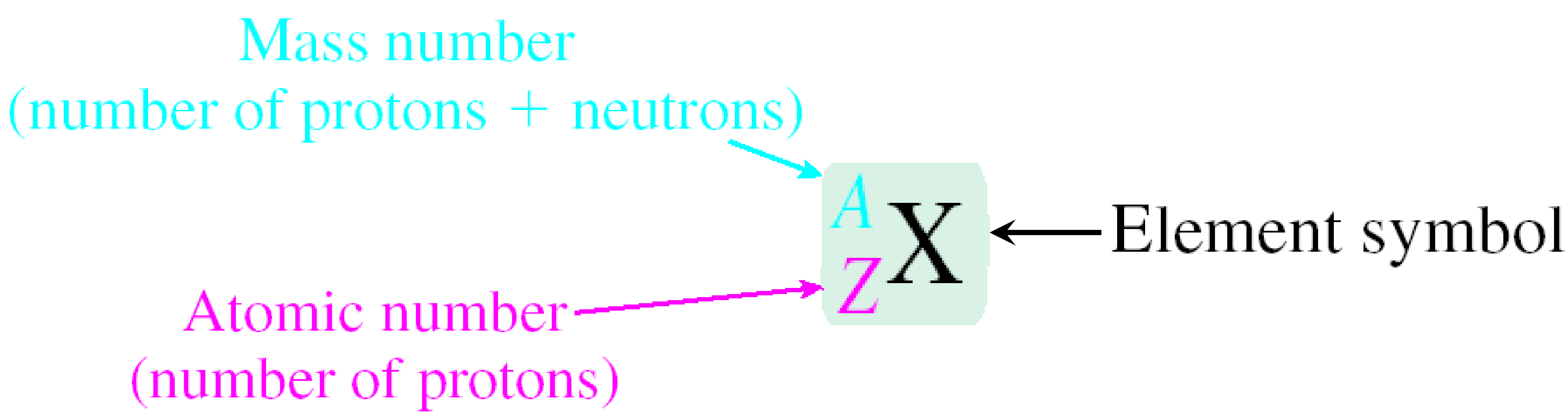
- Rutherford’s model left one problem:
  - If H has a mass of 1
  - Then He should have a mass of 2
  - But its mass is 4!
- 1932 - James Chadwick
  - Discovered the neutron
    - Third subatomic particle
    - Neutral charge

2.3 Atomic Number, Mass Number and Isotopes

- The chemical identity of an atom can be determined solely from its atomic number
- ***Atomic number (Z)*** - number of protons in the nucleus of each atom of an element
  - Also indicates number of *electrons* in the atom—since atoms are *neutral*

- ***Mass number (A)*** - total number of neutrons *and* protons present in the nucleus

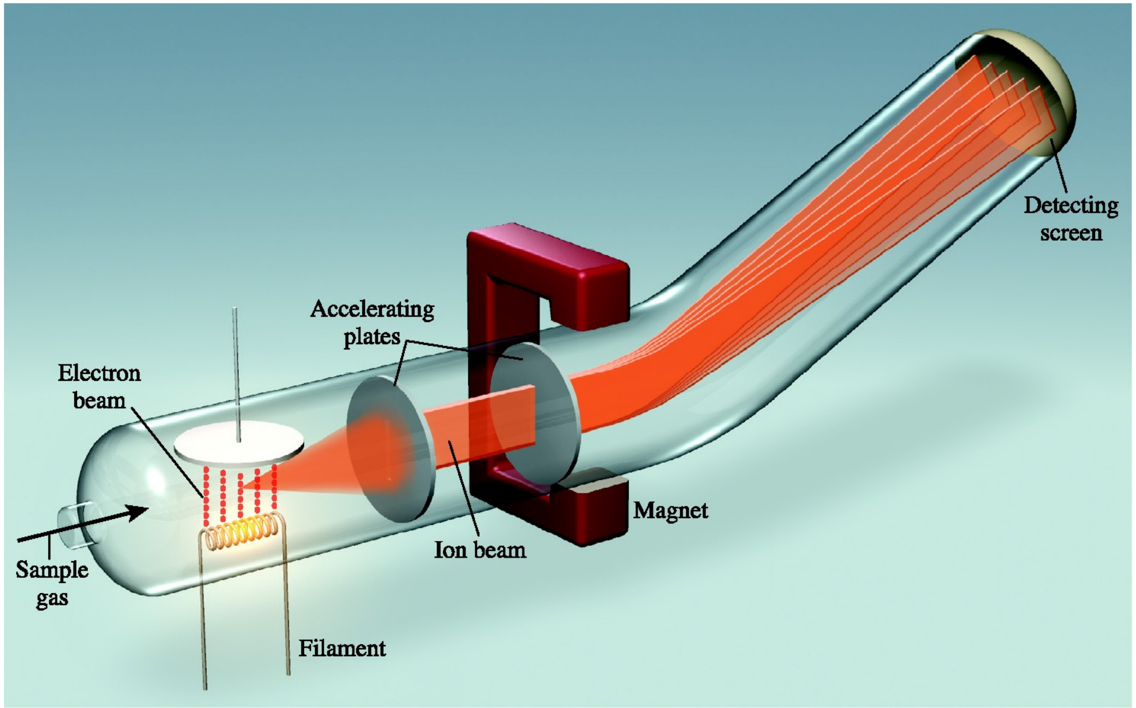
$$\text{mass number } (A) = \text{number of protons } (Z) + \text{number of neutrons}$$



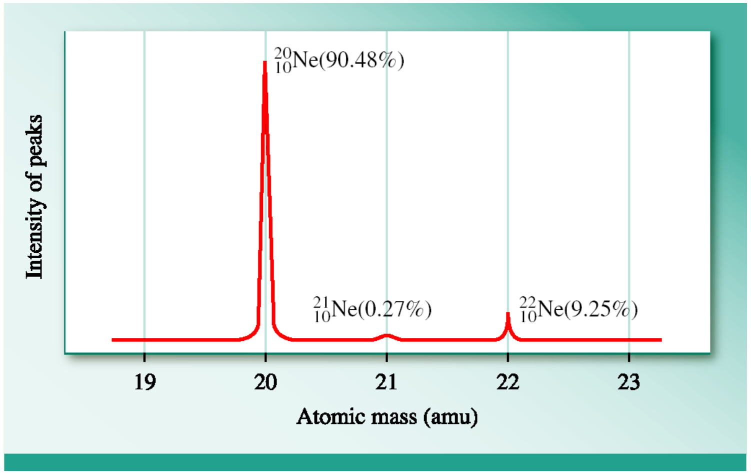
- Isotopes
  - All atoms are *not* identical (as had been proposed by Dalton)
  - Same atomic number (Z) but different mass numbers (A)
- Isotopes of Hydrogen
  - Hydrogen (protium) $^{1}_{1}\mathrm{H}$
  - Deuterium $^{2}_{1}\mathrm{H}$
  - Tritium $^{3}_{1}\mathrm{H}$

Mass Spectrometery

Sample In



Data Out



2.4 The 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

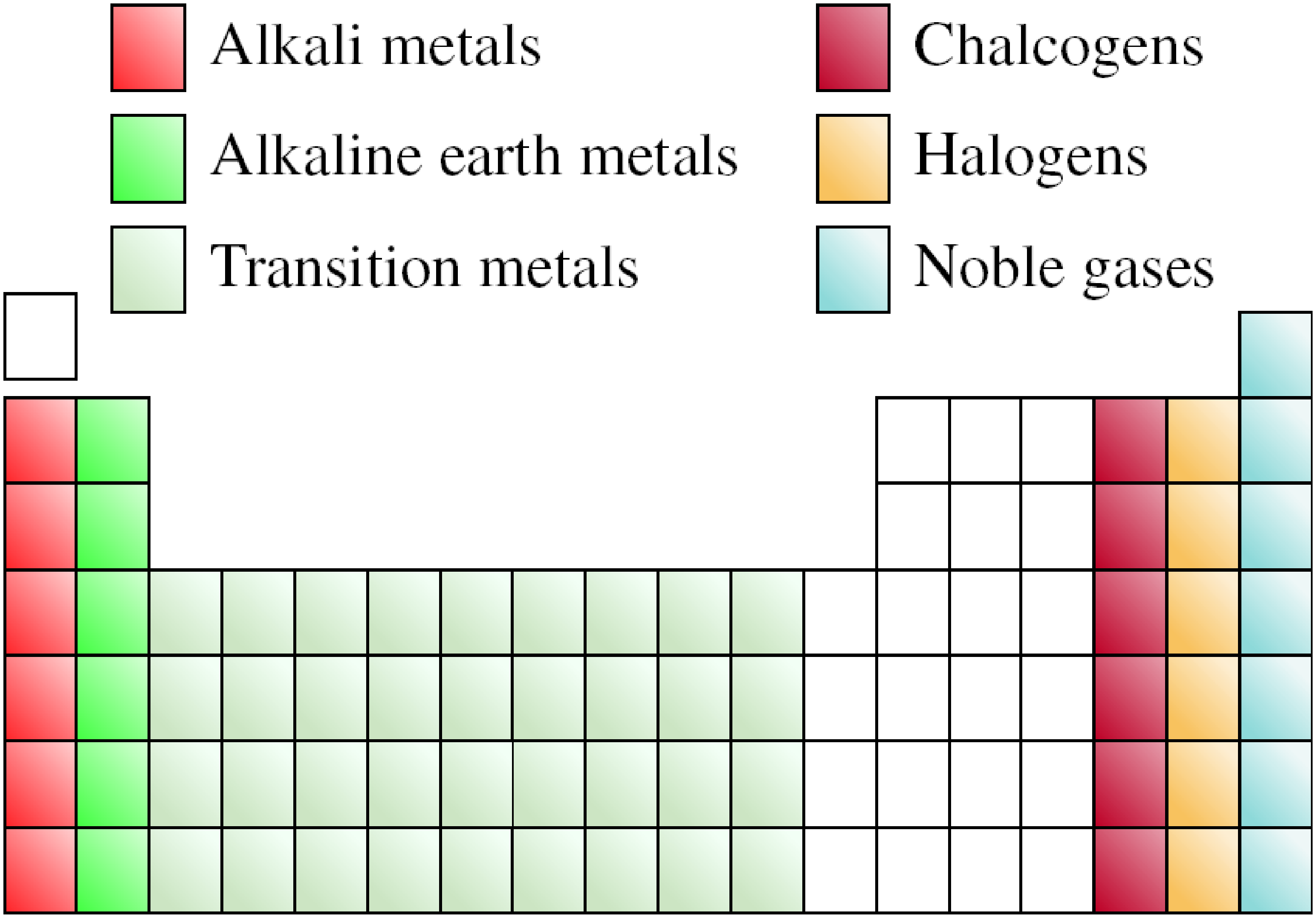
The Modern Periodic Table

1A 1	2A 2	3B 3	4B 4	5B 5	6B 6	7B 7	8B 8	9	10	1B 11	2B 12	3A 13	4A 14	5A 15	6A 16	7A 17	8A 18
1 H Hydrogen																	2 He Helium
3 Li Lithium	4 Be Beryllium											5 B Boron	6 C Carbon	7 N Nitrogen	8 O Oxygen	9 F Fluorine	10 Ne Neon
11 Na Sodium	12 Mg Magnesium											13 Al Aluminum	14 Si Silicon	15 P Phosphorus	16 S Sulfur	17 Cl Chlorine	18 Ar Argon
19 K Potassium	20 Ca Calcium	21 Sc Scandium	22 Ti Titanium	23 V Vanadium	24 Cr Chromium	25 Mn Manganese	26 Fe Iron	27 Co Cobalt	28 Ni Nickel	29 Cu Copper	30 Zn Zinc	31 Ga Gallium	32 Ge Germanium	33 As Arsenic	34 Se Selenium	35 Br Bromine	36 Kr Krypton
37 Rb Rubidium	38 Sr Strontium	39 Y Yttrium	40 Zr Zirconium	41 Nb Niobium	42 Mo Molybdenum	43 Tc Technetium	44 Ru Ruthenium	45 Rh Rhodium	46 Pd Palladium	47 Ag Silver	48 Cd Cadmium	49 In Indium	50 Sn Tin	51 Sb Antimony	52 Te Tellurium	53 I Iodine	54 Xe Xenon
55 Cs Cesium	56 Ba Barium	71 Lu Lutetium	72 Hf Hafnium	73 Ta Tantalum	74 W Tungsten	75 Re Rhenium	76 Os Osmium	77 Ir Iridium	78 Pt Platinum	79 Au Gold	80 Hg Mercury	81 Tl Thallium	82 Pb Lead	83 Bi Bismuth	84 Po Polonium	85 At Astatine	86 Rn Radon
87 Fr Francium	88 Ra Radium	103 Lr Lawrencium	104 Rf Rutherfordium	105 Db Dubnium	106 Sg Seaborgium	107 Bh Bohrium	108 Hs Hassium	109 Mt Meitnerium	110 Ds Darmstadtium	111 Rg Roentgenium	112 — -	113 — -	114 — -	115 — -	116 — -	117	118 — -

57 La Lanthanum	58 Ce Cerium	59 Pr Praseodymium	60 Nd Neodymium	61 Pm Promethium	62 Sm Samarium	63 Eu Europium	64 Gd Gadolium	65 Tb Terbium	66 Dy Dysprosium	67 Ho Holmium	68 Er Erbium	69 Tm Thulium	70 Yb Ytterbium
89 Ac Actinium	90 Th Thorium	91 Pa Protactinium	92 U Uranium	93 Np Neptunium	94 Pu Plutonium	95 Am Americium	96 Cm Curium	97 Bk Berkelium	98 Cf Californium	99 Es Einsteinium	100 Fm Fermium	101 Md Mendelevium	102 No Nobelium

- **Metals** - good conductors of heat and electricity (majority of elements on the table, located to the left of the stair step)
- **Nonmetals** - nonconductors (located in upper right-hand corner)
- **Metalloids** - in between metals and nonmetals (those that lie along the separation line)

Groups (Families) on the Periodic Table



2.5 The Atomic Mass Scale and Average atomic Mass

- ***Atomic mass*** is the mass of the atom in atomic mass units (amu)
- ***Atomic mass unit*** is defined as a mass exactly equal to one-twelfth the mass of one carbon-12 atom
- Carbon-12 (12 amu) provides the standard for measuring the atomic mass of the other elements

Why 1/12 the mass of Carbon-12? You might think that this would work with any particular isotope (e.g. 1/7 the mass of Lithium-7).

The problem is that there is a difference in the mass of the atom as a whole, and the sum of the masses of its individual protons, neutrons, and electrons.

When the particles bind together to form an atom, some of the mass is lost as “binding energy” according to:

$$E = mc^2$$

Different atoms have differing amounts of binding energy, so one particular isotope had to be selected as the standard.

- Average atomic mass
  - Masses on the periodic table are not whole numbers.
  - Mass spectrometer provides information about percentages of different isotopes of each element.
  - Periodic table mass is the weighted average of all of the isotopes of each element

Oxygen is the most abundant element both in the Earth's crust and in the human body. The atomic masses of its three stable isotopes, ^{16}O (99.757%), ^{17}O (0.038%), and ^{18}O (0.205%), are 15.9949 amu, 16.9999 amu, and 17.9999 amu, respectively. Calculate the average atomic mass of oxygen using the relative abundances given in parentheses.

Steps:

1. Convert each % into decimal abundance. (divide by 100)
2. Multiply mass of each isotope by its fractional abundance.
3. Add the contributions together.

$$(0.99757)(15.9949 \text{ amu}) +$$
$$(0.00038)(16.9999 \text{ amu}) +$$
$$(0.00205)(17.9999) = 15.999 \text{ amu}.$$

*To four significant figures, this is the same as the mass given in the periodic table in the book: 16.00

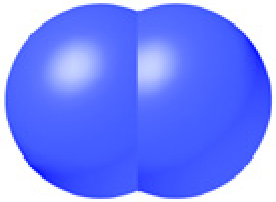
The atomic masses of the two stable isotopes of copper, ^{63}Cu (69.17%) and ^{65}Cu (30.83%), are 62.929599 amu and 64.927793 amu, respectively. Calculate the average atomic mass of copper.

(0.6917)(62.929599 amu) +
(0.3083)(64.927793 amu) + =
63.55 amu.

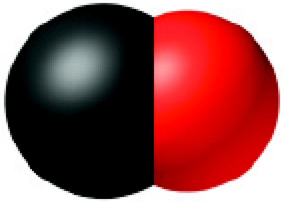
2.6 Molecules and Molecular Compounds

- ***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

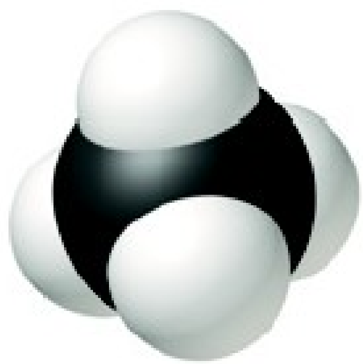
- **Diatomic molecules**:
  - Homonuclear (2 of the same atoms)
    - Examples: H_2, N_2, O_2, F_2, Cl_2, Br_2, and I_2



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



- **Polyatomic molecules**:
  - 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)



- ***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: $C_{12}H_{22}O_{11}$

- ***Allotrope*** - one of two or more distinct forms of an element
  - Examples: oxygen, O_2 and ozone, O_3; diamond and graphite (allotropic forms of carbon)
- ***Structural formula*** - shows the general arrangement of atoms within the molecule.

$$H{-}O{-}H$$

- Naming molecular compounds
  - Binary Molecular compounds
    - Composed of two nonmetals
    - Name the first element
    - Name the second element changing ending to “-ide”
    - Use prefixes to indicate number of atoms of each element

TABLE 2.2 Greek Prefixes

Prefix	Meaning	Prefix	Meaning
Mono-	1	Hexa-	6
Di-	2	Hepta-	7
Tri-	3	Octa-	8
Tetra-	4	Nona-	9
Penta-	5	Deca-	10

TABLE 2.3 Some Compounds Named Using Greek Prefixes

Compound	Name	Compound	Name
CO	Carbon monoxide	SO_3	Sulfur trioxide
CO_2	Carbon dioxide	NO_2	Nitrogen dioxide
SO_2	Sulfur dioxide	N_2O_5	Dinitrogen pentoxide

Name the following:

NO_2

nitrogen dioxide

N_2O_4

dinitrogen tetraoxide

Write formulas for the following:

Diphosphorus pentoxide

P_2O_5

Sulfur hexafluoride

SF_6

Common Names

B_2H_6 diborane

SiH_4 silane

NH_3 ammonia

PH_3 phosphine

H_2O water (vs. “dihydrogen monoxide)

H_2S hydrogen sulfide (vs. “dihydrogen sulfide”)

- ***Acid*** - a substance that produces hydrogen ions (H^+) when dissolved in water
- Binary acids:
  - Many have 2 names
    - Pure substance
    - Aqueous solution
  - Example: HCl, hydrogen chloride, when dissolved in water it is called hydrochloric acid

- Naming 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

Write formulas for the following:

Hydrochloric acid

$HCl(aq)$

Hydrofluoric acid

$HF(aq)$

- ***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
  - Many derivatives of alkanes are derived by replacing a hydrogen with one of the ***functional groups.***
    - Functional group determines chemical properties

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

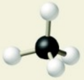
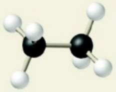
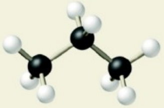
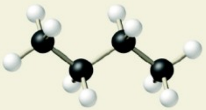
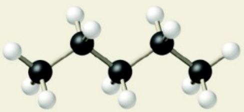
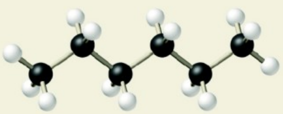
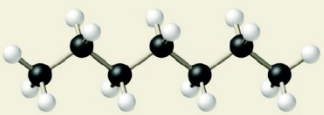
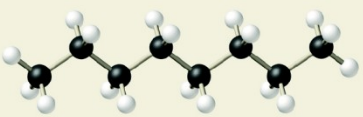
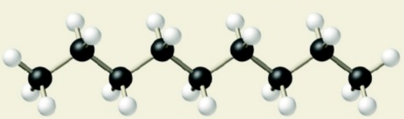
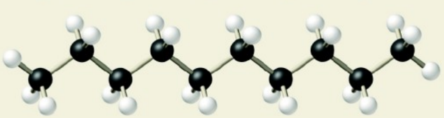
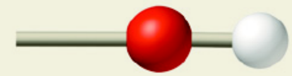
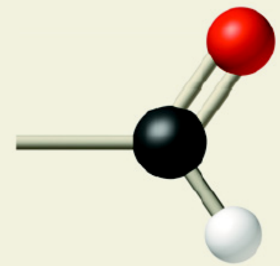
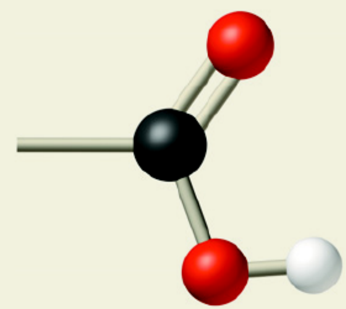
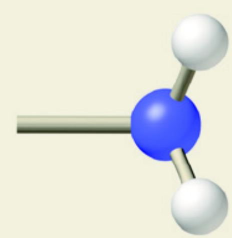
Formula	Name	Model
CH_4	Methane	
C_2H_6	Ethane	
C_3H_8	Propane	
C_4H_{10}	Butane	
C_5H_{12}	Pentane	
C_6H_{14}	Hexane	
C_7H_{16}	Heptane	
C_8H_{18}	Octane	
C_9H_{20}	Nonane	
$C_{10}H_{22}$	Decane	

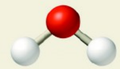
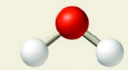
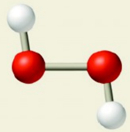
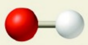
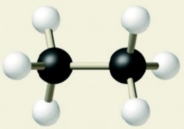
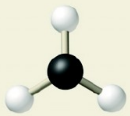
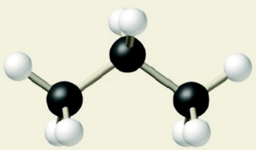
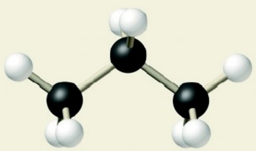
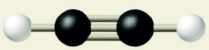
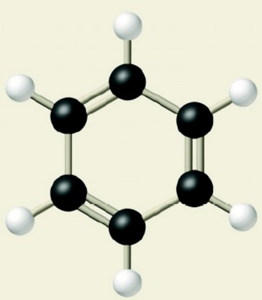
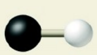
TABLE 2.6 Organic Functional Groups

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

- ***Empirical formulas*** *reveal the* elements present and in what whole-number ratio they are combined.

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

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		CH_3	
Propane	C_3H_8		C_3H_8	
Acetylene	C_2H_2		CH	
Benzene	C_6H_6		CH	

2.7 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

Common Monatomic Ions

1A 1	2A 2	3B 3	4B 4	5B 5	6B 6	7B 7	8B 8	9	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+}				

- Naming ions
  - Cations from A group metals
    - Name the element and add the word “ion”
    - Example: Na^{+}, sodium ion
  - Cations from transition metals with some exceptions
    - Name element
    - Indicate charge of metal with Roman numeral
    - Add word “ion”
    - Example: Cu^{2+} ,copper(II) ion

- Anions
  - Name the element and modify the ending to "-ide"
  - Example: Cl^-, chloride

- ***Polyatomic ions*** - ions that are a combination of two or more atoms
  - Notice similarities - number of oxygen atoms and endings for oxoanions
    - Nitrate, NO_3^- and nitrite, NO_2^-

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^-

- Ionic compounds - represented by empirical formulas
  - Compound formed is electrically neutral
  - Sum of the charges on the cation(s) and anion(s) in each formula unit must be zero
  - Examples:

Al^{3+} and O^{2-} Al_2O_3

Ca^{2+} and PO_4^{3-} $Ca_3(PO_4)_2$

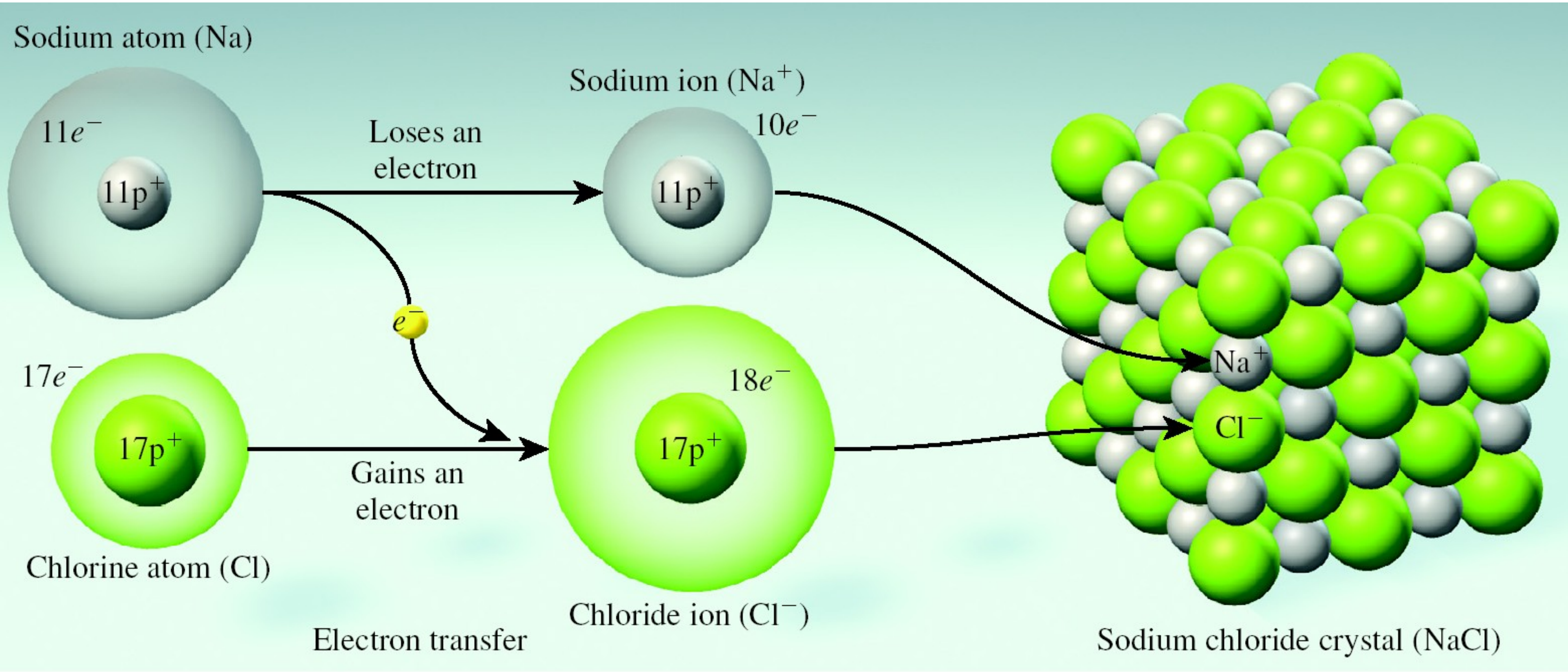
Al^{3+} O^{2-}

Al_2O_3

Ca^{2+} PO_4^{3-}

$Ca_3(PO_4)_2$

Formation of an Ionic Compound



Write empirical formulas for

- Aluminum and bromide

$AlBr_3$

- barium and phosphate

$Ba_3(PO_4)_2$

- Magnesium and nitrate

$Mg(NO_3)_2$

- Ammonium and sulfate

$(NH_4)_2SO_4$

- Naming ionic compounds
  - Name the cation
  - Name the anion
  - Check the name of cation
    - If it is a A group metal you are finished
    - If it is a transition metal, with some exceptions, add the appropriate Roman numeral to indicate the positive ionic charge

Write names for the following:

- $KMnO_4$
- $Sr_3(PO_4)_2$
- $Co(NO_3)_2$
- $FeSO_4$

Write names for the following:

- $KMnO_4$

  potassium permanganate
- $Sr_3(PO_4)_2$

  strontium phosphate
- $Co(NO_3)_2$

  cobalt(II) nitrate
- $FeSO_4$

  iron(II) sulfate

- Oxoacids
  - When writing formulas, add the number of H^+ ions necessary to balance the corresponding oxoanion's negative charge
  - Naming formulas
    - If the anion ends in "-ite" the acid ends with "-ous" acid
    - If the anion ends in "-ate" the acid ends in "-ic" acid

Name the following:

- H_2SO_3
- $HClO$
- H_3PO_4

Write formulas for the following:

- Perchloric acid
- Nitric acid

Name the following:

- H_2SO_3

  sulfurous acid

- HClO

  hypochlorous acid

- H_3PO_4

  phosphoric acid

Write formulas for the following:

- Perchloric acid

  $HClO_4$

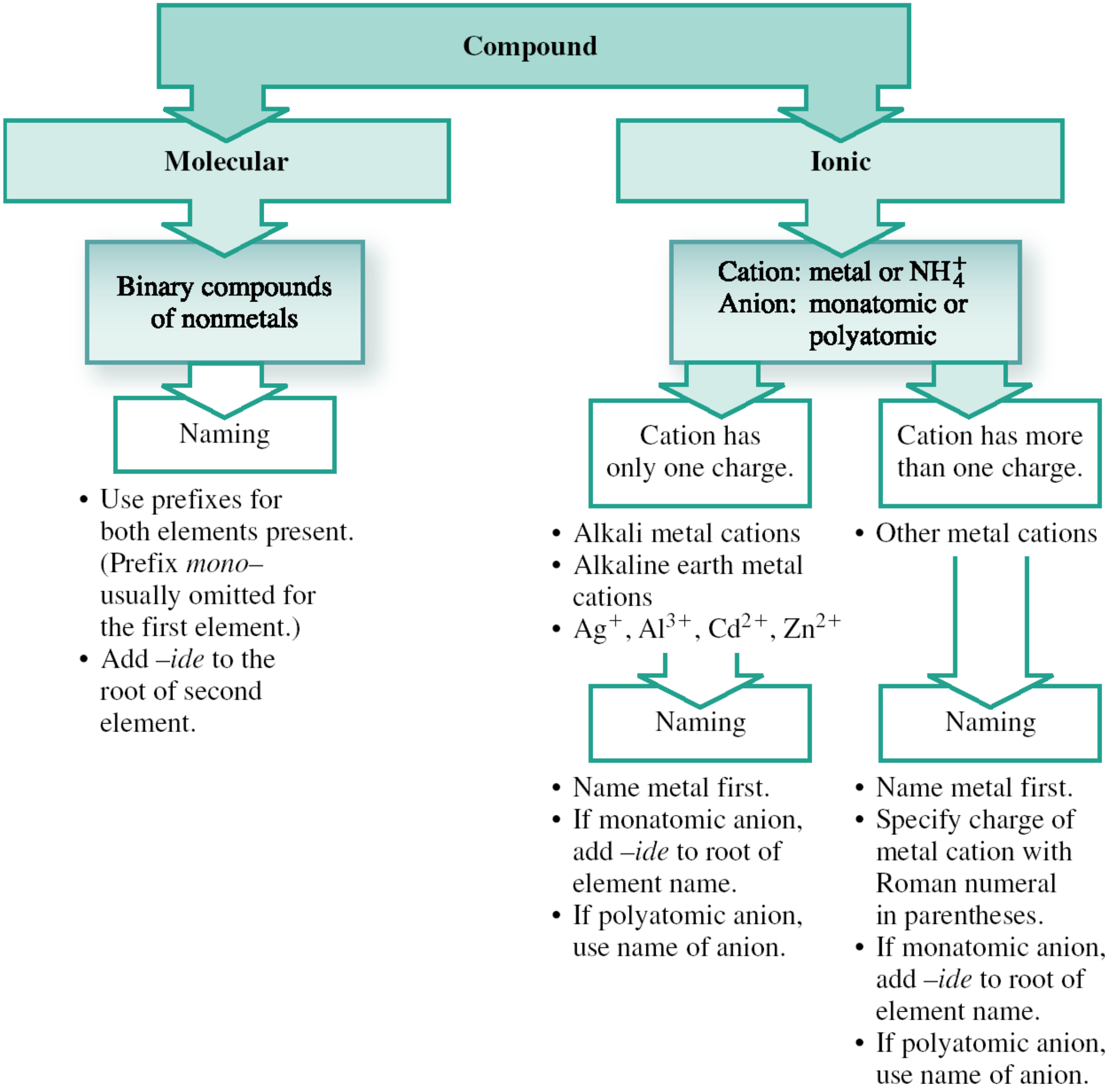
- Nitric acid

  HNO_3

- ***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

TABLE 2.10 Common and Systematic Names of Some Familiar Inorganic Compounds

Formula	Common Name	Systematic Name
H_2O	Water	Dihydrogen monoxide
NH_3	Ammonia	Trihydrogen nitride
CO_2	Dry ice	Solid carbon dioxide
NaCl	Salt	Sodium chloride
N_2O	Laughing gas	Dinitrogen monoxide
$CaCO_3$	Marble, chalk, limestone	Calcium carbonate
$NaHCO_3$	Baking soda	Sodium hydrogen carbonate
$MgSO_4 \cdot 7H_2O$	Epsom salt	Magnesium sulfate heptahydrate
$Mg(OH)_2$	Milk of magnesia	Magnesium hydroxide


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

Key Points

- Atomic theory - parts of the atom; theories; laws
- Types of radiation
- Atomic number and mass number
- Isotopes
- Periodic table; families and periods; metals, nonmetals and metalloids

Key Points

- Average atomic mass
- Naming and writing formulas for
  - Binary molecular compounds
  - Binary acids
  - Ionic compounds
  - Oxoacids
  - Hydrates