Chapter 1
The Atomic Nature of Matter

1-1 Chemistry: Science of Change
1-2 The Composition of Matter
1-3 The Atomic Theory of Matter
1-4 Chemical Formulas and Relative Atomic Masses
1-5 The Building Blocks of the Atom
1-6 Finding Atomic Masses the Modern Way
1-7 The Mole Concept: Counting and Weighing Atoms and Molecules
1-8 Finding Empirical and Molecular Formulas the Modern Way
1-9 Volume and Density
Chapter 2
Stoichiometry

• 2-1 Writing Balanced Chemical Equations
• 2-2 Using Balanced Chemical Equations
• 2-3 Limiting Reactant and Percentage Yield
• 2-4 The Stoichiometry of Reactions in Solution
• 2-5 the Scale of Chemical Processes
Limiting Reactant

\[ 2 \text{A} + \text{B} \rightarrow \text{A}_2\text{B} \]

1. Select any one product
2. Use the balanced equation to calculate the amount of the selected product that would form if the entire supply of the first reactant were used up.
3. Repeat with respect to every other reactant
4. Identify the limiting reactant as the reactant that gives the smallest amount to the selected product.
OFB Chapter 3
Chemical Periodicity and the Formation of Simple Compounds

3-1 Groups of Elements
3-2 The Periodic Table
3-3 Ions and Ionic Compounds
3-4 Covalent Bonding and Lewis Structures
3-5 Drawing Lewis Structures
3-6 Naming Compounds in which Covalent Bonding Occurs
3-7 The Shapes of Molecules
3-8 Elements Forming More than One Ion
Elements are classified as metal, non-metals, or semi-metals, and also fall into groups based on similarities in chemical and physical properties.
The Electronegativity (the power of an atom when in chemical combination to attract electrons to itself) is a periodic property.

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<th>0.98</th>
<th>1.57</th>
<th>0.93</th>
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</tr>
</tbody>
</table>
Lewis Structures: Ionic compounds

Whenever possible, the *valence electrons* in a compound are distributed in such a way that each main-group element in a molecule (except hydrogen) is surrounded by *eight electrons* (an octet of electrons). Hydrogen should have *two electrons* in such a structure.
Drawing Lewis Structures

1. Count up the **total number of valence electrons available** (symbolized by $A$) by first adding the group numbers of all the atoms present. If the species is a negative ion, *add* the absolute value of the total charge; if it is a positive ion, *subtract* it.

2. Calculate the **total number of electrons needed** ($N$) for each atom to have its *own* noble-gas set of electrons around it (two for hydrogen, eight for the elements from carbon on in the periodic table).

3. Subtract the number in step 1 from the number in step 2. This is the number of *shared* (or bonding) electrons present ($S$).

4. Assign two *bonding electrons* (as one shared pair) to each connection between two atoms in the molecule or ion.
Drawing Lewis Structures

5. If any of the electrons earmarked for sharing remain, assign them in pairs by making some of the bonds double or triple bonds. In some cases, there may be more than one way to do this. In general, double bonds form only between atoms of carbon, nitrogen, oxygen, and sulfur.

6. Assign the remaining electrons as lone pairs to the atoms, giving octets to all atoms except hydrogen.

7. Determine the formal charge on each atom, and write it next to that atom. Check that the formal charges add to give a correct total charge on the molecule or molecular ion.

   Formal charge = group number - lone-pair electrons - ½ (number of electrons in bonding pairs)
Chapter 4
Types of Chemical Reactions

• 4-1 Dissolution Reactions
  – Chapter 9

• 4-2 Precipitation Reactions
  – Chapter 9

• 4-3 Acids and Bases and Their Reactions
  – Chapter 8

• 4-4 Oxidation-Reduction Reactions
  – Chapter 12 & 13
1. **Dissolution**
   - Ionic Compounds in Water
   - Molecular Compounds in Water
   - Electrolytes and Non-Electrolytes
   - Solubilities (like dissolves like)
   - Predicting Dissolution Reactions

2. **Precipitation Reactions**
   - Ionic Equations and Net Ionic Equations
   - Predicting Precipitation Reactions

3. **Acids and Bases**
   - Arrhenius Acids and Bases Theory
   - Strong and Weak acid
   - Naming Acids
   - Weak Bases
   - Modifying the Arrhenius model
   - Acid-Base Titrations

4. **Oxidation-Reduction Reactions (Redox) Reactions**
   - Oxidizing and Reducing Agents
   - Oxidation Number
   - Types of Redox Reactions
     » Combination and Decomposition
     » Oxygenation
     » Hydrogenation
     » Displacement Reactions
     » Disproportionation
Oxidation-Reduction Reactions

**Redox**: an extensive and important class of reactions that is characterized by the transfer of electrons.

\[
\begin{align*}
2 \text{Mg}(s) + \text{O}_2(g) & \rightarrow 2 \text{MgO}(s) \\
\downarrow \text{loss} & \quad \uparrow \text{gain} \\
2 \times 2e^- & \equiv 1 \times 2 \times 2e^-
\end{align*}
\]

Magnesium is **oxidized**: it gives up electrons as the charge on its atoms increases from zero to +2.

Oxygen is **reduced**: it gains electrons as the charge on its atoms decreases from zero to -2 (i.e., becomes more negative).
Oxidation Numbers (also called oxidation states) are determined for the atoms in covalently bonded compounds by applying the following set of simple rules:

0. In the free elements, each atom has an oxidation number of Zero no matter how complicated the element.

E.g.,
The atoms of Na (s), Fe (s), K (s), Mg (s) all have zero oxidation states
The atoms of H₂, O₂, N₂, Cl₂ all have zero oxidation states
The atoms of S (s), S₈ (s), P₄ (s) all have zero oxidation states.
Oxidation Numbers (also called oxidation states) are determined for the atoms in covalently bonded compounds by applying the following set of simple rules:

1. The oxidation number of the atoms in a neutral molecule must all up to zero; those in an ion must add up to the charge on the ion.

2. Alkali metal (Group I) atoms have oxidation number +1, and alkaline earth (Group II) atoms have oxidation number +2 in their compounds; atoms of Group III elements usually have oxidation number +3 in their compounds.

3. Fluorine always has an oxidation number of -1 in its compounds. The other halogens have oxidation number -1 in their compounds, except in compounds with oxygen and with other halogens, in which they can have positive oxidation numbers.
Oxidation Numbers

1. The oxidation number of the atoms in a neutral molecule must all up to zero; those in an ion must add up to the charge on the ion.

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3. Fluorine always has an oxidation number of -1 in its compounds. The other halogens have oxidation number -1 in their compounds, except in compounds with oxygen and with other halogens, in which they can have positive oxidation numbers.

4. Hydrogen is assigned an oxidation number of +1 in its compounds, except in metal hydrides such as LiH, in which rule 2 take precedence and hydrogen has an oxidation number of -1.

5. Oxygen is assigned an oxidation number of -2 in compounds. There are two exceptions: in compounds with fluorine, rule 3 takes precedence, and in compounds that contain O—O bonds, rules 2 and 4 take precedence. Thus, the oxidation number of oxygen in OF₂ is +2; in peroxides (e.g., H₂O₂ and Na₂O₂), its oxidation number is -1, and in superoxides (e.g., KO₂), its oxidation number is -½.
3-8 Elements Forming more than One Ion

• Oxidation State (Oxidation Number)
  – Ionic vs Covalent bonding
  – **Not** formal electric charges, rather what the charge would be if the compound were ionic
  – Range from -3 to +7

• Examples
  1. CrO\textsubscript{3}
     \[ 0 = (\text{Cr} \times 1) + (\text{O} \times 3) \]
     O is -2 and Cr is +6
  2. TlCl\textsubscript{3}
     Cl is -1 and Tl is +3
  3. Mn\textsubscript{3}N\textsubscript{2}
     N is -3 and Mn is +2
  4. VCl\textsubscript{4}
     Cl is -1 and V is +4
  5. Mn\textsubscript{2}O\textsubscript{7}
     O is -2 Mn is +7
OFB Chapter 5

The Gaseous State

5-1 The Chemistry of Gases
5-2 Pressure and Boyle’s Law
5-3 Temperature and Charles’s Law
5-4 The Ideal Gas Law
5-5 Chemical Calculations for Gases
5-6 Mixtures of Gases
5-7 The Kinetic Theory of Gases
5-8 Real Gases
<table>
<thead>
<tr>
<th>Law</th>
<th>Equation</th>
<th>Notes</th>
</tr>
</thead>
<tbody>
<tr>
<td>Boyle’s Law</td>
<td>$V \propto P^{-1}$</td>
<td>$P_1V_1 = P_2V_2$</td>
</tr>
<tr>
<td></td>
<td></td>
<td>(at a fixed temperature)</td>
</tr>
<tr>
<td>Charles’ Law</td>
<td>$V \propto T$</td>
<td>$V_1 / V_2 = T_1 / T_2$</td>
</tr>
<tr>
<td></td>
<td></td>
<td>(at a fixed pressure)</td>
</tr>
<tr>
<td>Avogadro</td>
<td>$V \propto n$</td>
<td>$n = \text{number of moles}$</td>
</tr>
<tr>
<td></td>
<td></td>
<td>(at a fixed pressure and temperature)</td>
</tr>
</tbody>
</table>

$V \propto T$  $V \propto P^{-1}$  $V \propto n$  $V \propto nTP^{-1}$
The Ideal Gas Law

\[ V \propto nTP^{-1} \]

\[ V = RnTP^{-1} \]

or

\[ PV = nRT \]

ideal gas law
PV = \( nRT \)

ideal gas law

\[
R = 0.082 \text{ L atm mol}^{-1} \text{ K}^{-1}
\]

\[
R = 8.31 \text{ J mol}^{-1} \text{ K}^{-1}
\]
Gas Density and Molar Mass

\[ PV = nRT \]

\[ PV = \frac{m}{M} \cdot RT \]

Rearrange

\[ \frac{m}{V} = \frac{P}{RT} \cdot M \]

\[ \frac{m}{V} = d = \frac{P}{RT} \cdot M \]

\[ d = \frac{P}{RT} \cdot M \]

\[ M = d \cdot \frac{RT}{P} \]
Mixtures of Gases

Dalton’s Law of Partial Pressures

The total pressure of a mixture of gases equals the sum of the partial pressures of the

\[
\begin{align*}
P_A &= \frac{n_A RT}{V} \\
P_B &= \frac{n_B RT}{V} \\
P_A + P_B &= P_{\text{total}} = (n_A + n_B) \frac{RT}{V}
\end{align*}
\]
The Kinetic Theory of Gases

1. A pure gas consists of a large number of identical molecules separated by distances that are large compared with their size.

2. The molecules of a gas are constantly moving in random directions with a distribution of speeds.

3. The molecules of a gas exert no forces on one another except during collisions, so that between collisions they move in straight lines with constant velocities.

4. The collisions of the molecules with each other and with the walls of the container are elastic; no energy is lost during a collision.
OFB Chapter 6: Condensed Phases and Phase Transitions

6-1 Intermolecular Forces: Why Condensed Phases Exist
6-2 The Kinetic Theory of Liquids and Solids
6-3 Phase Equilibrium
6-4 Phase Transitions
6-5 Phase Diagrams
6-6 Colligative Properties of Solutions
6-7 Mixtures and Distillation
6-8 Colloidal Dispersions
Intermolecular Forces: Why Condensed Phases Exist

- **Intramolecular Forces**
  - Chemical bonds
    - Strong
    - Directional
    - Short Range (relative)

- **Intermolecular Forces**
  - Weaker than chemical bonds, usually much weaker
  - Less directional than covalent bonds, more directional than ionic bonds
  - Longer range than covalent bonds but at shorter range than ionic bonds

- **Condensed Phases**
  - Solids and Liquids
  - **Intermolecular forces**: mutual attractions hold the molecules closer together than gases

- **Potential Energy Curves**
  - Distinction between *intra*molecular and *inter*molecular Forces
Types of Non-Bonded (Intermolecular) Attractions

1. Dipole-Dipole Interactions
2. Ion-Dipole Interactions
3. Induced Dipole Attractions
4. Dispersive Forces
Kinetic Theory of Liquids and Solids

- **Intermolecular distances**
  - **Solids** about $3 \times 10^{-10} \text{ m} = 0.3 \text{ nm}$
  - **Liquids** about $5 \times 10^{-10} \text{ m} = 0.5 \text{ nm}$
  - **Gases** about $30 \times 10^{-10} \text{ m} = 3.0 \text{ nm}$
  - **Intramolecular bonds** about 0.05 to 0.25 nm
6-3 Phase Equilibrium

**Phase:** A sample of matter that is uniform throughout, both in its chemical constitution and in its physical state.

- Coexisting phase equilibrium
  - Vapor pressure, temperature dependent
Ideal Solutions and Raoult’s Law

1. **Lowering of Vapor Pressure**
   - Vapor Pressure of a solvent above a dilute solution is always less than the vapor pressure above the pure solvent.

2. **Elevation of Boiling Point** - normal BP is temp at which vapor pressure is 1 atm, so...
   - The boiling point of a solution of a non-volatile solute in a volatile solvent always exceeds the
Elevation of Boiling Point

\[ \Delta T_b = m K_b \]

Where \( m \) = molality

(moles of solute per kilogram of solvent)

The Effect of Dissociation

\[ \Delta T_b = i m K_b \]

\( i \) = the number of particles released into the solution per formula unit of solute

e.g., \( \text{NaCl} \) dissociates into \( i = 2 \)
e.g., \( \text{Na}_2\text{SO}_4 \) dissociates into \( i = 3 \)

\((2 \text{Na}^+ + 1 \text{SO}_4^{-2})\)
OFB Chapter 7
Chemical Equilibrium

7-1 Chemical Reactions in Equilibrium
7-2 Calculating Equilibrium Constants
7-3 The Reaction Quotient
7-4 Calculation of Gas-Phase Equilibrium
7-5 The Effect of External Stresses on Equilibria: Le Châtelier’s Principle
7-6 Heterogeneous Equilibrium
7-7 Extraction and Separation Processes
7-3 The Reaction Quotient

\[ aA + bB \overset{\text{forward}}{\rightleftharpoons} cC + dD \]

\[
\frac{(P_C)^c (P_D)^d}{(P_A)^a (P_B)^b} = K
\]

\[
\frac{(P_C)^c (P_D)^d}{(P_A)^a (P_B)^b} = Q
\]

\[ Q = K \]

Note that **K** (the Equilibrium Constant) uses *equilibrium partial pressures*

Note that **Q** (the reaction quotient) uses *prevailing partial pressures*, not necessarily at equilibrium
The Reaction Quotient

\[ \text{aA + bB} \xrightarrow{\text{forward}} \text{cC + dD} \xleftarrow{\text{reverse}} \]

\[ \frac{(P_C)^c (P_D)^d}{(P_A)^a (P_B)^b} = Q \]

If \( Q < K \), reaction proceeds in a forward direction (toward products);

If \( Q > K \), reaction proceeds in a backward direction (toward reactants);

If \( Q = K \), the reaction stops because the reaction is in equilibrium.
7-5 Effects of External Stresses on Equilibria: Le Châtelier’s Principle

A system in equilibrium that is subjected to a stress reacts in a way that counteracts the stress.

Le Châtelier’s Principle provides a way to predict the response of an equilibrium system to an external perturbation, such as…

1. Effects of Adding or Removing Reactants or Products
2. Effects of Changing the Volume (or Pressure) of the System
3. Effects of Changing the Temperature
If a forward reaction is **exothermic**, 

Then the reverse reaction must be **endothermic**

**Exothermic**: liberation of heat by a reaction

**Endothermic**: absorption of heat by a reaction
Chapter 8
Acid-Base Equilibria

• 8-1 Brønsted-Lowry Acids and Bases
• 8-2 Water and the pH Scale
• 8-3 The Strengths of Acids and Bases
• 8-4 Equilibria Involving Weak Acids and Bases
• 8-5 Buffer Solutions
• 8-6 Skip Acid-Base Titration Curves
• 8-7 Skip Polyprotic Acids
• 8-8 Lewis Acids and Bases
Acid and Base Definitions

1) **Arrhenius** *(Section 4.3)*
   - Acids are H\(^+\) donors
   - Bases are OH\(^-\) donors

2) **Broadened Definition** *(Section 4.3)*
   - Acids are substances that increase [H\(^+\)]
   - Bases are substances that increase [OH\(^-\)]

3) **Brønsted-Lowry** *(Section 8.1)*
   - Acids donate H\(^+\)
   - Bases accept H\(^+\)

4) **Lewis** *(Section 8.8)*
   - Acids are electron pair acceptors
   - Bases are electron pair donators
Autoionization of $\text{H}_2\text{O}$

$\text{H}_2\text{O} + \text{H}_2\text{O} \leftrightarrow \text{H}_3\text{O}^+ + \text{OH}^-$

**Point of View #1**

$\text{H}_2\text{O} + \text{H}_2\text{O} \leftrightarrow \text{H}_3\text{O}^+ + \text{OH}^-$

- Acid: $\text{H}_2\text{O}$
- Base: $\text{H}_2\text{O}$
- Conjugate acid of $\text{H}_2\text{O}$: $\text{H}_3\text{O}^+$
- Conjugate base of $\text{H}_2\text{O}$: $\text{OH}^-$

**Point of View #2**

$\text{H}_2\text{O} + \text{H}_2\text{O} \leftrightarrow \text{H}_3\text{O}^+ + \text{OH}^-$

- Acid: $\text{OH}^-$
- Base: $\text{H}_3\text{O}^+$
- Conjugate acid of $\text{OH}^-$: $\text{H}_2\text{O}$
- Conjugate base of $\text{H}_3\text{O}^+$: $\text{H}_2\text{O}$
The pH Function

\[ \text{pH} = -\log_{10}[\text{H}_3\text{O}^+] \]

- **pH < 7** acidic solution
  - \([\text{H}_3\text{O}^+] > [\text{OH}^-]\)
- **pH = 7** neutral solution
  - \([\text{H}_3\text{O}^+] = [\text{OH}^-]\)
- **pH > 7** basic solution
  - \([\text{H}_3\text{O}^+] < [\text{OH}^-]\)
$$pK_a = - \log_{10} K_a$$

Table 8-2

<table>
<thead>
<tr>
<th>Strong Acids</th>
<th>Weak Acids</th>
<th>Weak Conjugate Bases</th>
<th>Strong Conjugate Bases</th>
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</thead>
<tbody>
<tr>
<td><strong>Hydriodic</strong></td>
<td>HF</td>
<td>$F^-$</td>
<td>$F^-$</td>
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<tr>
<td><strong>Hydrobromic</strong></td>
<td>HBr</td>
<td>$Br^-$</td>
<td>$Br^-$</td>
</tr>
<tr>
<td><strong>Perchloric</strong></td>
<td>HClO₄</td>
<td>$ClO_4^-$</td>
<td>$ClO_4^-$</td>
</tr>
<tr>
<td><strong>Hydrochloric</strong></td>
<td>HCl</td>
<td>$Cl^-$</td>
<td>$Cl^-$</td>
</tr>
<tr>
<td><strong>Chloric</strong></td>
<td>HClO₃</td>
<td>$ClO_3^-$</td>
<td>$ClO_3^-$</td>
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<tr>
<td><strong>Sulfuric (1)</strong></td>
<td>H₂SO₄</td>
<td>$SO_4^{2-}$</td>
<td>$SO_4^{2-}$</td>
</tr>
<tr>
<td><strong>Nitric</strong></td>
<td>HNO₃</td>
<td>$NO_3^-$</td>
<td>$NO_3^-$</td>
</tr>
<tr>
<td><strong>Hydronium ion</strong></td>
<td>H₃O⁺</td>
<td>$H^+$</td>
<td>$H^+$</td>
</tr>
<tr>
<td><strong>Urea acidum ion</strong></td>
<td>(NH₂)₂CO (u)</td>
<td>$CO(NH_2)^-$</td>
<td>$CO(NH_2)^-$</td>
</tr>
<tr>
<td><strong>Iodic</strong></td>
<td>HIRO₄</td>
<td>$IO_4^-$</td>
<td>$IO_4^-$</td>
</tr>
<tr>
<td><strong>Oxalic (1)</strong></td>
<td>H₂C₂O₄</td>
<td>$H_2C_2O_4^-$</td>
<td>$H_2C_2O_4^-$</td>
</tr>
<tr>
<td><strong>Sulfurous (1)</strong></td>
<td>H₂SO₂</td>
<td>$HSO_2^-$</td>
<td>$HSO_2^-$</td>
</tr>
<tr>
<td><strong>Sulfuric (2)</strong></td>
<td>H₂SO₃</td>
<td>$HSO_3^-$</td>
<td>$HSO_3^-$</td>
</tr>
<tr>
<td><strong>Chlorous</strong></td>
<td>HClO₂</td>
<td>$ClO_2^-$</td>
<td>$ClO_2^-$</td>
</tr>
<tr>
<td><strong>Phosphoric (1)</strong></td>
<td>H₃PO₄</td>
<td>$H_3PO_4^-$</td>
<td>$H_3PO_4^-$</td>
</tr>
<tr>
<td><strong>Arsenic (1)</strong></td>
<td>H₃AsO₄</td>
<td>$H_3AsO_4^-$</td>
<td>$H_3AsO_4^-$</td>
</tr>
<tr>
<td><strong>Chloroacetic</strong></td>
<td>CH₂C₂COOH</td>
<td>$CH_2C_2COOH^-$</td>
<td>$CH_2C_2COOH^-$</td>
</tr>
<tr>
<td><strong>Hydrofluoric</strong></td>
<td>HF</td>
<td>$F^-$</td>
<td>$F^-$</td>
</tr>
<tr>
<td><strong>Nitrous</strong></td>
<td>HNO₂</td>
<td>$NO_2^-$</td>
<td>$NO_2^-$</td>
</tr>
<tr>
<td><strong>Formic</strong></td>
<td>HCO₂</td>
<td>$HCOO^-$</td>
<td>$HCOO^-$</td>
</tr>
<tr>
<td><strong>Benzoic</strong></td>
<td>C₆H₅CO₂</td>
<td>$C_6H_5COO^-$</td>
<td>$C_6H_5COO^-$</td>
</tr>
<tr>
<td><strong>Oxalic (2)</strong></td>
<td>H₂C₄O₄</td>
<td>$H_2C_4O_4^-$</td>
<td>$H_2C_4O_4^-$</td>
</tr>
<tr>
<td><strong>Hydrazoic</strong></td>
<td>NH₂</td>
<td>$N_2H_4^-$</td>
<td>$N_2H_4^-$</td>
</tr>
<tr>
<td><strong>Acetatic</strong></td>
<td>CH₃CO₂</td>
<td>$CH_3COO^-$</td>
<td>$CH_3COO^-$</td>
</tr>
<tr>
<td><strong>Propionic</strong></td>
<td>CH₃CH₂CO₂</td>
<td>$CH_3CH_2COO^-$</td>
<td>$CH_3CH_2COO^-$</td>
</tr>
<tr>
<td><strong>Pyridinium ion</strong></td>
<td>H₃C₅N⁺</td>
<td>$C_5H_5N^+$ (pyridine)</td>
<td>$C_5H_5N^+$</td>
</tr>
<tr>
<td><strong>Carbonic (1)</strong></td>
<td>H₂CO₃</td>
<td>$H_2CO_3^-</td>
<td>$H_2CO_3^-$</td>
</tr>
<tr>
<td><strong>Sulfurous (2)</strong></td>
<td>H₂SO₃</td>
<td>$SO_3^{2-}$</td>
<td>$SO_3^{2-}$</td>
</tr>
<tr>
<td><strong>Arsenic (2)</strong></td>
<td>H₃AsO₄</td>
<td>$H_3AsO_4^-$</td>
<td>$H_3AsO_4^-$</td>
</tr>
<tr>
<td><strong>Hydrosulfuric</strong></td>
<td>H₂S</td>
<td>$HS^-$</td>
<td>$HS^-$</td>
</tr>
<tr>
<td><strong>Phosphoric (2)</strong></td>
<td>H₃PO₄</td>
<td>$H_3PO_4^-$</td>
<td>$H_3PO_4^-$</td>
</tr>
<tr>
<td><strong>Hypochlorous</strong></td>
<td>HClO₂</td>
<td>$ClO^-,$</td>
<td>$ClO^-,$</td>
</tr>
<tr>
<td><strong>Hydrocyanic</strong></td>
<td>HCN</td>
<td>$CN^-$</td>
<td>$CN^-$</td>
</tr>
<tr>
<td><strong>Ammonium ion</strong></td>
<td>NH₄⁺</td>
<td>$NH_4^+$</td>
<td>$NH_4^+$</td>
</tr>
<tr>
<td><strong>Carbonic (2)</strong></td>
<td>H₂CO₂</td>
<td>$H_2CO_2^-</td>
<td>$H_2CO_2^-$</td>
</tr>
<tr>
<td><strong>Methylammonium ion</strong></td>
<td>CH₃NH₂⁺</td>
<td>$CH_3NH_2^+$</td>
<td>$CH_3NH_2^+$</td>
</tr>
<tr>
<td><strong>Arsenic (3)</strong></td>
<td>H₃AsO₄</td>
<td>$H_3AsO_4^-$</td>
<td>$H_3AsO_4^-$</td>
</tr>
<tr>
<td><strong>Hydrogen peroxide</strong></td>
<td>H₂O₂</td>
<td>$HO_2^-,$</td>
<td>$HO_2^-,$</td>
</tr>
<tr>
<td><strong>Phosphoric (3)</strong></td>
<td>H₃PO₄</td>
<td>$H_3PO_4^-$</td>
<td>$H_3PO_4^-$</td>
</tr>
<tr>
<td><strong>Water</strong></td>
<td>H₂O</td>
<td>$H_2O^-$</td>
<td>$H_2O^-$</td>
</tr>
</tbody>
</table>

**Known Scale of pKa**

-11 To +14

E.g., Acetone = 19, methane = 48, ethane = 50
Chapter 8
Acid-Base Equilibria

• **Base Strength**
  - strong acids have weak conjugate bases
  - weak acids have strong conjugate bases

The strength of a base is inversely related to the strength of its conjugate acid; the weaker the acid, the stronger its conjugate base, and vice versa

\[
K_a K_b = K_w
\]

\[
pK_a + pK_b = pK_w
\]
How Do Buffers Work?

$\text{HA} + \text{H}_2\text{O} \leftrightarrow \text{H}_3\text{O}^+ + \text{A}^-$

HA = generic acid

$K_a = \frac{[\text{H}_3\text{O}^+][\text{A}^-]}{[\text{HA}]}$

rearranged $K_a$ formula

$[\text{H}_3\text{O}^+] = K_a \frac{[\text{HA}]}{[\text{A}^-]}$

$\text{[H}^+]$ depends on $K_a$ and the ratio of acid to salt or $[\text{A}^-]$.

Thus if both conc. HA and A$^-$ are large then small additions of acid or base don’t change the ratio much.