Dalton's Atomic Theory:

- Each element is composed of extremely small particles called atoms.
- 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.
- Atoms of an element are not changed into atoms of a different element by chemical reactions.
- Atoms are neither created nor destroyed in chemical reactions.
- Atoms of more than one element combine in simple whole ratios to form Compounds.

Dalton’s Theory Explains

- Law of Constant Composition (law of definite proportions)
  In a compound the relative number and kinds of atoms are constant.
  ![Diagram of NH₃ formation]
  Ammonia always has 3 H and 1 N

- Law of Conservation of Mass
  Total mass of substances before chemical reaction is the same as the mass of substances after the reaction.
  ![Diagram of 3H₂ + N₂ → 2NH₃]
Law of Multible proportion
If two elements A and B combine to form more than one compound, the mass of B that combine with a given mass of A are in the ratio of small whole number.

**EXAMPLES!!!!!!!**

Example 1. The Law of Definite Proportions
Two samples of carbon dioxide are decomposed into their constituent elements. One sample produces 25.6 g of oxygen and 9.60 g of carbon, and the other produces 21.6 g of oxygen and 8.10 g of carbon. Show that these results are consistent with the law of definite proportions.

Example 2. Law of Multiple Proportions
Nitrogen forms several compounds with oxygen, including nitrogen dioxide and dinitrogen monoxide. Nitrogen dioxide contains 2.28 g oxygen to every 1.00 g nitrogen, while dinitrogen monoxide contains 0.570 g oxygen to every 1.00 g nitrogen. Show that these results are consistent with the law of multiple proportions.

Example 3. Law of Conservation of Mass
A 0.455 g sample of magnesium is allowed to burn in 2.315 g of oxygen gas. The sole product is magnesium oxide. After the reaction, no magnesium remains and the mass of unreacted oxygen is 2.015 g. What mass of magnesium oxide is produced?
Example: 7.12 g Mg is heated with 1.80 g of Br₂. All the Br₂ is used up, and 2.07 g magnesium bromide is only product. What mass of Magnesium remain unreacted?

Example: We wish to make exactly 2.00 g of MgO. What mass of Mg and Oxygen must we combine to do this?

2-2 Electrons and Other Discoveries in Atomic Physics

- Certain objects, in addition to mass, may display a property called electric charge, which can be either positive or negative.
- Object with similar charge (both positive or both negative) repel each other.
- Object with no charge exert no force on other. It is hung straight down due to the force of gravity.
- Object with opposite charge (+, -) attract each other.

\[ \text{FIGURE 2-4}\]

Forces between electrically charged objects
Atoms, Molecules, and Ions

**Discovery of Electrons**

- When Faraday passed electricity through glass tubes from which most of the air had been evacuated, he discovered cathode rays, a type of radiation emitted by the negative terminal, or cathode.

- The radiation crossed the evacuated tube to the positive terminal, or anode.

- Scientists found that cathode rays travel in straight lines and have properties that are independent of the cathode material (whether it is iron, platinum, and so on).
**Atoms, Molecules, and Ions**

**Discovery of Electrons**

- J. J. Thomson established the ratio of mass (m) to electric charge (e) for cathode rays, m/e.

- Thomson also concluded that cathode rays are negatively charged fundamental particles of matter found in all atoms.

- Cathode rays subsequently became known as electrons.

- The currently accepted value of the electronic charge e, expressed in coulombs is \(-1.6022 \times 10^{-19}\) C.

- By combining this value with an accurate value of the mass-to-charge ratio for an electron, we find that the mass of an electron is \(9.1094 \times 10^{-28}\) g.

  \[
  \text{Electron m/e} = -5.6857 \times 10^{-9} \text{ g coulomb}^{-1}
  \]

**X-Rays and Radioactivity**

- Radioactivity involves fundamental changes at the **subatomic** level – in radioactive decay, one element is changed into another, a process known as **transmutation**.

- Ernest Rutherford identified two types of radiation from radioactive materials, alpha (α) and beta (β).

- Alpha particles carry two fundamental units of **positive charge** and have essentially the same mass as helium atoms. In fact, alpha particles are identical to ions.

- Beta particles are **negatively charged particles** produced by changes occurring within the nuclei of radioactive atoms and have the same properties as electrons.

- A third form of radiation is not affected by electric or magnetic fields. This radiation, called gamma rays (γ), is **not made up of particles**; it is electromagnetic radiation of extremely high penetrating power.
The Structure of Atom

According to Rutherford nuclear theory of the atom:

- The nucleus is very small, dense, and positively charged.
- Electrons surround the nucleus.
- Most of the atom is empty space.

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 × 10^{-4}</td>
</tr>
</tbody>
</table>

The mass of an electron is so small we ignore it.

1 amu = 1.66054 × 10^{-24} g

Atoms are also extremely small in size!

- Diameter around 1×10^{-10} m and 5×10^{-10} m (100-500 pm)
- Non-SI unit used to express atomic dimension is the angstrom (Å)

1 Å = 10^{-10} m
Symbols of Elements

Elements are symbolized by one or two letters.

Atomic Number

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

Atomic Mass

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

ISOTOPES

- Atoms with identical atomic numbers but different mass number are called ISOTOPES.
- Isotopes have different numbers of neutrons.

<table>
<thead>
<tr>
<th></th>
<th>11C</th>
<th>12C</th>
<th>13C</th>
<th>14C</th>
</tr>
</thead>
<tbody>
<tr>
<td>#</td>
<td>5</td>
<td>6</td>
<td>7</td>
<td>8</td>
</tr>
<tr>
<td>Neutrons</td>
<td></td>
<td></td>
<td></td>
<td></td>
</tr>
</tbody>
</table>
Atomic Mass (Average Mass)

- Because in the real world elements exist as mixtures of isotopes.
- Average mass is calculated from the isotopes of an element weighted by their relative abundances.

“Natural abundance”

Example

<table>
<thead>
<tr>
<th>Isotope</th>
<th>Abundance</th>
<th>Atomic mass</th>
</tr>
</thead>
<tbody>
<tr>
<td>(^{24}\text{Mg})</td>
<td>78.99%</td>
<td>23.98504 amu</td>
</tr>
<tr>
<td>(^{25}\text{Mg})</td>
<td>10.00%</td>
<td>24.98584 amu</td>
</tr>
<tr>
<td>(^{26}\text{Mg})</td>
<td>11.01%</td>
<td>25.98259 amu</td>
</tr>
</tbody>
</table>

what is the average molecular mass of magnesium (Mg)?

ANS: 24.31

Periodic Properties of the Elements

Example:
Cooper has two naturally occurring isotopes; \(^{63}\text{Cu}\) with mass of 62.9396 amu and natural abundance of 69.17 % and \(^{65}\text{Cu}\) with mass of 64.9278 amu and a natural abundance of 30.83 %. Calculate the atomic mass of cooper. (ANS: 63.55)
Atoms, Molecules, and Ions

Periodic Table

- Elements are arranged in order of increasing atomic number.

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>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>
</tbody>
</table>

These four groups are known by their names.

- The rows on the periodic chart are periods.
- Columns are groups.
- Elements in the same group have similar chemical properties.

Nonmetals are on the upper right-hand corner 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.

Metals versus Nonmetals

**Metals**
- Have a shiny luster; various colors, although most are silverly
- Good conductors of heat and electricity
- Most metal oxides are ionic solids that are basic
- Tend to form cations in aqueous solution

**Nonmetals**
- Do not have a luster; various colors
- Poor conductors of heat and electricity
- Most nonmetal oxides are molecular substances that form acidic solutions
- Tend to form anions or oxyanions in aqueous solution
Atoms, Molecules, and Ions

Metals versus Nonmetals

• Metals tend to form cations.
• Nonmetals tend to form anions.

Metals

They tend to be lustrous, malleable, ductile, and good conductors of heat and electricity.

They tend to have low ionization energies so tend to form (+) ions. As a result they are oxidized (lose electrons) when they undergo chemical reactions.

2Ni(s) + O_2(g) → 2NiO(s)

Nonmetals

• These are dull, brittle substances that are poor conductors of heat and electricity.
• They tend to gain electrons in reactions with metals to acquire a noble gas configuration.
Nonmetals

- Substances containing only nonmetals are molecular compounds (For example the oxides, halides and hydrides of nonmetals are molecular compounds)
- Most nonmetal oxides are acidic.

Metalloids

- These have some characteristics of metals and some of nonmetals.
- For instance, silicon looks shiny, but is brittle and fairly poor conductor.

Group Trends

- Active Metals
  - Alkali metals (group 1A)
  - Alkaline earth metals (group 2A)
- Halogens
- Nobel gases

Alkali Metals

- Alkali metals are soft, metallic solids.
**Atoms, Molecules, and Ions**

**Alkali Metals**

- They are found only in compounds in nature, not in their elemental forms.
- They also have low ionization energies which reflect the relative ease with which its outer electron can be removed. As a result, the alkali metals are all very reactive, readily loose electrons to form ions carrying a 1+ charge.

**Alkaline Earth Metals**

- Alkaline earth metals have higher densities and melting points than alkali metals.
- Their ionization energies are low, but not as low as those of alkali metals. Consequently, the alkaline earth metals are less reactive than their alkali metal neighbors.

**Group VIIA: Halogens**

- The halogens are prototypical nonmetals.
Atoms, Molecules, and Ions

Group VIIA: Halogens
- They have large, negative electron affinities. 
  ➢ Therefore, they tend to oxidize other elements easily.
- They react directly with metals to form metal halides.
- Chlorine is added to water supplies to serve as a disinfectant.

Group VIIA: Noble Gases
- The noble gases have astronomical ionization energies.
- Their electron affinities are positive. 
  ➢ Therefore, they are relatively unreactive.
- They are found as monatomic gases.

Group VIIIA: Noble Gases
- Noble gases have completely filled s and p subshells. Because the noble gases possess such stable electron configuration, they are exceptionally unreactive.
Atoms, Molecules, and Ions

1 mole of any particle (atom, molecule or ion) contain $6.02 \times 10^{23}$ number of these particle.

1 mol C atom = $12 \text{ g C} = 6.02 \times 10^{23}$ C atoms

1 mol H$_2$O molecule = $6.02 \times 10^{23}$ H$_2$O molecules

1 mol of NO$_3^-$ ions = $6.02 \times 10^{23}$ NO$_3^-$ ions

Example:

How many copper atoms are in a copper penny with a mass of 310 g. (MW of Cu: 63.55 g/mol)

Ans: $2.94 \times 10^{22}$ Cu atoms

Avogadro’s Number

Avogadro’s Hypothesis and Avogadro Number

Molar Mass

- Molar mass is the mass of 1 mol of a substance (i.e., g/mol).
  
  - The molar mass of an element is the mass number for the element that we find on the periodic table.
  
  - The formula weight (in amu's) will be the same number as the molar mass (in g/mol).
Atoms, Molecules, and Ions

Using Moles

Moles provide a bridge from the molecular scale to the real-world scale.

Example:
An aluminium sphere $8.55 \times 10^{22}$ Al atoms. What is the radius of the sphere in centimeters?

(Density of Al is 2.7 g/cm$^3$) (MW of Al 26.98 g Al)(V=$\frac{4}{3}\pi r^3$)

Ans: 0.697 cm

Mole Relationships

<table>
<thead>
<tr>
<th>Name of Substance</th>
<th>Formula</th>
<th>Formula Weight (amu)</th>
<th>Molar Mass (g/mol)</th>
<th>Number and Kind of Particles in One Mole</th>
</tr>
</thead>
<tbody>
<tr>
<td>Atomic nitrogen</td>
<td>N</td>
<td>14.0</td>
<td>14.0</td>
<td>$6.02 \times 10^{23}$ N atoms</td>
</tr>
<tr>
<td>Molecular nitrogen</td>
<td>$N_2$</td>
<td>28.0</td>
<td>28.0</td>
<td>$\left{\begin{array}{l} 6.02 \times 10^{23} N_2 \text{ molecules} \ 2 \times 6.02 \times 10^{24} N \text{ atoms} \end{array}\right.$</td>
</tr>
<tr>
<td>Silver</td>
<td>Ag</td>
<td>107.9</td>
<td>107.9</td>
<td>$6.02 \times 10^{23}$ Ag atoms</td>
</tr>
<tr>
<td>Silver ions</td>
<td>$Ag^+$</td>
<td>107.9</td>
<td>107.9</td>
<td>$6.02 \times 10^{23}$ Ag$^+$ ions</td>
</tr>
<tr>
<td>Barium chloride</td>
<td>BaCl$_2$</td>
<td>208.2</td>
<td>208.2</td>
<td>$6.02 \times 10^{23}$ BaCl$_2$ units</td>
</tr>
<tr>
<td></td>
<td></td>
<td></td>
<td></td>
<td>$2 \times 6.02 \times 10^{23}$ Cl$^-$ ions</td>
</tr>
</tbody>
</table>

- One mole of atoms, ions, or molecules contains Avogadro's number of those particles.
- One mole of molecules or formula units contains Avogadro's number times the number of atoms or ions of each element in the compound.

2.57. Determine
(a) the number of moles of Zn in a 415.0 g sample of zinc metal. (Molar mass of Zn is 65.38 g/mol)

(b) the number of Cr atoms in 147.4 kg chromium. (Molar mass of Cr is 51.996 g/mol)

(c) the mass of a one-trillion-atom sample ($1.0 \times 10^{12}$) of metallic gold. (Molar mass of Au is 196.97 g/mol)

(d) the mass of one fluorine atom. (Molar mass of F is 18.998 g/mol)
Example:
Fill in the gaps in the following table, assuming each column represents a neutral atom.

<table>
<thead>
<tr>
<th>Symbol</th>
<th>$^64\text{Zn}$</th>
<th>$^{197}\text{Ru}$</th>
<th>$^{90}\text{Sr}$</th>
<th>$^{116}\text{Ag}$</th>
<th>$^{235}\text{U}$</th>
</tr>
</thead>
<tbody>
<tr>
<td>Protons</td>
<td>30</td>
<td>50</td>
<td>49</td>
<td>47</td>
<td>92</td>
</tr>
<tr>
<td>Neutrons</td>
<td></td>
<td>40</td>
<td>49</td>
<td></td>
<td></td>
</tr>
<tr>
<td>Electrons</td>
<td></td>
<td>38</td>
<td>47</td>
<td>108</td>
<td>235</td>
</tr>
<tr>
<td>Mass No</td>
<td></td>
<td></td>
<td></td>
<td></td>
<td></td>
</tr>
</tbody>
</table>

59. How many Cu atoms are present in a piece of sterling-silver jewelry weighing 33.24 g? (Sterling silver is a silver-copper alloy containing 92.5% Ag by mass.)

55. What is the total number of atoms in (a) 15.8 mol Fe; (b) 0.000467 mol Ag; (c) $8.5 \times 10^{-11}$ mol Na?

Example:
The percent natural abundance of 40 K isotope is 0.012 %. How many 40 K atoms do you ingest by drinking one cup of whole milk containing 371 mg K?
Example:
How many C atoms are there in a 1.00 carat diamond? Diamond is pure carbon, 1 carat=0.200g.