Solution Concentration: Molarity

**Molarity**: The number of moles of a substance dissolved in each liter of solution

**Solution**: A homogeneous mixture

**Solute**: The dissolved substance in a solution (substance you have less quantity of in solution)

**Solvent**: The major component in a solution (substance you have more of in solution)
Solution Concentration: Molarity

A measured number of moles of solute is placed in a volumetric flask.

Enough solvent is added to dissolve the solute by swirling.

Further solvent is added to reach the calibration mark on the neck of the flask, and the solution is mixed until uniform.
Definition of Molarity

Molarity ($M$)

- Commonly used expression for concentration
- Defined as **moles of solute per volume of solution in liters**

\[
M = \text{molarity} = \frac{\text{moles of solute}}{\text{liters of solution}}
\]

Also, \[
M = \frac{\text{moles of solute}}{1000 \text{ ml of solution}}
\]

Use if given volume of solution in mL
Definition of Molarity

\[ M = \text{molarity} = \frac{\text{moles of solute}}{\text{liters of solution}} \]

Example: What is the molarity if you put 1.00 mol of sodium chloride in enough water to make 1.00 L of solution?  

\[
\frac{1.00 \text{ mol}}{1.00 \text{ L}} = 1.00 \frac{\text{mol}}{\text{L}} \quad \text{or} \quad 1.00 \text{ M}
\]

End class D section 9/30
Solution Concentration: Molarity

- Calculate the molarity of a solution prepared by dissolving 1.56 g of gaseous HCl \( (36.46 \text{ g/mol}) \) in enough water to make 26.8 mL of solution

- Doc camera
Solution Concentration: Molarity

- Where are we going?
  - To find the molarity of HCl solution
- What do we know?
  - 1.56 g HCl
  - 26.8 mL solution
- What information do we need to find molarity?
  - Moles solute

\[
\text{Molarity} = \frac{\text{mol solute}}{\text{L solution}}
\]
Solution Concentration: Molarity

- How do we get there?
  - What are the moles of HCl (36.46 g/mol)?
    \[1.56 \text{ g HCl} \times \frac{1 \text{ mol HCl}}{36.46 \text{ g HCl}} = 4.28 \times 10^{-2} \text{ mol HCl}\]
  - What is the volume of solution (in liters)?
    \[26.8 \text{ mL} \times \frac{1 \text{ L}}{1000 \text{ mL}} = 2.68 \times 10^{-2} \text{ L}\]
Solution Concentration: Molarity

- What is the molarity of the solution?

  \[
  \text{Molarity} = \frac{4.28 \times 10^{-2} \ \text{mol HCl}}{2.68 \times 10^{-2} \ \text{L solution}} = 1.60 \ \text{M} \ \text{HCl}
  \]

- Reality check
  - The units are correct for molarity
Determining Moles of Solute in a Sample

- Use the definition of molarity as a conversion factor (to get moles of solute)

\[
\text{Liters of solution} \times \text{molarity} = \frac{\text{moles of solute}}{\text{liters of solution}} \times \text{liters of solution}
\]

\[
\text{Moles of solute} = \text{Liters of solution} \times \text{molarity}
\]
Solution Concentration: Molarity

- Calculate the number of moles of Cl\(^{-}\) ions in 1.75 L of 0.152 \(M\) \(NaCl\) (\(M = \text{moles} / \text{liter solution}\))

- On doc camera
HW: Solution Concentration: Molarity

How many grams of solute would you use to prepare 1.50 L of 0.250 M glucose, C$_6$H$_{12}$O$_6$?

Molar mass C$_6$H$_{12}$O$_6$ = 180.0 g/mol
Solution Concentration: Molarity

How many grams of solute would you use to prepare 1.50 L of 0.250 M glucose, C_6H_{12}O_6?

Molar mass C_6H_{12}O_6 = 180.0 g/mol

\[
1.50 \text{ L} \times \frac{0.250 \text{ mol}}{1 \text{ L}} \times \frac{180.0 \text{ g}}{1 \text{ mol}} = 67.5 \text{ g}
\]
Dilution

- Process of adding water to a concentrated (stock) solution

\[ M_i V_i = \text{moles initial} , M_f V_f = \text{moles final} \]

- dilution equation: \[ M_i V_i = M_f V_f \]

- as long as keep volume unit the same between initial and final, can have both mL or both L
Diluting Concentrated Solutions

The **volume** to be diluted is placed in an empty volumetric flask.

Solvent is added to a level just below the calibration mark, and the flask is shaken.

More solvent is added to reach the calibration mark, and the flask is again shaken.
Diluting Concentrated Solutions

How would you prepare 250.0 mL of 0.500 M from 18.0 M aqueous H$_2$SO$_4$?

\[ V_i = \frac{M_f \times V_f}{M_i} = \frac{0.500 \text{ M}}{18.0 \text{ M}} \times 250.0 \text{ mL} = 6.94 \text{ mL} \]

Add 6.94 mL 18.0 M sulfuric acid to enough water to make 250.0 mL of 0.500 M solution.
Describe how you would prepare 2.00 L of each of the following solutions:  (on doc camera)

a. 0.250 M NaOH from 1.00 M NaOH stock solution

b. 0.100 M K₂CrO₄ from 1.75 M K₂CrO₄ stock solution
HW: Diluting Concentrated Solutions [do (a) only for HW]

End 9/30 F section

- Describe how you would prepare 2.00 L of each of the following solutions: (on doc camera)
  a. 0.250 \( M \) NaOH from 1.00 \( M \) NaOH stock solution

  Add 500. mL (= 0.500 L) of the 1.00 \( M \) NaOH stock solution to a 2-L volumetric flask; fill to the mark with water

  b. 0.100 \( M \) \( K_2CrO_4 \) from 1.75 \( M \) \( K_2CrO_4 \) stock solution

  Add 114 mL (= 0.114 L) of the 1.75 \( M \) \( K_2CrO_4 \) stock solution to a 2-L volumetric flask; fill to the mark with water
**Electrolytes in Aqueous Solution**

**Electrolytes**: dissociate in water to produce conducting solutions of ions

\[ \text{H}_2\text{O} \quad \text{NaCl(s)} \rightarrow \text{Na}^+(aq) + \text{Cl}^-(aq) \]

**Nonelectrolytes**: do not dissociate to produce ions in water (covalent compounds)

\[ \text{H}_2\text{O} \quad \text{C}_{12}\text{H}_{22}\text{O}_{11}(s) \rightarrow \text{C}_{12}\text{H}_{22}\text{O}_{11}(aq) \]
Electrolytes in Aqueous Solution

A solution of NaCl conducts electricity because of the movement of charged particles (ions), thereby completing the circuit and allowing the bulb to light.

A solution of sucrose does not conduct electricity or complete the circuit because it contains no mobile charged particles. The bulb therefore remains dark.
Electrolytes in Aqueous Solution

**Strong Electrolytes:** completely dissociates in water (ionic compounds, strong acids, strong bases)

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

**Weak Electrolytes:** incompletely dissociates in water (weak acids, weak bases)

\[
\text{CH}_3\text{CO}_2\text{H}(aq) \rightleftharpoons \text{H}^+(aq) + \text{CH}_3\text{CO}_2^- (aq)
\]
Electrolytes in Aqueous Solution

### TABLE 4.1 Electrolyte Classification of Some Common Substances

<table>
<thead>
<tr>
<th>Strong Electrolytes</th>
<th>Weak Electrolytes</th>
<th>Nonelectrolytes</th>
</tr>
</thead>
<tbody>
<tr>
<td>HCl, HBr, HI</td>
<td>CH₃CO₂H</td>
<td>H₂O</td>
</tr>
<tr>
<td>HClO₄</td>
<td>HF</td>
<td>CH₃OH (methyl alcohol)</td>
</tr>
<tr>
<td>HNO₃</td>
<td>HCN</td>
<td>C₂H₅OH (ethyl alcohol)</td>
</tr>
<tr>
<td>H₂SO₄</td>
<td></td>
<td>C₁₂H₂₂O₁₁ (sucrose)</td>
</tr>
<tr>
<td>KBr</td>
<td></td>
<td>Most compounds of carbon (organic compounds)</td>
</tr>
<tr>
<td>NaCl</td>
<td></td>
<td></td>
</tr>
<tr>
<td>NaOH, KOH</td>
<td></td>
<td></td>
</tr>
<tr>
<td>Other soluble ionic compounds</td>
<td></td>
<td></td>
</tr>
</tbody>
</table>

**Strong Acids**: Hydrochloric acid, hydrobromic acid, hydroiodic acid, perchloric acid, nitric acid, sulfuric acid
# Electrolytes in Aqueous Solution

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**Ionic compounds**
# Electrolytes in Aqueous Solution

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</tr>
<tr>
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<td>HF</td>
<td>CH$_3$OH (methyl alcohol)</td>
</tr>
<tr>
<td>HNO$_3$</td>
<td>HCN</td>
<td>C$_2$H$_5$OH (ethyl alcohol)</td>
</tr>
<tr>
<td>H$_2$SO$_4$</td>
<td></td>
<td>C$<em>{12}$H$</em>{22}$O$_{11}$ (sucrose)</td>
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</table>

**Weak acids**
### Electrolytes in Aqueous Solution

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<td>NaOH, KOH</td>
<td>Other soluble ionic compounds</td>
<td></td>
</tr>
</tbody>
</table>

**Molecular (covalent) compounds**

End class 9/30 Monday G section
Types of Chemical Reactions in Aqueous Solution

Precipitation Reactions: soluble reactants give insoluble solid product (precipitates out)

\[ \text{Pb(NO}_3\text{)}_2(aq) + 2\text{KI}(aq) \rightarrow 2\text{KNO}_3(aq) + \text{PbI}_2(s) \]
Types of Chemical Reactions in Aqueous Solution

Oxidation–Reduction (Redox) Reactions: one or more electrons (negative charge) are transferred between reaction partners

\[
\text{Mg}(s) + 2\text{HCl}(aq) \rightarrow \text{MgCl}_2(aq) + \text{H}_2(g)
\]

\[
\begin{array}{ccccccc}
\text{zero} & +1 & -1 & +2 & -1 & \text{zero} \\
\hline
-2 \text{ electrons} & \hline
\end{array}
\]

\[2 (+1) \text{ electrons}\]
Types of Chemical Reactions in Aqueous Solution

Acid-Base Neutralization Reactions: acid reacts with base to give water & salt (salt is ionic compound, not always NaCl)

\[ \text{HCl}(aq) + \text{NaOH}(aq) \rightarrow \text{H}_2\text{O}(l) + \text{NaCl}(aq) \]

acid          base                     water         salt
Aqueous Reactions and Net Ionic Equations (different ways to write reactions)

**Molecular Equation**: write complete formulas as if molecules.

\[
Pb(\text{NO}_3)_2(\text{aq}) + 2\text{KI}(\text{aq}) \rightarrow 2\text{KNO}_3(\text{aq}) + \text{PbI}_2(\text{s})
\]

- **Strong electrolytes**
- **Precipitate**
**Ionic Equation**: write strong electrolytes as dissociated ions. (solid, liquid, gas, weak electrolyte compounds written as molecular formula-do NOT dissociate)

\[
Pb(NO_3)_2(aq) + 2KI(aq) \rightarrow Pb(NO_3)_2(s) + 2KNO_3(aq)
\]

\[
Pb^{2+}(aq) + 2NO_3^-(aq) + 2K^+(aq) + 2I^-(aq) \rightarrow 2K^+(aq) + 2NO_3^-(aq) + PbI_2(s)
\]
Aqueous Reactions and Net Ionic Equations
(different ways to write reactions)

**Spectator Ions**: ions that do not change during the reaction. (& just watch the reaction without doing anything)

\[
Pb^{2+}(aq) + 2NO_3^-(aq) + 2K^+(aq) + 2I^-(aq) \rightarrow 2K^+(aq) + 2NO_3^-(aq) + PbI_2(s)
\]
Aqueous Reactions and Net Ionic Equations (different ways to write reactions)

**Net Ionic Equation**: Only the ions undergoing change are shown. *(leave out spectators)*

\[ \text{Pb}^{2+}(aq) + 2\text{I}^-(aq) \rightarrow \text{PbI}_2(s) \]
# Precipitation Reactions and Solubility Guidelines

**Typo L⁺ is really Li⁺**

<table>
<thead>
<tr>
<th>Soluble Compounds</th>
<th>Common Exceptions</th>
</tr>
</thead>
<tbody>
<tr>
<td>L⁺, Na⁺, K⁺, Rb⁺, Cs⁺ (group 1A cations)</td>
<td>None</td>
</tr>
<tr>
<td>NH₄⁺ (ammonium ion)</td>
<td>None</td>
</tr>
<tr>
<td>Cl⁻, Br⁻, I⁻ (halide)</td>
<td>Halides of Ag⁺, Hg₂²⁺, Pb²⁺</td>
</tr>
<tr>
<td>NO₃⁻ (nitrate)</td>
<td>None</td>
</tr>
<tr>
<td>ClO₄⁻ (perchlorate)</td>
<td>None</td>
</tr>
<tr>
<td>CH₃CO₂⁻ (acetate)</td>
<td>None</td>
</tr>
<tr>
<td>SO₄²⁻ (sulfate)</td>
<td>Sulfates of Sr²⁺, Ba²⁺, Hg₂²⁺, Pb²⁺</td>
</tr>
</tbody>
</table>

**Insoluble Compounds**

<table>
<thead>
<tr>
<th>Insoluble Compounds</th>
<th>Common Exceptions</th>
</tr>
</thead>
<tbody>
<tr>
<td>CO₃²⁻ (carbonate)</td>
<td>Carbonates of group 1A cations, NH₄⁺</td>
</tr>
<tr>
<td>S²⁻ (sulfide)</td>
<td>Sulfides of group 1A cations, NH₄⁺, Ca²⁺, Sr²⁺, and Ba²⁺</td>
</tr>
<tr>
<td>PO₄³⁻ (phosphate)</td>
<td>Phosphates of group 1A cations, NH₄⁺</td>
</tr>
<tr>
<td>OH⁻ (hydroxide)</td>
<td>Hydroxides of group 1A cations, NH₄⁺, Ca²⁺, Sr²⁺, and Ba²⁺</td>
</tr>
</tbody>
</table>
HW: Are the following compounds soluble or insoluble in water?

- Li₂CO₃
- BaSO₄
- Mg(NO₃)₂
Write the molecular, ionic, and net ionic equations for the reaction that occurs when aqueous solutions of AgNO₃ and Na₂CO₃ are mixed.
Precipitation Reactions and Solubility Guidelines

Write the molecular, ionic, and net ionic equations for the reaction that occurs when aqueous solutions of AgNO$_3$ and Na$_2$CO$_3$ are mixed.

Write the chemical formulas of the products (use proper ionic rules). (exchange cation / anion partners)

\[
\text{AgNO}_3(\text{aq}) + \text{Na}_2\text{CO}_3(\text{aq}) \rightarrow \text{Ag}_2\text{CO}_3 + \text{NaNO}_3
\]

AB + CD $\rightarrow$ CB + AD

Cation (on left) Anion (on right)

double replacement reaction
Write the molecular, ionic, and net ionic equations for the reaction that occurs when aqueous solutions of AgNO₃ and Na₂CO₃ are mixed.

2. **Molecular Equation**: Balance the equation and predict the solubility of each possible product.

\[
2\text{AgNO}_3(aq) + \text{Na}_2\text{CO}_3(aq) \rightarrow \text{Ag}_2\text{CO}_3(s) + 2\text{NaNO}_3(aq)
\]

- insoluble (look up solubility chart)
- contains a group 1A cation (soluble from solubility chart)
Precipitation Reactions and Solubility Guidelines

Write the molecular, ionic, and net ionic equations for the reaction that occurs when aqueous solutions of AgNO₃ and Na₂CO₃ are mixed.

3. **Ionic Equation**: dissociate the soluble ionic compounds. Do **NOT** dissociate precipitates.

\[
2\text{AgNO}_3(aq) + \text{Na}_2\text{CO}_3(aq) \rightarrow \text{Ag}_2\text{CO}_3(s) + 2\text{NaNO}_3(aq)
\]

\[
2\text{Ag}^+(aq) + 2\text{NO}_3^-(aq) + 2\text{Na}^+(aq) + \text{CO}_3^{2-}(aq)
\]

End 10/2 G section
Precipitation Reactions and Solubility Guidelines

Write the molecular, ionic, and net ionic equations for the reaction that occurs when aqueous solutions of AgNO₃ and Na₂CO₃ are mixed.

4. **Net Ionic Equation**: Eliminate the spectator ions from the ionic equation.

\[ 2\text{Ag}^+(aq) + 2\text{NO}_3^-(aq) + 2\text{Na}^+(aq) + \text{CO}_3^{2-}(aq) \rightarrow \text{Ag}_2\text{CO}_3(s) + 2\text{Na}^+(aq) + 2\text{NO}_3^-(aq) \]

\[ 2\text{Ag}^+(aq) + \text{CO}_3^{2-}(aq) \rightarrow \text{Ag}_2\text{CO}_3(s) \]
HW: Precipitation Reaction For the following, Write out (a) molecular reaction (b) total ionic equation and (c) choose one of the following as the net ionic equation

Lead(II) nitrate + sodium chloride

- What is the net ionic equation for this reaction?

  a. $\text{Pb}^{2+} (aq) + 2\text{NO}_3^- (aq) \rightarrow \text{Pb(NO}_3\text{)}_2 (s)$

  b. $\text{Na}^{2+} (aq) + \text{Cl}^- (aq) \rightarrow \text{NaCl} (s)$

  c. $\text{Pb}^{2+} (aq) + 2\text{Cl}^- (aq) \rightarrow \text{PbCl}_2 (s)$

  d. $\text{Na}^+ (aq) + \text{NO}_3^- (aq) \rightarrow \text{NaNO}_3 (s)$
HW: Precipitation Reaction For the following, Write out (a) molecular reaction (b) total ionic equation and (c) choose one of the following as the net ionic equation

Lead(II) nitrate + sodium chloride

What is the net ionic equation for this reaction?

a. \( \text{Pb}^{2+} (aq) + 2\text{NO}_3^- (aq) \rightarrow \text{Pb(NO}_3\text{)}_2 (s) \)

b. \( \text{Na}^{2+} (aq) + \text{Cl}^- (aq) \rightarrow \text{NaCl} (s) \)

c. \( \text{Pb}^{2+} (aq) + 2\text{Cl}^- (aq) \rightarrow \text{PbCl}_2 (s) \) *

d. \( \text{Na}^+ (aq) + \text{NO}_3^- (aq) \rightarrow \text{NaNO}_3 (s) \)

End 10/4 F, G sect
End 10/7 D section
10/7 Monday was just a review day – no new material covered – only questions & reviewing quiz 2 answer keys (note D section had already seen the question on precipitation molecular, ionic and net ionic reaction on Thursday 10/3 their quiz date but we completed the HW in class on that HW problem for the D section on 10/7 – the answer for the precipitation problem was already posted on 10/5 before the D section did the HW problem in class for points)
Acids, Bases, and Neutralization Reactions

**Acid (Arrhenius):** dissociates in water to produce hydrogen ions, $H^+$

- **General Reaction:** $HA(aq) \rightarrow H^+(aq) + A^-(aq)$
- **Example:** $HCl(aq) \rightarrow H^+(aq) + Cl^-(aq)$

In water, acids produce **hydronium ions, $H_3O^+$:**

$$HCl(aq) + H_2O(aq) \rightarrow H_3O^+(aq) + Cl^-(aq)$$
Acids, Bases, and Neutralization Reactions

Base (Arrhenius): dissociates in water to produce hydroxide ions, OH⁻:

- General reaction: $\text{MOH}(aq) \rightarrow \text{M}^+(aq) + \text{OH}^-(aq)$
- Example: $\text{NaOH}(aq) \rightarrow \text{Na}^+(aq) + \text{OH}^-(aq)$

Ammonia is a weak base produces ammonium and hydroxide ions: (only weak base commonly seen)

$$\text{NH}_3(aq) + \text{H}_2\text{O}(aq) \rightarrow \text{NH}_4^+(aq) + \text{OH}^-(aq)$$
Strong acids and strong bases are strong electrolytes. Weak acids and weak bases are weak electrolytes. Know name, formula, whether strong/weak acid or base for *
(most group 1A & 2A hydroxides are strong bases if soluble)

<table>
<thead>
<tr>
<th>TABLE 4.3 Some Common Acids and Bases</th>
</tr>
</thead>
<tbody>
<tr>
<td>Strong acid</td>
</tr>
<tr>
<td></td>
</tr>
<tr>
<td></td>
</tr>
<tr>
<td></td>
</tr>
<tr>
<td></td>
</tr>
<tr>
<td>Strong base</td>
</tr>
<tr>
<td></td>
</tr>
<tr>
<td></td>
</tr>
<tr>
<td></td>
</tr>
</tbody>
</table>
Acids – characterized by \# H\(^+\) dissociation

\[ \text{H}_3\text{PO}_4 \rightarrow 3 \text{H}^+ + \text{PO}_4^{3-} \]  
triprotic

\[ \text{H}_2\text{SO}_4 \rightarrow 2 \text{H}^+ + \text{SO}_4^{2-} \]  
diprotic

\[ \text{HBr} \rightarrow \text{H}^+ + \text{Br}^- \]  
monoprotic

\[ \text{CH}_3\text{CO}_2\text{H} \rightarrow \text{H}^+ + \text{CH}_3\text{CO}_2^- \]  
monoprotic

(acetic acid – not 4 H\(^+\))

polyprotic acid = more than one hydrogen is acidic
Acids, Bases, and Neutralization Reactions

Naming Binary Acids (all halogens)

HCl  hydrochloric acid
HBr  hydrobromic acid
HF   hydrofluoric acid  (only weak halogen acid)
HI   hydroiodic acid
**Acids, Bases, and Neutralization Reactions**

These acid–base neutralization reactions are double-replacement reactions just like the precipitation reactions:

(Exchange cation / anion partners)

<table>
<thead>
<tr>
<th>Acid</th>
<th>Base</th>
<th>Salt</th>
<th>Water</th>
</tr>
</thead>
<tbody>
<tr>
<td>$\text{HA}(aq)$</td>
<td>$\text{MOH}(aq)$</td>
<td>$\text{MA}(aq)$</td>
<td>$\text{H}_2\text{O}(l)$</td>
</tr>
</tbody>
</table>

Double replacement reaction
Write the molecular, ionic, and net ionic equations for the reaction of aqueous HBr and aqueous Ba(OH)$_2$.

1. Write the chemical formulas of the products (use proper ionic rules for the salt to write neutral formula of the salt).

\[
\text{HBr}(aq) + \text{Ba(OH)}_2(aq) \rightarrow \text{HOH} + \text{BaBr}_2
\]

Acid  Base  Water  Salt
Write the **molecular**, ionic, and net ionic equations for the reaction of aqueous HBr and aqueous Ba(OH)$_2$.

2. **Molecular Equation**: Balance the equation and predict the solubility of the salt in the products. (from solubility rules table)

\[ 2\text{HBr}(aq) + \text{Ba(OH)}_2(aq) \rightarrow 2\text{H}_2\text{O}(l) + \text{BaBr}_2(aq) \]

Use the solubility rules.
Acids, Bases, and Neutralization Reactions

Write the molecular, ionic, and net ionic equations for the reaction of aqueous HBr and aqueous Ba(OH)$_2$.

3. **Ionic Equation**: Dissociate the strong acid and the soluble ionic compounds. (solid, liquid, gas compounds, weak acid written as molecular formula-do NOT dissociate)

\[
\begin{align*}
2\text{HBr}(aq) & \quad \text{Ba(OH)}_2(aq) \\
2\text{H}^+(aq) + 2\text{Br}^- (aq) + \text{Ba}^{2+}(aq) + 2\text{OH}^- (aq) & \quad \longrightarrow 2\text{H}_2\text{O}(l) + \text{Ba}^{2+}(aq) + 2\text{Br}^- (aq) \\
& \quad \text{BaBr}_2(aq)
\end{align*}
\]
Write the molecular, ionic, and net ionic equations for the reaction of aqueous HBr and aqueous Ba(OH)₂.

4. **Net Ionic Equation:** Eliminate the spectator ions from the ionic equation.

\[
2H^+(aq) + 2Br^- (aq) + Ba^{2+} (aq) + 2OH^- (aq) \\
\rightarrow 2H_2O(l) + Ba^{2+} (aq) + 2Br^- (aq)
\]

\[
2H^+(aq) + 2OH^- (aq) \rightarrow 2H_2O(l)
\]

Net ionic equation for acid/base neutralization is almost always this (unless weak undissociated acid/base)
Acids, Bases, and Neutralization Reactions

Write the molecular, ionic, and net ionic equations for the reaction of aqueous NaOH and aqueous HF.

1. Write the chemical formulas of the products (use proper ionic rules for the salt to write neutral ionic formulas).

\[
\text{HF}(aq) + \text{NaOH}(aq) \rightarrow \text{HOH} + \text{NaF}
\]

Acid \hspace{1cm} Base \hspace{1cm} Water \hspace{1cm} Salt
Write the molecular, ionic, and net ionic equations for the reaction of aqueous NaOH and aqueous HF.

2. **Molecular Equation**: Balance the equation and predict the solubility of the salt in the products.

\[
HF(aq) + NaOH(aq) \rightarrow H_2O(l) + NaF(aq)
\]

use the solubility rules.
Write the molecular, ionic, and net ionic equations for the reaction of aqueous NaOH and aqueous HF.

3. **Ionic Equation**: Dissociate the soluble ionic compounds. (HF is a weak acid so does not dissociate)

\[
\text{NaOH}(aq) \\
\text{HF}(aq) + \text{Na}^+(aq) + \text{OH}^- (aq) \rightarrow \text{H}_2\text{O}(l) + \text{Na}^+(aq) + \text{F}^- (aq) \\
\text{NaF}(aq)
\]
Write the molecular, ionic, and net ionic equations for the reaction of aqueous NaOH and aqueous HF.

4. **Net Ionic Equation**: Eliminate the spectator ions from the ionic equation.

\[
\text{HF}(aq) + \text{Na}^+(aq) + \text{OH}^-(aq) \rightarrow \text{H}_2\text{O}(l) + \text{Na}^+(aq) + \text{F}^-(aq)
\]

\[
\text{HF}(aq) + \text{OH}^-(aq) \rightarrow \text{H}_2\text{O}(l) + \text{F}^-(aq)
\]
Write the molecular, ionic, and net ionic equations for the reaction of aqueous \( \text{Ca} (\text{OH})_2 \) and aqueous \( \text{HI} \).

**molecular equation:**

\[
\text{Ca} (\text{OH})_2 + \text{HI} \rightarrow \text{________} + \text{________}
\]

**ionic equation**

net ionic equation
Solution Stoichiometry (already know how to do combining 2 things you know – stoichiometry & Molarity)

What volume of 0.250 M \( \text{H}_2\text{SO}_4 \) is needed to react with 50.0 mL of 0.100 M \( \text{NaOH} \)? 10.0 mL \( \text{H}_2\text{SO}_4 \)

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

End 10/11 F F section

What volume of 0.250 M \( \text{HCl} \) is needed to react with 50.0 mL of 0.100 M \( \text{NaOH} \)? 20.0 mL \( \text{HCl} \)

\[
\text{HCl}(aq) + \text{NaOH}(aq) \rightarrow \text{NaCl}(aq) + \text{H}_2\text{O}(l)
\]

End 10/10/19 D, G section
Solution Stoichiometry (short cut version)

What volume of 0.250 M H Cl is needed to react with 50.0 mL of 0.100 M NaOH?

\[ \text{H Cl}(aq) + \text{NaOH}(aq) \rightarrow \text{Na Cl}(aq) + \text{H}_2\text{O}(l) \]

Short cut: use \( M_{\text{acid}} V_{\text{acid}} = M_{\text{base}} V_{\text{base}} \)

only works for 1:1 acid/base reaction

\[(0.250 \text{ M HCl}) \times (V_{\text{HCl}}) = (0.100 \text{ M NaOH}) \times (50.0 \text{ mL NaOH})\]

\[ V_{\text{HCl}} = \frac{(0.100 \text{ M NaOH}) \times (50.0 \text{ mL NaOH})}{0.250 \text{ M H Cl}} = 20.0 \text{ mL HCl} \]
What volume of 2.25 M Ca(OH)$_2$ is needed to react with 25.0 mL of 0.100 M HCl?

$$2 \text{HCl}(aq) + \text{Ca(OH)}_2(aq) \rightarrow \text{CaCl}_2(aq) + 2\text{HOH}(l)$$
HW: Solution Stoichiometry

What volume of 2.25 M Ca(OH)\(_2\) is needed to react with 25.0 mL of 0.100 M HCl?

\[
2 \text{ HCl}(aq) + \text{Ca(OH)}_2(aq) \rightarrow \text{CaCl}_2(aq) + 2\text{HOH}(l)
\]

\[
25.0 \text{ mL} \times 0.100 \text{ moles H Cl} \times \frac{1 \text{ mol Ca(OH)}_2}{2 \text{ mol H Cl}} \times \frac{1000 \text{ mL Ca(OH)}_2}{1000 \text{ mL HCl soln}} = 0.556 \text{ mL Ca(OH)}_2 \text{ solution}
\]
Titration: A procedure for determining the concentration using a second solution with known concentration.

$$\text{HCl}(aq) + \text{NaOH}(aq) \rightarrow \text{NaCl}(aq) + \text{H}_2\text{O}(l)$$

**Short cut:** use\[ M_{\text{acid}}V_{\text{acid}} = M_{\text{base}}V_{\text{base}} \]

How can you tell when the reaction is complete? **Equivalence point** – point where you have the same amount $\text{H}^+$ and $\text{OH}^-$

**Indicator:** Used to mark the equivalence point changes color at the equivalence point.
Measuring the Concentration of a Solution: Titration

A measured volume of acid solution is placed in a flask, and phenolphthalein indicator is added. Base solution of known concentration is added from a buret until the indicator changes color. Reading the volume of base from the buret allows calculation of the acid concentration.
Oxidation–Reduction (Redox) Reactions

2Mg(s) + O₂(g) → 2MgO(s)
Oxidation–Reduction (Redox) Reactions

\[ 2P(s) + 3Br_2(l) \rightarrow 2PBr_3(l) \]
Oxidation–Reduction (Redox) Reactions

Rusting of iron: an oxidation of Fe

4Fe(s) + 3O₂(g) → 2Fe₂O₃(s)

Manufacture of iron: a reduction of Fe₂O₃

2Fe₂O₃(s) + 3C(s) → 4Fe(s) + 3CO₂(g)

Reactant A might be:
- a neutral atom,
- a monatomic ion,
- a polyatomic ion,
- or a molecule.
Oxidation–Reduction (Redox) Reactions

**Oxidation**: The loss of one or more electrons by a substance, whether element, compound, or ion.

**Reduction**: The gain of one or more electrons by a substance (memorize as add negative – reduced ox number).

**Oxidation–Reduction (Redox) Reaction**: Any process in which electrons are transferred from one substance to another.
Oxidation–Reduction (Redox) Reactions

Oxidation Number (State): A value that indicates whether an atom is neutral, electron-rich, or electron-poor (These rules are on the memorize list on departmental syllabus.)

**Rules for Assigning Oxidation Numbers**

1. An atom in its elemental state has an oxidation number of 0 (zero).

   ![Diagram of Oxidation Numbers](image)

   - Na
   - H₂
   - Br₂
   - S
   - Ne

   Oxidation number 0
Oxidation–Reduction (Redox) Reactions

2. An atom in a monatomic ion has an oxidation number identical to its charge.

(group 1A to 3A: group #) or

(group 7A to 5A: group # - 8)

<table>
<thead>
<tr>
<th>Group</th>
<th>Ion</th>
<th>Oxidation Number</th>
</tr>
</thead>
<tbody>
<tr>
<td>1A</td>
<td>Na⁺</td>
<td>+1</td>
</tr>
<tr>
<td>2A</td>
<td>Ca²⁺</td>
<td>+2</td>
</tr>
<tr>
<td>3A</td>
<td>Al³⁺</td>
<td>+3</td>
</tr>
<tr>
<td>7A</td>
<td>Cl⁻</td>
<td>−1</td>
</tr>
<tr>
<td>6A</td>
<td>O²⁻</td>
<td>−2</td>
</tr>
<tr>
<td>5A</td>
<td>N³⁻</td>
<td>−3</td>
</tr>
</tbody>
</table>
Oxidation–Reduction (Redox) Reactions

3. An atom in a polyatomic ion or in a molecular compound usually has the same oxidation number it would have if it were a monatomic ion.

a) Hydrogen can be either +1 or –1.

\[
\begin{align*}
\left( \text{H} - \text{O} \right)^{1-} & \\
\text{H} - \text{Ca} - \text{H} & \\
\text{+1} & \quad \text{+2} & \quad \text{-1} & \quad \text{-1}
\end{align*}
\]

b) Oxygen usually has an oxidation number of –2.

\[
\begin{align*}
\text{H} - \text{O} - \text{H} & \\
\text{H} - \text{O} - \text{O} - \text{H} & \\
\text{+1} & \quad \text{-2} & \quad \text{+1} & \quad \text{+1} & \quad \text{-1} & \quad \text{-1} & \quad \text{+1}
\end{align*}
\]
3. c) Halogens *usually* have an oxidation number of $-1$. 

\[
\begin{align*}
\text{H} \rightarrow \text{Cl} & \quad & \text{Cl} \rightarrow \text{O} \rightarrow \text{Cl} \\
+1 \quad -1 & \quad & +1 \quad -2 \quad +1
\end{align*}
\]
4. The sum of the oxidation numbers is 0 for a neutral compound and is equal to the net charge for a polyatomic ion.

\[
\text{Cr}_2\text{O}_7^{2-} \quad 2(?) + 7(-2) = -2 \text{ (net charge)} \\
\text{Cr} = ? = +6
\]

\[
\text{H}_2\text{SO}_4 \quad 2(+1) + (?) + 4(-2) = 0 \text{ (net charge)} \\
\text{S} = ? = +6
\]
What is the oxidation state of the following.

a. Na  element zero

b. H₂  element zero

c. Na Cl  \( \text{Na}^{+1}(\text{gp} \ 1\Lambda) \ 7-8 = -1 \ (\text{gp} \ 7\Lambda) \)

d. HNO₃  \( \text{H}^{+1}(\text{gp} \ 1\Lambda), \ \text{O}^{-2} \ (\text{gp} \ 6\Lambda) \)

\[
\text{ox state } N = \text{variable } N, \text{ calculate }
\]

\[
\text{zero } = +1 + N + 3*(-2)
\]

\[
\text{algebra } N = +6 - 1 = 5
\]
HW: What is the oxidation state of the following.

a. Cl₂

b. Fe

c. Al Cl₃

d. PO₄⁻³
Identifying Redox Reactions

**Oxidizing Agent (is reduced)**
Causes oxidation
Gains one or more electrons
Undergoes reduction
Oxidation number of atom decreases.
(becomes more negative)

**Reducing Agent (is oxidized)**
Causes reduction
Loses one or more electrons
Undergoes oxidation
Oxidation number of atom increases.
(becomes more positive)
Identifying Redox Reactions
this chapter - only responsible for recognizing oxidation/reduction & oxidizing agent & reducing agent – NOT for balancing redox equations (chapter 18)

Oxidizing Agent

Reducing Agent

\[
4\text{Fe}(s) + 3\text{O}_2(g) \rightarrow 2\text{Fe}_2\text{O}_3(s)
\]
Identifying Redox Reactions

Oxidizing Agent

Reduction

\[ 2\text{Fe}_2\text{O}_3(s) + 3\text{C}(s) \rightarrow 4\text{Fe}(s) + 3\text{CO}_2(g) \]

Reduction

Oxidation

Reducing Agent
HW: Oxidation or Reduction
(reducing agent / oxidizing agent)

(a) Assign ox states of Zn & H
(b) assign oxidation or reduction to brackets

\[
\begin{align*}
\text{Zn} + 2 \text{H Cl} & \rightarrow \text{ZnCl}_2 + \text{H}_2 \\
\end{align*}
\]
HW: Oxidation or Reduction
(reducing agent / oxidizing agent)

(a) Assign ox states of Zn & H
(b) assign oxidation or reduction to brackets

\[
\text{Zn} + 2 \text{HCl} \rightarrow \text{ZnCl}_2 + \text{H}_2
\]

- 2e per Zn oxidation (Zn is reducing agent)

+ 1e per H reduction (H is oxidizing agent)

10/17 R - D section
Iron loses electron to copper

\[ \text{Fe}(s) + \text{Cu}^{2+}(aq) \rightarrow \text{Fe}^{2+}(aq) + \text{Cu}(s) \]
The Activity Series of the Elements

Fe(s) + Cu\(^{2+}\)(aq) → Fe\(^{2+}\)(aq) + Cu(s)

<table>
<thead>
<tr>
<th>Oxidation Reaction</th>
<th>Elements that are higher up in the table (reaction goes →) are more likely to be oxidized (lose e).</th>
</tr>
</thead>
<tbody>
<tr>
<td>Strongly reducing</td>
<td>These elements react rapidly with aqueous H(^+) ions (acid) or with liquid H(_2)O to release H(_2) gas.</td>
</tr>
<tr>
<td></td>
<td>Li → Li(^+) + e(^-)</td>
</tr>
<tr>
<td></td>
<td>K → K(^+) + e(^-)</td>
</tr>
<tr>
<td></td>
<td>Ba → Ba(^{2+}) + 2 e(^-)</td>
</tr>
<tr>
<td></td>
<td>Ca → Ca(^{2+}) + 2 e(^-)</td>
</tr>
<tr>
<td></td>
<td>Na → Na(^+) + e(^-)</td>
</tr>
<tr>
<td></td>
<td>Mg → Mg(^{2+}) + 2 e(^-)</td>
</tr>
<tr>
<td></td>
<td>Al → Al(^{3+}) + 3 e(^-)</td>
</tr>
<tr>
<td></td>
<td>Mn → Mn(^{2+}) + 2 e(^-)</td>
</tr>
<tr>
<td></td>
<td>Zn → Zn(^{2+}) + 2 e(^-)</td>
</tr>
<tr>
<td></td>
<td>Cr → Cr(^{3+}) + 3 e(^-)</td>
</tr>
<tr>
<td></td>
<td>Fe → Fe(^{2+}) + 2 e(^-)</td>
</tr>
<tr>
<td>Weakly reducing</td>
<td>These elements react with aqueous H(^+) ions or with steam to release H(_2) gas.</td>
</tr>
<tr>
<td></td>
<td>Co → Co(^{2+}) + 2 e(^-)</td>
</tr>
<tr>
<td></td>
<td>Ni → Ni(^{2+}) + 2 e(^-)</td>
</tr>
<tr>
<td></td>
<td>Sn → Sn(^{2+}) + 2 e(^-)</td>
</tr>
<tr>
<td></td>
<td>H(_2) → 2 H(^+) + 2 e(^-)</td>
</tr>
<tr>
<td></td>
<td>Cu → Cu(^{2+}) + 2 e(^-)</td>
</tr>
<tr>
<td>Weakly reducing</td>
<td>These elements react with aqueous H(^+) ions to release H(_2) gas.</td>
</tr>
<tr>
<td></td>
<td>Ag → Ag(^+) + e(^-)</td>
</tr>
<tr>
<td></td>
<td>Hg → Hg(^{2+}) + 2 e(^-)</td>
</tr>
<tr>
<td></td>
<td>Pt → Pt(^{2+}) + 2 e(^-)</td>
</tr>
<tr>
<td></td>
<td>Au → Au(^{3+}) + 3 e(^-)</td>
</tr>
<tr>
<td>These elements do not react with aqueous H(^+) ions to release H(_2).</td>
<td>Thus, any element higher in the activity series will reduce (reducing agent) (lose e to) the ion of any element lower in the activity series.</td>
</tr>
</tbody>
</table>
The Activity Series of the Elements

Cu(s) + 2Ag^+(aq) \rightarrow Cu^{2+}(aq) + 2Ag(s) \quad (a)

2Ag(s) + Cu^{2+}(aq) \rightarrow 2Ag^+(aq) + Cu(s) \quad (b)

Which one of these reactions will occur?

<table>
<thead>
<tr>
<th>TABLE 4.5</th>
<th>A Partial Activity Series of the Elements</th>
</tr>
</thead>
<tbody>
<tr>
<td>Oxidation Reaction</td>
<td>Li \rightarrow Li^+ + e^-</td>
</tr>
<tr>
<td>Strongly reducing</td>
<td>K \rightarrow K^+ + e^-</td>
</tr>
<tr>
<td></td>
<td>Ba \rightarrow Ba^{2+} + 2e^-</td>
</tr>
<tr>
<td></td>
<td>Ca \rightarrow Ca^{2+} + 2e^-</td>
</tr>
<tr>
<td></td>
<td>Na \rightarrow Na^+ + e^-</td>
</tr>
<tr>
<td></td>
<td>Mg \rightarrow Mg^{2+} + 2e^-</td>
</tr>
<tr>
<td></td>
<td>Al \rightarrow Al^{3+} + 3e^-</td>
</tr>
<tr>
<td></td>
<td>Mn \rightarrow Mn^{2+} + 2e^-</td>
</tr>
<tr>
<td></td>
<td>Zn \rightarrow Zn^{2+} + 2e^-</td>
</tr>
<tr>
<td></td>
<td>Cr \rightarrow Cr^{3+} + 3e^-</td>
</tr>
<tr>
<td></td>
<td>Fe \rightarrow Fe^{2+} + 2e^-</td>
</tr>
<tr>
<td></td>
<td>Co \rightarrow Co^{2+} + 2e^-</td>
</tr>
<tr>
<td></td>
<td>Ni \rightarrow Ni^{2+} + 2e^-</td>
</tr>
<tr>
<td></td>
<td>Sn \rightarrow Sn^{2+} + 2e^-</td>
</tr>
<tr>
<td></td>
<td>H_2 \rightarrow 2H^+ + 2e^-</td>
</tr>
<tr>
<td></td>
<td>Cu \rightarrow Cu^{2+} + 2e^-</td>
</tr>
<tr>
<td></td>
<td>Ag \rightarrow Ag^+ + e^-</td>
</tr>
<tr>
<td></td>
<td>Hg \rightarrow Hg^{2+} + 2e^-</td>
</tr>
<tr>
<td></td>
<td>Pt \rightarrow Pt^{2+} + 2e^-</td>
</tr>
<tr>
<td></td>
<td>Au \rightarrow Au^{3+} + 3e^-</td>
</tr>
</tbody>
</table>

Cu is above Ag in activity series
Cu \rightarrow Ag \leftarrow
The Activity Series of the Elements

Cu(s) + 2Ag⁺(aq) ⟷ Cu²⁺(aq) + 2Ag(s) (a) *

2Ag(s) + Cu²⁺(aq) ⟷ 2Ag⁺(aq) + Cu(s) (b)

Which one of these reactions will occur?

Cu is above Ag in activity series
Cu → Ag ←
The Activity Series of the Elements

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

copper is above silver in activity series chart

(copper loses electron and silver gets electron)
Redox Titrations

\[ 5\text{H}_2\text{C}_2\text{O}_4(aq) + 2\text{MnO}_4^-(aq) + 6\text{H}^+(aq) \rightarrow 10\text{CO}_2(g) + 2\text{Mn}^{2+}(aq) + 8\text{H}_2\text{O}(l) \]
A solution is prepared with 0.2585 g of oxalic acid, $\text{H}_2\text{C}_2\text{O}_4$. 22.35 mL of an unknown solution of potassium permanganate is needed to titrate the solution. What is the concentration of the potassium permanganate solution? (already know how to do this: if given balanced reaction, just stoichiometry & molarity problem)

$$5\text{H}_2\text{C}_2\text{O}_4(aq) + 2\text{MnO}_4^-(aq) + 6\text{H}^+(aq) \rightarrow 10\text{CO}_2(g) + 2\text{Mn}^{2+}(aq) + 8\text{H}_2\text{O}(l)$$
Redox Titrations

5H₂C₂O₄(aq) + 2MnO₄⁻(aq) + 6H⁺(aq) → 10CO₂(g) + 2Mn²⁺(aq) + 8H₂O(l)

Moles of H₂C₂O₄ available:

0.2585 g H₂C₂O₄ × \frac{1 \text{ mol}}{90.04 \text{ g}} = 0.002871 \text{ mol H₂C₂O₄}

Moles of KMnO₄ reacted:

0.002871 \text{ mol H₂C₂O₄} × \frac{2 \text{ mol KMnO₄}}{5 \text{ mol H₂C₂O₄}} = 0.001148 \text{ mol KMnO₄}
Redox Titrations

\[ 5\text{H}_2\text{C}_2\text{O}_4(aq) + 2\text{MnO}_4^-(aq) + 6\text{H}^+(aq) \rightarrow 10\text{CO}_2(g) + 2\text{Mn}^{2+}(aq) + 8\text{H}_2\text{O}(l) \]

Concentration of KMnO\(_4\) solution:

\[
\frac{0.001148 \text{ mol KMnO}_4}{22.35 \text{ mL}} \times \frac{1000 \text{ mL}}{1 \text{ L}} = 0.05136 \text{ M KMnO}_4
\]