NET IONIC EQUATIONS

A balanced chemical equation can describe all chemical reactions, an example of such an equation is:

\[ \text{NaCl} + \text{AgNO}_3 \rightarrow \text{AgCl} + \text{NaNO}_3 \]

In this case, the simple formulas of the various reactants and products are present and an arrow separates them.

In some cases, additional information may supplement the original equation. An example of this situation would be:

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

The additional symbolism in this example gives information on the form of the chemicals present. The notation (aq) designates which materials are present as solutes in an aqueous solution, and the notation (s) indicates that one substance is present as a solid (not in solution) precipitate.

Both of the preceding examples are Molecular Equations. A molecular equation simply lists the formulas of the materials present without any consideration of their actual form. In this type of equation, treat everything as if it were a molecule.

In some cases, it is better to re-examine molecular equations and consider a better description of what is really happening. For the above reaction the result is:

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

In this form, you should represent all strong electrolytes as separate ions. This is a Total or Complete Ionic Equation.

While a complete ionic equation may be a better description of what is happening, it tends to say too much and overlook the underlying cause of the reaction. Note that in the above example two ions (Na\(^+\text{(aq)}\) and \(\text{NO}_3^-(\text{aq})\)) do not undergo any apparent change. Ions that do not change in a reaction are Spectator Ions. Spectator ions are present in the same form throughout the reaction; since they do not change, they have no direct bearing on the reaction. To look at the true reaction one must focus on the non-spectator ions. A complete ionic equation without the spectator ions is a Net Ionic Equation. For the above reaction the net ionic equation is:

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

Net ionic equations have the advantage of concentrating on the actual cause of the reaction. They also allow one to summarize a large number of related reactions. For example, if KCl, LiCl or even CaCl\(_2\) replaced the NaCl, a new molecular equation would be required in every case, however; only one net ionic equation applies in these cases.
To produce a net ionic equation it is important to be able to identify the types of materials present in the reaction. The choices are **nonelectrolytes**, **weak electrolytes**, and **strong electrolytes**. In most cases, the distinction between a nonelectrolyte and a weak electrolyte is not important, as net ionic equations treat both types in the same way.

Nonelectrolytes are substances that go into solution as simple molecules. Most compounds containing only nonmetals (other than acids or bases) are nonelectrolytes. When deriving an ionic equation, from a molecular equation, nonelectrolytes are not separated into ions.

Electrolytes are materials that produce ions in solution. A strong electrolyte is 100% ionized (dissociated) in solution. In the above example, NaCl is a strong electrolyte. For this reason there is no NaCl in the solution, there are only Na⁺(aq) and Cl⁻(aq) ions.

A weak electrolyte, like a strong electrolyte, produces ions in solution. However, unlike a strong electrolyte only part of the material is ionized. A weak electrolyte is any electrolyte that is less than 100% dissociated. A weak electrolyte in solution is present both as undissociated molecules and as separated ions. In most cases, the fact that some of the molecules dissociate is ignored when converting a molecular equation to an ionic equation. In this way, nonelectrolytes and weak electrolytes appear the same in a net ionic equation.

The key to deriving net ionic equations from molecular equations is the identification of the strong electrolytes. There are three types of compounds that may be strong electrolytes:

1. **Strong Acids** (HCl, HBr, HI, HNO₃, H₂SO₄, HClO₃, and HClO₄)
2. **Strong Bases** ((Li, Na, K, Rb, Cs)OH, (Ca, Sr, Ba)(OH)₂)
3. **Soluble Salts** (any ionic compound that according to the solubility rules is soluble)

While there are other strong acids and bases, these are the only types normally encountered in introductory level chemistry courses. It is appropriate to consider all other acids or bases as being weak. When a molecular equation is converted to an ionic equation, all strong acids and bases are separated into their component ions (technically, only one of the H⁺ ions in H₂SO₄ is strong, and so the separation should be to H⁺ and HSO₄⁻). Do not separate any other acid or base.

The first thing a student must do, before attempting to write a net ionic equation, is to learn basic nomenclature. For example, the student must know that Na₂SO₄ contains Na⁺ ions and SO₄²⁻ ions, and not ions such as Na₂²⁺. The student must also know the strong electrolytes (strong acids, strong bases and soluble ionic salts). Any attempt to devise a net ionic equation without learning these two concepts can be very frustrating.
The following example will illustrate the various points required to derive a net ionic equation:

\[
\text{Na}_2\text{SO}_4 + \text{BaCl}_2 \rightarrow \text{BaSO}_4 + 2 \text{NaCl}
\]

The ions present are Na\(^+\), SO\(_4^{2-}\), Ba\(^{2+}\), and Cl\(^-\). The basic formulas and the assignment of the charges depend on the nomenclature rules. According to the solubility rules, all the compounds, in this example, except BaSO\(_4\) are soluble in water. In some cases, as a reminder, the student may wish to add the notation “(s)” after the formula of the barium sulfate. This symbolism indicates that this formula will remain unchanged throughout the remainder of the following process.

Separate all the soluble materials into their constituent ions. This gives a total ionic equation. Leave all insoluble materials unchanged.

\[
2 \text{Na}^+ + \text{SO}_4^{2-} + \text{Ba}^{2+} + 2 \text{Cl}^- \rightarrow \text{BaSO}_4(s) + 2 \text{Na}^+ + 2 \text{Cl}^-
\]

Do not forget to change subscripts into coefficients (Na\(_2\)SO\(_4\) and BaCl\(_2\)). Do not break polyatomic ions into their constituent atoms (SO\(_4^{2-}\)). Do not forget that a coefficient applies to everything in the formula that follows (2 NaCl).

The preceding equation is converted to a net ionic equation by eliminating 2 Na\(^+\) and 2 Cl\(^-\) from each side of the reaction arrow. These ions appear in identical form on each side of the arrow. If there had been 1 Na\(^+\) on one side, and 2 Na\(^+\) on the opposite side, then it would have been appropriate to eliminate 1 Na\(^+\) from each side. The unchanged ions are the spectator ions. After elimination of the spectator ions, the remaining equation is the net ionic equation. The net ionic equation for this reaction is:

\[
\text{SO}_4^{2-} + \text{Ba}^{2+} \rightarrow \text{BaSO}_4(s)
\]

You should check the following worked examples. Try to predict what the result of the next step will be before you look ahead. If your prediction is incorrect, make sure you understand the reason before proceeding. Assume all reactions take place in aqueous solution.

\[
\begin{align*}
2 \text{Al} + 6 \text{HBr} & \rightarrow 3 \text{H}_2 + 2 \text{AlBr}_3 \\
\text{ZnO} + 2 \text{HClO}_3 & \rightarrow \text{Zn(ClO}_3)_2 + \text{H}_2\text{O} \\
\text{Fe(OH)}_3 + 3 \text{HCl} & \rightarrow \text{FeCl}_3 + 3 \text{H}_2\text{O} \\
2 \text{NaCl} + (\text{NH}_4)_2\text{SO}_4 & \rightarrow \text{Na}_2\text{SO}_4 + 2 \text{NH}_4\text{Cl} \\
\text{Na}_2\text{CO}_3 + 2 \text{HNO}_3 & \rightarrow 2 \text{NaNO}_3 + \text{CO}_2 + \text{H}_2\text{O} \\
\text{MgBr}_2 + \text{Ca(OH)}_2 & \rightarrow \text{Mg(OH)}_2 + \text{CaBr}_2 \\
2 \text{HI} + \text{Ba(OH)}_2 & \rightarrow \text{BaI}_2 + 2 \text{H}_2\text{O}
\end{align*}
\]
These molecular equations change to the following total ionic equations:

2 Al + 6 H⁺ + 6 Br⁻ → 3 H₂ + 2 Al³⁺ + 6 Br⁻
HBr is a strong acid and AlBr₃ is soluble (Rule 1)
Al is a metal (neutral). H₂ is a molecular gas.
Spectator ions: Br⁻

ZnO + 2 H⁺ + 2 ClO₃⁻ → Zn²⁺ + 2 ClO₃⁻ + H₂O
HClO₃ is a strong acid and Zn(ClO₃)₂ is soluble (Rule 1)
ZnO is insoluble (Rule 3). H₂O is a molecular liquid.
Spectator ions: ClO₃⁻

Fe(OH)₃ + 3 H⁺ + 3 Cl⁻ → Fe³⁺ + 3 Cl⁻ + 3 H₂O
HCl is a strong acid and FeCl₃ is soluble (Rule 1)
Fe(OH)₃ is not a strong base. H₂O is a molecular liquid.
Spectator ions: Cl⁻

2 Na⁺ + 2 Cl⁻ + 2 NH₄⁺ + SO₄²⁻ → 2 Na⁺ + SO₄²⁻ + 2 NH₄⁺ + 2 Cl⁻
NaCl, (NH₄)₂SO₄, Na₂SO₄, and NH₄Cl are soluble (Rule 1 for all)
Spectator ions: Na⁺, NH₄⁺, Cl⁻, and SO₄²⁻
All ions are spectator ion, so there is no reaction

2 Na⁺ + CO₃²⁻ + 2 H⁺ + 2 NO₃⁻ → 2 Na⁺ + 2 NO₃⁻ + CO₂ + H₂O
Na₂CO₃ and NaNO₃ are soluble (Rule 1), and HNO₃ is a strong acid
CO₂ and H₂O are molecular compounds
Spectator ions: Na⁺ and NO₃⁻

Mg²⁺ + 2 Br⁻ + Ca²⁺ + 2 OH⁻ → Mg(OH)₂ + Ca²⁺ + 2 Br⁻
Ca(OH)₂ is a strong base, and MgBr₂ and CaBr₂ are soluble (Rule 1)
Mg(OH)₂ is not a strong base
Spectator ions: Br⁻ and Ca²⁺

2 H⁺ + 2 I⁻ + Ba²⁺ + 2 OH⁻ → Ba²⁺ + 2 I⁻ + 2 H₂O
HI is a strong acid, Ba(OH)₂ is a strong base and BaI₂ is soluble (Rule 1)
H₂O is a molecular liquid
Spectator ions: Ba²⁺ and I⁻
These total ionic equations become the following net ionic equations:

\[
\begin{align*}
2 \text{Al} + 6 \text{H}^+ & \rightarrow 3 \text{H}_2 + 2 \text{Al}^{3+} \\
\text{ZnO} + 2 \text{H}^+ & \rightarrow \text{Zn}^{2+} + \text{H}_2\text{O} \\
\text{Fe(OH)}_3 + 3 \text{H}^+ & \rightarrow \text{Fe}^{3+} + 3 \text{H}_2\text{O} \\
2 \text{NaCl} + (\text{NH}_4)_2\text{SO}_4 & \rightarrow \text{No Reaction (NR)}
\end{align*}
\]

\[
\begin{align*}
\text{CO}_3^{2-} + 2 \text{H}^+ & \rightarrow \text{CO}_2 + \text{H}_2\text{O} \\
\text{Mg}^{2+} + 2 \text{OH}^- & \rightarrow \text{Mg(OH)}_2 \\
2 \text{H}^+ + 2 \text{OH}^- & \rightarrow 2 \text{H}_2\text{O}
\end{align*}
\]

For the final (correct) answer, you must divide by 2.

\[
\text{H}^+ + \text{OH}^- \rightarrow \text{H}_2\text{O}
\]

In some equations, the entire reaction is not given. In these cases, it will be necessary to predict the parts of the equations that are not given. Initially, these predictions depend upon nomenclature.

The first step in predicting the results of a reaction is to determine which ions, if any, are present and assign charges to these ions. (Note: The ions may or may not separate in later steps.) The next step is to rearrange the ions so that the ions pair differently on the opposite side of the reaction arrow. The compounds on the opposite side of the arrow must contain both a cation and an anion. For example:

\[
\text{Al}_2(\text{SO}_4)_3 + \text{Pb(NO}_3)_2 \rightarrow
\]

The ions present are Al\(^{3+}\), SO\(_4^{2-}\), Pb\(^{2+}\), and NO\(_3^-\). The possible combinations on the opposite side are Al\(^{3+}\) with NO\(_3^-\) and Pb\(^{2+}\) with SO\(_4^{2-}\). The other combinations are either already present or involve two cations (Al\(^{3+}\) and Pb\(^{2+}\)) or two anions (SO\(_4^{2-}\) and NO\(_3^-\)) together. The total charge for a compound must be zero, the combination of two cations or of two anions is not possible. Do not forget to take the charges of the ions into account when deciding what formulas belong on the right side. The above equation becomes:

\[
\text{Al}_2(\text{SO}_4)_3 + \text{Pb(NO}_3)_2 \rightarrow \text{PbSO}_4 + \text{Al(NO}_3)_3
\]

This balances to

\[
\text{Al}_2(\text{SO}_4)_3 + 3 \text{Pb(NO}_3)_2 \rightarrow 3 \text{PbSO}_4 + 2 \text{Al(NO}_3)_3
\]
This is the molecular equation for the reaction. Based on the previous discussion, this converts to the following complete ionic and net ionic equations:

$$2 \text{Al}^{3+} + 3 \text{SO}_4^{2-} + 3 \text{Pb}^{2+} + 6 \text{NO}_3^- \rightarrow 3 \text{PbSO}_4 + 2 \text{Al}^{3+} + 6 \text{NO}_3^-$$

$$\text{SO}_4^{2-} + \text{Pb}^{2+} \rightarrow \text{PbSO}_4$$

You should check the following worked examples. Try to predict what the result of the next step will be before you look ahead. If your prediction is incorrect, make sure you understand the reason before proceeding. Assume all reactions take place in aqueous solution.

- **CoS + HCl** → CoCl₂ + H₂S
  Ions present on left: Co²⁺, S²⁻, H⁺ and Cl⁻

- **Na₂CO₃ + SrCl₂** → SrCO₃ + NaCl
  Ions present on left: Na⁺, CO₃²⁻, Sr²⁺, and Cl⁻

- **K₃PO₄ + HI** → KI + H₃PO₄
  Ions present on left: K⁺, PO₄³⁻, H⁺, and I⁻

- **Ba(OH)₂ + HClO₃** → Ba(ClO₃)₂ + H₂O
  Ions present on left: Ba²⁺, OH⁻, H⁺, and ClO₃⁻

- **NaBr + (NH₄)₃PO₄** → NH₄Br + Na₃PO₄
  Ions present on left: Na⁺, Br⁻, NH₄⁺, and PO₄³⁻
Ca(C$_2$H$_3$O$_2$)$_2$ + FeSO$_4$ → Fe(C$_2$H$_3$O$_2$)$_2$ + CaSO$_4$
Ions present on left: Ca$^{2+}$, C$_2$H$_3$O$_2^-$, Fe$^{2+}$, and SO$_4^{2-}$

Zn(NO$_3$)$_2$ + H$_2$S → ZnS + HNO$_3$
Ions present on left: Zn$^{2+}$, NO$_3^-$, H$^+$, and S$^{2-}$

Cr(OH)$_3$ + HClO$_4$ → Cr(ClO$_4$)$_3$ + H$_2$O
Ions present on left: Cr$^{3+}$, OH$^-$, H$^+$, and ClO$_4^-$

These equations balance to:

CoS + 2 HCl → CoCl$_2$ + H$_2$S

Na$_2$CO$_3$ + SrCl$_2$ → SrCO$_3$ + 2 NaCl

K$_3$PO$_4$ + 3 HI → 3 KI + H$_3$PO$_4$

Ba(OH)$_2$ + 2 HClO → Ba(ClO$_3$)$_2$ + 2 H$_2$O

3 NaBr + (NH$_4$)$_3$PO$_4$ → 3 NH$_4$Br + Na$_3$PO$_4$

Ca(C$_2$H$_3$O$_2$)$_2$ + FeSO$_4$ → Fe(C$_2$H$_3$O$_2$)$_2$ + CaSO$_4$

Zn(NO$_3$)$_2$ + H$_2$S → ZnS + 2 HNO$_3$

Cr(OH)$_3$ + 3 HClO$_4$ → Cr(ClO$_4$)$_3$ + 3 H$_2$O

The total ionic equations are:

CoS + 2 H$^+$ + 2 Cl$^-$ → Co$^{2+}$ + 2 Cl$^-$ + H$_2$S

2 Na$^+$ + CO$_3^{2-}$ + Sr$^{2+}$ + 2 Cl$^-$ → SrCO$_3$ + 2 Na$^+$ + 2 Cl$^-$

3 K$^+$ + PO$_4^{3-}$ + 3 H$^+$ + 3 I$^-$ → 3 K$^+$ + 3 I$^-$ + H$_3$PO$_4$

Ba$^{2+}$ + 2 OH$^-$ + 2 H$^+$ + 2 ClO$_3^-$ → Ba$^{2+}$ + 2 ClO$_3^-$ + 2 H$_2$O

3 Na$^+$ + 3 Br$^-$ + 3 NH$_4^+$ + PO$_4^{3-}$ → 3 NH$_4^+$ + 3 Br$^-$ + 3 Na$^+$ + PO$_4^{3-}$

Ca$^{2+}$ + 2 C$_2$H$_3$O$_2^-$ + Fe$^{2+}$ + SO$_4^{2-}$ → Fe$^{2+}$ + 2 C$_2$H$_3$O$_2^-$ + CaSO$_4$

Zn$^{2+}$ + 2 NO$_3^-$ + H$_2$S → ZnS + 2 H$^+$ + 2 NO$_3^-$

Cr(OH)$_3$ + 3 H$^+$ + 3 ClO$_4^-$ → Cr$^{3+}$ + 3 ClO$_4^-$ + 3 H$_2$O
The final net ionic equations are:

\[
\begin{align*}
    \text{CoS} + 2 \text{H}^+ & \rightarrow \text{Co}^{2+} + \text{H}_2\text{S} \\
    \text{CO}_3^{2-} + \text{Sr}^{2+} & \rightarrow \text{SrCO}_3 \\
    \text{PO}_4^{3-} + 3 \text{H}^+ & \rightarrow \text{H}_3\text{PO}_4 \\
    \text{OH}^- + \text{H}^+ & \rightarrow \text{H}_2\text{O} \\
    \text{No Reaction} \\
    \text{Ca}^{2+} + \text{SO}_4^{2-} & \rightarrow \text{CaSO}_4 \\
    \text{Zn}^{2+} + \text{H}_2\text{S} & \rightarrow \text{ZnS} + 2 \text{H}^+ \\
    \text{Cr(OH)}_3 + 3 \text{H}^+ & \rightarrow \text{Cr}^{3+} + 3 \text{H}_2\text{O}
\end{align*}
\]
The final category, where net ionic equations are often seen, are redox equations. In these redox reactions, oxidation (a loss of electrons) accompanies a reduction (a gain of electrons). These processes are coupled so that the electrons lost are exactly equal to the electrons gained. One of the simplest types of redox reactions utilizes an activity series. An activity series lists reactions showing how various metals and hydrogen oxidize in aqueous solution. Elements at the beginning of the series are more active than elements below. The more active elements have a stronger tendency to oxidize than the less active elements. The less active elements tend to reduce instead of oxidize. The reduction reactions are the reverse of the oxidation reactions given in the activity series table.

A typical activity series table appears below:

<table>
<thead>
<tr>
<th>Activity Series of Metals</th>
<th>Continued from bottom of preceding column</th>
</tr>
</thead>
<tbody>
<tr>
<td>Li $\rightarrow$ Li$^+$  + e$^-$</td>
<td>Ni $\rightarrow$ Ni$^{2+}$  + 2 e$^-$</td>
</tr>
<tr>
<td>K $\rightarrow$ K$^+$  + e$^-$</td>
<td>Sn $\rightarrow$ Sn$^{2+}$  + 2 e$^-$</td>
</tr>
<tr>
<td>Ba $\rightarrow$ Ba$^{2+}$  + 2 e$^-$</td>
<td>Pb $\rightarrow$ Pb$^{2+}$  + 2 e$^-$</td>
</tr>
<tr>
<td>Sr $\rightarrow$ Sr$^{2+}$  + 2 e$^-$</td>
<td>H$_2$ $\rightarrow$ 2 H$^+$  + 2 e$^-$</td>
</tr>
<tr>
<td>Ca $\rightarrow$ Ca$^{2+}$  + 2 e$^-$</td>
<td>Cu $\rightarrow$ Cu$^{2+}$  + 2 e$^-$</td>
</tr>
<tr>
<td>Na $\rightarrow$ Na$^+$  + e$^-$</td>
<td>Ag $\rightarrow$ Ag$^+$  + e$^-$</td>
</tr>
<tr>
<td>Mg $\rightarrow$ Mg$^{2+}$  + 2 e$^-$</td>
<td>Hg $\rightarrow$ Hg$^{2+}$  + 2 e$^-$</td>
</tr>
<tr>
<td>Al $\rightarrow$ Al$^{3+}$  + 3 e$^-$</td>
<td>Pd $\rightarrow$ Pd$^{2+}$  + 2 e$^-$</td>
</tr>
<tr>
<td>Mn $\rightarrow$ Mn$^{2+}$  + 2 e$^-$</td>
<td>Pt $\rightarrow$ Pt$^{2+}$  + 2 e$^-$</td>
</tr>
<tr>
<td>Zn $\rightarrow$ Zn$^{2+}$  + 2 e$^-$</td>
<td>Au $\rightarrow$ Au$^{3+}$  + 3 e$^-$</td>
</tr>
<tr>
<td>Cr $\rightarrow$ Cr$^{3+}$  + 3 e$^-$</td>
<td>Fe $\rightarrow$ Fe$^{2+}$  + 2 e$^-$</td>
</tr>
<tr>
<td>Cd $\rightarrow$ Cd$^{2+}$  + 2 e$^-$</td>
<td>Co $\rightarrow$ Co$^{2+}$  + 2 e$^-$</td>
</tr>
<tr>
<td>V $\rightarrow$ V$^{3+}$  + 3 e$^-$</td>
<td></td>
</tr>
</tbody>
</table>

A problem of this type might appear as: “Write the balanced molecular and net ionic equation for the reaction of chromium metal with sulfuric acid.” The solution to this problem would begin as:

\[
\text{Cr} + \text{H}_2\text{SO}_4 \rightarrow
\]

This partial equation contains Cr, H$^+$, and SO$_4^{2-}$. Both Cr and H$^+$ appear in the activity series table. The reactions involving these ions and their relative positions noted. From the above table this gives:

\[
\begin{align*}
\text{Cr} & \rightarrow \text{Cr}^{3+}  + 3 \text{e}^- \\
\text{H}_2 & \rightarrow 2 \text{H}^+  + 2 \text{e}^- 
\end{align*}
\]

There are two important facts from the table: the Cr is above the H$^+$ on the table; the Cr and the H$^+$ are on opposite sides of the reaction arrows. Since Cr is higher on the table, it will oxidize. (Note: only
Cr, not Cr\(^{3+}\), undergoes oxidation. Thus if Cr\(^{3+}\) were present instead of Cr, this discussion is no longer valid.) If a substance undergoes oxidation then something must also undergo reduction. In this problem, the only option for reduction is the H\(^+\). The sulfate ion is not important to this problem since it does not appear in the activity series table.

Written as a reduction, the H\(^+\) reaction becomes:

\[
2 \text{H}^+ + 2 e^- \rightarrow \text{H}_2
\]

This reaction combines with the Cr reaction.

\[
2 (\text{Cr} \rightarrow \text{Cr}^{3+} + 3 e^-) \\
3 (2 \text{H}^+ + 2 e^- \rightarrow \text{H}_2)
\]

The numbers (2 and 3) in front of the parentheses are multipliers to make sure that the total electrons on the right side are identical to the total electrons on the left side. Multiplying gives the following reactions:

\[
2 \text{Cr} \rightarrow 2 \text{Cr}^{3+} + 6 e^- \\
6 \text{H}^+ + 6 e^- \rightarrow 3 \text{H}_2
\]

These add to give:

\[
2 \text{Cr} + 6 \text{H}^+ + 6 e^- \rightarrow 2 \text{Cr}^{3+} + 3 \text{H}_2 + 6 e^- 
\]

The equal numbers of electrons on each side cancel to give the net ionic equation:

\[
2 \text{Cr} + 6 \text{H}^+ \rightarrow 2 \text{Cr}^{3+} + 3 \text{H}_2 
\]

The presence of 6 H\(^+\) requires 3 H\(_2\)SO\(_4\). The three additional SO\(_4^{2-}\) ions in the sulfuric acid require an equal number on the right side to keep both sides balanced. Adding three sulfate ions to each side give the total ionic equation:

\[
2 \text{Cr} + 6 \text{H}^+ + 3 \text{SO}_4^{2-} \rightarrow 2 \text{Cr}^{3+} + 3 \text{SO}_4^{2-} + 3 \text{H}_2 
\]

Combing ions gives the molecular equation:

\[
2 \text{Cr} + 3 \text{H}_2\text{SO}_4 \rightarrow \text{Cr}_2(\text{SO}_4)_3 + 3 \text{H}_2 
\]
You should check the following worked examples. Try to predict what the result of the next step will be before you look ahead. If your prediction is incorrect, make sure you understand the reason before proceeding. Assume all reactions take place in aqueous solution.

\[
\begin{align*}
\text{HCl} + \text{Ni} & \rightarrow \\
\text{CuBr}_2 + \text{Zn} & \rightarrow \\
\text{Fe} + \text{H}_2\text{SO}_4 & \rightarrow \\
\text{Mg} + \text{HC}_2\text{H}_3\text{O}_2 & \rightarrow \\
\text{H}_2 + \text{FeCl}_2 & \rightarrow \\
\text{Al}_2(\text{SO}_4)_3 + \text{Fe} & \rightarrow \\
\text{AgNO}_3 + \text{Zn} & \rightarrow \\
\text{Li} + \text{H}_2\text{O} & \rightarrow 
\end{align*}
\]

The elements and cations present are given below, and their relative positions. Note: The activity series does not include anions.

\[
\begin{align*}
\text{HCl} + \text{Ni} & \rightarrow \text{H}^+ \text{ below Ni} \text{—reaction possible} \\
\text{CuBr}_2 + \text{Zn} & \rightarrow \text{Cu}^{2+} \text{ below Zn} \text{—reaction possible} \\
\text{Fe} + \text{H}_2\text{SO}_4 & \rightarrow \text{H}^+ \text{ below Fe} \text{—reaction possible} \\
\text{Mg} + \text{HC}_2\text{H}_3\text{O}_2 & \rightarrow \text{H}^+ \text{ below Mg} \text{—reaction possible} \\
\text{H}_2 + \text{FeCl}_2 & \rightarrow \text{Fe}^{2+} \text{ above } \text{H}_2 \text{—no reaction possible (NR)} \\
\text{Al}_2(\text{SO}_4)_3 + \text{Fe} & \rightarrow \text{Al}^{3+} \text{ above Fe} \text{—no reaction possible (NR)} \\
\text{AgNO}_3 + \text{Zn} & \rightarrow \text{Ag}^+ \text{ below Zn} \text{—reaction possible} \\
\text{Li} + \text{H}_2\text{O} & \rightarrow \text{H}^+ \text{ below Li} \text{—reaction possible} \\
\text{( Treat } \text{H}_2\text{O}, \text{ in this case, as } \text{H}^+ + \text{OH}^- \text{)}
\end{align*}
\]
The following equations derive from the activity series. Note that it is necessary to reverse some of the reactions. In addition, it is necessary to use multipliers in some cases. As always, you should try to understand each one before moving on the next.

1. \[
\text{HCl} + \text{Ni} \rightarrow 2 \text{H}^+ + 2 \text{e}^- \rightarrow \text{H}_2 \\
\text{Ni} \rightarrow \text{Ni}^{2+} + 2 \text{e}^-
\]

2. \[
\text{CuBr}_2 + \text{Zn} \rightarrow \text{Cu}^{2+} + 2 \text{e}^- \rightarrow \text{Cu} \\
\text{Zn} \rightarrow \text{Zn}^{2+} + 2 \text{e}^-
\]

3. \[
\text{Fe} + \text{H}_2\text{SO}_4 \rightarrow 2 \text{H}^+ + 2 \text{e}^- \rightarrow \text{H}_2 \\
\text{Fe} \rightarrow \text{Fe}^{2+} + 2 \text{e}^-
\]

4. \[
\text{Mg} + \text{HC}_2\text{H}_3\text{O}_2 \rightarrow 2 \text{H}^+ + 2 \text{e}^- \rightarrow \text{H}_2 \\
\text{Mg} \rightarrow \text{Mg}^{2+} + 2 \text{e}^-
\]

5. \[
\text{H}_2 + \text{FeCl}_2 \rightarrow \text{NR}
\]

6. \[
\text{Al}_2(\text{SO}_4)_3 + \text{Fe} \rightarrow \text{NR}
\]

7. \[
\text{AgNO}_3 + \text{Zn} \rightarrow 2(\text{Ag}^{+} + \text{e}^- \rightarrow \text{Ag}) \\
\text{Zn} \rightarrow \text{Zn}^{2+} + 2 \text{e}^-
\]

8. \[
\text{Li} + \text{H}_2\text{O} \rightarrow 2 \text{H}^+ + 2 \text{e}^- \rightarrow \text{H}_2 \\
2(\text{Li} \rightarrow \text{Li}^{+} + \text{e}^-)
\]

The above equations now combine to give the following equations. Some of these equations are net ionic equations, but some are not.

\[
\text{HCl} + \text{Ni} \rightarrow 2 \text{H}^+ + \text{Ni} \rightarrow \text{H}_2 + \text{Ni}^{2+}
\]

\[
\text{CuBr}_2 + \text{Zn} \rightarrow \text{Cu}^{2+} + \text{Zn} \rightarrow \text{Cu} + \text{Zn}^{2+}
\]

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

\[
\text{Mg} + \text{HC}_2\text{H}_3\text{O}_2 \rightarrow 2 \text{H}^+ + \text{Mg} \rightarrow \text{H}_2 + \text{Mg}^{2+}
\]
\[
\begin{align*}
H_2 + FeCl_2 & \rightarrow \text{NR} \\
Al_2(SO_4)_3 + Fe & \rightarrow \text{NR} \\
AgNO_3 + Zn & \rightarrow \ \\
2Ag^+ + Zn & \rightarrow \ 2Ag + Zn^{2+} \\
Li + H_2O & \rightarrow \ \\
2H^+ + 2Li & \rightarrow \ H_2 + 2Li^+ \\
\end{align*}
\]

The molecular equations are:

\[
\begin{align*}
HCl + Ni & \rightarrow \text{net ionic} \\
2H^+ + Ni & \rightarrow \text{net ionic} \\
2HCl + Ni & \rightarrow \ H_2 + NiCl_2 \text{ molecular} \\
CuBr_2 + Zn & \rightarrow \text{net ionic} \\
Cu^{2+} + Zn & \rightarrow \text{net ionic} \\
CuBr_2 + Zn & \rightarrow \text{net ionic} \\
\end{align*}
\]

\[
\begin{align*}
2H^+ + Fe & \rightarrow \text{net ionic} \\
H_2SO_4 + Fe & \rightarrow \text{net ionic} \\
\end{align*}
\]

\[
\begin{align*}
Mg + HC_2H_3O_2 & \rightarrow \text{net ionic} \\
2H^+ + Mg & \rightarrow \text{net ionic} \\
2HC_2H_3O_2 + Mg & \rightarrow \text{net ionic} \\
\end{align*}
\]

\[
\begin{align*}
H_2 + FeCl_2 & \rightarrow \text{NR} \\
Al_2(SO_4)_3 + Fe & \rightarrow \text{NR} \\
AgNO_3 + Zn & \rightarrow \text{net ionic} \\
2AgNO_3 + Zn & \rightarrow \text{net ionic} \\
Li + H_2O & \rightarrow \text{net ionic} \\
2H^+ + 2Li & \rightarrow \text{net ionic} \\
2H_2O + 2Li & \rightarrow \text{net ionic} \\
\end{align*}
\]

(\text{Remember H}_2\text{O is H}^+\text{ and OH}^-)
Common errors:

1. There are three types of equations involved: Molecular, Total Ionic and Net Ionic. A problem may ask for one or more of these. Many students lose points because either they give the wrong type of equation, or they do not give all the types the problem requests.

2. The term “Ionic” in both Total Ionic and Net Ionic equations means that there are ions present. Ions have charges. Many students lose points because they do not show any charges. While not everything will have a charge, there must be at least two species with charges present in any ionic equation. Conversely, Molecular equations must never show any charges.

3. Many students forget that charges as well as atoms must balance. The total of charges on each side of the reaction arrow must be identical. Many students turn in papers with answers such as: a. +1 - 2 = 0, b. 0 = 4(+1), and c. +2 - 2(-1) = +2.

4. There is no acceptable reason to ignore simple nomenclature. The total charge of all the ions in a compound must be 0. There are absolutely no exceptions.

5. Compounds and elements do not have charges. For example, if you to use aluminum in a reaction, it must appear as Al not as Al$^{3+}$. You can prove that aluminum is Al and not Al$^{3+}$ by picking up a piece of aluminum foil. If it were really Al$^{3+}$, you would get a shock greater than a normal bolt of lightning. Al$^{3+}$ may be present in compounds or in a solution, but nowhere else.

6. Finally, the most common mistake on a recent exam: The formula for water is H$_2$O. The only other compound of hydrogen and oxygen is hydrogen peroxide (H$_2$O$_2$), and it seldom appears in this type of problem.