Elements can be characterized as metals, non-metals and metalloids. The periodic table is arranged according to properties of elements. Mendeleev and Meyer independently arranged known elements according to their properties. In the periodic table elements are arranged as periodic functions of their atomic number.

Columns -> groups -> similar chemical and physical properties
Rows -> periods -> properties change progressively across

Group IA – alkali metals (Li, Na, K, Rb, Cs, Fr)
Group IIA – alkaline (Be, Mg, Ca, Sr, Ba, Ra)
Group VIIA – halogens (F, Cl, Br, I, At)
VIIIA – Rare or noble gases (He, Ne, Ar, Kr, Xe, Rn)

Physical properties of metals vs. non-metals (pg. 127)

<table>
<thead>
<tr>
<th>Metals</th>
<th>Nonmetals</th>
</tr>
</thead>
<tbody>
<tr>
<td>High electrical conductivity</td>
<td>Poor electrical conductivity</td>
</tr>
<tr>
<td>High thermal conductivity</td>
<td>Good heat insulators (except diamond)</td>
</tr>
<tr>
<td>Metallic gray and silver luster (except Ag or Au)</td>
<td>No metallic luster</td>
</tr>
<tr>
<td>Almost all are solid</td>
<td>Solids, liquids, or gases</td>
</tr>
<tr>
<td>Malleable</td>
<td>Brittle in solid state</td>
</tr>
<tr>
<td>Ductile</td>
<td>Nonductile</td>
</tr>
</tbody>
</table>

Some chemical properties of metals vs. non-metals (pg. 127)

<table>
<thead>
<tr>
<th>Metals</th>
<th>Nonmetals</th>
</tr>
</thead>
<tbody>
<tr>
<td>Outer shell contain few electrons (3 or less)</td>
<td>Outer shell contains 4 or more electrons</td>
</tr>
<tr>
<td>Form cations by losing electrons</td>
<td>Form anions by gaining electrons</td>
</tr>
<tr>
<td>Form ionic compounds with nonmetals</td>
<td>Form ionic compounds with metals and molecular compounds with nonmetals</td>
</tr>
<tr>
<td>Solid state characterized by metallic bonding</td>
<td>Covalently bonded molecules, noble gases are monatomic</td>
</tr>
</tbody>
</table>

Being familiar with the arrangement of the periodic table enables us to understand reactions that elements undergo.

Aqueous Solutions (already familiar with descriptions of aqueous solutions, percent solute, molar)

Solution implies complete dissolution of solute, characterized by how they conduct electricity:
Strong electrolytes - conduct electricity well
Weak electrolytes – conduct electricity poorly
Non-electrolytes – do not conduct electricity
Dissociation: \( \text{NaCl(s)} \rightarrow \text{Na}^+(aq) + \text{Cl}^-(aq) \) Ionic compounds
Ionization: \( \text{HCl}(g) \rightarrow \text{H}^+(aq) + \text{Cl}^-(aq) \)  

**Molecular compounds**

**Examples of strong electrolytes** *(Table 4-5, 4-6, 4-7)*

Strong acids (HCl, H\(_2\)SO\(_4\))
Strong bases (NaOH, KOH)
Soluble salts (NaCl, KCl)

How do we define acids, bases, and salts?
Acid – donates \( \text{H}^+ \) when ionized
Base – donates \( \text{OH}^- \) when ionized
Salt – ionic substances that do not contain either \( \text{H}^+ \) or \( \text{OH}^- \)

Strong acids:

\[
\text{HCl} \rightarrow \text{H}^+(aq) + \text{Cl}^-(aq) \quad \text{completely ionized}
\]

Weak acids:

\[
\text{CH}_3\text{COOH} \rightleftharpoons \text{H}^+(aq) + \text{CH}_3\text{COO}^-(aq) \quad \text{reversible}
\]

Reversible Reactions:

\[
\text{H}^+(aq) + \text{Cl}^- \rightarrow \text{No Reaction (The reverse of the ionization of HCl)}
\]

\[
\text{H}^+(aq) + \text{CH}_3\text{COO}^-(aq) \quad \text{CH}_3\text{COOH} \quad \text{reversible}
\]

Strong Bases, Insoluble bases, Weak bases:

\[
\text{NaOH}(s) \rightarrow \text{Na}^+(aq) + \text{HO}^-(aq) \quad \text{to completion (strong base)}
\]

\[
\text{NH}_3(aq) + \text{H}_2\text{O} \rightarrow \text{NH}_4^+(aq) + \text{HO}^-(aq) \quad \text{reversible (weak base)}
\]

\[
\text{Cu(OH)}_2, \text{Zn(OH)}_2.. \quad \text{(insoluble in water)}
\]

**Reactions in aqueous solutions:** *(Table 4-8 shows solubility guidelines for Common ionic compounds in water)*

**Formula unit equations:** show complete formula for all compounds

\[
2\text{AgNO}_3(aq) + \text{Cu}(s) \rightarrow 2\text{Ag(s)} + \text{Cu(NO}_3)_2(aq) \quad \text{(both silver nitrate and copper nitrate are soluble ionic compounds)}
\]

**Total ionic equations:** shows the predominant form of each substance when in contact with aqueous solution

\[
2[\text{Ag}^+(aq) + \text{NO}_3^-(aq)] + \text{Cu}(s) \rightarrow 2\text{Ag(s)} + [\text{Cu}^{2+}(aq) + 2\text{NO}_3^-(aq)]
\]

**Net ionic equations:** show only species that react, ignore “spectator ions”
\[
2\text{Ag}^+(aq) + \text{Cu}(s) \rightarrow 2\text{Ag}(s) + \text{Cu}^{2+}(aq)
\]

Net ionic equations focus on what occurs in a chemical reaction in aqueous solutions

**Oxidation Numbers**

For a binary compound AX, the oxidation number is the number of electrons gained or lost by an atom of the element when it forms the compound. It is sometimes referred to as the oxidation state.

Oxidation numbers (Table 4-10) are used to track electron transfer in oxidation-reduction (redox) reactions. Oxidation numbers are more important in binary ionic compounds than molecular compounds. Some abbreviated rules that apply are below (full list on page 138):

1. Oxidation number of free element is zero, even polyatomic ones (S\(_8\), P\(_4\),...).
2. Oxidation number of an element that is a single monatomic ion is its charge (Na\(^+\) (+1); O\(^2-\), (-2)).
3. Sum of oxidation numbers of all atoms is zero.
4. In polyatomic ion, the sum of the oxidation numbers of constituent ions is equal to its charge.
5. Oxygen usually has a –2 oxidation number except peroxides (-1) and superoxides (-1/2).
6. Hydrogen is usually +1 except in hydrides where it is –1.
7. The position of the element on the periodic table determines its oxidation number:
   a. Group IA is always +1
   b. Group IIA is always +2
   c. Group IIIA is almost always +3
   d. Group VA is –3 in binary compounds with metals, H\(^+\) and NH\(_4^+\). There are exceptions and other rules apply.
   e. Group VIA below oxygen is –2 in binary compounds with metals, H\(^+\), and NH\(_4^+\). There are exceptions.
   f. Group VIIA is –1 in binary compounds with metals, H\(^+\), and NH\(_4^+\). There are exceptions.

Example 4-4 Determine the oxidation numbers of nitrogen in the following: (a) N\(_2\)O\(_4\); (b) NH\(_3\); (c) HNO\(_3\); (d) NO\(_3^-\); (e) N\(_2\)

**Naming inorganic compounds**

Binary compounds (name the more metallic first, the less metallic second). The less metallic is named by adding –ide to the stem. (pg. 140)

Binary ionic (metal cations, nonmetal anions) – cation named first, the anion second (KBr- potassium bromide, NaCl- sodium chloride).
Method is good for elements with one oxidation other than zero. For elements with more than one oxidation number (transition metals, others...) use roman numerals.

\[
\text{Cu}_2\text{O} \, (+1) \quad \text{copper(I) oxide} \quad \text{CuF}_2 \, (+2) \quad \text{copper(II) fluoride}
\]
Method can be applied to any binary compound of metal and nonmetal. An older system uses –ous and –ic to distinguish between lower and higher oxidation numbers

CuCl (+1) cuprous chloride CuCl₂ (+2) cupric chloride

Limited since only can distinguish between two different oxidation numbers for a metal

Pseudobinary compounds (named as though were binary compounds)

NH₄I ammonium iodide Ca(CN)₂ calcium cyanide

Table 4-11 (List of common cations and anions)

**Binary molecular compounds** (Mostly two nonmetals bonded together)

Use *Greek* and *Latin* prefixes instead of Roman numerals and suffixes.

Examples: SO₂ – sulfur dioxide; SO₃ – sulfur trioxide; As₄O₆ – tetraarsenic hexoxide

Learn the common prefixes (pg. 142)- 2 (di), 3 (tri), 4(tetra), etc….

**Binary acids**

HCl – Hydrogen chloride hydrochloric acid, HCl (aq)

named as typical binary compound in pure form, named as –ic acids when dissolved in water

Ternary acids and their salts

Usually composed of hydrogen, oxygen and another nonmetal. In some cases the designation of –ic at the ending is arbitrary, however –ous at the ending always has an oxidation state of 2 lower than in the –ic. (page 144)

Examples of acids:

H₂SO₄ (sulfuric acid, S(+6)) H₂SO₃ (sulfurous acid (+4))

HNO₃ (nitric acid (+5)) HNO₂ (nitrous acid (+3))

Examples of salts:

Acids with –ic endings produce –ate anions (SO₄²⁻ - sulfate; Na₂SO₄ – sodium sulfate)
Acids with –ous endings produce –ite anions (SO₃²⁻ - sulfite; Na₂SO₃ – sodium sulfite)
Prefixes per- and hypo- are retained (ClO⁻ - hypochlorite; NaClO- sodium hypochlorite)
Chemical Reactions

Oxidation-Reduction (Redox)

Oxidation originally referred to the combination of a substance with oxygen. Results in an increase in the oxidation state of the element and a loss of electrons.

\[4\text{Fe(s)} + 3\text{O}_2 \rightarrow \text{Fe}_2\text{O}_3\] Fe (0) to Fe (+3) – formation of rust

Combustion

\[\text{C(s)} + \text{O}_2 \rightarrow \text{CO}_2\] C (0) to C (+4)

\[2\text{CO (g)} + \text{O}_2\text{ (g)} \rightarrow 2\text{CO}_2\] C (+2) to C (+4)

\[\text{C}_3\text{H}_8\text{(g)} + 5\text{O}_2\text{(g)} \rightarrow 3\text{CO}_2\text{(g)} + 4\text{H}_2\text{O(g)}\] C (-8/3) to C (+4)

Reduction originally described the removal of oxygen from a compound. Results in a decrease in oxidation number of the element and a gain of electrons

\[\text{WO}_3\text{(s)} + 3\text{H}_2\text{(g)} \rightarrow \text{W(s)} + 3\text{H}_2\text{O(g)}\] W (+6) to W (0)

Oxidation and reductions occur simultaneously -> redox reactions

Oxidizing agents – oxidize other substances, contain atoms that are reduced, gain or appear to gain electrons.
Reducing agents – reduce other substances, contain atoms that are oxidized, lose or appear to lose electrons.

\[2\text{Fe(s)} + 3\text{Cl}_2\text{(g)} \rightarrow 2\text{FeCl}_3\text{(s)}\] Fe (0 to +3); Cl (0 to –1)

\[2\text{FeBr}_3\text{(aq)} + 3\text{Cl}_2 \rightarrow 2\text{FeCl}_3\text{(aq)} + 3\text{Br}_2\text{(l)}\] Br (-1 to 0); Cl (0 to –1)

Example 4-5: Write each of the following formula unit equations as net ionic equations. Which ones are redox reactions? If redox, identify oxidizing agent, reducing agent, species oxidized, species reduced.

(a) \[2\text{AgNO}_3\text{(aq)} + \text{Cu(s)} \rightarrow \text{Cu(NO}_3\text{)_2(aq)} + 2\text{Ag(s)}\]

(b) \[4\text{KClO}_3\text{(s)} \rightarrow \text{KCl(s)} + 3\text{KClO}_4\text{(s)}\]

(c) \[3\text{AgNO}_3\text{(aq)} + K_3\text{PO}_4\text{(aq)} \rightarrow \text{Ag}_3\text{PO}_4\text{(s)} + 3\text{KNO}_3\text{(aq)}\]

Combination Reactions (Most are red-ox reactions) – 2 or more substances combine to form a compound.
There are many types:
1 element + 1 element -> compound
metal + non-metal -> binary ionic compound (2M + X₂ -> 2M⁺X⁻ (s)), M = Li, Na, K, Rb, Cs, and X = F, Cl, Br, I.

This also occurs with group II metals (M + X₂ -> MX₂).

Non-metal + non-metal -> binary covalent compound
P₄(s) + 6 Cl₂ -> 4PCl₃
P₄(s) + 10Cl₂ -> 4PCl₅

Compound + element -> compound
PCl₃ + Cl₂ -> PCl₅(s)

Compound + compound -> compound
CaO + CO₂ -> CaCO₃(s)

**Decomposition Reactions** (reverse of combination) – can form 2 elements, 1 element and 1 or more compound, or 2 or more compounds.

Compound -> elements: 2H₂O -> 2H₂ + O₂ or 2HgO -> 2Hg + O₂

Compound -> compd. + element: KClO₃(s) -> KCl +3O₂

Compound -> 2 compounds: CaCO₃ -> CaO + CO₂

**Displacement Reactions**
Active metal + salt of less active metal -> less active metal + salt of active metal
Zn(s) + CuSO₄(aq) -> ZnSO₄(aq) + Cu(s)

Active metal + non-oxidizing acid -> Hydrogen + salt of Acid
Zn(s) + H₂SO₄

**Metathesis Reactions** – a reaction in which the positive ion and the negative ion change partners
AX + BY -> AY + BX or AgNO₃(aq) + NaCl -> AgCl(s) + NaNO₃ (AgCl precipitates)

Metathesis reactions produce either a weak or non-electrolyte such as H₂O or an insoluble solid.

Acid-Base neutralization: HCl + NaOH -> H₂O + NaCl or 2HI + Ca(OH)₂ -> 2H₂O + CaI₂

Precipitation reaction: Pb(NO₃)₂ + K₂CrO₄ -> PbCrO₄ + 2KNO₃

Chapter 4 Homework: 34, 35, 41, 42, 45, 46, 51, 52, 69, 76, 91, 92, 112