Chapter 2: Structure of the Atom

Read: BLB Ch. 2.1–2.4; Ch. 3.4
HW: BLB 2:4, 23, 25, 31; Packet 2:1–3
Know: the mole
atomic structure—protons, neutrons, electrons
atomic weights

KEY QUESTIONS: THE BIG PICTURE
What is the structure of an atom?
What are the properties of atoms?
Remember: structure affects function!
Where are the electrons located?
What is the energy of an electron?

History of Atomic Structure

- Ancient Greeks
- J.J. Thompson (early 1900s)
  Experiments with Cathode Ray Tubes
  • Discovery of electron
  • Charge-to-mass ratio
- Millikan - Charge of electron
- Becquerel - Radioactivity
- Rutherford (1910)
  Scattering of α particles by metal foil
  ⇒ dense, positively charged nucleus
Nuclear Atom

Rutherford concludes:
massive nucleus with electrons orbiting around it.

Modern View of atom

- An atom is almost entirely empty space.

1 nm = $10^{-9}$ m, $10$ Å = 1 nm

Atom Diameter = 0.1-0.5 nm
Nuclear diameter = $10^{-5}$ nm
Ratio of nuclear size to atomic radius: $10^{-4}$
Carbon atom diameter = 0.154 nm
Components of the Atom

<table>
<thead>
<tr>
<th></th>
<th>Mass (amu)</th>
<th>charge</th>
</tr>
</thead>
<tbody>
<tr>
<td>proton</td>
<td>p</td>
<td>~1</td>
</tr>
<tr>
<td>electron</td>
<td>e or e^-</td>
<td>~0</td>
</tr>
<tr>
<td>neutron</td>
<td>n</td>
<td>~1</td>
</tr>
</tbody>
</table>

neutral atom:

- $p$:
- $e$:
- $n$:

$^{19}\text{F}$

ion:

- $p$:
- $e$:
- $n$:

$^{39}\text{K}^+$

Element: defined by number of protons

Isotope: atoms of a given element with different numbers of neutrons

The Mole and Avogadro’s Number

Mole: connection between microscopic (what we don’t see: atoms) and macroscopic (what we see)

Laboratory-sized sample

1 mole of anything (atoms, molecules, green peas) = Avogadro’s number of them

1 mole $^{12}\text{C}$ atoms has $6.02 \times 10^{23}$ $^{12}\text{C}$ atoms.
1 mole $\text{CO}_2$ molecules = $6.02 \times 10^{23}$ $\text{CO}_2$ molecules
1 mole peas = $6.02 \times 10^{23}$ peas

Experimentally, 1 mole of $^{12}\text{C}$ has a mass of 12 g.
Atomic Mass and Molar Mass

- **Atomic mass (or weight):** in units of amu (atomic mass units); what one atom weighs. For a molecule or compound, we refer to a formula weight or molecular weight.

- **Molar mass:** in units of g/mol; what one mole of that atom (or molecule) weighs.

**Example:**

- 1 atom of $^{16}$O has an atomic weight of **16 amu**.

- 1 mol of $^{16}$O atoms has a molar mass of **16 g**.

- 1 molecule of NaOH has a mass of **40 amu**, and 1 mol of NaOH has a molar mass of **40 g**.

The Atomic Mass Scale

- **By definition:**
  - Mass of one atom of $^{12}$C = exactly 12 amu
  - Mass of one MOLE of $^{12}$C atoms = exactly 12 g
  - 1 $^{12}$C atom weighs $1.992648 \times 10^{-22}$ g
  
  $\Rightarrow 1.66054 \times 10^{-24} \text{ g} = 1 \text{ amu}$

**How many amu in one g?**

- **Using atomic mass units:**
  - One atom of $^1$H weighs $1.6735 \times 10^{-24}$ g
  - One atom of $^{16}$O weighs $2.6560 \times 10^{-23}$ g.

<table>
<thead>
<tr>
<th>Average Atomic Masses</th>
</tr>
</thead>
<tbody>
<tr>
<td>Relative atomic mass = average masses of isotopes:</td>
</tr>
<tr>
<td>Naturally occurring carbon consists of 98.892% $^{12}$C + 1.108% $^{13}$C.</td>
</tr>
</tbody>
</table>

**Average mass of C:**

**Average atomic mass is also known as Atomic weight**

**Atomic weights are listed on the periodic table.**
Chapter 5; Intro to Thermochemistry

- **Read:** BLB Chap. 5.1–5.3
- **HW:** BLB 5:3, 25, 28

- **Know:**
  - kinetic & potential energy
  - conservation of energy
  - energy: system & surroundings
  - State functions
  - What is enthalpy?

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**Energy and Thermodynamics**

- **Thermodynamics:** the study of energy and its transformations
- **Thermochemistry:** the study of the *relationship* between chemical reactions and energy changes

<table>
<thead>
<tr>
<th>Type</th>
<th>Example</th>
<th>Formula/Description</th>
</tr>
</thead>
<tbody>
<tr>
<td><strong>Kinetic</strong></td>
<td>mass in a place where force can act</td>
<td>$1/2mv^2$</td>
</tr>
<tr>
<td><strong>Potential</strong></td>
<td>moving charge</td>
<td>$Q_1Q_2/d$ (electrostatic)</td>
</tr>
<tr>
<td><strong>Light</strong></td>
<td>photons</td>
<td></td>
</tr>
<tr>
<td><strong>Chemical</strong></td>
<td>molecules moving uniformly</td>
<td>$\text{bonds}$</td>
</tr>
<tr>
<td><strong>Sound</strong></td>
<td>molecules moving randomly</td>
<td>$mgh$, where $g = 9.8 \text{ m/s}^2$</td>
</tr>
<tr>
<td><strong>Gravitational</strong></td>
<td></td>
<td></td>
</tr>
</tbody>
</table>
Units of energy

- SI unit: Joule (J)
- \(1 \text{ J} = 1 \text{ kg m}^2/\text{s}^2\)

[KE of a 2 kg object moving at 1 m per second using \(1/2 \text{mv}^2\), or PE of 100 g at height = 1 m using mgh]

In chemistry (!!!)

- kilojoule \(1 \text{ kJ} = 10^3 \text{ J}\)
- calorie \(1 \text{ cal} = 4.184 \text{ J}\)
- Calorie (dietary) \(1 \text{ Cal} = 10^3 \text{ cal} = 1 \text{ kcal}\)

First Law of Thermodynamics

- Energy can be converted from one form to another, but the total energy . . .

“Law of conservation of energy”

- all energy lost by a system under observation must be gained by the surroundings (and vice versa)

* during energy conversion, some heat is always produced
Internal Energy

**Internal energy:** total energy (E) associated with a system; the sum of all sources of kinetic & potential energy.

“The capacity to do work or transfer heat.”

**Quantitative definition:**

\[
\Delta E = E_{\text{final}} - E_{\text{initial}}
\]

\[
\Delta E = q + w
\]

- **E is a state function** (path independent; more about this later)

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Exothermic and Endothermic

**Exothermic energy**

System $\rightarrow$ Surroundings $\Delta E = -$

**Endothermic energy**

Surroundings $\rightarrow$ System $\Delta E = +$
State Functions

- a function whose value does not depend on the Pathway used to get to the present state:

  \[ \Delta \text{elevation} \]

- State Function: only depends on the current state (composition, T, P): does not depend on past history

\[ \text{Chemical Reaction} \]

\[ \begin{align*}
    \text{A} & \rightarrow \text{B} \\
    \text{A} & \rightarrow \text{C} \\
    \text{A} & \rightarrow \text{D} \\
\end{align*} \]

State functions are written as capital letters
\((E, H, P, V, T, S...)\)

q and w are not state functions
but \(\Delta E (= q + w)\) is a state function

Changes in state functions are path-independent:

\[ \Delta E \]

Enthalpy \((\Delta H)\)
\[ \Delta H = q_p \]
Defined as Energy transferred when P is constant

Chemical reactions are commonly run at constant pressure (atmospheric pressure); often use energy and enthalpy interchangeably.
At this point be sure you know:

• Difference between kinetic and potential energy (can you make up your own examples?)

• Know units of energy and order of magnitude of different kinds of energies (photon energy, etc.)

• Be watching for “electronic energy”, i.e., the energy of an electron.

• What is the First Law of Thermodynamics?

• Watch for how the potential energy of electrons gets converted to kinetic energy (a photon).

• What is meant by the system? The surroundings?

• What does the “sign” of energy mean?
  + means energy into the system (endothermic)
  – means energy out of the system (exothermic)

• Notice State functions (like P, T V) as we encounter them and keep the definition of state function in mind. Why is it useful to know that these are state functions?

• Remember that Enthalpy is another word for energy: specifically it is an energy measured when “the system” is at constant pressure, but for now, just know that it is a form of energy (heat).