Chapter 2

Atoms, Ions, and Molecules
The Building Blocks of Matter
Chapter Outline

2.1 The Rutherford Model of Atomic Structure
2.2 Nuclides and Their Symbols
2.3 Navigating the Periodic Table
2.4 The Masses of Atoms, Ions, and Molecules
2.5 Moles and Molar Mass
2.6 Making Elements
2.7 Artificial Nuclides
Electrons

- J. J. Thomson (1897)
  - Beam from cathode ray tube was deflected toward positively charged plate
  - Atoms contain negatively charged particles with a constant mass-to-charge ratio

Figure 2.2
Robert Millikan (1909)

- Determined the mass and charge of an electron with his oil-droplet experiment
- $e^- = -1.602 \times 10^{-19} \text{ C}$
- $m_e = 9.109 \times 10^{-28} \text{ g}$
Thomson’s Plum-Pudding Model

- Plum-Pudding Model:
  - $e^-$ distributed throughout diffuse, positively charged sphere
Henri Becquerel (1896)

- Some materials produce invisible radiation, consisting of charged particles

Radioactivity

- Spontaneous emission of high-energy radiation and particles
  - Beta particles ($\beta$, high-energy electrons)
  - Alpha particles ($\alpha$, +2 charge, mass = He nucleus)
Rutherford’s Experiment:
- Bombarded a thin gold foil with $\alpha$ particles to test Thomson’s model of the atom.
b) Expected results from “plum-pudding” model.  
c) Actual results.
The Nucleus:
• Positively charged center of an atom, containing nearly all of the atom’s mass
• About 1/10,000 the size of the atom
• Consists of two types of particles
  • Proton: Positively charged subatomic particle
    – Number defines the element
  • Neutron: Electrically neutral subatomic particle
Electrons:
- Negatively charged particles
- Roughly 2000 times smaller mass than protons
- Located outside of the nucleus in orbitals or “electron clouds”
- Outer electrons define the radius of the atom.
- Electrons, and the nucleus, are much smaller than the atom itself, so most of the atom is empty space
The Nuclear Atom

• Ions:
  • Atoms are electrically neutral
    • This means that they have the same number of electrons as protons
  • Ions are formed when atoms gain or lose electrons (which are negatively charged)
    • Cations have lost electrons, so have a positive charge
    • Anions have gained electrons, so have a negative charge
Nucleus: Protons (+ charge) plus neutrons (neutral)
Atomic Mass Units

- Atomic Mass Units (amu)
  - Unit used to express the relative masses of atoms and subatomic particles
  - Equal to 1/12 of a carbon atom:
    - 6 protons + 6 neutrons = 12 amu
  - 1 amu = 1 dalton (Da)
## Subatomic Particles

### TABLE 2.1 Properties of Subatomic Particles

<table>
<thead>
<tr>
<th>Particle</th>
<th>Symbol</th>
<th>Mass (amu)</th>
<th>Mass Number</th>
<th>Mass (kg)</th>
<th>Charge (relative value)</th>
<th>Charge (C)</th>
</tr>
</thead>
<tbody>
<tr>
<td>Neutron</td>
<td>(^1_n)</td>
<td>1.00867</td>
<td>1</td>
<td>(1.67493 \times 10^{-27})</td>
<td>0</td>
<td>0</td>
</tr>
<tr>
<td>Proton</td>
<td>(^1_p)</td>
<td>1.00728</td>
<td>1</td>
<td>(1.67262 \times 10^{-27})</td>
<td>1+</td>
<td>(+1.602 \times 10^{-19})</td>
</tr>
<tr>
<td>Electron</td>
<td>(^0_e) or (^0_\beta)</td>
<td>(5.485799 \times 10^{-4})</td>
<td>0</td>
<td>(9.10938 \times 10^{-31})</td>
<td>1−</td>
<td>(−1.602 \times 10^{-19})</td>
</tr>
</tbody>
</table>
2.1 The Rutherford Model of Atomic Structure
2.2 Nuclides and Their Symbols
2.3 Navigating the Periodic Table
2.4 The Masses of Atoms, Ions, and Molecules
2.5 Moles and Molar Mass
2.6 Making Elements
2.7 Artificial Nuclides
Ne gas ions of different masses strike detector in different locations
Isotopes

• Positive Ray Analyzer Results:
  • Two different kinds of neon gas atoms existed
    • 90% = 20 amu
    • 10% = 22 amu
  • Aston proposed theory of “isotopes”
• Isotopes: Atoms of an element containing the same # of protons but different # of neutrons
• Nuclide: A specific isotope of an element
Symbols of Nuclides

Atomic Mass ($A$) = total number of “nucleons” (protons, neutrons) in the nucleus

Elemental Symbol = a one- or two-letter symbol to identify the type of atom

Atomic Number ($Z$) = the number of protons in the nucleus; determines the identity of the element
Symbols of Nuclides

- $Z = \text{atomic \#} = \# \text{ of protons} = p$
- $A = \text{mass \#} = p + n$
- Isotopes are denoted using the chemical symbol, X, or the element name:
  
  $^{Z}A_{X} = ^{A}X = X-A = \text{element name-A}$

So a carbon isotope with 6 neutrons could be written as:

$^{12}_{6}C = ^{12}C = C-12 = \text{Carbon-12}$
Practice: Isotopic Symbols

• Use the format $^AX$ to write the symbol for the nuclides having 28 protons and 31 neutrons

• Collect and Organize:
• Analyze:
• Solve:
• Think about It:
Practice: Identifying Atoms and Ions

• Complete the missing information in the table.

<table>
<thead>
<tr>
<th>Symbol</th>
<th>$^{23}\text{Na}$</th>
<th>?</th>
<th>?</th>
<th>?</th>
</tr>
</thead>
<tbody>
<tr>
<td>Number of Protons</td>
<td>?</td>
<td>39</td>
<td>?</td>
<td>79</td>
</tr>
<tr>
<td>Number of Neutrons</td>
<td>?</td>
<td>50</td>
<td>?</td>
<td>?</td>
</tr>
<tr>
<td>Number of Electrons</td>
<td>?</td>
<td>?</td>
<td>50</td>
<td>?</td>
</tr>
<tr>
<td>Mass Number</td>
<td>?</td>
<td>?</td>
<td>118</td>
<td>197</td>
</tr>
</tbody>
</table>

• Collect and Organize:
• Analyze:
• Solve:
• Think about It:
Chapter Outline

2.1 The Rutherford Model of Atomic Structure
2.2 Nuclides and Their Symbols
2.3 Navigating the Periodic Table
2.4 The Masses of Atoms, Ions, and Molecules
2.5 Moles and Molar Mass
2.6 Making Elements
2.7 Artificial Nuclides
Mendeleev’s Periodic Table

Dmitri Mendeleev (1872)

- Ordered elements by atomic mass
- Arranged elements in columns based on similar chemical and physical properties
- Left open spaces in the table for elements not yet discovered

Figure 2.9
The Modern Periodic Table

• Also based on a classification of elements in terms of their physical and chemical properties.

• Horizontal rows: Called **periods** (1 → 7)
  Columns: Contain elements of the same family or **group** (1 → 18)

• Several groups have names as well as numbers.
Groups of Elements

- Group 1: Alkali metals
- Group 2: Alkaline earth metals
- Group 17: Halogens
- Group 18: Noble gases
Elements in the same group have similar properties.
<table>
<thead>
<tr>
<th>Main group elements</th>
<th>Transition metals</th>
</tr>
</thead>
<tbody>
<tr>
<td>(representative elements)</td>
<td></td>
</tr>
</tbody>
</table>
Broad Categories of Elements

- **Metals (left side and bottom of the table)**
  - Shiny solids; conduct heat and electricity; are malleable and ductile
- **Nonmetals (right side and top of the table)**
  - Solids (brittle), liquids and gases; nonconductors
- **Metalloids (between metals/nonmetals)**
  - Shiny solids (like metals); brittle (like nonmetals); semiconductors
Chapter Outline

2.1 The Rutherford Model of Atomic Structure
2.2 Nuclides and Their Symbols
2.3 Navigating the Periodic Table
2.4 The Masses of Atoms, Ions, and Molecules
2.5 Moles and Molar Mass
2.6 Making Elements
2.7 Artificial Nuclides
Most of the mass of an atom is in the nucleus, but the actual atomic mass is not exactly the sum of the masses of the nucleons (p + n).

- The electrons do have a small mass.
- Some mass is lost in the energy binding nucleons together ($E = mc^2$)

The periodic table shows the average atomic masses, in amu.

These masses are the weighted averages of the masses of all of the naturally occurring isotopes.
Average Atomic Mass

• **Average Atomic Mass:**
  • Weighted average of masses of all isotopes of an element
  • Calculated by multiplying the natural percent abundance of each isotope by its mass in amu and then summing these products

• **Natural Abundance:**
  • Proportion of a particular isotope, usually expressed as a percentage, relative to all the isotopes for that element in a natural sample
  • Assumes the same percentages over the surface of the Earth
Average atomic mass = sum over all isotopes of the mass of each isotope times its fractional abundance (percentage/100):

$$\text{Avg. at. mass} = \sum \left( \frac{\text{atomic mass}}{100\%} \right) \left( \frac{\text{percent abundance}}{100\%} \right)$$

where “Σ” means to sum over all isotopes present.
Average atomic mass of neon:

\[
(19.9924 \times 0.904838) + (20.99395 \times 0.002696) + (21.9914 \times 0.092465) = 20.1797 \text{ amu}
\]
Lithium has two naturally occurring isotopes:
Li-6  6.015 amu    7.42% abundance
Li-7  7.016 amu    92.58% abundance

So, for a standard sample of lithium:

\[
\text{avg. at. mass} = \frac{(6.015 \text{ amu})(7.42\%)}{100\%} + \frac{(7.016 \text{ amu})(92.58\%)}{100\%}
\]

\[
= 0.446 \text{ amu} + 6.50 \text{ amu}
\]

\[
= 6.94 \text{ amu}
\]
Chapter Outline

2.1 The Rutherford Model of Atomic Structure
2.2 Nuclides and Their Symbols
2.3 Navigating the Periodic Table
2.4 The Masses of Atoms, Ions, and Molecules
2.5 Moles and Molar Mass
2.6 Making Elements
2.7 Artificial Nuclides
The Mole

• A “mole” is a unit for a specific number:
  • 1 dozen items = 12 items
  • 1 mole particles = $6.022 \times 10^{23}$ particles (also known as Avogadro’s number)
• Exactly 12 grams of carbon-12 will contain exactly 1 mole of atoms
• Convenient unit for expressing macroscopic quantities (atoms or molecules) involved in macroscopic processes we observe
Moles as Conversion Factor

- To convert between number of particles and an equivalent number of moles

\[
\text{Number of particles} \times \frac{1 \text{ mole}}{6.022 \times 10^{23} \text{ particles}} = \text{Number of moles}
\]

\[
\text{Number of moles} \times \frac{6.022 \times 10^{23} \text{ particles}}{1 \text{ mole}} = \text{Number of particles}
\]
2 gloves—1 pair of gloves
12 rolls—1 dozen rolls
144 pencils—1 gross of pencils
500 sheets of paper—1 ream of paper
6.022 × 10^{23} iron atoms—1 mole of iron atoms
To convert between moles and numbers:

Use the conversion factor:

1 mole = 6.022E23

For example:

1.204E24 atoms \( \times \left( \frac{1 \text{ mole atoms}}{6.022\text{E23 atoms}} \right) = 2.000 \text{ moles atoms} \)

2.50 moles atoms \( \times \left( \frac{6.022\text{E23 atoms}}{1 \text{ mole atoms}} \right) = 1.506\text{E24 atoms} \)
Moles as Conversion Factor

1.5 mole molecules \( \left( \frac{6.022\text{E}23 \text{ molecules}}{1 \text{ mole molecules}} \right) \) = 9.03E23 molecules

1.5 mole \( H_2O \) \( \left( \frac{6.022\text{E}23 \text{ } H_2O}{1 \text{ mole } H_2O} \right) \) = 9.03E23 \( H_2O \) (molecules)

3.01E23 \( H_2O \) \( \left( \frac{1 \text{ mole } H_2O}{6.022\text{E}23 \text{ } H_2O} \right) \) = 0.500 mole \( H_2O \)
Molar Mass

• Molar Mass:
  • The mass (in grams) of one mole of the substance (atom, molecule or formula unit)
    • 1 atom of He = 4.003 amu
    • Mass of 1 mole of He atoms = 4.003 g
  • The molar mass ($M$) of He is 4.003 g/mol
Molar Mass

• Molar Mass:
  • The mass (in grams) of one mole of the substance (atom, molecule or formula unit)
  • The molar mass ($M$) of He is 4.003 g/mol

• The average mass on PT gives:
  • mass, in amu, of one atom of the element
    • 1 atom of He = 4.003 amu He
  • mass, in grams, of one mole of atoms of the element
    • 1 mole of He atoms = 4.003 g He
Molar Mass

To find the moles of hydrogen atoms in 1.01 grams of H:

\[ 1.01 \text{ g } H \left( \frac{1 \text{ mole } H}{1.01 \text{ g } H} \right) = 1.00 \text{ mole } H \]

or for 12.5 grams of hydrogen atoms:

\[ 12.5 \text{ g } H \left( \frac{1 \text{ mole } H}{1.01 \text{ g } H} \right) = 12.4 \text{ mole } H \]
Or to find the mass of 2.5 moles of helium:

\[
2.5 \text{ moles } He \left( \frac{4.00 \, \text{g } He}{1 \, \text{mole } He} \right) = 10. \, \text{g } He
\]

or for 2.5 moles of lead:

\[
2.5 \text{ moles } Pb \left( \frac{207.20 \, \text{g } Pb}{1 \, \text{mole } Pb} \right) = 518.0 \, \text{g } Pb
\]

Note these two examples have the same number of atoms, but different masses.
Molecular Mass/Formula Mass

• Molecular Mass:
  • Mass of one molecule of a molecular compound
  • Sum of the atomic masses of the atoms in that compound
  • Example: For one molecule of $\text{CO}_2$
    $\text{CO}_2 = \text{C} + 2 \text{O}$
    $= 12.01 \text{ amu} + 2(16.00\text{ amu})$
    $= 44.01 \text{ amu/molecule}$

• Formula Mass:
  • Mass in atomic mass units of one formula unit of an ionic compound (e.g., NaCl)
Molecular Mass/Formula Mass

So, to find the mass of 1.5 moles of CO$_2$:

$$1.5 \text{ moles } \text{CO}_2 \left( \frac{44.01 \text{ g CO}_2}{1 \text{ mole } \text{CO}_2} \right) = 66.02 \text{ g CO}_2$$

or to find the moles of molecules of 85.0 grams of CO$_2$:

$$85.0 \text{ g CO}_2 \left( \frac{1 \text{ mole } \text{CO}_2}{44.01 \text{ g CO}_2} \right) = 1.93 \text{ mole } \text{CO}_2$$
Moles, Mass, and Particles

- **MASS of element**
  - Multiply by molar mass, $M$
  - Divide by $M$

- **MOLES of element**
  - Divide by Avogadro’s number, $N_A$
  - Multiply by $N_A$

- **ATOMS of element**

- **MASS of compound**
  - Multiply by $M$
  - Divide by $M$

- **MOLES of compound**
  - Divide by $N_A$
  - Multiply by $N_A$

- **MOLECULES (or formula units) of compound**

- **Chemical formula**

© 2014 W. W. Norton Co., Inc.
Moles, Mass, and Particles

grams use molar mass

moles use Avogadro's number

molecules numbers (#)
Practice: Mole Calculations #1

a) How many moles of K atoms are present in 19.5 g of potassium?

b) How many formula units are present in 5.32 moles of baking soda (NaHCO$_3$)?

- Collect and Organize:
- Analyze:
- Solve:
- Think about It:
How many moles are present in 58.4 g of chalk (CaCO$_3$)?

- Collect and Organize:
- Analyze:
- Solve:
- Think about It:
The uranium used in nuclear fuel exists in nature in several minerals. Calculate how many moles of uranium are found in 100.0 grams of carnotite, $K_2(\text{UO}_2)_2(\text{VO}_4)_2\cdot3\text{H}_2\text{O}$.

- Collect and Organize:
- Analyze:
- Solve:
- Think about It:
Chapter Outline

2.1 The Rutherford Model of Atomic Structure
2.2 Nuclides and Their Symbols
2.3 Navigating the Periodic Table
2.4 The Masses of Atoms, Ions, and Molecules
2.5 Moles and Molar Mass
2.6 Making Elements
2.7 Artificial Nuclides
Big Bang Theory

H and He atoms in stars fuse to form heavier elements.

Subatomic particles fuse to form H and He nuclei.

Existence of subatomic particles.
Nucleosynthesis

- Nucleosynthesis:
  - Energy from Big Bang transformed into matter (more details of this matter/energy relationship in Chapter 21)
  - Fusing of fundamental/subatomic particles (protons/neutrons) created atomic nuclei

\[
\begin{align*}
\frac{1}{1} p + \frac{1}{0} n & \rightarrow \frac{2}{1} d \\
2 \frac{2}{1} d & \rightarrow \frac{4}{2} \alpha
\end{align*}
\]
Nuclear Binding Energies

• The stability of a nucleus is proportional to its binding energy (BE)
  • $E = (\Delta m)c^2$
  • $\Delta m$ = mass defect of the nucleus (in kg).
  • $c$ = speed of light ($2.998 \times 10^8$ m/s$^2$)

• Mass defect ($\Delta m$) – difference between the mass of the stable nucleus and the masses of the individual nucleons that compose it.
Stability of Nuclei

- Stability: Proportional to BE/# of nucleons
- $^{56}\text{Fe} = \text{most stable nucleus}$
Stellar Nucleosynthesis

- High density and temperature in stars caused additional fusion reactions to create elements heavier than H, He:

\[ 3 \, ^4_2\alpha \rightarrow {^{12}_6}\text{C} \]

\[ {^{12}_6}\text{C} + ^4_2\alpha \rightarrow {^{16}_8}\text{O} \]

- Stellar core forms shells of heavier elements produced from fusion of lighter elements
Alpha (\(\alpha\)) Decay

- **Alpha Decay:** Nuclear reaction in which an unstable nuclide spontaneously emits an alpha particle
  - \(\alpha\) particle = He nucleus

- **Example:**
  \[
  {^{238}_{92}\text{U}} \rightarrow {^{234}_{90}\text{Th}} + {^{4}_{2}\alpha}
  \]
Beta Decay

- **Beta Decay:** Spontaneous ejection of a \( \beta \)-particle (electron) by a neutron-rich nucleus

\[
\begin{align*}
1_0 ^0n & \rightarrow 1_1 ^1p + 0_{-1} ^{-1}e
\end{align*}
\]

- **Example:** \( ^{14}_6 \text{C} \rightarrow ^{14}_7 \text{N} + \beta^0 \)

- (Note: mass and charge balance!)
Chapter Outline

2.1 The Rutherford Model of Atomic Structure
2.2 Nuclides and Their Symbols
2.3 Navigating the Periodic Table
2.4 The Masses of Atoms, Ions, and Molecules
2.5 Moles and Molar Mass
2.6 Making Elements
2.7 Artificial Nuclides
Radioactive radon-222 decays with a loss of one $\alpha$ particle. Write the balanced equation for this decay.

- Collect and Organize:
- Analyze:
- Solve:
- Think about It:
The Big Bang resulted in the formation of extremely tiny particles called quarks.

Click here to launch the ChemTours website
This concludes the Lecture PowerPoint presentation for Chapter 2.