1. Consider the following solutes in water:
   \[ \text{C}_6\text{H}_{12}\text{O}_6 \quad \text{BaSO}_4 \quad \text{HNO}_3 \quad (\text{NH}_4)_2\text{CO}_3 \quad \text{CH}_3\text{COOH} \]
   a. Identify each as either: strong electrolyte (SE), weak electrolyte (WE) or non-electrolyte (NE)
   b. If all solutions are 1.0 M, put in order of lowest to highest conductivity.

2. What is the final chloride ion concentration if 59.3 g NaCl is added to 400. mL of 0.50 M barium chloride?

3. What volume of 4.20 M HCl is required to prepare 12.0 L of 0.210 M HCl?

4. If 100. mL of 0.200 M silver nitrate is added to 200. mL of 0.150 M barium chloride...
   a. Write the balanced molecular equation, complete ionic equation and the net ionic equation. (use the solubility rules to help you)
b. What will be the mass of precipitate formed?

c. Calculate the concentration of each ion remaining in solution after the above reaction is complete.

5. What volume of 0.0500 M Ba(OH)_2 is required to neutralize exactly 22.00 mL of 0.113 M H_3PO_3?

6. Write the balanced molecular, complete ionic and net ionic equations for the reaction of hydrochloric acid and barium hydroxide.
7. If 0.47 g of barium hydroxide are added to 100.0 mL of 0.040 M hydrochloric acid, what are the concentrations of each ion remaining in solution? Is the solution acidic, basic or neutral? (use the above reaction)

8. Assign oxidation states to each atom in the following:
   a. NH₄⁺
   b. Na₂C₂O₄
   c. PbSO₃

9. For each of the following, indicate whether it is a redox reaction. If it is, identify the substance oxidized and reduced, the oxidizing and reducing agents and the number of electrons transferred.
   a. NaHCO₃ (aq) + HBr (aq) \rightarrow H₂O (l) + CO₂ (g) + NaBr (aq)
   b. CH₄ (g) + 2O₂ (g) \rightarrow CO₂ (g) + 2H₂O (l)
10. Balance the following redox reactions:

a. $\text{MnO}_4^- + \text{Fe}^{2+} \rightarrow \text{Mn}^{2+} + \text{Fe}^{3+}$ (Acidic)

b. $\text{CrI}_3 + \text{Cl}_2 \rightarrow \text{CrO}_4^{2-} + \text{IO}_4^- + \text{Cl}^-$ (basic)
### Solubility of Ionic Compounds in Water

<table>
<thead>
<tr>
<th>Anion</th>
<th>Soluble*</th>
<th>Slightly Soluble*</th>
<th>Insoluble*</th>
</tr>
</thead>
<tbody>
<tr>
<td>NO$_3^-$ (nitrate)</td>
<td>All</td>
<td>—</td>
<td>—</td>
</tr>
<tr>
<td>CH$_3$COO$^-$ (acetate)</td>
<td>Most</td>
<td>—</td>
<td>Be(CH$_3$COO)$_2$</td>
</tr>
<tr>
<td>F$^-$ (fluoride)</td>
<td>Group I, AgF, BeF$_2$</td>
<td>SrF$_2$, BaF$_2$, PbF$_2$</td>
<td>MgF$_2$, CaF$_2$</td>
</tr>
<tr>
<td>Cl$^-$ (chloride)</td>
<td>Most</td>
<td>PbCl$_2$</td>
<td>AgCl, HgCl$_2$</td>
</tr>
<tr>
<td>Br$^-$ (bromide)</td>
<td>Most</td>
<td>PbBr$_2$, HgBr$_2$</td>
<td>AgBr, HgBr$_2$</td>
</tr>
<tr>
<td>SO$_4^{2-}$ (sulfate)</td>
<td>Most</td>
<td>CaSO$_4$, Ag$_2$SO$_4$</td>
<td>BaSO$_4$, SrSO$_4$, PbSO$_4$, HgSO$_4$</td>
</tr>
<tr>
<td>S$^{2-}$ (sulfide)</td>
<td>Group I and II, (NH$_4$)$_2$S</td>
<td>—</td>
<td>Most</td>
</tr>
<tr>
<td>CO$_3^{2-}$ (carbonate)</td>
<td>Group I, (NH$_4$)$_2$CO$_3$</td>
<td>—</td>
<td>Most</td>
</tr>
<tr>
<td>SO$_3^{2-}$ (sulfite)</td>
<td>Group I, (NH$_4$)$_2$SO$_3$</td>
<td>—</td>
<td>Most</td>
</tr>
<tr>
<td>PO$_4^{3-}$ (phosphate)</td>
<td>Group I, (NH$_4$)$_2$PO$_4$</td>
<td>—</td>
<td>Most</td>
</tr>
<tr>
<td>OH$^-$ (hydroxide)</td>
<td>Group I, Ba(OH)$_2$, NH$_4$OH</td>
<td>Sr(OH)$_2$, Ca(OH)$_2$</td>
<td>Most</td>
</tr>
</tbody>
</table>

*Soluble* dissolves to the extent of > 10 g/L.  *Slightly Soluble*: 0.1 to 10 g/L.  *Insoluble*: < 0.1 g/L.

### PERIODIC TABLE

<table>
<thead>
<tr>
<th>Period</th>
<th>1A</th>
<th>2A</th>
<th>3A</th>
<th>4A</th>
<th>5A</th>
<th>6A</th>
<th>7A</th>
<th>8A</th>
</tr>
</thead>
<tbody>
<tr>
<td>1</td>
<td>1H</td>
<td></td>
<td></td>
<td></td>
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<tr>
<td>2</td>
<td>3Li</td>
<td>4Be</td>
<td>5B</td>
<td>6C</td>
<td>7N</td>
<td>8O</td>
<td>9F</td>
<td>10Ne</td>
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<td>36As</td>
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<td>38Br</td>
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<td>42Zr</td>
<td>43Nb</td>
<td>44Mo</td>
<td>45Tc</td>
<td>46Ru</td>
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<td>48Pd</td>
<td>49Ag</td>
<td>50Cd</td>
<td>51In</td>
<td>52Sn</td>
<td>53Sb</td>
<td>54Te</td>
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<tr>
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<td>55121I</td>
<td>56B</td>
<td>57Cs</td>
<td>132Bi</td>
<td>137La-Lu</td>
<td>177Re</td>
<td>178Os</td>
<td>183Ir</td>
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<td>217Hg</td>
<td>219Tl</td>
<td>223Fr</td>
<td>226Ra-Ra-Lr</td>
<td>265Rh</td>
<td>266Ir</td>
<td>268Os</td>
<td>272Ir</td>
</tr>
</tbody>
</table>

**Notes:**
- **1A**: 1H
- **2A**: 3Li, 4Be, 5B, 6C, 7N, 8O, 9F, 10Ne
- **3A**: 11Na, 13Al, 14Si, 15P, 16S, 17Cl, 18Ar
- **4A**: 23K, 24Ca, 25Sc, 26Ti, 27V, 28Cr, 29Mn, 30Fe
- **5A**: 31Ni, 32Cu, 33Zn, 34Ga, 35Ge, 36As, 37Se, 38Br
- **6A**: 39Kr, 40Sr, 41Y, 42Zr, 43Nb, 44Mo, 45Tc, 46Ru
- **7A**: 47Rh, 48Pd, 49Ag, 50Cd, 51In, 52Sn, 53Sb, 54Te
- **8A**: 55121I, 56B, 57Cs, 132Bi, 137La-Lu, 177Re, 178Os, 183Ir

**Atomic Numbers:**
- **18**: 6O
- **20**: 8F
- **35**: 55Br
- **56**: 132Ba
- **87**: 223Fr
- **88**: 226Ra
- **114**: 269Un
- **118**: 282Un

**Elements:**
- **He**: Helium
- **Be**: Beryllium
• Rules for assigning oxidation states:

1. Elements = 0
2. The sum of oxidation states for a molecule or ion = the charge on molecule or ion
3. Group 1 Metals (Li, Na, K, ...) = +1
4. Group 2 Metals (Be, Mg, Ca...) = +2
5. Fluorine = -1
6. Hydrogen = +1
7. Oxygen = -2 (except for peroxides: $\text{O}_2^2\rightarrow -1$)
8. Halogens (Cl, Br, I) = -1

• Definitions: (remember “OIL RIG” or “LEO GER”)

Oxidation - increase in oxidation state (or loss of electrons)

Reduction - decrease in oxidation state (or gain of electrons)

Oxidizing agent - the reactant being reduced

Reducing agent - the reactant being oxidized

• Rules for balancing redox reactions by the half-reaction method:

1. Write the oxidation and reduction half-reactions

2. For each half-reaction:
   a. balance all elements besides O and H
   b. balance O using H$_2$O
   c. balance H using H$^+$
   d. balance the charge using electrons (add electrons to the more positive side)

3. Equalize the numbers of electrons in both half-reactions by multiplying one or both reactions

4. Add the half-reactions and simplify by canceling anything present on both sides

5. If the reaction is in basic solution, add the number of OH$^-$ ions equal to the number of H$^+$ ions to both sides
   a. H$^+$ + OH$^-$ forms H$_2$O
   b. cancel the H$_2$O that appears on both sides