Chapter 5
Chemical Reactions

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Soluble Ionic Compounds (Table 5.1)

ALL
- ammonium and group 1A (Na⁺, K⁺... and NH₄⁺ salts)
- nitrates (NO₃⁻)
- acetates (CH₃COO⁻)
- chlorates (ClO₃⁻)
- perchlorates (ClO₄⁻)

MOST
- chlorides, bromides and iodides
  - (not AgX, Hg₂X₂, and PbX₂; X = Cl⁻, Br⁻, I⁻)
- sulfates (SO₄²⁻)
  - (not CaSO₄, SrSO₄, BaSO₄, and PbSO₄)

Insoluble Ionic Compounds (Table 5.1)

ALL
- phosphates (PO₄³⁻)
- carbonates (CO₃²⁻)
- oxalates (C₂O₄²⁻)
- oxides (O²⁻)
- sulfides (S²⁻)
  - MgS, CaS & BaS are slightly soluble.
- hydroxides (OH⁻)
  - Sr, Ba & Ca are slightly soluble

Except group 1A, NH₄⁺

Aqueous Solubility of Ionic Compounds

(a) Nitrates (soluble)
- AgNO₃
- Cu(NO₃)₂
- Cu(OH)₂
- AgOH

(b) Hydroxides (insoluble)
- Cu(OH)₂
- AgOH

(c) Sulfides
- CaS, PbS, (NH₄)₂S
- Sb₂S₃

Exchange Reactions: Precipitation

Three types of exchange reactions:
- Form a precipitate
  - an insoluble ionic compound
    - AgNO₃(aq) + KCl(aq) → KNO₃(aq) + AgCl(s)
  - Form a molecular compound
    - Often water
      - HCl(aq) + NaOH(aq) → NaCl(aq) + H₂O(l)
  - Form a molecular gas
    - 2 HCl(aq) + Na₂S(aq) → 2 NaCl(aq) + H₂S(g)
When ionic solutions mix, a precipitate may form:

\[ \text{AgNO}_3(\text{aq}) + \text{NaCl}(\text{aq}) \rightarrow \text{AgCl(s)} + \text{NaNO}_3(\text{aq}) \]

\[ \text{KNO}_3(\text{aq}) + \text{NaCl}(\text{aq}) \rightarrow \text{No reaction} \]

A reaction occurs if a product is insoluble.

The solubility rules help predict reactions.

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Net Ionic Equations

Soluble ionic compounds fully dissociate:

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

\[ \text{KCl}(\text{aq}) \rightarrow \text{K}^{+}(\text{aq}) + \text{Cl}^{-}(\text{aq}) \]

On mixing:

\[ \text{Ag}^{+}(\text{aq}) + \text{NO}_3^{-}(\text{aq}) + \text{K}^{+}(\text{aq}) + \text{Cl}^{-}(\text{aq}) \rightarrow \text{AgCl(s)} + \text{K}^{+}(\text{aq}) + \text{NO}_3^{-}(\text{aq}) \]

\[ \text{K}^{+} \text{ and } \text{NO}_3^{-} \text{ appear on both sides} \]

- \text{NO}_3^{-} \text{ and } \text{K}^{+} \text{ are spectator ions.}
- \text{They are not directly involved in the reaction.}

Net ionic equation: \[ \text{Ag}^{+}(\text{aq}) + \text{Cl}^{-}(\text{aq}) \rightarrow \text{AgCl(s)} \]

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Precipitation Reactions

\[ \text{Na}_2\text{SO}_4(\text{aq}) \text{ and } \text{Ba(NO}_3)_2(\text{aq}) \text{ are mixed. Will they react?} \]

Products? \[ \text{Na}^{+} \rightarrow \text{SO}_4^{2-} \]

\[ \text{Ba}^{2+} \rightarrow \text{NO}_3^{-} \]

\[ \text{NaNO}_3 \text{ and } \text{BaSO}_4. \text{ Insoluble?} \]

Yes. \( \text{BaSO}_4 \) – A reaction occurs:

\[ \text{Na}_2\text{SO}_4(\text{aq}) + \text{Ba(NO}_3)_2(\text{aq}) \rightarrow \text{BaSO}_4(\text{s}) + 2 \text{ NaNO}_3(\text{aq}) \]

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Net Ionic Equations

General rules:

1. Write a balanced equation.
2. Solubility?
3. Dissociate soluble compounds.
4. Write the complete ionic equation.
5. Cancel ions appearing on both sides (spectators).
6. Check the charges are balanced.

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Net Ionic Equations

Aqueous solutions of \( \text{(NH}_4)_2\text{S} \) and \( \text{Hg(NO}_3)_2 \) react to give \( \text{HgS} \) and \( \text{NH}_4\text{NO}_3 \).

(a) Write the overall balanced equation
(b) Name each compound
(c) Write the net ionic equation
(d) What type of reaction is this?

\[ \text{(NH}_4)_2\text{S}(\text{aq}) + \text{Hg(NO}_3)_2(\text{aq}) \rightarrow \text{HgS} + 2 \text{NH}_4\text{NO}_3 \]

Product Solubility

HgS – insoluble
(All sulfides – except group 1A, 2A and \( \text{NH}_4^{+} \) salts)

\( \text{NH}_4\text{NO}_3 \) – soluble
(All ammonium salts; all nitrates)

(b) name each compound

ammonium sulfide mercury(II) sulfide
mercury(II) nitrate ammonium nitrate
(c) Write a net ionic equation

\[(\text{NH}_4)_2\text{S(aq)} + \text{Hg(NO}_3\text{)}_2(\text{aq}) \rightarrow \text{HgS(s)} + 2 \text{NH}_4\text{NO}_3(\text{aq})\]

\[2\text{H}^+(\text{aq}) + \text{S}^2(\text{aq}) + \text{Hg}^{2+}(\text{aq}) + 2\text{NO}_3(\text{aq}) \rightarrow \text{HgS(s)} + 2\text{NH}_4^+(\text{aq}) + 2\text{NO}_3(\text{aq})\]

\[\text{S}^2(\text{aq}) + \text{Hg}^{2+}(\text{aq}) \rightarrow \text{HgS(s)}\]

(d) Type of reaction?
- Exchange reaction
- Precipitation reaction

**Acids**

Increase the concentration of H\(^+\) ions in water.
- Protons (H\(^+\)) always combine with water to form H\(_3\)O\(^+\) (hydronium ion).
- Sour tasting.
- Change the color of pigments (indicators)
  § Litmus, phenolphthalein...

**Strong acids** dissociate (>99%) in water (strong electrolytes).

**Weak acids** partially dissociate (weak electrolytes).

**Bases**

Increase the concentration of OH\(^-\) (hydroxide ion) in water. Bases:
- Counteract an acid (neutralize an acid).
- Change an indicator’s color (phenolphthalein…).
- Have a bitter taste.
- Feel slippery.

Bases can be “strong” or “weak”.
- \(\text{NaOH(s)} \xrightarrow{\text{H}_2\text{O(ℓ)}} \text{Na}^+(\text{aq}) + \text{OH}^-\text{(aq)}\) strong
- \(\text{NH}_3(\text{aq}) + \text{H}_2\text{O(ℓ)} \xrightarrow{\text{H}_3\text{O}^+\text{(aq)}} \text{NH}_4^+(\text{aq}) + \text{OH}^-\text{(aq)}\) weak

**Common Acids and Bases**

<table>
<thead>
<tr>
<th>Strong Acid</th>
<th>Strong Base</th>
</tr>
</thead>
<tbody>
<tr>
<td>HCl</td>
<td>LiOH</td>
</tr>
<tr>
<td>HBr</td>
<td>NaOH</td>
</tr>
<tr>
<td>HI</td>
<td>KOH</td>
</tr>
<tr>
<td>HNO(_3)</td>
<td>Ca(OH)(_2)</td>
</tr>
<tr>
<td>H(_2)SO(_4)</td>
<td>Ba(OH)(_2)</td>
</tr>
<tr>
<td>HClO(_3)</td>
<td>Sr(OH)(_2)</td>
</tr>
<tr>
<td>HF</td>
<td>Weak Acid</td>
</tr>
<tr>
<td>H(_2)PO(_4)</td>
<td>NH(_3)</td>
</tr>
<tr>
<td>CH(_3)COOH</td>
<td>CH(_3)NH(_2)</td>
</tr>
</tbody>
</table>

<table>
<thead>
<tr>
<th>Weak Acid</th>
<th>Weak Base</th>
</tr>
</thead>
<tbody>
<tr>
<td>Hydrofluoric acid</td>
<td>Ammonia</td>
</tr>
<tr>
<td>Phosphoric acid</td>
<td>Methylamine</td>
</tr>
</tbody>
</table>

**Neutralization Reactions**

acid + base → salt + water.

\[\text{salt} = \text{ionic compound made from an acid anion and base cation.}\]

- \(\text{HX(aq)} + \text{MOH(aq)} \rightarrow \text{MX(aq)} + \text{H}_2\text{O(ℓ)}\)
- \(\text{HBr(aq)} + \text{KOH(aq)} \rightarrow \text{KBr(aq)} + \text{H}_2\text{O(ℓ)}\)
- \(\text{H}_3\text{PO}_4(\text{aq}) + 3 \text{NaOH(aq)} \rightarrow \text{Na}_3\text{PO}_4(\text{aq}) + 3 \text{H}_2\text{O(ℓ)}\)

Neutralizations are exchange reactions.
Net Ionic Equations for Acid-Base Reactions

**Strong Acid + Strong Base**
Overall: \( HX + MOH \rightarrow MX + H_2O \)

**Full ionic:**
\( H^+(aq) + X^-(aq) + M^+(aq) + OH^-(aq) \rightarrow M^+(aq) + X^-(aq) + H_2O(l) \)

**net ionic:**
\( H^+(aq) + OH^-(aq) \rightarrow H_2O(l) \)

**Weak Acid + Strong Base**
Similar: \( HA + MOH \rightarrow MA + H_2O \)

but the weak-acid remains undissociated:

**Full ionic**
\( HA(aq) + M^+(aq) + OH^-(aq) \rightarrow M^+(aq) + A^-(aq) + H_2O(l) \)

**Net ionic:**
\( HA(aq) + OH^-(aq) \rightarrow A^-(aq) + H_2O(l) \)

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Gas-Forming Exchange Reactions

**Metal carbonate/acid exchange**
Overall:
\( CaCO_3(s) + 2 HCl(aq) \rightarrow CaCl_2(aq) + H_2CO_3(aq) \)

\( H_2CO_3(aq) \rightarrow H_2O(l) + CO_2(g) \)

often written:
\( CaCO_3(s) + 2 HCl(aq) \rightarrow CaCl_2(aq) + H_2O(l) + CO_2(g) \)

**Net ionic:**
\( CaCO_3(s) + 2 H^+(aq) \rightarrow Ca^{2+}(aq) + H_2O(l) + CO_2(g) \)

**Gas-Forming Exchange Reactions**

**Metal hydrogen carbonate/acid exchange**
Overall:
\( NaHCO_3(aq) + HCl(aq) \rightarrow NaCl(aq) + H_2O(l) + CO_2(g) \)

**Net ionic**
\( HCO_3^-(aq) + H^+(aq) \rightarrow H_2O(l) + CO_2(g) \)

**Gas-Forming Exchange Reactions**

**Metal sulfite/acid exchange**
Similar. Overall:
\( Na_2SO_3(aq) + 2 HCl(aq) \rightarrow 2NaCl(aq) + H_2SO_3(aq) \)

\( H_2SO_3(aq) \rightarrow H_2O(l) + SO_2(g) \)

or
\( Na_2SO_3(aq) + 2 HCl(aq) \rightarrow 2NaCl(aq) + H_2O(l) + SO_2(g) \)

**net ionic:**
\( SO_3^{2-}(aq) + 2 H^+(aq) \rightarrow H_2O(l) + SO_2(g) \)

**Gas-Forming Exchange Reactions**

**Metal sulfide/acid exchange**
Overall:
\( Na_2S(aq) + 2 HCl(aq) \rightarrow 2NaCl(aq) + H_2S(g) \)

**net ionic:**
\( S^{2-}(aq) + 2 H^+(aq) \rightarrow H_2S(g) \)
Oxidation-Reduction Reactions

Oxidation
Originally: add oxygen.
\[ 2 \text{Cu(s)} + \text{O}_2(g) \rightarrow 2 \text{CuO(s)} \]
\[ 2 \text{CO(g)} + \text{O}_2(g) \rightarrow 2 \text{CO}_2(g) \]
Cu and CO are oxidized. \( \text{O}_2 \) is the oxidizing agent for both.

Reduction
Originally: reduce ore to metal; reverse of oxidation.
\[ \text{CuO(s)} + \text{H}_2(g) \rightarrow \text{Cu(s)} + \text{H}_2\text{O} \]
\[ \text{SnO}_2(s) + 2 \text{C(s)} \rightarrow \text{Sn(s)} + 2 \text{CO}(g) \]
CuO and SnO\(_2\) are reduced. H\(_2\) and C are the reducing agents.

Redox Reactions and Electron Transfer

Loss of electrons is oxidation

Gain of electrons is reduction

Oxidation is loss
Reduction is gain

Oil rig

Common Oxidizing and Reducing Agents

<table>
<thead>
<tr>
<th>Oxidizing Agent</th>
<th>Reaction Product</th>
</tr>
</thead>
<tbody>
<tr>
<td>( \text{O}_2 ) (oxygen)</td>
<td>( \text{O}^2- ) (oxide ion)</td>
</tr>
<tr>
<td>( \text{H}_2\text{O}_2 ) (hydrogen peroxide)</td>
<td>( \text{H}_2\text{O}(l) )</td>
</tr>
<tr>
<td>( \text{F}_2, \text{Cl}_2, \text{Br}_2, \text{I}_2 ) (halogen)</td>
<td>( \text{F}^-, \text{Cl}^-, \text{Br}^-, \text{I}^- )</td>
</tr>
<tr>
<td>( \text{HNO}_3 ) (nitric acid)</td>
<td>( \text{NO}, \text{NO}_2^- )</td>
</tr>
<tr>
<td>( \text{Cr}_2\text{O}_7^{2-} ) (dichromate ion)</td>
<td>( \text{Cr}^{3+} ) (chromium(III) ion)</td>
</tr>
<tr>
<td>( \text{MnO}_4^- ) (permanganate ion)</td>
<td>( \text{Mn}^{2+} ) (manganese(II) ion)</td>
</tr>
</tbody>
</table>

<table>
<thead>
<tr>
<th>Reducing Agent</th>
<th>Reaction Product</th>
</tr>
</thead>
<tbody>
<tr>
<td>( \text{H}_2 ) (hydrogen)</td>
<td>( \text{H}^+ ) or ( \text{H}_2\text{O} )</td>
</tr>
<tr>
<td>C</td>
<td>( \text{CO} ) and ( \text{CO}_2 )</td>
</tr>
<tr>
<td>M (metal: Na, K, Fe…….)</td>
<td>( M^{n+} ) (Na(^+), K(^+), Fe(^{3+})…….)</td>
</tr>
</tbody>
</table>

Oxidation-Reduction Reactions

In all cases:
- If something is oxidized, something must be reduced.
- Oxidation - reduction = redox.
- Redox reactions move e\(^-\).

\[ 2 \text{Ag}^+(aq) + \text{Cu(s)} \rightarrow 2 \text{Ag(s)} + \text{Cu}^{2+}(aq) \]

Here:
- Cu changes to \( \text{Cu}^{2+} \).
- Cu loses 2 e\(^-\); each Ag\(^+\) gains one e\(^-\).
- Ag\(^+\) is reduced (ore turned to metal).
- Gain of e\(^-\) = reduction (so, loss of e\(^-\) = oxidation).

Redox Reactions and Electron Transfer

<table>
<thead>
<tr>
<th>Loss of electrons is oxidation</th>
<th>Gain of electrons is reduction</th>
</tr>
</thead>
<tbody>
<tr>
<td>Leo says ger</td>
<td></td>
</tr>
</tbody>
</table>

Oxidation is loss
Reduction is gain

Oxidation Numbers & Redox Reactions

Oxidation number
Comparing the charge of an uncombined atom with its actual or relative charge in a compound.

General rules:
- Pure element = 0.
- Monatomic ion = charge of ion.
- Some elements have the same oxidation number in almost all their compounds.
- The sum of the oxidation numbers of all atoms in any species = the charge of the species.
### Oxidation Numbers & Redox Reactions

<table>
<thead>
<tr>
<th>Element</th>
<th>Ox#</th>
<th>Exceptions</th>
</tr>
</thead>
<tbody>
<tr>
<td>F</td>
<td>−1</td>
<td>None</td>
</tr>
<tr>
<td>O</td>
<td>−2</td>
<td>Metal peroxides (MO₂) = -1, Halogen oxides (OXₙ) (varies)</td>
</tr>
<tr>
<td>Cl, Br, I</td>
<td>−1</td>
<td>Interhalogens (strongest = -1)</td>
</tr>
<tr>
<td>H</td>
<td>+1</td>
<td>Metal hydrides (e.g. NaH) = -1</td>
</tr>
</tbody>
</table>

### Oxidation Numbers & Redox Reactions

**Find the oxidation number for all elements in SO₃²⁻**

- O = -2 (not a peroxide, not bonded to halogen)

Sulfur?

- Sum of the oxidation numbers = charge = -2
  - (ox. no. for S) + 3(ox. no. for O) = -2
  - (ox. no. for S) + 3(-2) = -2
  - (ox. no. for S) = 6 = -2
  - Sulfur = +4

### Oxidation Numbers & Redox Reactions

**Compound**

<table>
<thead>
<tr>
<th>Compound</th>
<th>Known</th>
<th>Unknown</th>
</tr>
</thead>
<tbody>
<tr>
<td>SO₂</td>
<td>O = -2</td>
<td>S = +4</td>
</tr>
<tr>
<td>SO₄²⁻</td>
<td>O = -2</td>
<td>S = +6</td>
</tr>
<tr>
<td>NH₄⁺</td>
<td>H = +1</td>
<td>N = -3</td>
</tr>
<tr>
<td>NO₂⁻</td>
<td>O = -2</td>
<td>N = +3</td>
</tr>
<tr>
<td>NO₃⁻</td>
<td>O = -2</td>
<td>N = +5</td>
</tr>
<tr>
<td>OF₂</td>
<td>F = -1</td>
<td>O = +2</td>
</tr>
<tr>
<td>ClF₅</td>
<td>F = -1</td>
<td>Cl = +5</td>
</tr>
<tr>
<td>KMnO₄</td>
<td>K = +1</td>
<td>Mn = +7</td>
</tr>
<tr>
<td>H₂O₂</td>
<td>H = +1</td>
<td>O = -1</td>
</tr>
</tbody>
</table>

### Oxidation Numbers & Redox Reactions

**A more complex example:**

$$\text{Cu(s)} + 4 \text{H}^+(aq) + 2 \text{NO}_3^-(aq) \rightarrow \text{Cu}^{2+}(aq) + 2 \text{NO}_2(g) + 2 \text{H}_2\text{O}(l)$$

- Cu is oxidized (ox. no. ↑; loss of e⁻).
- H is unchanged.
- O is unchanged.
- N is reduced (ox. no. ↓; gain of e⁻).

### Oxidation Numbers & Redox Reactions

**Exchange reactions are not redox**

- no change in oxidation state occurs.

**e.g.**

$$\text{AgNO}_3(aq) + \text{KCl}(aq) \rightarrow \text{KNO}_3(aq) + \text{AgCl}(s)$$

- NO₃⁻: N is in a +5, and O is in a -2 ox. state.
Redox:

\[
\text{Fe(s) + CuSO}_4(\text{aq}) \rightarrow \text{FeSO}_4(\text{aq}) + \text{Cu(s)}
\]

\[
\begin{align*}
0 & +2 & +2 & 0 \\
+1 & +2 & 0
\end{align*}
\]

Not all metals can displace another from its salts:

\[
\text{Cu(s) + ZnSO}_4(\text{aq}) \rightarrow \text{no reaction}
\]

An activity series was developed…
Molarity

\[ \text{Molarity} = \frac{\text{moles solute}}{\text{liters of solution}} = \frac{\text{mol}}{\text{L}} \]

- \( V \) of solution not solvent.
- Shorthand: \([\text{NaOH}] = 1.00 \text{ M}\)

Brackets [ ] represent “molarity of”

Molarity

Calculate the molarity of sodium sulfate in a solution that contains 36.0 g of \( \text{Na}_2\text{SO}_4 \) in 750.0 mL of solution.

\[
[\text{Na}_2\text{SO}_4] = \frac{36.0 \text{ g}}{142.0 \text{ g/mol}} = 0.2534 \text{ mol} \\
[\text{Na}_2\text{SO}_4] = \frac{0.2534 \text{ mol}}{0.7500 \text{ L}} \quad \text{(Unit change!)} \\
[\text{Na}_2\text{SO}_4] = 0.338 \text{ mol/L} = 0.338 \text{ M}
\]

Molarity

6.37 g of \( \text{Al(NO}_3\text{)}_3 \) are dissolved to make a 250. mL aqueous solution. Calculate (a) \([\text{Al(NO}_3\text{)}_3]\) (b) \([\text{Al}^{3+}]\) and \([\text{NO}_3^-]\).

(a) \( \text{Al(NO}_3\text{)}_3 \)

\[
\frac{\text{FM}}{\text{g/mol}} = 26.98 + 3(14.00) + 9(16.00) = 213.0 \text{ g/mol} \\
\quad n_{\text{Al(NO}_3\text{)}_3} = \frac{6.37 \text{ g}}{213.0 \text{ g/mol}} = 2.991 \times 10^{-2} \text{ mol} \\
\quad [\text{Al(NO}_3\text{)}_3] = \frac{2.991 \times 10^{-2} \text{ mol}}{0.250 \text{ L}} = 0.120 \text{ M}
\]

(b) Molarity of \( \text{Al}^{3+}, \text{NO}_3^-? \)

\[
\quad 1 \text{ Al(NO}_3\text{)}_3(\text{aq}) \rightarrow \text{Al}^{3+}(\text{aq}) + 3 \text{ NO}_3^-(\text{aq}) \\
\quad 1 \text{ Al(NO}_3\text{)}_3 \equiv \text{1 Al}^{3+} \quad \text{1 Al(NO}_3\text{)}_3 \equiv 3 \text{ NO}_3^- \\
\quad [\text{Al}^{3+}] = \frac{0.120 \text{ M Al(NO}_3\text{)}_3}{1 \text{ Al(NO}_3\text{)}_3} = 0.120 \text{ M Al}^{3+} \\
\quad [\text{NO}_3^-] = \frac{0.120 \text{ M Al(NO}_3\text{)}_3}{3 \text{ NO}_3^-} = 0.360 \text{ M NO}_3^-
\]

Molarity

6.37 g of \( \text{Al(NO}_3\text{)}_3 \) in a 250. mL aqueous solution. Calculate (a) the molarity of the \( \text{Al(NO}_3\text{)}_3 \), (b) the molar concentration of \( \text{Al}^{3+} \) and \( \text{NO}_3^- \) ions in solution.

(a) \( \text{Al(NO}_3\text{)}_3 \)

\[
\quad n_{\text{Al(NO}_3\text{)}_3} = \frac{6.37 \text{ g}}{213.0 \text{ g/mol}} = 0.0299 \text{ mol} \\
\quad [\text{Al(NO}_3\text{)}_3] = \frac{0.0299 \text{ mol}}{0.250 \text{ L}} = 0.120 \text{ M}
\]

(b) Molarity of \( \text{Al}^{3+}, \text{NO}_3^-? \)

\[
\quad \text{Al(NO}_3\text{)}_3(\text{aq}) \rightarrow \text{Al}^{3+}(\text{aq}) + 3 \text{ NO}_3^-(\text{aq}) \\
\quad 1 \text{ Al(NO}_3\text{)}_3 \equiv 1 \text{ Al}^{3+} \quad \text{1 Al(NO}_3\text{)}_3 \equiv 3 \text{ NO}_3^- \\
\quad [\text{Al}^{3+}] = \frac{0.120 \text{ M Al(NO}_3\text{)}_3}{1 \text{ Al(NO}_3\text{)}_3} = 0.120 \text{ M Al}^{3+} \\
\quad [\text{NO}_3^-] = \frac{0.120 \text{ M Al(NO}_3\text{)}_3}{3 \text{ NO}_3^-} = 0.360 \text{ M NO}_3^-
\]

Solution Preparation

Solutions are prepared either by:

1. Diluting a more concentrated solution.
2. Dissolving a measured amount of solute and diluting to a fixed volume.

Solution Preparation by Dilution

\[
M_{\text{conc}}V_{\text{conc}} = M_{\text{dil}}V_{\text{dil}}
\]

Example

Commercial concentrated sulfuric acid is 17.8 M. If 75.0 mL of this acid is diluted to 1.00 L, what is the final concentration of the acid?

\[
M_{\text{conc}} = 17.8 \text{ M} \\
V_{\text{conc}} = 75.0 \text{ mL} \\
M_{\text{dil}} = ? \\
V_{\text{dil}} = 1000. \text{ mL}
\]

\[
M_{\text{dil}} = \frac{M_{\text{conc}}V_{\text{conc}}}{V_{\text{dil}}} = \frac{17.8 \text{ M} \times 75.0 \text{ mL}}{1000. \text{ mL}} = 1.34 \text{ M}
\]
Solution Preparation from Pure Solute

Prepare a 0.5000 M solution of potassium permanganate in a 250.0 mL volumetric flask.

Mass of KMnO₄ required

\[ n_{\text{KMnO}_4} = [\text{KMnO}_4] \times V \]
\[ = 0.5000 \text{ M} \times 0.2500 \text{ L} \quad (M \equiv \text{mol/L}) \]
\[ = 0.1250 \text{ mol KMnO}_4 \]

Mass of KMnO₄ = 0.1250 mol x 158.03 g/mol
\[ = 19.75 \text{ g} \]

Solution Preparation from Pure Solute

• Weigh exactly 19.75 g of pure KMnO₄
• Transfer it to a volumetric flask.

• Rinse all the solid from the weighing dish into the flask.
• Fill the flask \( \approx \frac{1}{3} \) full.
• Swirl to dissolve the solid.
• Fill the flask to the mark on the neck.
• Shake to thoroughly mix.

Molarity and Reactions in Aqueous Solution

\[ n_A = [A] \times V \]

[product] = \( n_{\text{product}} / \) (total volume).

Molarity and Reactions in Aqueous Solution

What volume, in mL, of 0.0875 M H₂SO₄ is required to neutralize 25.0 mL of 0.234 M NaOH?

\[ \text{H}_2\text{SO}_4(\text{aq}) + 2 \text{NaOH}(\text{aq}) \rightarrow \text{Na}_2\text{SO}_4(\text{aq}) + 2 \text{H}_2\text{O}(\ell) \]

\[ n_{\text{NaOH}} = 0.0250 \text{ L} \times 0.234 \text{ mol/L} = 5.850 \times 10^{-3} \text{ mol} \]

\[ 2 \text{NaOH} \equiv 1 \text{H}_2\text{SO}_4 \]
\[ n_{\text{H}_2\text{SO}_4} = 5.850 \times 10^{-3} \text{ mol NaOH} \times \frac{1 \text{ H}_2\text{SO}_4}{2 \text{NaOH}} \]
\[ = 2.925 \times 10^{-3} \text{ mol} \]

Molarity and Reactions in Aqueous Solution

A 4.554 g mixture of oxalic acid, H₂C₂O₄ and NaCl was neutralized by 29.58 mL of 0.550M NaOH. What was the weight % of oxalic acid in the mixture?

\[ \text{H}_2\text{C}_2\text{O}_4(\text{aq}) + 2 \text{NaOH(\text{aq})} \rightarrow \text{Na}_2\text{C}_2\text{O}_4(\text{aq}) + 2 \text{H}_2\text{O}(\ell) \]

\[ n_{\text{NaOH}} = 0.02958 \text{ L} \times 0.550 \text{ mol/L} = 0.01627 \text{ mol} \]

\[ 1 \text{H}_2\text{C}_2\text{O}_4 \equiv 2 \text{NaOH} \]

Oxalic acid reacted = \[ \frac{1 \text{ H}_2\text{C}_2\text{O}_4}{2 \text{NaOH}} = 8.135 \times 10^{-3} \text{ mol} \]
Molarity and Reactions in Aqueous Solution

A 4.554 g H₂C₂O₄/NaCl mixture ... Wt % of oxalic acid in the mixture?

Mass of acid consumed, \( m_{\text{acid}} \)

\[
= 8.135 \times 10^{-3} \text{ mol} \times (90.04 \text{ g/mol acid}) \\
= 0.7324 \text{ g}
\]

Weight % = \( \frac{m_{\text{acid}}}{\text{sample mass}} \times 100\% \)

Weight % = \( \frac{0.7324 \text{ g}}{4.554 \text{ g}} \times 100\% = 16.08\% \)

Molarity and Reactions in Aqueous Solution

25.0 mL of 0.234 M FeCl₃ and 50.0 mL of 0.453 M NaOH are mixed. Which reactant is limiting? How many moles of Fe(OH)₃ will form?

\[
\text{FeCl}_3(\text{aq}) + 3 \text{NaOH}(\text{aq}) \rightarrow 3 \text{NaCl (aq)} + \text{Fe(OH)}_3(\text{s})
\]

\[
\begin{align*}
\text{n}_{\text{FeCl}_3} &= 0.0250 \text{ L} \times 0.234 \text{ mol/L} = 0.005850 \text{ mol} \\
\text{n}_{\text{NaOH}} &= 0.0500 \text{ L} \times 0.453 \text{ mol/L} = 0.02265 \text{ mol}
\end{align*}
\]

FeCl₃ is limiting; 0.00585 mol Fe(OH)₃ produced.

Molarity and Reactions in Aqueous Solution

FeCl₃(aq) + 3 NaOH(aq) → 3 NaCl (aq) + Fe(OH)₃(s)

\[
0.00585 \text{ mol FeCl}_3 \rightarrow 0.00585 \text{ mol Fe(OH)}_3 \\
0.02265 \text{ mol NaOH} \rightarrow 0.00755 \text{ mol Fe(OH)}_3
\]

FeCl₃ is limiting; 0.00585 mol Fe(OH)₃ produced.

Aqueous Solution Titrations

**Titration** = volume-based method used to determine an unknown concentration.

A *standard* solution (known concentration) is added to a solution of unknown concentration.

- Monitor the volume added.
- Add until *equivalence* is reached — stoichiometrically equal moles of reactants added.
- An indicator monitors the end point.

Often used to determine acid or base concentrations.

Aqueous Solution Titrations

**Buret** = volumetric glassware used for titrations.

Slowly add standard solution.

End point: indicator changes color.

Determine \( V_{\text{t} \text{ritrant added}} \).