Atoms, Elements, and the Periodic Table

**Element**: a pure substance that cannot be broken down into simpler substances by a chemical reaction.

- Each element is identified by a one- or two-letter symbol.
- Elements are arranged specifically in the periodic table.
- The position of an element in the periodic table tells us much about its chemical properties.

### TABLE 2.1 Common Elements and Their Symbols

<table>
<thead>
<tr>
<th>Element</th>
<th>Symbol</th>
<th>Element</th>
<th>Symbol</th>
</tr>
</thead>
<tbody>
<tr>
<td>Bromine</td>
<td>Br</td>
<td>Magnesium</td>
<td>Mg</td>
</tr>
<tr>
<td>Calcium</td>
<td>Ca</td>
<td>Manganese</td>
<td>Mn</td>
</tr>
<tr>
<td>Carbon</td>
<td>C</td>
<td>Molybdenum</td>
<td>Mo</td>
</tr>
<tr>
<td>Chlorine</td>
<td>Cl</td>
<td>Nitrogen</td>
<td>N</td>
</tr>
<tr>
<td>Chromium</td>
<td>Cr</td>
<td>Oxygen</td>
<td>O</td>
</tr>
<tr>
<td>Cobalt</td>
<td>Co</td>
<td>Phosphorus</td>
<td>P</td>
</tr>
<tr>
<td>Copper</td>
<td>Cu</td>
<td>Potassium</td>
<td>K</td>
</tr>
<tr>
<td>Fluorine</td>
<td>F</td>
<td>Sodium</td>
<td>Na</td>
</tr>
<tr>
<td>Hydrogen</td>
<td>H</td>
<td>Sulfur</td>
<td>S</td>
</tr>
<tr>
<td>Iodine</td>
<td>I</td>
<td>Zinc</td>
<td>Zn</td>
</tr>
<tr>
<td>Lead</td>
<td>Pb</td>
<td></td>
<td></td>
</tr>
</tbody>
</table>
Composition and Structure of the Atom

Atom: basic unit of an element; smallest unit that retains chemical properties of an element

Subatomic particles: Small particles that are the building blocks from which atoms are made

- Protons
- Neutrons
- Electrons

Protons: Positively charged, high mass particle

Neutrons: Neutral (no) charge, high mass

Electrons: Negative charge (same amount of charge as proton), small mass

Nucleus: Central core of atom; contains the protons and neutrons; contains most of the mass of the atom

Outer Region: Area other than the nucleus; contains the electrons

Table 2.3

<table>
<thead>
<tr>
<th>Subatomic Particle</th>
<th>Charge</th>
<th>Mass (g)</th>
<th>Mass (amu)</th>
</tr>
</thead>
<tbody>
<tr>
<td>Proton</td>
<td>+1</td>
<td>$1.6726 \times 10^{-24}$</td>
<td>1</td>
</tr>
<tr>
<td>Neutron</td>
<td>0</td>
<td>$1.6749 \times 10^{-24}$</td>
<td>1</td>
</tr>
<tr>
<td>Electron</td>
<td>-1</td>
<td>$9.1093 \times 10^{-31}$</td>
<td>Negligible</td>
</tr>
</tbody>
</table>
Isotopes

Atoms are electrically neutral; they have a net charge of zero

# protons = # electrons

Atomic number: Number of protons in the nucleus of an atom; represented by the letter Z

Atomic number = # protons = # electrons

Number of neutrons in the nucleus can vary

Isotopes: Atoms of the same element that have different numbers of neutrons in their nuclei

Mass Number: Number of protons plus the number of neutrons in the nucleus of an atom; represented by the letter A

Example: 3 kinds of Hydrogen atoms
All have one proton and one electron
Atomic number for all three isotopes = 1
Mass number varies with number of neutrons

Hydrogen-1: one proton and no neutrons
Hydrogen-2: one proton and 1 neutron
Hydrogen-3: one proton and 2 neutrons

Isotope Symbol Notation: A\text{E}

E = Elemental Symbol
A = Mass number (number of protons plus number of neutrons)
Z = Atomic number (number of protons)

Practice: An atom has 7 protons and 8 neutrons

Atomic number = 7
Mass number = 15
Isotope symbol = ^{15}\text{A}
Relative Masses of Atoms and Molecules

Atomic mass unit (amu): Unit used to express mass of an atom relative to the mass of a hydrogen atom; one amu = the mass of 1 H atom

Atomic weight: Mass of an average atom of an element, expressed in atomic mass units

Example: Hydrogen atomic weight = 1.01 amu
Carbon atomic weight = 12.01 amu
Carbon atoms are 12 times as massive as hydrogen atoms

The atomic weight is the weighted average of the masses of the naturally occurring isotopes of a particular element reported in atomic mass units.

From the periodic table:

<table>
<thead>
<tr>
<th>atomic number (Z)</th>
<th>element symbol</th>
<th>atomic weight (amu)</th>
</tr>
</thead>
<tbody>
<tr>
<td>6</td>
<td>C</td>
<td>12.01</td>
</tr>
</tbody>
</table>

Practice: Ca atomic weight = 40.08 u
Ne atomic weight = 20.18 u

How many Neon atoms are required to give the same mass as one calcium atom?

Dalton’s Atomic Theory
( Mostly true, with a few modifications)

• All matter consists of atoms
• Atoms cannot be created or destroyed, divided* or changed* to a different atom
• Atoms of the same type of element identical properties

• Atoms of different elements have different properties

• Atoms of different elements combine to form stable compounds

• Chemical change (reactions) involved joining, separating or rearranging atoms

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**Elements and the Periodic Table**

**Dimitri Mendeleev, 1834-1907**

**Periodic Table:** Table of all elements arranged in order of increasing atomic number; elements with similar properties occur at regular (periodic) intervals

**Group (family):** Vertical column of elements that have similar chemical properties

**Period:** Horizontal row of elements

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**The Periodic Table**

[Image of the Periodic Table]
**Metals:** Elements in the left 2/3 of the periodic table; conduct heat and electricity; have a metallic shine; tend to lose electrons during reactions

**Nonmetals:** Elements in the right 1/3 of the periodic table; brittle, powdery solids or gases; tend to gain electrons during reactions

**Metalloids:** In-between metals and nonmetals; may exhibit properties of both; include, Si, Ge, As, Sb, Te, Po and At

**Representative elements:** followed by letter A; typical elements that follow trends

**Transition elements:** followed by letter B; do not necessarily follow all trends

**Alkali metals:** group IA

**Alkaline earth metals:** group IIA

**Halogens:** group VIIA

**Noble Gases:** group VIIIA; gases at room temperature; non-reactive with other substances
Atoms can gain or lose electrons!

Octet rule: atoms are most stable when their valence shell is full; they will gain or lose electrons to achieve this noble-gas configuration (ionization)

Ion: Atom with a positive or negative charge, due to the loss or gain of electrons

Examples: Na → Na⁺ + e⁻ \hspace{1cm} Cl + e⁻ → Cl⁻

Cations: positive ions; atoms that have lost electrons
Anions: negative ions; atoms that have gained electrons

Trends in the Periodic Table

Atomic size: Increases from top to bottom of each group
Decreases from left to right across a period
Atomic Size

- The size of atoms decreases down a column, as the valence e\(^-\) are farther from the nucleus.
- The size of atoms decreases across a row, as the number of protons in the nucleus increases.
- The increasing number of protons pulls the e\(^-\) closer to the nucleus, making the atoms smaller.

Ionization energy: Energy required to remove an electron from an atom

- Decreases from top to bottom of a group
- Increases from left to right across a period

Electron affinity: Energy released when an electron is added to an atom

- Decreases from top to bottom of a group
- Increases from left to right across a period

Ionization Energy

\[
\text{Na} + \text{energy} \rightarrow \text{Na}^+ + \text{e}^- \]

- Ionization energies decrease down a column as the valence e\(^-\) get farther away from the positively charged nucleus.
- Ionization energies increase across a row as the number of protons in the nucleus increases.
Tom Lehrer’s Element Song

Electron Arrangement in the Periodic Table

- Electrons fill shells and orbitals in a particular order
- Lowest energy (closest to the nucleus) shells and orbitals are filled first
- No more than 2 electrons can occupy any given orbital

Valence Shell: Outermost (highest energy) shell of an element that contains electrons

Valence electron: electron(s) located in the valence shell

- Elements derive their properties from their valence electrons
- Similar properties stem from identical number of valence electrons
- Periodic Table is arranged according to number and arrangement of electrons
- All elements in a specific period (row) of periodic table have the same valence energy level (shell)
- All elements in a specific group of periodic table have same number of electrons in valence shell; the Roman numeral of the Group indicates the number of Valence electrons
Example: Group IA has 1 electron in valence shell; Group IIIA has 3 electrons in valence shell

Practice: How many valence electrons in Mg? O? Ne?

Electron Configuration
- Shells (n=1, n=2...n=7)
- Subshells (s, p, d, f, g, h, etc.)
- Orbitals (volume of space within subshell)
  Valence shell: Outermost shell containing electrons
  Valence electrons: Electrons in the outermost shells

Group number indicates # valence electrons
Period number indicates which shell is valence shell

Recall, the shells/subshells fill from lowest energy to highest energy
A shell must be full before electrons will start filling the next higher shell
Each shell has unique number of subshells, equal to shell number
Example: for n=1, there is 1 subshell (s)
         for n=2, there are 2 subshells (s, p)
         for n=3 there are 3 subshells (s, p, d)
### Principle energy level (n) Possible subshells

<table>
<thead>
<tr>
<th>n</th>
<th>Subshells</th>
</tr>
</thead>
<tbody>
<tr>
<td>1</td>
<td>1s</td>
</tr>
<tr>
<td>2</td>
<td>2s, 2p</td>
</tr>
<tr>
<td>3</td>
<td>3s, 3p, 3d</td>
</tr>
<tr>
<td>4</td>
<td>4s, 4p, 4d, 4f</td>
</tr>
</tbody>
</table>

Each shell has a maximum capacity for electrons, equal to $2(n^2)$, where $n$ is the shell number.

**Example:**
- Shell #1: $n=1$, $2(1^2) = 2$ electrons, maximum
- Shell #2: $n=2$, $2(2^2) = 8$ electrons, maximum
- Shell #3: $n=3$, $2(3^2) = 18$ electrons, maximum

Each shell contains subshells (s, p, d, f, g, h) that contain orbitals. Maximum electrons in any given orbital is 2.

Each orbital is a volume of space (shape) in which electrons are most likely to be found.

- S orbital is a sphere
- P orbitals are dumbbell-shaped
- Other orbitals are more complex, and we will not discuss them (much)
Each shell (n level) has only 1 s orbital (1s, 2s, 3s, etc.)
- There are 3 p orbitals in each shell #2 and higher
- There are 5 d orbitals in shell 3 and higher
- There are 7 f orbitals in shell 4 and higher

<table>
<thead>
<tr>
<th>Subshell of orbitals</th>
<th>Maximum Number of Electrons</th>
</tr>
</thead>
<tbody>
<tr>
<td>s</td>
<td>2</td>
</tr>
<tr>
<td>p</td>
<td>6</td>
</tr>
<tr>
<td>d</td>
<td>10</td>
</tr>
<tr>
<td>f</td>
<td>14</td>
</tr>
</tbody>
</table>

**Electronic Structure**

**Subshells and Orbitals**

**TABLE 2.4** Orbitals and Electrons Contained in the Principal Energy Levels (n = 1–4)

<table>
<thead>
<tr>
<th>Shell</th>
<th>Orbitals</th>
<th>Electrons in Each Subshell</th>
<th>Maximum Number of Electrons</th>
</tr>
</thead>
<tbody>
<tr>
<td>1</td>
<td>1s</td>
<td>2</td>
<td>2</td>
</tr>
<tr>
<td>2</td>
<td>2s 2p</td>
<td>2 x 2 = 6</td>
<td>6</td>
</tr>
<tr>
<td>3</td>
<td>3s 3p 3d</td>
<td>2 x 2 = 6</td>
<td>18</td>
</tr>
<tr>
<td>4</td>
<td>4s 4p 4d</td>
<td>2 x 2 = 6</td>
<td>32</td>
</tr>
<tr>
<td></td>
<td>4f 5g</td>
<td>2 x 2 = 6</td>
<td></td>
</tr>
<tr>
<td></td>
<td>5f 6g</td>
<td>2 x 2 = 6</td>
<td></td>
</tr>
<tr>
<td></td>
<td>7f 8g</td>
<td>2 x 2 = 6</td>
<td></td>
</tr>
</tbody>
</table>
To write the electron configuration for an atom:

1. Obtain total number of electrons in the atom (from periodic table)
2. Begin filling lowest shells and subshells first (max. 2 per orbital)
3. Recall: 1 s orbital per subshell, 3 p orbitals, 5 d orbitals, 7 f orbitals;
   all shells have s, n=2 and higher have p, n=3 and higher have d, n=4
   and higher have f orbitals
4. The maximum number of electrons per orbital is 2; the maximum
   number of electrons per shell is \(2(n^2)\)

**Examples:**

- H: \(1s^1\)
- C: \(1s^2\ 2s^2\ 2p^2\)
- He: \(1s^2\)
- N: \(1s^2\ 2s^2\ 2p^3\)
- Li: \(1s^2\ 2s^1\)
- Ne: \(1s^2\ 2s^2\ 2p^6\)
- Be: \(1s^2\ 2s^2\)
- Na: \(1s^2\ 2s^2\ 2p^6\ 3s^1\)
- B: \(1s^2\ 2s^2\ 2p^1\)
- Al: \(1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^1\)

**Abbreviated electron configuration**

Use the preceding noble gas as a starting point, and write only
the additional electron configuration for the particular element

**Example:** Na has 11 electrons

Na: \(1s^2\ 2s^2\ 2p^6\ 3s^1\)

Its preceding noble gas is Neon: \(1s^2\ 2s^2\ 2p^6\)

Shorthand for Na is: \([\text{Ne}]3s^1\)

Shorthand for Ca is: \([\text{Ar}]4s^2\)

\([\text{Ne}]3s^2\) is Mg

\([\text{Ar}]4s^1\) is K

**Writing Orbital Diagrams**

Same idea as electron configuration, except we use arrows to represent
individual electrons, and lines to represent individual orbitals

See board and experiment 6 in lab manual for examples
Summary of Electron Configuration

• Shells, subshells and orbitals
• Filling order (low energy to high)
• Writing electron configurations longhand and shorthand
• Writing electron orbital diagrams

Summary

Atoms are made of protons, neutrons and electrons

Periodic table is arranged in order of increasing atomic size

Group (column) number = # valence electrons

Period (row) number = shell number of outermost energy level

Electrons reside in specific shells, subshells and orbitals, and fill these shells in a specific order (lowest energy first)

Atoms interact and undergo chemical reactions to achieve a stable (full) valence shell

Electrons change orbit only by absorbing or releasing energy:

adding energy $\rightarrow$ bumps $e^-$ to next energy level up (further from nucleus)

dropping to lower energy level $\rightarrow$ releases energy as light
Energy is Gained or Released as Electrons Change Orbitals

**Ground state:** an atom’s electrons are in the lowest possible energy levels

**Excited state:** an atom’s electrons are in a higher energy state, having absorbed energy to move the electrons to a level further from the nucleus

The energy absorbed or released by electrons as they move between levels (orbitals) is a type of electromagnetic radiation (visible light)