Atoms, Molecules and Ions

The theory that atoms are the fundamental building blocks of matter re-emerged in the early 19th century, championed by John Dalton.

Each element is composed of extremely small particles called atoms

**Dalton’s Postulates:**

All atoms of a given element are identical to one another in mass and other properties, but the atoms of one element are different from the atoms of all other elements.

**Law of Constant Composition**

- Also known as the law of definite proportions.
- The elemental composition of a pure substance never varies.
- The total mass of substances present at the end of a chemical process is the same as the mass of substances present before the process took place.

- Protons were discovered by Rutherford in 1919.
- Neutrons were discovered by James Chadwick in 1932.

- Protons and electrons are the only particles that have a charge.
- Protons and neutrons have essentially the same mass.
- The mass of an electron is so small we ignore it.

**The Electron**

- Streams of negatively charged particles were found to emanate from cathode tubes.
- J. J. Thompson is credited with their discovery (1897).
- Thompson measured the charge/mass ratio of the electron to be $1.76 \times 10^8$ coulombs/g.
Millikan Oil Drop Experiment
Once the charge/mass ratio of the electron was known, determination of either the charge or the mass of an electron would yield the other.

Fig. 1  Millikan Oil Drop Experiment

Robert Millikan (University of Chicago) determined the charge on the electron in 1909.

RADIOACTIVITY:
The spontaneous emission of radiation by an atom.
First observed by Henri Becquerel.
Also studied by Marie and Pierre Curie
• Three types of radiation were discovered by Ernest Rutherford:
  ➢ $\alpha$ particles
  ➢ $\beta$ particles
  ➢ $\gamma$ rays

Fig. 2 Emission of radiation by an atom
Structure of the Atom around 1900

- “Plum pudding” model, put forward by Thompson.
- Positive sphere of matter with negative electrons imbedded in it.

**Discovery of the Nucleus**

Ernest Rutherford shot $\alpha$ particles at a thin sheet of gold foil and observed the pattern of scatter of the particles.

**Fig. 4 Rutherford’s Gold Foil Experiment**
THE NUCLEAR ATOM

Since some particles were deflected at large angles, Thompson’s model could not be correct

- Rutherford postulated a very small, dense nucleus with the electrons around the outside of the atom.
- Most of the volume of the atom is empty space.

![Diagram of an atom](image)

Fig. 5 Structure of an atom postulated by Rutherford

Other Subatomic Particles

- Protons were discovered by Rutherford in 1919.
- Neutrons were discovered by James Chadwick in 1932.

Subatomic Particles

- Protons and electrons are the only particles that have a charge.
- Protons and neutrons have essentially the same mass.
- The mass of an electron is so small we ignore it.

<table>
<thead>
<tr>
<th>Particle</th>
<th>Charge</th>
<th>Mass (amu)</th>
</tr>
</thead>
<tbody>
<tr>
<td>Proton</td>
<td>Positive (1+)</td>
<td>1.0073</td>
</tr>
<tr>
<td>Neutron</td>
<td>None (neutral)</td>
<td>1.0087</td>
</tr>
<tr>
<td>Electron</td>
<td>Negative (1−)</td>
<td>$5.486 \times 10^{-4}$</td>
</tr>
</tbody>
</table>
SYMBOLS OF ELEMENTS
Elements are symbolized by one or two letters.

Mass number (number of protons plus neutrons)

$^{12}_6\text{C}$ ← Symbol of element

ATOMIC NUMBER
All atoms of the same element have the same number of protons: The atomic number (Z)

Mass number (number of protons plus neutrons)

$^{12}_6\text{C}$ ← Symbol of element

ATOMIC MASS

The mass of an atom in atomic mass units (amu) is the total number of protons and neutrons in the atom.

Mass number (number of protons plus neutrons)

$^{12}_6\text{C}$ ← Symbol of element
The atom consists of: a positively charged nucleus and negatively charged electrons.

Hydrogen is the simplest atom that has 1 proton and 1 electron. The nucleus of the hydrogen atom is its nucleus, but the nucleus of another atom consists of protons and neutrons.

All atoms are neutral e.g. H, Li, Be, C, N, O, F, Ne i.e. the positive charge of the nucleus is balanced by the negative charge of a number of electrons. In a neutral atom: no. of protons = no. of electrons

<table>
<thead>
<tr>
<th>Subatomic particle</th>
<th>Mass (kg)</th>
<th>Mass (amu)</th>
<th>Charge</th>
</tr>
</thead>
<tbody>
<tr>
<td>proton</td>
<td>$1.67262 \times 10^{-27}$</td>
<td>1.0073</td>
<td>+1</td>
</tr>
<tr>
<td>neutron</td>
<td>$1.67493 \times 10^{-27}$</td>
<td>1.0087</td>
<td>0</td>
</tr>
<tr>
<td>electron</td>
<td>$0.00091 \times 10^{-27}$</td>
<td>0.00055</td>
<td>-1</td>
</tr>
</tbody>
</table>

**ISOTOPES:**

Atoms of the same element with different masses. Isotopes have different numbers of neutrons.

e.g.
1. 4 isotopes of carbon are: $\overset{11}{6}C$, $\overset{12}{6}C$, $\overset{13}{6}C$, $\overset{14}{6}C$
2. 3 isotopes of H are: $\overset{1}{1}H$, $\overset{2}{1}H$, $\overset{3}{1}H$
   hydrogen, deuterium and tritium respectively
3. 2 isotopes of Uranium are: $\overset{235}{92}U$, $\overset{238}{92}U$

Note: Mass number (A) = Atomic no. (Z) + no of neutrons
SAMPLE EXERCISE

Determining the Number of Subatomic Particles in Atoms

1. How many protons, neutrons, and electrons are in:
   (a) an atom of $^{197}\text{Au}$ (b) an atom of strontium-90?

Solution (a)
The superscript 197 is the mass number, the sum of the number of protons plus the number of neutrons. $197 - 79 = \mathbf{118 \text{ neutrons}}$.

(b) The atomic number of strontium is 38. Thus, all atoms of this element have 38 protons and 38 electrons. The strontium-90 isotope has $90 - 38 = \mathbf{52 \text{ neutrons}}$.

2. How many protons, neutrons, and electrons are in:
   (a) A $^{138}\text{Ba}$ atom, (b) an atom of phosphorus-31?

   \textbf{Answer:} \quad (a) 56 protons, 56 electrons, and 82 neutrons;
   
   (b) 15 protons, 15 electrons, and 16 neutrons.

3. Magnesium has three isotopes, with mass numbers 24, 25, and 26. (a) Write the complete chemical symbol (superscript and subscript) for each of them. (b) How many neutrons are in an atom of each isotope?

   \textbf{Solution:}

   (a) Magnesium has atomic number 12, and so all atoms of magnesium contain 12 protons and 12 electrons. The three isotopes are therefore represented by $__________$

   (b) The number of neutrons in each isotope is the mass number minus the number of protons. The numbers of neutrons in an atom of each isotope are therefore $____, ____,$ and $____'$ respectively.

4. Give the complete chemical symbol for the atom that contains 82 protons, 82 electrons, and 126 neutrons.

   \textbf{Answer:} $^{208}_{82}\text{Pb}$
THE ATOMIC MASS SCALE

Each 100 g of H₂O contains 11.1 g of hydrogen and 88.9 g of oxygen. Thus water contains 88.9/11.1 = 8 times as much oxygen, by mass, as hydrogen.

Water contains 2 hydrogen atoms and 1 oxygen atom, therefore scientists concluded that an oxygen atom must have 2 x 8 = 16 times as much mass as a hydrogen atom.

Hydrogen is the lightest atom, was arbitrarily assigned a relative mass of 1 (no units). And atomic masses were determined relative to hydrogen. Thus oxygen was assigned a mass of 16 units.

The masses of individual atoms can be determined with a high degree of accuracy. For example, the ¹H atom has a mass of 1.6735 x 10⁻²⁴ g and the ¹⁶O atom has a mass of 2.6560 x 10⁻²³ g.

(unified atomic mass unit) 1 u = 1.66054 x 10⁻²⁴ g and 1 g = 6.02214 x 10²³ u

The unified atomic mass unit is defined by assigning a mass of exactly 12 u to an atom of the ¹²C isotope (6 protons + 6 neutrons)

In these units, an ¹H atom = 1.0078 u and an ¹⁶O atom = 15.9949 u

Atomic and molecular masses can be measured with great accuracy with a mass spectrometer.
AVERAGE ATOMIC MASSES

- Because in the real world we use large amounts of atoms and molecules, we use average masses in calculations.
- Average mass is calculated from the isotopes of an element weighted by their relative abundances.

Most elements occur in nature as mixed isotopes.

The average atomic mass of an element can be determined by making use of its various isotopic masses and their relative percentage abundances.

For example: Naturally occurring carbon is composed of 98.93 % $^{12}$C + 1.07 % $^{13}$C. The masses of these isotopes are 12 u (exactly) and 13.00335 u, respectively.

The average atomic mass of carbon calculated from their fractional abundances:

$$\frac{98.93}{100} \times 12\ u + \frac{1.07}{100} \times 13.00335\ u = 12.01\ u$$
Practice Exercise 1.

**ISOTOPES – FRACTIONAL ABUNDANCE**

**Problem**

The element boron consists of two isotopes, $^{10}\text{B}$ and $^{11}\text{B}$. Their masses, based on the carbon scale, are 10.01 and 11.01, respectively.

The abundance of $^{10}\text{B}$ is 20.0% and the abundance of $^{11}\text{B}$ is 80.0%. What is the atomic mass of boron?

**Solution:**

The percentages of multiple isotopes must add up to 100%. Apply the following equation to the problem:

\[
\text{atomic mass} = [(\text{atomic mass } X_1) \times (\% \text{ of } X_1)/100] + [(\text{atomic mass } X_2) \times (\% \text{ of } X_2)/100] + ...
\]

where \(X\) is an isotope of the element and \(\% \text{ of } X\) is the abundance of the isotope \(X\).

Substitute the values for boron in this equation:

\[
\text{atomic mass of } B = (\text{atomic mass of } ^{10}\text{B} \times % \text{ of } ^{10}\text{B}/100) + \\
(\text{atomic mass of } ^{11}\text{B} \times % \text{ of } ^{11}\text{B}/100)
\]

\[
\text{atomic mass of } B = (10.01 \times 20.0/100) + (11.01 \times 80.0/100)
\]

atomic mass of \(B\) = 2.00 + 8.81

atomic mass of \(B\) = 10.81

**Answer:** The atomic mass of boron is 10.81.
Practice exercise 2.

Calculating % Natural abundance of Isotopes

Problem

Lithium has 2 natural isotopes, $^6$Li and $^7$Li. They have atomic masses of 6.0151 amu and 7.0160 amu respectively. Calculate the percentage natural abundance of these 2 isotopes if the average atomic mass of Li is 6.941 amu.

Solution

$$6.941 = 6.0151 \times (x) + 7.0160 \times (1-x)$$

So we get

$$6.0151x - 7.0160x = 6.941 - 7.0160$$

$$1.0009x = 0.075$$

$$x = 0.0749$$

So the 6.0151 amu has a 7.49 % abundance and

the 7.0160 amu has a 92.51 % abundance
Practice Exercise 3.

Naturally occurring chlorine is 75.78% $^{35}\text{Cl}$, which has an atomic mass of 34.969 amu, and 24.22% $^{37}\text{Cl}$, which has an atomic mass of 36.966 amu. Calculate the average atomic mass (that is, the atomic weight) of chlorine.

**Solution** The average atomic mass is found by multiplying the abundance of each isotope by its atomic mass and summing these products. Because 75.78% = 0.7578 and 24.22% = 0.2422, we have

\[
\text{Average atomic mass} = (0.7578)(34.969 \text{ amu}) + (0.2422)(36.966 \text{ amu})
\]
\[
= 26.50 \text{ amu} + 8.953 \text{ amu}
\]
\[
= 35.45 \text{ amu}
\]

This answer makes sense: The average atomic mass of Cl is between the masses of the two isotopes and is closer to the value of $^{35}\text{Cl}$, which is the more abundant isotope.

Practice Exercise 4.

Three isotopes of silicon occur in nature: $^{28}\text{Si}$ (92.23%), which has an atomic mass of 27.97693 amu; $^{29}\text{Si}$ (4.68%), which has an atomic mass of 28.97649 amu; and $^{30}\text{Si}$ (3.09%), which has an atomic mass of 29.97377 amu. Calculate the atomic weight of silicon.

*Answer:* 28.09 amu
The Periodic Table

Is a chart in which elements having similar chemical and physical properties are grouped together. A systematic catalog of elements.

Elements are arranged by atomic no. in horizontal rows called **periods** and in vertical columns called **groups**

Elements are divided into 3 groups:
(a) **metals** – good conductors of heat and electricity.
(b) **metalloids** – (semi conductors) has properties that are intermediate between those of metals and non-metals.
(c) **non-metals** - poor conductors of heat and electricity

---

<table>
<thead>
<tr>
<th>Period</th>
<th>1A</th>
<th>2A</th>
<th>3A</th>
<th>4A</th>
<th>5A</th>
<th>6A</th>
<th>7A</th>
</tr>
</thead>
<tbody>
<tr>
<td>1</td>
<td>1 H</td>
<td>2</td>
<td>3 Li</td>
<td>4 Be</td>
<td>5</td>
<td>6</td>
<td>7</td>
</tr>
<tr>
<td>2</td>
<td>8B</td>
<td>9</td>
<td>10</td>
<td>11</td>
<td>12</td>
<td>13 Al</td>
<td>14 Si</td>
</tr>
<tr>
<td>3</td>
<td>19 K</td>
<td>20 Ca</td>
<td>21 Sc</td>
<td>22 Ti</td>
<td>23 V</td>
<td>24 Cr</td>
<td>25 Mn</td>
</tr>
<tr>
<td>4</td>
<td>37 Rb</td>
<td>38 Sr</td>
<td>39 Y</td>
<td>40 Zr</td>
<td>41 Nb</td>
<td>42 Mo</td>
<td>43 Tc</td>
</tr>
<tr>
<td>5</td>
<td>55 Cs</td>
<td>56 Ba</td>
<td>57 La</td>
<td>58 Ce</td>
<td>59 Pr</td>
<td>60 Nd</td>
<td>61 Pm</td>
</tr>
<tr>
<td>6</td>
<td>89 Ac</td>
<td>90 Th</td>
<td>91 Pa</td>
<td>92 U</td>
<td>93 Np</td>
<td>94 Pu</td>
<td>95 Am</td>
</tr>
</tbody>
</table>

**Fig. 7** A periodic table of elements showing the division of elements into metals, metalloids and non-metals
Periodicity

When one looks at the chemical properties of elements, one notices a repeating pattern of reactivities.

Groups

<table>
<thead>
<tr>
<th>Group</th>
<th>Name</th>
<th>Elements</th>
</tr>
</thead>
<tbody>
<tr>
<td>1A</td>
<td>Alkali metals</td>
<td>Li, Na, K, Rb, Cs, Fr</td>
</tr>
<tr>
<td>2A</td>
<td>Alkaline earth metals</td>
<td>Be, Mg, Ca, Sr, Ba, Ra</td>
</tr>
<tr>
<td>6A</td>
<td>Chalcogens</td>
<td>O, S, Se, Te, Po</td>
</tr>
<tr>
<td>7A</td>
<td>Halogens</td>
<td>F, Cl, Br, I, At</td>
</tr>
<tr>
<td>8A</td>
<td>Noble gases (or rare gases)</td>
<td>He, Ne, Ar, Kr, Xe, Rn</td>
</tr>
<tr>
<td>5A</td>
<td>pnicogens</td>
<td>N, P, As, Sb, Bi</td>
</tr>
</tbody>
</table>

- Nonmetals are on the right side of the periodic table (with the exception of H).
- Metalloids border the stair-step line (with the exception of Al and Po).
- Metals are on the left side of the chart.

Practice Exercise:

1. Which two of the following elements would you expect to show the greatest similarity in chemical and physical properties: B, Ca, F, He, Mg, P?
Solution Elements that are in the same group of the periodic table are most likely to exhibit similar chemical and physical properties. We therefore expect that Ca and Mg should be most alike because they are in the same group (2A, the alkaline earth metals).

2. Locate Na (sodium) and Br (bromine) on the periodic table. Give the atomic number of each, and label each a metal, metalloid, or nonmetal.

Answer: Na, atomic number 11, is a metal; Br, atomic number 35, is a nonmetal.

Molecules and Ions

Molecules is an aggregate of at least two atoms in a definite arrangement held together by chemical forces (also called chemical bonds)

Diatomice molecules e.g. H₂, N₂, O₂, Br₂, Cl₂
Polyatomic molecules (more than two atoms) e.g. O₃, NH₃, H₂O

Fig. 8 Examples of some simple molecules
These seven elements occur naturally as molecules containing two atoms:

\[ \begin{array}{c}
1 & 2 & 3 & 4 & 5 & 6 & 7 \\
H & O & N & F & Cl & Br & I
\end{array} \]

Fig. 9 Naturally occurring diatomic molecules

Counting Atoms and Molecules

**Molecular formula:** shows the exact number of atoms of each element in the smallest unit of a substance.

- \( \text{H}_2 \) is the molecular formula for hydrogen gas.
- \( \text{O}_2 \) is the molecular formula for oxygen gas.
- \( \text{H}_2\text{O} \) is the molecular formula for water.

**Empirical formula:** tells us which elements are present and the simplest whole number ratio of their atoms.

- e.g. \( \text{HO} \) is the empirical formula of \( \text{H}_2\text{O}_2 \)
- \( \text{NH}_2 \) is the empirical formula of \( \text{N}_2\text{H}_4 \)

The molecular and empirical formulas in \( \text{H}_2\text{O}, \ \text{NH}_3, \ \text{CO}_2, \ \text{CH}_4 \) are one and the same thing.

**Practice exercise**

1. Write the empirical formulas for the following molecules:
   (a) glucose, molecular formula is \( \text{C}_6\text{H}_{12}\text{O}_6 \);
   (b) nitrous oxide, a substance used as an anesthetic and commonly called laughing gas, whose molecular formula is \( \text{N}_2\text{O} \).

**Solution**

(a) The resultant empirical formula for glucose is \( \text{CH}_2\text{O} \).
(b) The empirical formula for nitrous oxide is the same as its molecular formula, \( \text{N}_2\text{O} \).
TYPES OF FORMULAS

- Structural formulas show the order in which atoms are bonded.
- Perspective drawings also show the three-dimensional array of atoms in a compound.

IONS

Ionization or a group of atoms that have a net positive or negative charge.

Cation formation

- e^-

\[
\text{e.g. Na atom} \\
11 \text{ protons} \\
11 \text{ electrons} \\
1s^22s^22p^63s^1 \\
\longrightarrow \\
\text{Na}^+ \text{ ion (cation)} \\
11 \text{ protons} \\
10 \text{ electrons} \\
1s^22s^22p^6 = [\text{Ne}] \text{ noble gas configuration}
\]
Anion formation

\[ \text{e.g. Cl atom} \quad + \quad e^- \quad \rightarrow \quad \text{Cl}^{-} \text{ion (anion)} \]

17 protons 17 protons
17 electrons 18 electrons
1s\(^2\)2s\(^2\)2p\(^6\)3s\(^2\)3p\(^5\) 1s\(^2\)2s\(^2\)2p\(^6\)3s\(^2\)3p\(^6\) = [Ar] noble gas configuration

When atoms lose or gain electrons, they become ions.
- Cations are positive and are formed by elements on the left side of the periodic chart.
- Anions are negative and are formed by elements on the right side of the periodic chart.

**Ionic Bonds:** Ionic compounds (such as NaCl) are generally formed between metals and nonmetals.
Writing Formulas

Because compounds are electrically neutral, one can determine the formula of a compound this way:

- The charge on the cation becomes the subscript on the anion.
- The charge on the anion becomes the subscript on the cation.
- If these subscripts are not in the lowest whole-number ratio, divide them by the greatest common factor.

Practice exercise

1. Give the chemical symbol, including mass number, for each of the following ions:
   (a) The ion with 22 protons, 26 neutrons, and 19 electrons;
   (b) The ion of sulfur that has 16 neutrons and 18 electrons.

Solution

(a) The number of protons (22) is the atomic number of the element, therefore element is titanium (Ti). The mass number of this isotope is $22 + 26 = 48$ (the sum of the protons and neutrons). Because the ion has three more protons than electrons, it has a net charge of $3^+$. Thus, the symbol for the ion is $^{48}\text{Ti}^{3+}$.

(b) By referring to a periodic table sulfur (S) has an atomic number of 16. Thus, each atom or ion of sulfur must contain 16 protons. We are told that the ion also has 16 neutrons, meaning the mass number of the ion is $16 + 16 = 32$. Because the ion has 16 protons and 18 electrons, its net charge is $2^-$. Thus, the symbol for the ion is $^{32}\text{S}^{2-}$. 
2. How many protons and electrons does the Se$^{2-}$ ion possess?
   \textit{Answer:} 34 protons and 36 electrons

3. Predict the charge expected for the most stable ion of barium and for the most stable ion of oxygen.

   \textbf{Solution}
   Barium has atomic number 56. The nearest noble gas is xenon, atomic number 54. Barium can attain a stable arrangement of 54 electrons by losing two of its electrons, forming the \textbf{Ba}$^{2+}$ \textit{cation}.
   Oxygen has atomic number 8. The nearest noble gas is neon, atomic number 10. Oxygen can attain this stable electron arrangement by gaining two electrons, thereby forming the \textbf{O}$^{2-}$ \textit{anion}.
   Therefore forming \textit{BaO}

4. Predict the charge expected for the most stable ion of aluminum and for the most stable ion of fluorine.

   \textit{Answer:} Al = 3+ and F = 1–
# Common Cations

<table>
<thead>
<tr>
<th>Charge</th>
<th>Formula</th>
<th>Name</th>
<th>Formula</th>
<th>Name</th>
</tr>
</thead>
<tbody>
<tr>
<td>1+</td>
<td>H⁺</td>
<td>Hydrogen ion</td>
<td>NH₄⁺</td>
<td>Ammonium ion</td>
</tr>
<tr>
<td></td>
<td>Li⁺</td>
<td>Lithium ion</td>
<td>Cu⁺</td>
<td>Copper(I) or cuprous ion</td>
</tr>
<tr>
<td></td>
<td>Na⁺</td>
<td>Sodium ion</td>
<td></td>
<td></td>
</tr>
<tr>
<td></td>
<td>K⁺</td>
<td>Potassium ion</td>
<td></td>
<td></td>
</tr>
<tr>
<td></td>
<td>Cs⁺</td>
<td>Cesium ion</td>
<td></td>
<td></td>
</tr>
<tr>
<td></td>
<td>Ag⁺</td>
<td>Silver ion</td>
<td></td>
<td></td>
</tr>
<tr>
<td>2+</td>
<td>Mg²⁺</td>
<td>Magnesium ion</td>
<td>Co²⁺</td>
<td>Cobalt(II) or cobaltous ion</td>
</tr>
<tr>
<td></td>
<td>Ca²⁺</td>
<td>Calcium ion</td>
<td>Cu²⁺</td>
<td>Copper(II) or cupric ion</td>
</tr>
<tr>
<td></td>
<td>Sr²⁺</td>
<td>Strontium ion</td>
<td>Fe²⁺</td>
<td>Iron(II) or ferrous ion</td>
</tr>
<tr>
<td></td>
<td>Ba²⁺</td>
<td>Barium ion</td>
<td>Mn²⁺</td>
<td>Manganese(II) or manganous ion</td>
</tr>
<tr>
<td></td>
<td>Zn²⁺</td>
<td>Zinc ion</td>
<td>Hg₂⁺</td>
<td>Mercury(I) or mercurous ion</td>
</tr>
<tr>
<td></td>
<td>Cd²⁺</td>
<td>Cadmium ion</td>
<td>Hg²⁺</td>
<td>Mercury(II) or mercuric ion</td>
</tr>
<tr>
<td></td>
<td></td>
<td></td>
<td>Ni²⁺</td>
<td>Nickel(II) or nickleous ion</td>
</tr>
<tr>
<td></td>
<td></td>
<td></td>
<td>Pb²⁺</td>
<td>Lead(II) or plumbous ion</td>
</tr>
<tr>
<td></td>
<td></td>
<td></td>
<td>Sn²⁺</td>
<td>Tin(II) or stannous ion</td>
</tr>
<tr>
<td>3+</td>
<td>Al³⁺</td>
<td>Aluminum ion</td>
<td>Cr³⁺</td>
<td>Chromium(III) or chromic ion</td>
</tr>
<tr>
<td></td>
<td></td>
<td></td>
<td>Fe³⁺</td>
<td>Iron(III) or ferric ion</td>
</tr>
</tbody>
</table>

*The most common ions are in boldface.*
Common Anions

<table>
<thead>
<tr>
<th>Charge</th>
<th>Formula</th>
<th>Name</th>
<th>Formula</th>
<th>Name</th>
</tr>
</thead>
<tbody>
<tr>
<td>1−</td>
<td>H(^{-})</td>
<td>Hydride ion</td>
<td>C(_2)H(_3)O(_2)(^{-})</td>
<td>Acetate ion</td>
</tr>
<tr>
<td></td>
<td>F(^{-})</td>
<td>Fluoride ion</td>
<td>ClO(_3)(^{-})</td>
<td>Chlorate ion</td>
</tr>
<tr>
<td></td>
<td>Cl(^{-})</td>
<td>Chloride ion</td>
<td>ClO(_4)(^{-})</td>
<td>Perchlorate ion</td>
</tr>
<tr>
<td></td>
<td>Br(^{-})</td>
<td>Bromide ion</td>
<td>NO(_3)(^{-})</td>
<td>Nitrate ion</td>
</tr>
<tr>
<td></td>
<td>I(^{-})</td>
<td>Iodide ion</td>
<td>MnO(_4)(^{-})</td>
<td>Permanganate ion</td>
</tr>
<tr>
<td></td>
<td>CN(^{-})</td>
<td>Cyanide ion</td>
<td></td>
<td></td>
</tr>
<tr>
<td></td>
<td>OH(^{-})</td>
<td>Hydroxide ion</td>
<td></td>
<td></td>
</tr>
<tr>
<td>2−</td>
<td>O(_2)(^{-})</td>
<td>Oxide ion</td>
<td>CO(_3)(^{-})</td>
<td>Carbonate ion</td>
</tr>
<tr>
<td></td>
<td>O(_2)(^{2-})</td>
<td>Peroxide ion</td>
<td>CrO(_4)(^{2-})</td>
<td>Chromate ion</td>
</tr>
<tr>
<td></td>
<td>S(^{2-})</td>
<td>Sulfide ion</td>
<td>Cr₂O(_7)(^{2-})</td>
<td>Dichromate ion</td>
</tr>
<tr>
<td>3−</td>
<td>N(_3)(^{-})</td>
<td>Nitride ion</td>
<td>PO(_4)(^{3-})</td>
<td>Phosphate ion</td>
</tr>
</tbody>
</table>

*The most common ions are in boldface.

**PRACTICE EXERCISE**

1. **Which of the following compounds would you expect to be ionic:**
   N\(_2\)O, Na\(_2\)O, CaCl\(_2\), SF\(_4\)?

   **Solution:** Na\(_2\)O and CaCl\(_2\) are ionic compounds because they are composed of a metal combined with a nonmetal. The other two compounds, composed entirely of nonmetals, are predicted (correctly) to be molecular compounds.

2. **Which of the following compounds are molecular:** CBr\(_4\), FeS, P\(_4\)O\(_6\), PbF\(_2\)?

   **Answer:** CBr\(_4\) and P\(_4\)O\(_6\)

3. **What are the empirical formulas of the compounds formed by:**

   - **(a)** Al\(^{3+}\) and Cl\(^{-}\) ions,
   - **(b)** Al\(^{3+}\) and O\(^{2-}\) ions,
   - **(c)** Mg\(^{2+}\) and NO\(_3\)\(^{-}\) ions?
Solution

(a) Three Cl\(^{-}\) ions are required to balance the charge of one Al\(^{3+}\) ion. Thus, the formula is AlCl\(_3\).

(b) Two Al\(^{3+}\) ions are required to balance the charge of three O\(^{2-}\) ions (that is, the total positive charge is 6+ and the total negative charge is 6–). Thus, the formula is Al\(_2\)O\(_3\).

(c) Two NO\(_3^{-}\) ions are needed to balance the charge of one Mg\(^{2+}\). Thus, the formula is Mg(NO\(_3\))\(_2\). In this case the formula for the entire polyatomic ion NO\(_3^{-}\) must be enclosed in parentheses so that it is clear that the subscript 2 applies to all the atoms of that ion.

4. Write the empirical formulas for the compounds formed by the following ions:

   (a) Na\(^{+}\) and PO\(_4^{3-}\), (b) Zn\(^{2+}\) and SO\(_4^{2-}\), (c) Fe\(^{3+}\) and CO\(_3^{2-}\).

   \textit{Answers:} (a) Na\(_3\)PO\(_4\), (b) ZnSO\(_4\), (c) Fe\(_2\)(CO\(_3\))\(_3\)

NAMING OF COMPOUNDS

First distinguish between \textbf{organic} and \textbf{inorganic} compounds. Organic compounds: contain carbon, usually in combination with H,N,O,S e.g. CH\(_4\), CH\(_3\)CH\(_2\)CH\(_2\)CH\(_3\), C\(_6\)H\(_5\)NH\(_2\), CH\(_3\)CH\(_2\)OH

All other compounds are considered \textbf{inorganic}.

Exceptions: some carbon containing compounds are inorganic e.g. CO, CO\(_2\), CS\(_2\), CN\(^{-}\), CO\(_3^{2-}\), HCO\(_3^{-}\).

Inorganic compounds are classified into four categories;

1. Ionic compounds
2. Molecular compounds
3. Acids and bases
4. Hydrates
Inorganic Nomenclature (naming)

- Write the name of the cation.
- If the anion is an element, change its ending to -ide; if the anion is a polyatomic ion, simply write the name of the polyatomic ion.
- If the cation can have more than one possible charge, write the charge as a Roman numeral in parentheses.

  e.g. NaCl is sodium chloride, KBr is potassium bromide
  ZnI₂ is zinc iodide, Al₂O₃ is aluminium oxide

The ending -ide is also used for certain anions
e.g. OH⁻ = hydroxide as in LiOH – lithium hydroxide
e.g. CN⁻ = cyanide as KCN – potassium cyanide

The -ide nomenclature of some common anions

<table>
<thead>
<tr>
<th>Group 4A</th>
<th>element</th>
<th>anion</th>
<th>formula</th>
</tr>
</thead>
<tbody>
<tr>
<td>C</td>
<td>C</td>
<td>carbide</td>
<td>(C⁺), C₂²⁻</td>
</tr>
<tr>
<td>Si</td>
<td>Si</td>
<td>silicide</td>
<td>Si⁴⁺</td>
</tr>
</tbody>
</table>

<table>
<thead>
<tr>
<th>Group 5A</th>
<th>element</th>
<th>anion</th>
<th>formula</th>
</tr>
</thead>
<tbody>
<tr>
<td>N</td>
<td>N</td>
<td>nitride</td>
<td>N³⁻</td>
</tr>
<tr>
<td>P</td>
<td>P</td>
<td>phosphide</td>
<td>P³⁻</td>
</tr>
</tbody>
</table>

<table>
<thead>
<tr>
<th>Group 6A</th>
<th>element</th>
<th>anion</th>
<th>formula</th>
</tr>
</thead>
<tbody>
<tr>
<td>O</td>
<td>O</td>
<td>oxide</td>
<td>O²⁻</td>
</tr>
<tr>
<td>S</td>
<td>S</td>
<td>sulfide</td>
<td>S²⁻</td>
</tr>
<tr>
<td>Se</td>
<td>Se</td>
<td>selenide</td>
<td>Se²⁻</td>
</tr>
<tr>
<td>Te</td>
<td>Te</td>
<td>telluride</td>
<td>Te²⁻</td>
</tr>
</tbody>
</table>

<table>
<thead>
<tr>
<th>Group 7A</th>
<th>element</th>
<th>anion</th>
<th>formula</th>
</tr>
</thead>
<tbody>
<tr>
<td>F</td>
<td>F</td>
<td>fluoride</td>
<td>F⁻</td>
</tr>
<tr>
<td>Cl</td>
<td>Cl</td>
<td>chloride</td>
<td>Cl⁻</td>
</tr>
<tr>
<td>Br</td>
<td>Br</td>
<td>bromide</td>
<td>Br⁻</td>
</tr>
<tr>
<td>I</td>
<td>I</td>
<td>iodide</td>
<td>I⁻</td>
</tr>
</tbody>
</table>
Metals can form more than one type of cation

"-ous" and "-ic" nomenclature

e.g.  \( \text{Fe}^{2+} = \text{ferrous (less positive charges)} \)
     \( \text{FeCl}_2 = \text{ferrous chloride} \)

     \( \text{Fe}^{3+} = \text{ferric (more positive charges)} \)
     \( \text{FeCl}_3 = \text{ferric chloride} \)

     \( \text{CuCl} = \text{cuprous chloride (Cu}^+\) \)
     \( \text{CuCl}_2 = \text{cupric chloride (Cu}^{2+}\) \)

Some metals can assume 3 different charges. Use Roman numerals to designate the different cations – called the **STOCK SYSTEM**

e.g.  \( \text{Mn}^{2+} - \text{MnO} \text{ manganese(II) oxide} \)
     \( \text{Mn}^{3+} - \text{Mn}_2\text{O}_3 \text{ manganese(III) oxide} \)
     \( \text{Mn}^{4+} - \text{MnO}_2 \text{ manganese(IV) oxide} \)

**Patterns in Oxyanion Nomenclature**

When there are two oxyanions involving the same element:

- The one with fewer oxygens ends in -ite
  - \( \text{NO}_2^- : \text{nitrite}; \text{SO}_3^{2-} : \text{sulfite} \)
- The one with more oxygens ends in -ate
  - \( \text{NO}_3^- : \text{nitrate}; \text{SO}_4^{2-} : \text{sulfate} \)
**Molecular Compounds**

Made up of discrete molecular units. Many molecular compounds are binary compounds. The ending on the more electronegative element is changed to *-ide.*

**e.g.**  
HCl = hydrogen chloride,  HBr = hydrogen bromide  
SiC = silicon carbide

Other examples,

- CO = carbon monoxide,  CO\(_2\) = carbon dioxide  
- SO\(_2\) = sulfur dioxide,  SO\(_3\) = sulfur trioxide  
- NO\(_2\) = nitrogen dioxide  CCl\(_4\) = carbon tetrachloride

<table>
<thead>
<tr>
<th>Prefix</th>
<th>Meaning</th>
</tr>
</thead>
<tbody>
<tr>
<td>Mono-</td>
<td>1</td>
</tr>
<tr>
<td>Di-</td>
<td>2</td>
</tr>
<tr>
<td>Tri-</td>
<td>3</td>
</tr>
<tr>
<td>Tetra-</td>
<td>4</td>
</tr>
<tr>
<td>Penta-</td>
<td>5</td>
</tr>
<tr>
<td>Hexa-</td>
<td>6</td>
</tr>
<tr>
<td>Hepta-</td>
<td>7</td>
</tr>
<tr>
<td>Octa-</td>
<td>8</td>
</tr>
<tr>
<td>Nona-</td>
<td>9</td>
</tr>
<tr>
<td>Deca-</td>
<td>10</td>
</tr>
</tbody>
</table>

If the prefix ends with *a* or *o* and the name of the element begins with a vowel, the two successive vowels are often elided into one:

N\(_2\)O\(_5\) = dinitrogen pentoxide, N\(_2\)O\(_4\) = dinitrogen tetroxide
Exceptions to the use of Greek prefixes for molecular compounds containing hydrogen:

B₂H₆ - diborane
CH₄ - methane
SiH₄ - silane
NH₃ - ammonia
PH₃ - phosphine
H₂O - water
H₂S - hydrogen sulfide

**Acid Nomenclature**

Oxoacids: are acids that contain H, O and another element

Formulas: H-first, central element, then O

e.g.  HNO₃ – nitric acid
     H₂CO₃ – carbonic acid
     H₂SO₄ – sulfuric acid
     HClO₃ – perchloric acid
• If the anion in the acid ends in -ide, change the ending to -ic acid and add the prefix hydro-:
  - HCl: hydrochloric acid
  - HBr: hydrobromic acid
  - HI: hydroiodic acid

• If the anion in the acid ends in -ate, change the ending to -ic acid:
  - HClO₃: chloric acid
  - HClO₄: perchloric acid

• If the anion in the acid ends in -ite, change the ending to -ous acid:
  - HClO: hypochlorous acid
  - HClO₂: chlorous acid

Patterns in Oxyanion Nomenclature
ACIDS AND ACID ANHYDRIDES

<table>
<thead>
<tr>
<th>Acid anhydride</th>
<th>Acid</th>
<th>Strong acid</th>
<th>Weak acid</th>
<th>Stable acid</th>
</tr>
</thead>
<tbody>
<tr>
<td>1. SO$_3$</td>
<td>H$_2$SO$_4$</td>
<td>✓</td>
<td></td>
<td>✓</td>
</tr>
<tr>
<td>2. SO$_2$</td>
<td>H$_2$SO$_3$</td>
<td></td>
<td>✓</td>
<td></td>
</tr>
<tr>
<td>3. P$_2$O$_5$</td>
<td>H$_3$PO$_4$</td>
<td>✓</td>
<td></td>
<td>✓</td>
</tr>
<tr>
<td>4. N$_2$O$_5$</td>
<td>HNO$_3$</td>
<td>✓</td>
<td></td>
<td>✓</td>
</tr>
<tr>
<td>5. N$_2$O$_3$</td>
<td>HNO$_2$</td>
<td></td>
<td>✓</td>
<td></td>
</tr>
<tr>
<td>6. HCl</td>
<td></td>
<td>✓</td>
<td></td>
<td>✓</td>
</tr>
<tr>
<td>7. acetic anhydride</td>
<td>CH$_3$COOH</td>
<td>✓</td>
<td>✓</td>
<td>✓</td>
</tr>
</tbody>
</table>

HYDRATES

These are compounds that have a specific number of water molecules attached to them or number of waters of hydration.

e.g. CuSO$_4$.5H$_2$O
FeSO$_4$.6H$_2$O
LiCl.H$_2$O
BaCl$_2$.2H$_2$O
MgSO$_4$.7H$_2$O
CaCl$_2$.5H$_2$O
(NH$_4$)$_2$SO$_4$.FeSO$_4$.6H$_2$O

\[
\text{MgSO}_4 \text{ (anhydrous)} \xrightleftharpoons{+7\text{H}_2\text{O}}{-7\text{H}_2\text{O}} \text{ MgSO}_4.7\text{H}_2\text{O} \text{ (hydrous)}
\]
Practice Exercise

1. Based on the formula for the sulfate ion, predict the formula for:
   (a) the selenate ion
   (b) the selenite ion. (Sulfur and selenium are both members of group 6A and form analogous oxyanions.)

Solution:
(a) The sulfate ion is $\text{SO}_4^{2-}$. The analogous selenate ion is therefore $\text{SeO}_4^{2-}$.
(b) The ending -ite indicates an oxyanion with the same charge but one O atom fewer than the corresponding oxyanion that ends in -ate. Thus, the formula for the selenite ion is $\text{SeO}_3^{2-}$.

2. The formula for the bromate ion is analogous to that for the chlorate ion. Write the formula for the hypobromite and perbromate ions.

Answer: $\text{BrO}^-$ and $\text{BrO}_4^-$

3. Name the following compounds: (a) $\text{K}_2\text{SO}_4$, (b) $\text{Ba(OH)}_2$, (c) $\text{FeCl}_3$.

Solution
(a) The cation in this compound is $\text{K}^+$ and the anion is $\text{SO}_4^{2-}$—we have the name of the compound, **potassium sulfate**.

(b) In this case the compound is composed of $\text{Ba}^{2+}$ and $\text{OH}^-$ ions. $\text{Ba}^{2+}$ is the barium ion and $\text{OH}^-$ is the hydroxide ion. Thus, the compound is called **barium hydroxide**.

(c) You must determine the charge of Fe in this compound because an iron atom can form more than one cation. Because the compound contains three $\text{Cl}^-$ ions, the cation must be $\text{Fe}^{3+}$ which is the iron(III), or ferric, ion. The $\text{Cl}^-$ ion is the chloride ion. Thus, the compound is **iron(III) chloride or ferric chloride**.
4. Name the following compounds: (a) NH₄Br, (b) Cr₂O₃, (c) Ca(NO₃)₂.

**Answers:**

(a) ammonium bromide, (b) chromium(III) oxide, (c) calcium(II) nitrate

5. Write the chemical formulas for the following compounds: (a) potassium sulfide, (b) calcium hydrogen carbonate, (c) nickel(II) perchlorate.

**Answers:** (a) K₂S, (b) Ca(HCO₃)₂, (c) Ni(ClO₄)₂

6. Name the following acids: (a) HCN, (b) HNO₃, (c) H₂SO₄, (d) H₂SO₃.

**Solution:**

(a) The anion from which this acid is derived is CN⁻ the cyanide ion. Because this ion has an -ide ending, the acid is given a hydro-prefix and an -ic ending: hydrocyanic acid. Only water solutions of HCN are referred to as hydrocyanic acid: The pure compound, which is a gas under normal conditions, is called hydrogen cyanide. Both hydrocyanic acid and hydrogen cyanide are extremely toxic.

(b) Because NO₃⁻ is the nitrate ion, HNO₃ is called nitric acid (the -ate ending of the anion is replaced with an -ic ending in naming the acid).

(c) Because SO₄²⁻ is the sulfate ion, H₂SO₄ is called sulfuric acid.

(d) Because SO₃²⁻ is the sulfite ion, H₂SO₃ is sulfurous acid (the -ite ending of the anion is replaced with an -ous ending).

7. Give the chemical formulas for (a) hydrobromic acid, (b) carbonic acid.

**Answers:** (a) HBr, (b) H₂CO₃