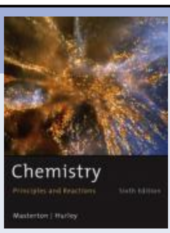

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Chapter 2
Atoms, Molecules, and Ions

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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
- 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

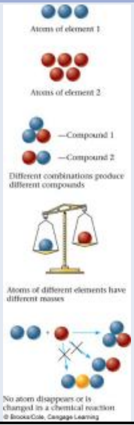
The Structure of Matter

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

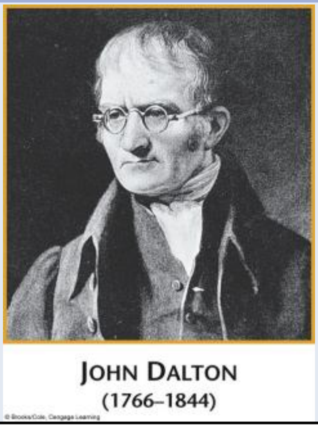
Atoms and Atomic Theory

- 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 element 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

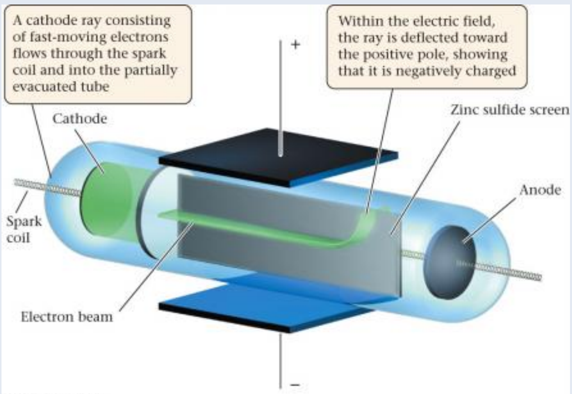
- J.J. Thomson, Cavendish Laboratories, Cambridge, England
- Ernest Rutherford
  - McGill University, Canada
  - Manchester and Cambridge Universities, England

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



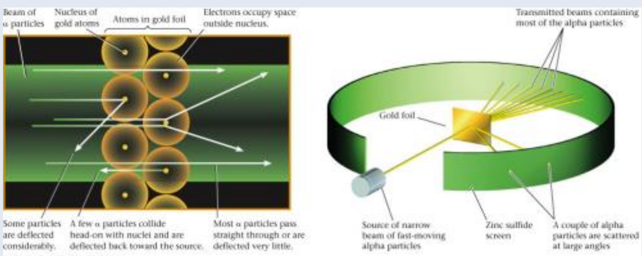
The Electron and the Atom

- 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 minus their 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

Figure 2.4 – 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 (Chapter 3).

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-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

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

Note that some of the ice is at the bottom of the glass – this is ^2H_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.

Strategy Remember the definitions of atomic number and mass number and where they appear in the nuclear symbol.

SOLUTION

(a) The mass number is $27 + 33 = 60$. The nuclear symbol is $^{60}_{27}Co$.
(b) The number of protons is given by the atomic number (left subscript) and is 38. The mass number (left superscript) is 90. The number of neutrons is $90 - 38 = 52$.
(c) The number of protons (31) identifies the element as gallium. Its mass number is 69 (31 protons + 38 neutrons). Its nuclear symbol is $^{69}_{31}Ga$.

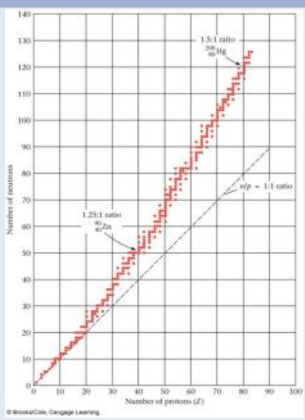
Radioactivity

- Radioactive isotopes are unstable
  - These isotopes decay over time
  - Emit other particles and are transformed into other elements
  - Radioactive decay is ***not*** a chemical process!
- Particles emitted
  - High speed electrons: β (beta) particles
  - Alpha (α) particles: helium nuclei
  - Gamma (γ) rays: high energy light

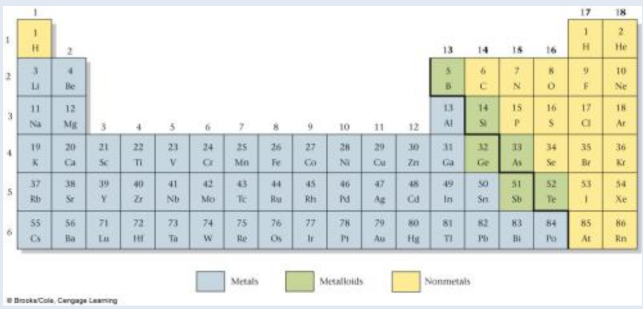
Nuclear Stability

- Nuclear stability depends on the neutron/proton ratio
  - For light elements, n/p is approximately 1
  - For heavier elements, n/p is approximately 1.4/1
- The belt of stability

Figure 2.5 – The Nuclear Belt of Stability



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

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

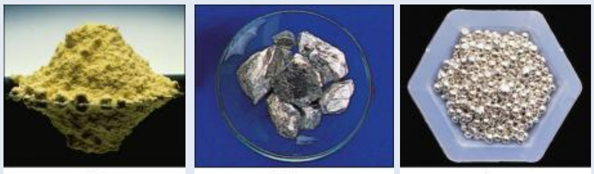
- Dmitri Mendeleev, 1836-1907
- Arranged elements by chemical properties
  - Left space for elements unknown at the time
  - Predicted detailed properties for elements as yet unknown
    - Sc, Ga, Ge
    - By 1886, all these elements had been discovered, and with properties similar to those he predicted

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)

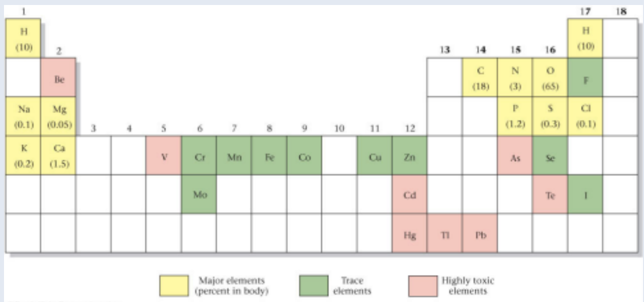


Sulfur (a) Antimony (b) Silver (c)

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 2.8 – Biologically Important and Toxic Elements

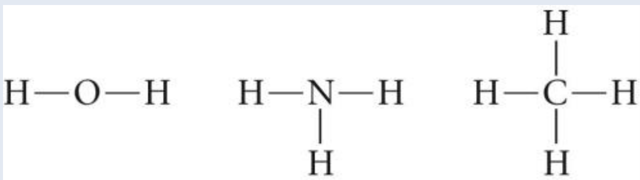


Molecule

- 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

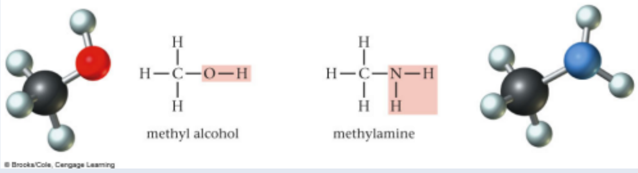
Structural Formulas

- Structural formulas show the bonding patterns within the molecule

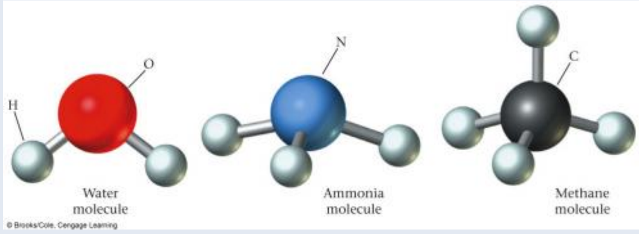


Structural Formulas

- 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 2.2

Example 2.2 Give the molecular formulas of (a) ethyl alcohol and (b) ethylamine.

Strategy To find the molecular formulas, simply add up the atoms of each type and use the sums as subscripts in the formulas.

SOLUTION

(a) C_2H_6O (b) C_2H_7N

Reality Check 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 structural formulas.

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 2.3

Example 2.3 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.

Strategy The atomic number gives directly the number of protons. The number of neutrons is found in the usual way by subtracting the atomic number from the mass number. The number of electrons can be deduced from the number of protons by taking into account the charge of the ion.

SOLUTION

(a) Number of protons = 13; number of neutrons = 27 − 13 = 14; number of electrons = 13 − 3 = 10.
(b) For sulfur, $Z = 16$, so there must be 16 protons. The mass number is 16 + 16 = 32. The charge on the ion is 16 − 18 (number of protons − number of electrons) = −2. The nuclear symbol is then $^{32}_{16}S^{2-}$.
(c) Since the element has a +2 charge and 38 electrons, it must have 40 protons. The element with that atomic number is zirconium. Its mass number is 40 + 51 = 91. Its nuclear symbol is $^{91}_{40}Zr^{2+}$.

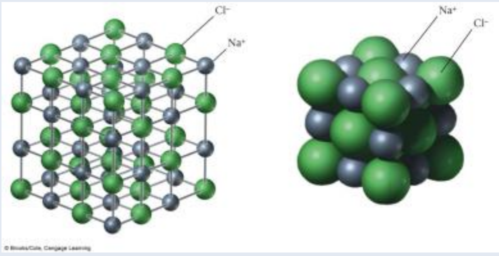
Reality Check A cation always contains more protons than electrons; the reverse is true for an anion.

Polyatomic Ions

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

Ionic Compounds

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



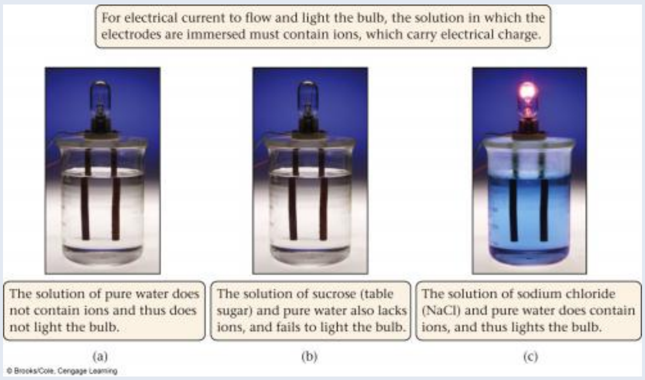
Forces Between Ions

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

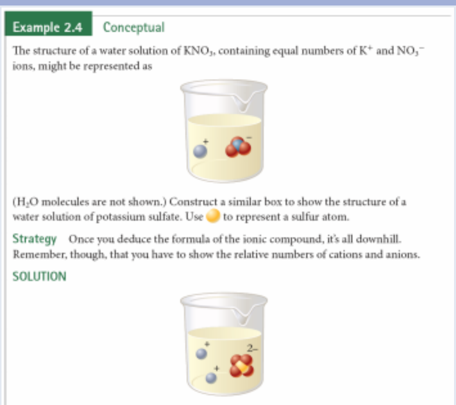
Solutions of Ionic Compounds

- 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 2.12 – Electrical Conductivity



Example 2.4



Example 2.4 Conceptual

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

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

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 2.5

Example 2.5 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 First identify the elements. Then (using Table 2.2 as necessary) identify the charges of the cation and anion. Finally, balance positive and negative charges.

SOLUTION

(a) Barium is in Group 2, so it forms a +2 ion. Iodine in Group 17 forms the I^- ion. The formula is then BaI_2.
(b) The cation is Cu^+, the anion O^{2-}. Two cations are required to balance one anion. Hence, the formula is Cu_2O.
(c) The element is Sr (strontium), which is in Group 2 and so forms the Sr^{2+} ion. Nitrogen as an ion has a −3 charge. The formula thus is Sr_3N_2.
(d) From Table 2.2, the ions are NH_4^+ and PO_4^{3-}. The formula is therefore $(NH_4)_3PO_4$.

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*chlorate
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 2.6

Example 2.6 Name the following ionic compounds:

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

Strategy To name an ionic compound, you must know the rules for naming individual ions, as discussed previously.

SOLUTION

(a) calcium sulfide (b) aluminum nitrate (c) 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 names
  - H_2O, water
  - H_2O_2, hydrogen peroxide

Common Molecular Compounds

To illustrate these rules, consider the names of the several oxides of nitrogen:

N_2O_5	dinitrogen *penta*oxide	N_2O_3	dinitrogen *tri*oxide
N_2O_4	dinitrogen *tetra*oxide	NO	nitrogen ox*ide*
NO_2	nitrogen *di*oxide	N_2O	*di*nitrogen ox*ide*

When the prefixes tetra, penta, hexa, . . . are followed by the letter "o," the *a* is often dropped. For example, N_2O_5 is often referred to as *dinitrogen pentoxide.*

Many of the best-known binary compounds of the nonmetals have acquired common names. These are widely—and in some cases exclusively—used. Examples include

H_2O	water	PH_3	phosphine
H_2O_2	hydrogen peroxide	AsH_3	arsine
NH_3	ammonia	NO	nitric oxide
N_2H_4	hydrazine	N_2O	nitrous oxide
C_2H_2	acetylene	CH_4	methane

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 2.7

Example 2.7 Give the names of the following molecules:

(a) SO_2 (b) SO_3 (c) PCl_3 (d) Cl_2O_7

Strategy Start with the prefix denoting the number of atoms (if there is more than one) of the first element followed by the name of that element. Repeat for the second element, ending with the suffix -*ide*.

SOLUTION

(a) sulfur dioxide (b) sulfur trioxide
(c) phosphorus trichloride (d) 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

Most acids contain oxygen in addition to hydrogen atoms. Such species are referred to as *oxoacids.* Two oxoacids that you are likely to encounter in the general chemistry laboratory are

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 2.8

Example 2.8 Give the names of

(a) $HCl(g)$ (b) $HNO_2(aq)$ (c) $H_2SO_3(aq)$ (d) $HIO(aq)$

Strategy In (a), note that the specie is a gas, not an acid. In (b) and (c), refer back to Table 2.3 for the name of the oxoanion. In (d), reason by analogy with chlorine oxoacids.

SOLUTION

(a) hydrogen chloride gas (b) nitrous acid
(c) sulfurous acid (d) hypoiodous acid

Key Concepts

1. Relate a nuclear symbol to the numbers of protons and neutrons in the nucleus,
2. Relate elements and the periodic table.
3. Relate structural, condensed, and molecular formulas.
4. Relate the ionic charge to the number of electrons.
5. Write the formula of ionic compounds from the charges on the ions.

Key Concepts

6. Relate names to formulas for
   - Ionic compounds
   - Binary molecular compounds
   - Oxoanions and oxoacids