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Chapter 2
Atoms, Molecules, and Ions

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Learning a Language

- When learning a new language:
  - Start with the alphabet
  - Then, form words
  - Finally, form more complex structures such as sentences
- Chemistry has an alphabet and a language; in this chapter, the fundamentals of the language of chemistry will be introduced

Outline

- Atoms and Atomic Theory
- Components of the Atom
- Quantitative Properties of the Atom
- Introduction to the Periodic Table
- Molecules and Ions
- Formulas of Ionic Compounds
- Names of Compounds

The Language of Chemistry

- This chapter introduces the fundamental language of chemistry
  - Atoms, molecules and ions
  - Formulas
  - Names

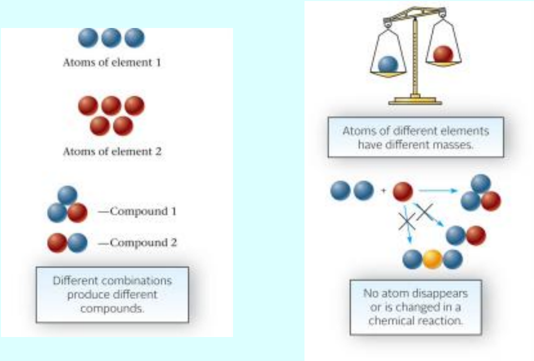
The Structure of Matter

- Atoms
  - Composed of electrons, protons and neutrons
- Molecules
  - Combinations of atoms
- Ions
  - Charged particles

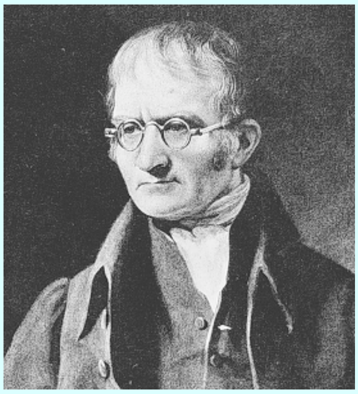
Atoms and Atomic Theory

- An element is composed of tiny particles called atoms
  - All atoms of the same element have the same chemical properties
- In an ordinary chemical reaction
  - There is a change in the way atoms are combined with each other
  - Atoms are not created or destroyed
- Compounds are formed when two or more atoms of different elements combine

Figure 2.1 - John Dalton and Atomic Theory



John Dalton



Fundamental Laws of Matter

- There are three fundamental laws of matter
  - Law of conservation of mass
    - Matter is conserved in chemical reactions
  - Law of constant composition
    - Pure water has the same composition everywhere
  - Law of multiple proportions
    - Compare Cr_2O_3 to CrO_3
    - The ratio of Cr:O between the two compounds is a small whole number

Figure A – The Law of Multiple Proportions

Two different oxides of chromium

Components of the Atom

- Atomic theory raised more questions than it answered
  - Could atoms be broken down into smaller particles
  - 100 years after atomic theory was proposed, the answers were provided by experiment

Fundamental Experiments

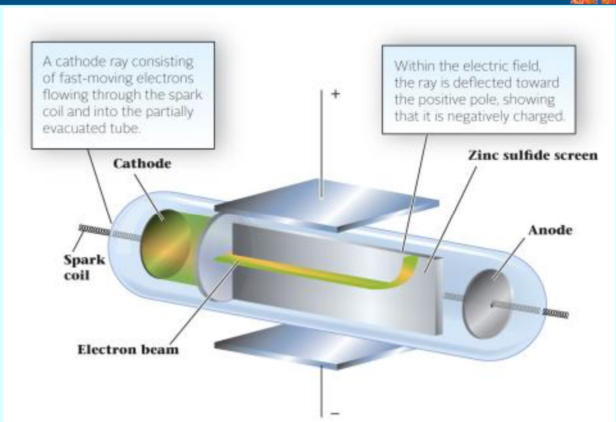
- J.J. Thomson, Cavendish Laboratories, Cambridge, England
- Ernest Rutherford
  - McGill University, Canada
  - Manchester and Cambridge Universities, England

Figure 2.2 – J.J. Thomson and Ernest Rutherford

Electrons

- First evidence for subatomic particles came from the study of the conduction of electricity by gases at low pressures
  - J.J. Thomson, 1897
  - Rays emitted were called cathode rays
  - Rays are composed of negatively charged particles called electrons
  - Electrons carry unit negative charge (-1) and have a very small mass (1/2000 the lightest atomic mass)

Figure 2.3 – Cathode Ray Apparatus



The Electron and the Atom

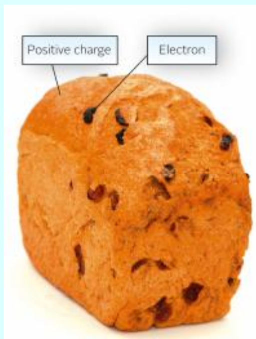
- Every atom has at least one electron
- Atoms are known that have one hundred or more electrons
- There is one electron for each positive charge in an atom
- Electrical neutrality is maintained

Protons and Neutrons – The Nucleus

- Ernest Rutherford, 1911
- Bombardment of gold foil with α particles (helium atoms without electrons)
  - Expected to see the particles pass through the foil
  - Found that some of the alpha particles were deflected by the foil
  - Led to the discovery of a region of heavy mass at the center of the atom

The Plum-Pudding Model

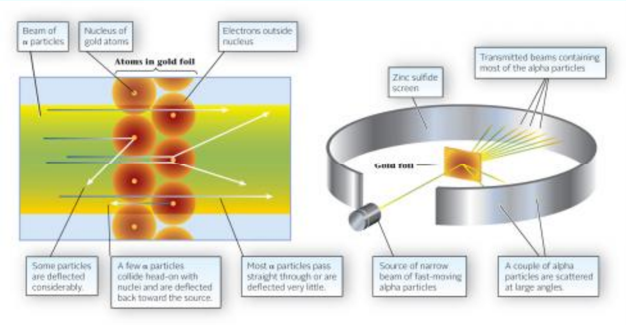
- J.J. Thomson proposed the atom as a positively charged sphere
  - Within the sphere are electrons
  - Plum-pudding or raisin-bread model



Rutherford's Model

- Rutherford's experiment revealed a small, dense core with positive charge
  - Electrons are outside this core
  - Most of the atom is empty space

Figure 2.5 – Rutherford Backscattering



Nuclear Particles

1. Protons
   - Mass nearly equal to the H atom
   - Positive charge
2. Neutrons
   - Mass slightly greater than that of the proton
   - No charge

Mass and the Atom

- More than 99.9% of the atomic mass is concentrated in the nucleus
- The volume of the nucleus is much smaller than the volume of the atom

Table 2.1 – Subatomic Particles

TABLE 2.1 Properties of Subatomic Particles

Particle	Location	Relative Charge	Relative Mass*
Proton	Nucleus	+1	1.00728
Neutron	Nucleus	0	1.00867
Electron	Outside nucleus	−1	0.00055

*These are expressed in atomic mass units.

Terminology

- Atomic number, Z
  - Number of protons in the atom
- Mass number, A
  - Number of protons plus number of neutrons

Isotopes

- Isotopes are two atoms of the same element
  - Same atomic number
  - Different mass numbers
    - Number of neutrons is A minus Z
    - Number of neutrons differs between isotopes

Nuclear symbolism

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

- A is the mass number
- Z is the atomic number
- X is the chemical symbol

Isotopes of hydrogen

- ${}^{1}H$, ${}^{2}H$, ${}^{3}H$
  - Hydrogen, deuterium, tritium
  - Different masses

Note that some of the ice is at the bottom of the glass – this is ${}^{2}H_2O$

Example 2.1

EXAMPLE 2.1

a. An isotope of cobalt (Co, $Z = 27$) is used in radiation therapy for cancer. This isotope has 33 neutrons in its nucleus. What is its nuclear symbol?

b. One of the most harmful components of nuclear waste is a radioactive isotope of strontium, ${}^{90}_{38}Sr$; it can be deposited in your bones, where it replaces calcium. How many protons are in the nucleus of Sr-90? How many neutrons?

c. Write the nuclear symbol for the element used in diagnostic bone scans. It has 31 protons and 38 neutrons.

(a)

ANALYSIS	
Information given:	Z (27); number of neutrons, n (33)
Asked for:	nuclear symbol

STRATEGY

1. Note that Z stands for the atomic number or the number of protons p^+.
2. Recall that a nuclear symbol is written ${}^{A}_{Z}X$ where A stands for the number of neutrons (n) plus protons (p^+).

SOLUTION	
nuclear symbol	$Z = p^+ = 27$; $A = p^+ + n = 27 + 33$; ${}^{A}_{Z}Co = {}^{60}_{27}Co$

Example 2.1, (Cont'd)

(b)

ANALYSIS	
Information given:	nuclear symbol: ${}^{90}_{38}Sr$
Asked for:	p^+; n

SOLUTION	
protons	$Z = p^+ = 38$
neutrons	$A = p^+ + n = 90$; $90 = 38 + n$; $n = 90 - 38 = 52$

(c)

ANALYSIS	
Information given:	$p^+ = 31$; $n = 38$
Information implied:	identity of the element
Asked for:	nuclear symbol

SOLUTION	
nuclear symbol	$Z = p^+ = 31$ (placed on bottom left of element) $A = p^+ + n = 31 + 38 = 69$ The element (X) is gallium identified by its atomic number Z. nuclear symbol: ${}^{A}_{Z}X = {}^{69}_{31}Ga$

Atomic Masses: The Carbon-12 Scale

- The atomic mass of an element indicates how heavy, on average, an atom of an element is when compared to an atom of another element
- Unit is the atomic mass unit (amu)
- Mass of one ${}^{12}C$ atom = 12 amu (exactly)
- Note that ${}^{12}C$ and C-12 mean the same thing

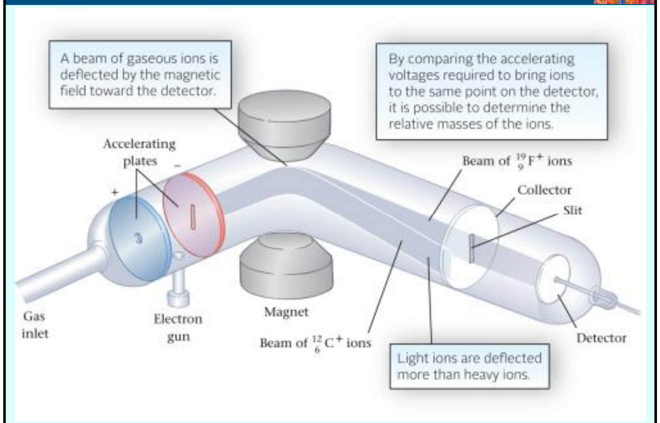
Determining Atomic Masses

- Atomic masses can be determined to highly precise values by using a mass spectrometer
- The mass spectrometer separates matter based on its mass and charge
- The resulting data is plotted with abundance on the y-axis and mass on the x-axis

Mass Spectrometry

- Atoms are ionized at low pressure in the gas phase
- The cations that form are accelerated toward a magnetic field
- The extent to which the cation beam is deflected is inversely related to the mass of the cation

Figure 2.6: The Mass Spectrometer



Fluorine

- Atomic fluorine exists as a single isotope
- Mass of F is exactly 19.00 amu

Carbon

- ^{12}C has a mass of exactly 12.000 amu
- Carbon in the periodic table has a mass of 12.011 amu
  - Why isn't it exactly 12.000?
  - Why are most atomic masses not whole numbers?

Isotopes

- Recall that an isotope is an atom with the same number of protons
  - Therefore, the same element
- Different mass number
  - Therefore, a different number of neutrons

Averages

- Take the simple average of two numbers
- Add the numbers and divide by 2
  - Each number counts 50%
  - (10 + 15) / 2 = 12.5
  - (0.5 X 10) + (0.5 X 15) = 12.5
- A simple average is simply a weighted average where each contributor counts 50%

Isotopic Abundance

- To determine the mass of an element, we must know the mass of each isotope and the atom percent of the isotopes (***isotopic abundance***)
- The mass spectrometer can determine the isotopic abundance

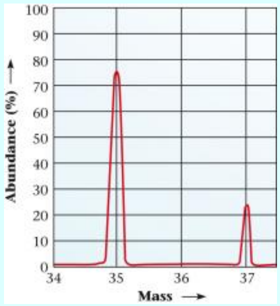
Chlorine

- There are two isotopes of chlorine

	Atomic mass	Abundance
Cl-35	34.97 amu	75.53%
Cl-37	36.97 amu	24.47%

Figure 2.7 – Mass Spectrum of Cl

- The mass spectrum is a plot of abundance vs. mass
- The area under the peak in the mass spectrogram gives the isotopic abundance



Average Atomic Masses

- The atomic mass quoted in the periodic table is the weighted average of the atomic masses of all isotopes of that element

Weighted Averages

- For an element with two isotopes, Y_1 and Y_2:

$$\text{atomic mass Y} = (\text{atomic mass } Y_1)\frac{\%Y_1}{100} + (\text{atomic mass } Y_2)\frac{\%Y_2}{100}$$

- For chlorine:

$$\textit{atomic mass Cl} = 34.97\ \textit{amu}\left(\frac{75.53}{100}\right) + 36.97\ \textit{amu}\left(\frac{24.47}{100}\right)$$

$$\textit{atomic mass Cl} = 35.46\ \textit{amu}$$

Relative Atomic Masses

- The mass of an element is indicated below the symbol for the element
- Consider He
  - On average, one He atom weighs 4.003 amu
  - This is about 1/3 the mass of a carbon-12 atom

$$\frac{4.003\ \text{amu}}{12.00\ \text{amu}} = 0.3336$$

Example 2.2

EXAMPLE 2.2

Bromine is a red-orange liquid with an average atomic mass of 79.90 amu. Its name is derived from the Greek word *bromos* (βρομος), which means stench. It has two naturally occurring isotopes: Br-79 (78.92 amu) and Br-81 (80.92 amu). What is the abundance of the heavier isotope?

ANALYSIS

Information given:	Br-81 mass (80.92 amu); Br-79 mass (78.92 amu) average atomic mass (79.90)
Asked for:	abundance of Br-81

STRATEGY

1. All abundances must add up to 100%
2. Recall the formula relating abundance and atomic mass (Equation 2.1)

$$\text{atomic mass Y} = (\text{atomic mass } Y_1 \times \frac{\%\ Y_1}{100\%}) + (\text{atomic mass } Y_2 \times \frac{\%\ Y_2}{100\%}) + \cdots$$

continued

Example 2.2, (Cont'd)

SOLUTION

1. % abundances	Br-81: x; Br-79: $100 - x$
2. Substitute into Equation 2.1.	$79.90\ \text{amu} = 78.92\ \text{amu}\left(\frac{100-x}{100}\right) + 80.92\ \text{amu}\left(\frac{x}{100}\right)$
3. Solve for x.	$79.90 = 0.7892(100 - x) + 0.8092\,x$ $79.90 = 78.92 - 0.7892\,x + 0.8092\,x$ $x = 49\%$

END POINT

The atomic mass of Br, 79.90, is just about halfway between the masses of the two isotopes, 78.92 and 80.92. So, it is reasonable that it should contain nearly equal amounts of the two isotopes.

Masses of Individual Atoms

- For many purposes, relative masses are not sufficient
- It is necessary to know the mass of an atom in grams so the quantity of matter can be determined by weighing
- The number that converts the mass of an atom in atomic mass units to the mass of a collection of atoms in grams is called ***Avogadro's Number***

Masses of Atoms in Grams

- It is usually sufficient to know the relative masses of atoms
  - One He atom is about four times as heavy as one H atom
  - Therefore
    - The mass of 100 He atoms is about four times the mass of 100 H atoms
    - The mass of a million He atoms is about four times the mass of a million H atoms

Example 2.3

EXAMPLE 2.3

Consider arsenic (As), a favorite poison used in crime stories. This element is discussed at the end of Chapter 1. Taking Avogadro's number to be 6.022×10^{23}, calculate

(a) the mass of an arsenic atom.
(b) the number of atoms in a ten-gram sample of arsenic.
(c) the number of protons in 0.1500 lb of arsenic.

(a)

ANALYSIS

Information given:	Avogadro's number (6.022×10^{23})
Information implied:	atomic mass
Asked for:	mass of an arsenic atom

STRATEGY

Change atoms to grams (atoms → g) by using the conversion factor

$$\frac{6.022 \times 10^{23}\ \text{atoms}}{\text{atomic mass}}$$

SOLUTION

mass of an As atom	$1\ \text{atom As} \times \frac{74.92\ \text{g As}}{6.022 \times 10^{23}\ \text{atoms As}} = 1.244 \times 10^{-22}\ \text{g}$

(b)

ANALYSIS

Information given:	mass of sample (10.00 g) from (a) mass of one As atom (1.244×10^{-22} g/atom)
Asked for:	number of atoms in a 10-gram sample

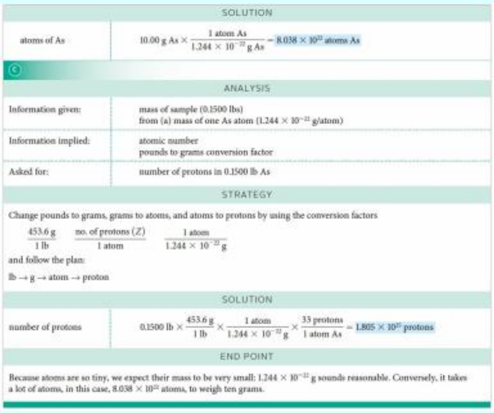
STRATEGY

Change grams to atoms (g → atom) by using the conversion factor

$$\frac{1\ \text{atom}}{1.244 \times 10^{-22}\ \text{g}}$$

continued

Example 2.3, (Cont'd)



SOLUTION	
atoms of As	$10.00\ \text{g As} \times \frac{1\ \text{atom As}}{1.244 \times 10^{-22}\ \text{g As}} = 8.038 \times 10^{22}\ \text{atoms As}$

ANALYSIS	
Information given:	mass of sample (0.1500 lbs) from (a) mass of one As atom (1.244×10^{-22} g/atom)
Information implied:	atomic number pounds to grams conversion factor
Asked for:	number of protons in 0.1500 lb As

STRATEGY

Change pounds to grams, grams to atoms, and atoms to protons by using the conversion factors

$$\frac{453.6\ \text{g}}{1\ \text{lb}} \quad \frac{\text{no. of protons}\ (Z)}{1\ \text{atom}} \quad \frac{1\ \text{atom}}{1.244 \times 10^{-22}\ \text{g}}$$

and follow the plan:

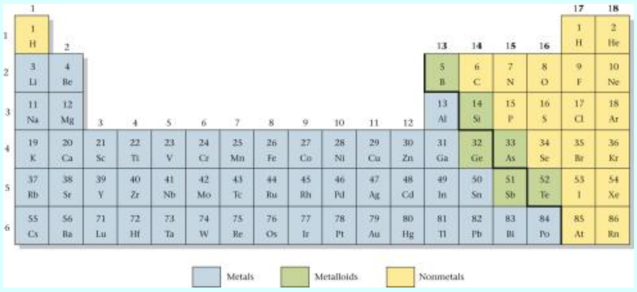
$\text{lb} \rightarrow \text{g} \rightarrow \text{atom} \rightarrow \text{proton}$

SOLUTION	
number of protons	$0.1500\ \text{lb} \times \frac{453.6\ \text{g}}{1\ \text{lb}} \times \frac{1\ \text{atom}}{1.244 \times 10^{-22}\ \text{g}} \times \frac{33\ \text{protons}}{1\ \text{atom As}} = 1.805 \times 10^{25}\ \text{protons}$

END POINT

Because atoms are so tiny, we expect their mass to be very small: 1.244×10^{-22} g sounds reasonable. Conversely, it takes a lot of atoms, in this case, 8.038×10^{22} atoms, to weigh ten grams.

Introduction to the Periodic Table



Periods and Groups

- Horizontal rows are ***periods***
  - First period is H and He
  - Second period is Li-Ne
  - Third period is Na-Ar
- Vertical columns are ***groups***
  - IUPAC convention: use numbers 1-18

Blocks in the Periodic Table

- Main group elements
  - 1, 2, 13-18
- Transition metals
  - 3-12
- Post-transition metals
  - Elements in groups 13-15 to the right of the transition metals
  - Ga, In, Tl, Sn, Pb, Bi, Po

Families with Common Names

- Alkali Metals, Group 1
- Alkaline Earth Metals, Group 2
- Halogens, Group 17
- Noble Gases, Group 18

Importance of Families

- Elements within a family have similar chemical properties
  - Alkali metals are all soft, reactive metals
  - Noble gases are all relatively unreactive gases; He, Ne and Ar do not form compounds

Arrangement of Elements

- Periods
  - Arranged by increasing atomic number
- Families
  - Arranged by chemical properties

Mendeleev

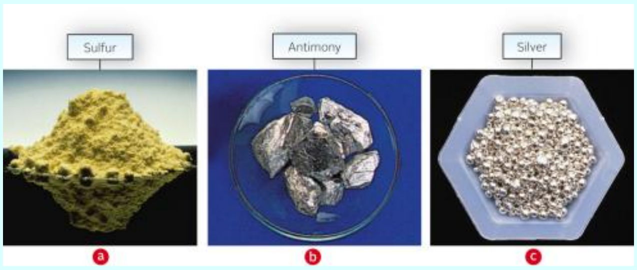
- Dmitri Mendeleev, 1836-1907
- Arranged elements by chemical properties
  - Left spaces for elements yet unknown
  - Predicted detailed properties for elements as yet unknown
    - Sc, Ga, Ge
    - By 1886, all these elements had been discovered, all with properties similar to those he predicted

Metals and Nonmetals

- Diagonal line starting with B separates the metals from the nonmetals
  - Elements along this diagonal have some of the properties of metals and some of the properties of nonmetals
  - Metalloids
    - B, Si, Ge, As, Sb, Te

A Look at the Sulfur Group

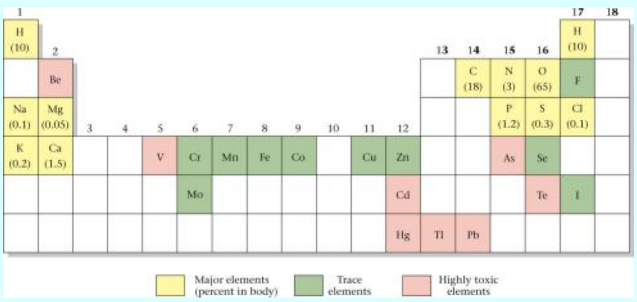
- Sulfur (nonmetal), antimony (metalloid) and silver (metal)



Biological View of the Periodic Table

- "Good guys"
  - Essential to life
  - Carbon, hydrogen, oxygen, sulfur and others
- "Bad guys"
  - Toxic or lethal
  - Some elements are essential but become toxic at higher concentrations
    - Selenium

Figure 2.9 – Biologically Important and Toxic Elements

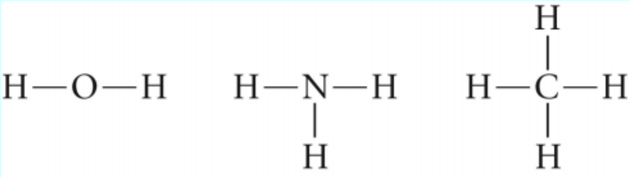


Molecules

- Two or more atoms may combine to form a molecule
  - Atoms involved are often nonmetals
  - Covalent bonds are strong forces that hold the atoms together
- Molecular formulas
  - Number of each atom is indicated by a subscript
  - Examples
    - Water, H_2O
    - Ammonia, NH_3

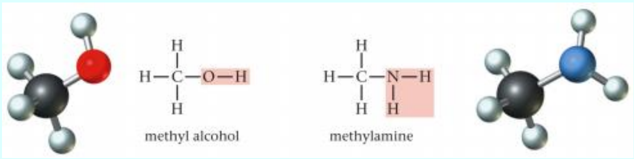
Structural Formulas

- Structural formulas show the bonding patterns within the molecule

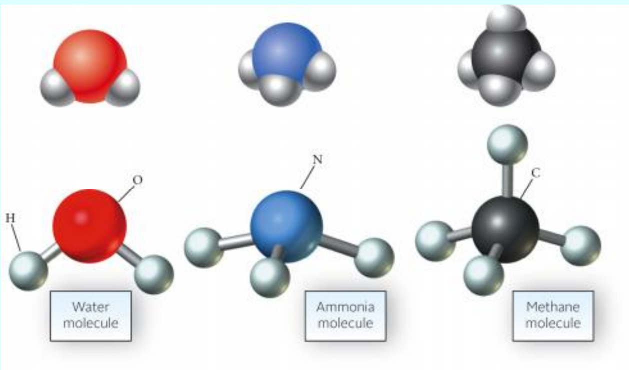


Structural Formulas

- Condensed structural formulas suggest the bonding pattern and highlight specific parts of a molecule, such as the reactive group of atoms



Ball and Stick Models



Example 2.4

EXAMPLE 2.4

Give the molecular formulas of ethyl alcohol, CH_3CH_2OH, and ethylamine, $CH_3CH_2NH_2$.

ANALYSIS

Information given:	structural formula
Asked for:	molecular formula

STRATEGY

Add up the atoms of each element and use the sums as the subscripts for the element.

SOLUTION

ethyl alcohol	C: 1 + 1 = 2; H: 3 + 2 + 1 = 6; O: 1 molecular formula: C_2H_6O
ethylamine	C: 1 + 1 = 2; H: 3 + 2 + 2 = 7; N: 1 molecular formula: C_2H_7N

END POINT

Note that although molecular formulas give the composition of the molecule, they reveal nothing about the way the atoms fit together. In that sense they are less useful than the structural formulas.

Molecular Elements

- Some elements exist as molecules, including such common ones as oxygen, hydrogen and nitrogen

Group 15	Group 16	Group 17
		H_2 (g)
N_2 (g)	O_2 (g)	F_2 (g)
P_4 (s)	S_8 (s)	Cl_2 (g)
		Br_2 (l)
		I_2 (s)

Ions

- When atoms or molecules lose or gain electrons, they form charged particles called ***ions***
  - $Na \rightarrow Na^+ + e^-$
  - $O + 2e^- \rightarrow O^{2-}$
- Positively charged ions are called ***cations***
- Negatively charged ions are called ***anions***
- There is ***no change in the number of protons in the nucleus when an ion forms.***

Example 2.5

EXAMPLE 2.5

Answer the questions below about the ions described.

(a) Aluminum is found in rubies and sapphires. How many protons, neutrons, and electrons are in this aluminum ion: ${}^{27}_{13}Al^{3+}$?

(b) Sulfur is present in an ore called chalcocite. The ion in the ore has 16 neutrons and 18 electrons. Write the nuclear symbol for the ion.

(c) An element found more abundantly in the sun and meteorites than on earth has an ion with a +2 charge. It has 38 electrons and 51 neutrons. Write its nuclear symbol.

(a)

ANALYSIS

Information given:	nuclear symbol and charge
Information implied:	A, Z
Asked for:	p^+, n, e^-

STRATEGY (FOR ALL PARTS)

1. Recall the placement of Z and A in the nuclear symbol.
2. $Z = p^+$; $A = p^+ + n$; $e^- = p^+ - \text{charge}$

SOLUTION

p^+, e^-	${}^{27}_{13}Al^{3+}$: $Z = p^+ = 13$; $e^- = p^+ - (\text{charge}) = 13 - (+3) = 10$;
n	$A = n + p^+$; $27 = n + 13$; $n = 14$

continued

Example 2.5 (Cont'd)

(b)

ANALYSIS

Information given:	e^-, n
Information implied:	Z
Asked for:	nuclear symbol for S

SOLUTION

Z	From the periodic table, sulfur has atomic number 16 so $Z = 16 = p^+$.
A and charge	$A = p^+ + n = 16 + 16 = 32$; charge $= p^+ - e^- = 16 - 18 = -2$
Nuclear symbol	${}^{32}_{16}S^{2-}$

(c)

ANALYSIS

Information given:	e^-, n, charge
Asked for:	nuclear symbol

SOLUTION

Z, element's identity	$Z = p^+ = e^- + \text{charge} = 38 + 2 = 40$; the element is Zr.
A	$A = p^+ + n = 40 + 51 = 91$ nuclear symbol: ${}^{A}_{Z}X = {}^{91}_{40}Zr^{2+}$

END POINT

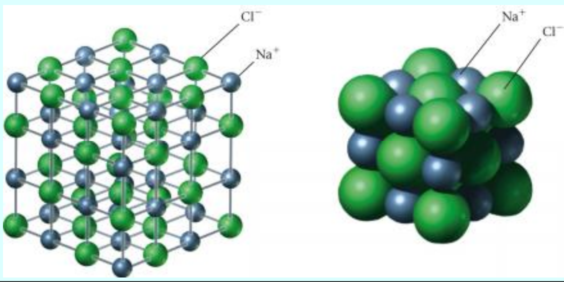
A cation always contains more protons than electrons; the reverse is true of an anion.

Polyatomic Ions

- Groups of atoms may carry a charge; these are the polyatomic ions
  - OH^-, hydroxide ion
  - NH_4^+, ammonium ion

Ionic Compounds

- Compounds can form between anions and cations
- Sodium chloride, NaCl
  - Sodium cations and chloride anions ions associate into a continuous network



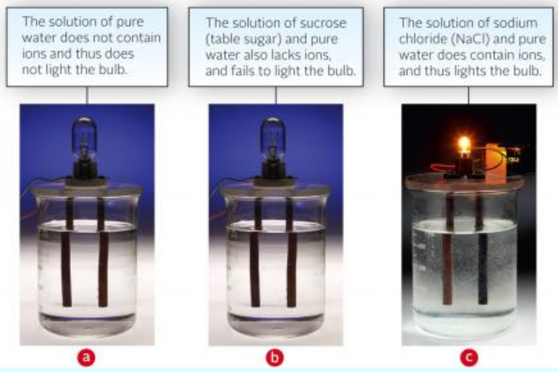
Forces Between Ions

- Ionic compounds are held together by strong forces called ionic bonds
  - Electrostatic attraction of + and – for each other
  - Compounds are usually solids at room temperature
    - High melting points
    - May be water-soluble

Solutions of Ionic Compounds

- When an ionic compound dissolves in water, the ions are released from each other
  - Presence of ions in the solution leads to electrical conductivity
    - Strong electrolytes
- When molecular compounds dissolve in water, no ions are formed
  - Without ions, solution does not conduct electricity
    - Nonelectrolytes

Figure 2.13 – Electrical Conductivity



Example 2.6

EXAMPLE 2.4 **CONCEPTUAL**

The structure of a water solution of KNO_3, containing equal numbers of K^+ and NO_3^- ions, may be represented as

(H_2O molecules are not shown.) Construct a similar beaker to show the structure of a water solution of potassium sulfate, K_2SO_4. Use to represent a sulfur atom.

continued

Example 2.6 (Cont'd)

STRATEGY

Once you deduce the formula of the ionic compound, it's all downhill. Remember, though, that you have to show the relative numbers of cations and anions.

SOLUTION

Formulas of Ionic Compounds

- Charge balance
  - Each positive charge must have a negative charge to balance it
  - Calcium chloride, $CaCl_2$
    - Ca^{2+}
    - Two Cl^- ions are required for charge balance

Noble Gas Connections

- Atoms that are close to a noble gas (group 18) form ions that contain the same number of electrons as the neighboring noble gas atom
- Applies to Groups 1, 2, 16 and 17, plus Al (Al^{3+}) and N (N^{3-})

Group	No. of Electrons in Atom	Charge of Ion Formed	Examples
1	1 more than noble-gas atom	+1	Na^+, K^+
2	2 more than noble-gas atom	+2	Mg^{2+}, Ca^{2+}
16	2 less than noble-gas atom	−2	O^{2-}, S^{2-}
17	1 less than noble-gas atom	−1	F^-, Cl^-

Cations of Transition and Post-Transition Metals

- Iron
  - Commonly forms Fe^{2+} and Fe^{3+}
- Lead
  - Commonly forms Pb^{2+} and Pb^{4+}

Polyatomic Ions

- There are only two common polyatomic cations
  - NH_4^+ and Hg_2^{2+}
  - All other common polyatomic ions are anions

Table 2.2 – Polyatomic ions

TABLE 2.2 Some Common Polyatomic Ions

+1	−1	−2	−3
NH_4^+ (ammonium)	OH^- (hydroxide)	CO_3^{2-} (carbonate)	PO_4^{3-} (phosphate)
Hg_2^{2+} (mercury I)	NO_3^- (nitrate)	SO_4^{2-} (sulfate)	
	ClO_3^- (chlorate)	CrO_4^{2-} (chromate)	
	ClO_4^- (perchlorate)	$Cr_2O_7^{2-}$ (dichromate)	
	CN^- (cyanide)	HPO_4^{2-} (hydrogen phosphate)	
	$C_2H_3O_2^-$ (acetate)		
	MnO_4^- (permanganate)		
	HCO_3^- (hydrogen carbonate)		
	$H_2PO_4^-$ (dihydrogen phosphate)		

Example 2.7

EXAMPLE 2.7

Predict the formula of the ionic compound

(a) formed by barium with iodine.

(b) containing a transition metal with a +1 charge in period 4 and Group 11 and oxide ions.

(c) containing an alkaline earth in period 5 and nitrogen.

(d) containing ammonium and phosphate ions.

STRATEGY

1. Recall charge of metals: group 1 (+1); group 2 (+2); Al (+3)
2. Recall charge of nonmetals: group 16 −2; group 17: −1; N: −3
3. The formula has to be electrically neutral.
4. Use Table 2.2 for polyatomic ions.

continued

Example 2.7 (Cont'd)

SOLUTION

(a) Charges	Ba is in group 2; thus its charge is +2. I is in group 17, thus its charge is −1.
electrical neutrality	$Ba^{2+}I^{1-}$; $2I^-$ are needed. The formula is BaI_2.
(b) Charges	Period 4, group 4 is Cu with a given charge of +1. O is in group 16, thus its charge is −2.
electrical neutrality	$Cu^{1+}O^{2-}$; 2 Cu^{1+} are needed. The formula is Cu_2O.
(c) Charges	The alkaline earth (group 2) in period 5 is Sr; thus its charge is +2. N in an ionic compound is always −3.
electrical neutrality	$Sr^{2+}N^{3-}$; 3 Sr^{2+} and 2 N^{3-} are needed. The formula is Sr_3N_2.
(d) Charges	Ammonium is a polyatomic ion: NH_4^+ (Table 2.2). Phosphate is a polyatomic ion: PO_4^{3-} (Table 2.2).
electrical neutrality	NH_4^+ PO_4^{3-}; 3 NH_4^+ are needed. The formula is $(NH_4)_3PO_4$.

END POINT

To be able to write the formulas of compounds, you must know the symbols of the elements. You must also know the symbols and charges of the polyatomic ions listed in Table 2.2. Learn them soon!

Names of Compounds - Cations

- Monatomic cations take the name from the metal from which they form
  - Na^+, sodium ion
  - K^+, potassium ion
- If more than one charge is possible, a Roman numeral is used to denote the charge
  - Fe^{2+} iron(II) ion
  - Fe^{3+} iron(III) ion

Names of Compounds - Anions

- Monatomic anions are named by adding *–ide* to the stem of the name of the element from which they form
  - Oxygen becomes oxide, O^{2-}
  - Sulfur becomes sulfide, S^{2-}
- Polyatomic ions are given special names (see table 2.3, p. 39)

Oxoanions

- When a nonmetal forms two oxoanions
  - *-ate* is used for the one with the larger number of oxygens
  - *-ite* is used for the one with the smaller number of oxygens
- When a nonmetal forms more than two oxoanions, prefixes are used
  - *per-* (largest number of oxygens)
  - *hypo-* (smallest number of oxygens)

Ionic Compounds

- Combine the name of the cation with name of the anion
  - $Cr(NO_3)_3$, chromium(III) nitrate
  - $SnCl_2$, tin(II) chloride

Table 2.3 – Oxoanions of Nitrogen, Sulfur and Chlorine

TABLE 2.3 Oxoanions of Nitrogen, Sulfur, and Chlorine

Nitrogen	Sulfur	Chlorine
		ClO_4^- *per*chlor*ate*
NO_3^- nitr*ate*	SO_4^{2-} sulf*ate*	ClO_3^- chlor*ate*
NO_2^- nitr*ite*	SO_3^{2-} sulf*ite*	ClO_2^- chlor*ite*
		ClO^- *hypo*chlor*ite*

Example 2.8

EXAMPLE 2.8

Name the following ionic compounds:

(a) CaS (b) $Al(NO_3)_3$ (c) $FeCl_2$

STRATEGY

Recall symbols for elements, symbols for polyatomic ions (Table 2.2), and suffixes for nonmetals.

SOLUTION

(a) CaS	Ca = calcium; S = sulfur → sulfide; calcium sulfide
(b) $Al(NO_3)_3$	Al^{3+} = aluminum; NO_3^- = nitrate; aluminum nitrate
(c) $FeCl_2$	Fe^{2+} = iron, which is a transition metal so (II) should be written after the name of the metal; Cl^- = chlorine → chloride; iron(II) chloride

Binary Molecular Compounds

- Unlike ionic compounds, there is no simple way to deduce the formula of a binary molecular compound
- Systematic naming
  1. The first word is the name of the first element in the formula, with a Greek prefix if necessary
  2. The second word consists of
     - The appropriate Greek prefix
     - The stem of the name of the second element
     - The suffix -ide

Some Examples

- Binary nonmetallic compounds
  - N_2O_5, dinitrogen pentaoxide
  - N_2O_4, dinitrogen tetraoxide
  - NO_2, nitrogen dioxide
  - N_2O_3, dinitrogen trioxide
  - NO, nitrogen oxide
  - N_2O, dinitrogen oxide

Common Molecular Compounds

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Table 2.4 - Greek Prefixes

TABLE 2.4 Greek Prefixes Used in Nomenclature

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

*The prefix mono (1) is seldom used.

Example 2.9

EXAMPLE 2.9

Give the names of the following molecules:

(a) SF_4 (b) PCl_3 (c) N_2O_3 (d) Cl_2O_7

STRATEGY

1st element: subscript → prefix (Table 2.4) + element name

2nd element: subscript → prefix + element name ending in *ide*

continued

Example 2.9 (Cont'd)

SOLUTION

(a) SF_4	S: subscript = 1; no prefix = sulfur F: subscript = 4 = tetra; F = fluorine ⟶ fluoride; sulfur tetrafluoride
(b) PCl_3	P: subscript = 1; no prefix = phosphorus Cl: subscript = 3 = tri; Cl = chlorine ⟶ chloride; phosphorus trichloride
(c) N_2O_3	N: subscript = 2 = di; dinitrogen O: subscript = 3 = tri; O = oxygen ⟶ oxide; dinitrogen trioxide
(d) Cl_2O_7	Cl: subscript = 2 = di; dichlorine O: subscript = 7 = hepta; O = oxygen ⟶ oxide; dichlorine heptaoxide

Acids

- Acids ionize to form H^+ ions
- Hydrogen and chlorine
  - As a molecule, HCl is hydrogen chloride
  - When put in water, HCl is hydrochloric acid

Common Acids

Pure Substance		Water Solution	
HCl(*g*)	Hydrogen chloride	$H^+(aq)$, $Cl^-(aq)$	Hydrochloric acid
HBr(*g*)	Hydrogen bromide	$H^+(aq)$, $Br^-(aq)$	Hydrobromic acid
HI(*g*)	Hydrogen iodide	$H^+(aq)$, $I^-(aq)$	Hydriodic acid

Oxoacids

- Two common oxoacids
  - HNO_3, nitric acid
  - H_2SO_4, sulfuric acid

Oxoacids of Chlorine

HNO_3 nitric acid H_2SO_4 sulfuric acid

The names of oxoacids are simply related to those of the corresponding oxoanions. The -*ate* suffix of the anion is replaced by -*ic* in the acid. In a similar way, the suffix -*ite* is replaced by the suffix -*ous*. The prefixes *per*- and *hypo*- found in the name of the anion are retained in the name of the acid.

ClO_4^-	*per*chlor*ate* ion	$HClO_4$	*per*chlor*ic* acid
ClO_3^-	chlor*ate* ion	$HClO_3$	chlor*ic* acid
ClO_2^-	chlor*ite* ion	$HClO_2$	chlor*ous* acid
ClO^-	*hypo*chlor*ite* ion	HClO	*hypo*chlor*ous* acid

Example 2.10

EXAMPLE 2.10

Give the names of

(a) HCl(*g*) (b) $HNO_2(aq)$ (c) $H_2SO_4(aq)$ (d) HIO(*aq*)

(a)

STRATEGY

Gases follow the rules for naming binary molecules.

SOLUTION

HCl(*g*)	No prefixes for both elements (subscripts are both 1). H = hydrogen; Cl = chlorine → chloride; hydrogen chloride

(b)

STRATEGY

Name of oxoanion (Table 2.3) (change *ite* to *ous*) + the word "acid."

SOLUTION

$HNO_2(aq)$	NO_2^- = nitrite → nitrous + acid; nitrous acid

Example 2.10, (Cont'd)

(c)

STRATEGY

Name of oxoanion (Table 2.3) (change *ate* to *ic*) + the word "acid."

SOLUTION

$H_2SO_4(aq)$	SO_4^{2-} = sulfate → sulfuric + acid; sulfuric acid

(d)

STRATEGY

Name of oxoanion (Table 2.3) (change *ite* to *ous*) + the word "acid." The oxoanion's name is analogous to the naming of the chlorine oxoanions.

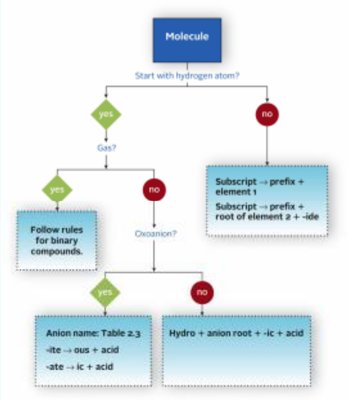
SOLUTION

HIO(*aq*)	IO^- = hypoiodite → hypoiodous + acid; hypoiodous acid Hypoiodous is analogous to hypochlorous, the name for ClO^-.

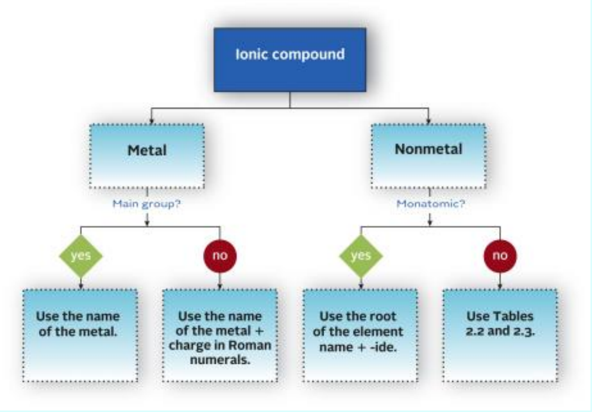
END POINT

You need to learn the names of the oxoanions listed in Table 2.3.

Flow Chart for Naming Molecular Compounds



Flow Chart for Naming Ionic Compounds



Key Concepts

1. Relate a nuclear symbol to the numbers of protons and neutrons in the nucleus,
2. Relate atomic mass, isotopic abundance and average mass of an element
3. Relate atomic mass to Avogadro's Number.
4. Relate elements and the periodic table.
5. Relate structural, condensed and molecular formulas.
6. Relate the ionic charge to the number of electrons.
7. Predict the formulas of ionic compounds from charge of ions
8. Relate names to formulas (ionic, molecular and oxoacids and oxoanions).