WRITING CHEMICAL EQUATIONS & PREDICTING CHEMICAL REACTIONS

Writing Chemical Equations

Chemical equations are symbolic representations of chemical and physical processes that use standardized notation to communicate information in an efficient manner. These equations follow the general form:

\[ \text{Reactants} \rightarrow \text{Products} \]

where the arrow indicates some change that occurs as a result of the process.

The first step in writing a chemical equation is to write a detailed description of the process that was observed, including the substances involved and their physical states. For example:

*Aluminum metal dissolves in an aqueous solution of sulfuric acid to form hydrogen gas bubbles and a solution of aluminum sulfate.*

The reactants and products are then identified along with their states of matter:

*Aluminum (s) + Sulfuric acid (aq) → Hydrogen (g) + Aluminum sulfate (aq)*

Each substance is then represented by its correct chemical formula. For ionic compounds (such as aluminum sulfate), the subscripts in the formula are determined by charge balancing (two Al\(^{3+}\) are needed to balance three SO\(_4^{2-}\)):

\[ \text{Al (s) + H}_2\text{SO}_4 (aq) \rightarrow \text{H}_2 (g) + \text{Al}_2(\text{SO}_4)_3 (aq) \]

Finally, the reaction equation must be properly balanced so that the same number of atoms of each elemental type appear on both sides of the equation (i.e., conservation of atoms):

\[ 2 \text{Al (s)} + 3 \text{H}_2\text{SO}_4 (aq) \rightarrow 3 \text{H}_2 (g) + \text{Al}_2(\text{SO}_4)_3 (aq) \]

In some cases, additional information may be included in the reaction equation. For example, an important catalyst, solvent, or other necessary reaction conditions may be indicated above or below the reaction arrow. Also, the energetics of the reaction may be indicated by including the term *energy* as a reactant or product, or alternatively, as we will typically do in this class, by sometimes indicating the *energy change* (\(\Delta H\)) next to the reaction. Remember that positive energy change indicates that an input of energy is required for the reaction to proceed (*endothermic*), while negative energy change indicates that energy (usually in the form of heat) is produced as a by-product of the reaction (*exothermic*):

\[ 2 \text{Al (s)} + 3 \text{H}_2\text{SO}_4 (aq) \rightarrow 3 \text{H}_2 (g) + \text{Al}_2(\text{SO}_4)_3 (aq) \quad \Delta H < 0 \]

*Chemists usually express energy in a certain form known as enthalpy (H), as we will discuss in Ch 5.*
Practice Writing Chemical Equations

For each of the following, write a balanced chemical reaction equation (including states of matter) that accurately represents the process described:

1) When dissolved beryllium chloride reacts with dissolved silver nitrate in water, aqueous beryllium nitrate and silver chloride powder are made.

\[
\text{BeCl}_2\text{(aq)} + 2 \text{AgNO}_3\text{(aq)} \rightarrow \text{Be(NO}_3\text{)}_2\text{(aq)} + 2\text{AgCl (s)}
\]

2) When isopropanol (C₃H₈O) burns in oxygen, carbon dioxide, water, and heat are produced.

\[
2 \text{C}_3\text{H}_8\text{O (l)} + 9 \text{O}_2\text{(g)} \rightarrow 6 \text{CO}_2\text{(g)} + 8 \text{H}_2\text{O (g)} \quad \Delta H < 0
\]

3) When dissolved sodium hydroxide reacts with sulfuric acid (H₂SO₄), aqueous sodium sulfate, water, and heat are formed.

\[
2 \text{NaOH}\text{(aq)} + \text{H}_2\text{SO}_4\text{(aq)} \rightarrow \text{Na}_2\text{SO}_4\text{(aq)} + 2 \text{H}_2\text{O (l)} \quad \Delta H < 0
\]

4) When fluorine gas is put into contact with calcium metal at high temperatures, calcium fluoride powder is created in an exothermic reaction.

\[
\text{F}_2\text{(g)} + \text{Ca (s)} \rightarrow \text{CaF}_2\text{(s)} \quad \Delta H < 0
\]

5) When sodium metal reacts with iron (II) chloride, iron metal and sodium chloride are formed.

\[
2 \text{Na (s)} + \text{FeCl}_2\text{(s)} \rightarrow \text{Fe (s)} + 2 \text{NaCl (s)}
\]
Visualizing Chemical Reactions

Chemical processes can be visualized at the atomic level and represented using pictorial schematics. Such visualizations can help to communicate and reinforce important chemical concepts of both a qualitative and quantitative nature, including stoichiometric relationships.

Consider as an example, for the hypothetical elements A and B, that excess elemental substance A₂ reacts with a limiting amount of elemental substance B₂, both in the gas phase, to produce the compound AB₂, which is a liquid under these conditions. We could represent this process with the following balanced reaction equation:

\[
A_2 (g) + 2 B_2 (g) \rightarrow 2 AB_2 (l)
\]

We can also represent this process schematically as shown below. Note that physical states (gas, liquid) are clearly represented in the drawing, along with various quantitative relationships (molecular formulas, reaction stoichiometry, limiting/excess reactant).

Practice Visualizing Chemical Reactions

1) Use the boxes at right to create a detailed pictorial schematic of the chemical process represented by the reaction equation shown below:

\[
A_2B (g) + 2 C (s) \rightarrow AC_2 (g) + AB (g)
\]

2) Write a balanced reaction equation for the process that is pictured at right:

\[
2 AB (g) + B_2 (g) \rightarrow 2 AB_2 (g)
\]

Note simplification of B₂ to one side of equation, and use of lowest whole number ratios for coefficients.
Predicting Chemical Reactions

The ultimate goal of scientific understanding is the ability to predict the outcomes of our actions. For chemistry, this usually means the ability to predict the outcomes of chemical and physical processes under various conditions.

One major aid in predicting the outcome of chemical reactions is the classification of reactions into various types based on similarities. One classification scheme groups reactions according to similarities in the general patterns of atomic rearrangement, as described below (note that the letters can represent single atoms or groups of atoms):

**Synthesis**

\[ A + B \rightarrow C \]

Examples:
- \( 4 \text{ Fe (s)} + 3 \text{ O}_2 (g) \rightarrow 2 \text{ Fe}_2\text{O}_3 (s) \)
- \( \text{CO}_2 (g) + \text{H}_2\text{O (l)} \rightarrow \text{H}_2\text{CO}_3 (aq) \)

**Decomposition**

\[ A \rightarrow B + C \]

Examples:
- \( 2 \text{ H}_2\text{O (l)} \rightarrow 2 \text{ H}_2 (g) + \text{O}_2 (g) \)
- \( \text{CaCO}_3 (s) \rightarrow \text{CaO (s)} + \text{CO}_2 (g) \)

**Single Displacement**

\[ A + BC \rightarrow AC + B \]

Examples:
- \( \text{C (s)} + \text{Fe}_2\text{O}_3 (s) \rightarrow \text{CO (g)} + \text{Fe (s)} \)
- \( \text{Zn (s)} + \text{CuCl}_2 (aq) \rightarrow \text{ZnCl}_2 (aq) + \text{Cu (s)} \)

**Metathesis (Exchange)**

\[ AB + CD \rightarrow AD + CB \]

Examples:
- \( \text{HF (aq)} + \text{NaOH (aq)} \rightarrow \text{H}_2\text{O (l)} + \text{NaF (aq)} \)
- \( \text{FeCl}_2 (aq) + \text{Na}_2\text{S (aq)} \rightarrow \text{FeS (s)} + \text{NaCl (aq)} \)

Other classification schemes group reactions according to more specific chemical changes that occur and/or more specific types of compounds that are involved. Two very important classes of reactions that we will study in this course are oxidation-reduction (‘redox’) reactions and aqueous double-displacement (ionic metathesis) reactions, both of which are described in more detail below.
Aqueous Metathesis (Double-Displacement) Reactions

All aqueous double displacement reactions follow the general pattern:

\[
AX + BY \rightarrow AY + BX
\]

where A and B are cations and X and Y are anions. For the reaction to result in a net change in the solution, ions must be removed from solution. Three different processes can lead to such ion removal and therefore provide a driving force for reaction:

1. Formation of a precipitate (insoluble or slightly soluble solid)
2. Formation of a weak or non-electrolyte (e.g., a weak acid or H₂O)
3. Formation of a gas that escapes from solution

It is customary to simplify metathesis reaction equations by expressing them in net ionic form. This involves first writing the complete ionic equation by expressing all soluble (aqueous) ionic compounds as ions, and then removing any spectator ions that appear on both sides of the reaction equation. Remember that insoluble solids, weak or non-electrolytes, and gases do not exist as ions in solution (and therefore should not be expressed as ions in the complete and net ionic equations).

Predicting Precipitation Reactions

Note that in order to predict precipitation reactions, you must be familiar with the solubility guidelines for ionic compounds in water. Refer to the table in your text.

Example: barium chloride + silver nitrate

1. Identify symbols & charges: \( \text{Ba}^{2+}, \; \text{Cl}^{-}, \; \text{Ag}^{+}, \; \text{NO}_3^{-} \)
2. Write formulas of reactants: \( \text{BaCl}_2 \text{(aq)}, \; \text{AgNO}_3 \text{(aq)} \)
3. Write formulas of products: \( \text{Ba(NO}_3)_2 \text{(?)}, \; \text{AgCl (?) (swap ions)} \)
4. Determine solubilities: \( \text{Ba(NO}_3)_2 \text{(aq)}, \; \text{AgCl (s)} \)
5. Write and balance eqn: \( \text{BaCl}_2 \text{(aq)} + 2 \text{AgNO}_3 \text{(aq)} \rightarrow \text{Ba(NO}_3)_2 \text{(aq)} + 2 \text{AgCl (s)} \)
6. Write complete ionic eqn: \( \text{Ba}^{2+} + 2 \text{Cl}^{-} + 2 \text{Ag}^{+} + 2 \text{NO}_3^{-} \rightarrow \text{Ba}^{2+} + 2 \text{NO}_3^{-} + 2 \text{AgCl (s)} \)
7. Write net ionic eqn: \( \text{Ag}^{+} + \text{Cl}^{-} \rightarrow \text{AgCl (s)} \)
Practice Predicting Precipitation Reactions

Write molecular, complete ionic, and net ionic equations for each of the following:

1) lead(II) nitrate + potassium iodide $\rightarrow$

\[
Pb(NO_3)_2 (aq) + 2 KI (aq) \rightarrow PbI_2 (s) + 2 KNO_3 (aq)
\]

\[
Pb^{2+} + 2 NO_3^- + 2 K^+ + 2 I^- \rightarrow PbI_2 (s) + 2 K^+ + 2 NO_3^-
\]

\[
Pb^{2+} + 2 I^- \rightarrow PbI_2 (s)
\]

2) copper(II) sulfate + barium chloride $\rightarrow$

\[
CuSO_4 (aq) + BaCl_2 (aq) \rightarrow CuCl_2 (aq) + BaSO_4 (s)
\]

\[
Cu^{2+} + SO_4^{2-} + Ba^{2+} + 2Cl^- \rightarrow CuCl_2 (s) + Ba^{2+} + SO_4^{2-}
\]

\[
Ba^{2+} + SO_4^{2-} \rightarrow BaSO_4 (s)
\]

3) iron(II) sulfate + sodium carbonate $\rightarrow$

\[
FeSO_4 (aq) + Na_2CO_3 (aq) \rightarrow FeCO_3 (s) + Na_2SO_4 (aq)
\]

\[
Fe^{2+} + SO_4^{2-} + 2Na^+ + CO_3^{2-} \rightarrow FeCO_3 (s) + 2Na^+ + SO_4^{2-}
\]

\[
Fe^{2+} + CO_3^{2-} \rightarrow FeCO_3 (s)
\]

4) potassium bromide + sodium nitrate $\rightarrow$

\[
KBr (aq) + NaNO_3 (aq) \rightarrow KNO_3 (aq) + NaBr (aq)
\]

\[
K^+ + Br^- + Na^+ + NO_3^- \rightarrow K^+ + NO_3^- + Na^+ + Br^-
\]

**No Reaction**

5) nickel(II) chloride + potassium hydroxide $\rightarrow$

\[
NiCl_2 (aq) + 2 KOH (aq) \rightarrow 2 KCl (aq) + Ni(OH)_2 (s)
\]

\[
Ni^{2+} + 2 Cl^- + 2 K^+ + 2 OH^- \rightarrow 2 K^+ + 2 Cl^- + Ni(OH)_2 (s)
\]

\[
Ni^{2+} + 2 OH^- \rightarrow Ni(OH)_2 (s)
\]
Predicting Acid-Base Reactions

Remember that an acid is any substance that can act as an H⁺ donor, while a base is any substance that can act as an H⁺ acceptor. Strong acids are those that dissociate (donate their H⁺) almost completely in water, while weak acids remain almost completely un-ionized (i.e., in molecular form).

The neutralization of an acid with a strong base will result in the formation of water (a non-electrolyte), and hence represents a metathesis reaction:

Molecular: \[ \text{HCl (aq)} + \text{NaOH (aq)} \rightarrow \text{H₂O (l)} + \text{NaCl (aq)} \]

Complete Ionic: \[ \text{H}^+ + \text{Cl}^- + \text{Na}^+ + \text{OH}^- \rightarrow \text{H₂O (l)} + \text{Na}^+ + \text{Cl}^- \]

Net Ionic: \[ \text{H}^+ + \text{OH}^- \rightarrow \text{H₂O (l)} \]

Note that there is often no visible evidence that such a reaction occurred. Since acid-base neutralizations are exothermic, the solution warms slightly, but this may be difficult to detect in practice.

A strong acid can also react to form a weak acid, which is a weak electrolyte. This also provides a driving force for ion removal from the solution, and hence results in a metathesis reaction:

Molecular: \[ \text{HCl (aq)} + \text{NaF (aq)} \rightarrow \text{HF (aq)} + \text{NaCl (aq)} \]

Complete Ionic: \[ \text{H}^+ + \text{Cl}^- + \text{Na}^+ + \text{F}^- \rightarrow \text{HF (aq)} + \text{Na}^+ + \text{Cl}^- \]

Net Ionic: \[ \text{H}^+ + \text{F}^- \rightarrow \text{HF (aq)} \]

In order to recognize whether a substance is a weak acid, it is helpful to memorize the 7 strong acids: HCl, HBr, HI, HNO₃, H₂SO₄, HClO₄, HClO₃. All other acids are weak.

Predicting Gas-Forming Acid-Base Reactions

Some weak acids that form in metathesis reactions decompose to form gases that are only slightly soluble in water. Other weak acids are themselves slightly soluble gases. In either case, as the gases leave the solution, they provide a driving force for further reaction that removes ions from solution:

- \( \text{NH}_4\text{OH} \) unstable, breaks down into \( \text{H}_2\text{O} \) and \( \text{NH}_3 \) (g)
- \( \text{H}_2\text{CO}_3 \) unstable, breaks down into \( \text{H}_2\text{O} \) and \( \text{CO}_2 \) (g)
- \( \text{H}_2\text{SO}_3 \) unstable, breaks down into \( \text{H}_2\text{O} \) and \( \text{SO}_2 \) (g)
- \( \text{H}_2\text{S} \) slightly soluble gas
Practice Predicting Acid-Base Reactions

Write molecular, complete ionic, and net ionic equations for each of the following:

1) sulfuric acid + potassium hydroxide →

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

\[ 2 \text{H}^+ + \text{SO}_4^{2-} + 2 \text{K}^+ + 2 \text{OH}^- \rightarrow 2 \text{H}_2\text{O}(\ell) + 2 \text{K}^+ + \text{SO}_4^{2-} \]

\[ 2 \text{H}^+ + 2 \text{OH}^- \rightarrow 2 \text{H}_2\text{O}(\ell), \text{ which reduces to } \text{H}^+ + \text{OH}^- \rightarrow \text{H}_2\text{O}(\ell) \]

2) sodium carbonate + hydrochloric acid →

\[ \text{Na}_2\text{CO}_3(\text{aq}) + 2 \text{HCl}(\text{aq}) \rightarrow 2 \text{NaCl}(\text{aq}) + \text{H}_2\text{O}(\ell) + \text{CO}_2(\text{g}) \]

\[ 2 \text{Na}^+ + \text{CO}_3^{2-} + 2 \text{H}^+ + 2 \text{Cl}^- \rightarrow 2 \text{Na}^+ + 2 \text{Cl}^- + \text{H}_2\text{O}(\ell) + \text{CO}_2(\text{g}) \]

\[ \text{CO}_3^{2-} + 2 \text{H}^+ \rightarrow \text{H}_2\text{O}(\ell) + \text{CO}_2(\text{g}) \]

3) hydrobromic acid + sodium acetate →

\[ \text{HBr}(\text{aq}) + \text{NaCH}_3\text{CO}_2(\text{aq}) \rightarrow \text{HCH}_3\text{CO}_2(\text{aq}) + \text{NaBr}(\text{aq}) \]

\[ \text{H}^+ + \text{Br}^- + \text{Na}^+ + \text{CH}_3\text{CO}_2^- \rightarrow \text{HCH}_3\text{CO}_2(\text{aq}) + \text{Na}^+ + \text{Br}^- \]

\[ \text{H}^+ + \text{CH}_3\text{CO}_2^- \rightarrow \text{HCH}_3\text{CO}_2(\text{aq}) \]

4) ammonium chloride + sodium hydroxide →

\[ \text{NH}_4\text{Cl}(\text{aq}) + \text{NaOH}(\text{aq}) \rightarrow \text{NaCl}(\text{aq}) + \text{H}_2\text{O}(\ell) + \text{NH}_3(\text{g}) \]

\[ \text{NH}_4^+ + \text{Cl}^- + \text{Na}^+ + \text{OH}^- \rightarrow \text{Na}^+ + \text{Cl}^- + \text{H}_2\text{O}(\ell) + \text{NH}_3(\text{g}) \]

\[ \text{NH}_4^+ + \text{OH}^- \rightarrow \text{H}_2\text{O}(\ell) + \text{NH}_3(\text{g}) \]

5) potassium phosphate + hydrochloric acid →

\[ \text{K}_3\text{PO}_4(\text{aq}) + 3 \text{HCl}(\text{aq}) \rightarrow \text{H}_3\text{PO}_4(\text{aq}) + 3 \text{KCl}(\text{aq}) \]

\[ 3 \text{K}^+ + \text{PO}_4^{3-} + 3 \text{H}^+ + 3 \text{Cl}^- \rightarrow \text{H}_3\text{PO}_4(\text{aq}) + 3 \text{K}^+ + 3 \text{Cl}^- \]

\[ 3 \text{H}^+ + \text{PO}_4^{3-} \rightarrow \text{H}_3\text{PO}_4(\text{aq}) \]
Redox Reactions

Redox, or oxidation-reduction, reactions involve the transfer of electrons from one substance to another. In such reactions, one substance is always oxidized (loses the electrons), while another is reduced (gains the electrons). These reactions can only happen spontaneously in one direction—that is, for any given pair of substances, one will have a stronger attraction for the electrons than the other. The relative ranking of different elements according to their ease of oxidation (likelihood of giving up electrons) is known as the Activity Series.

For example, since Zn is a more active metal (i.e., more easily oxidized) than Cu,

\[
\text{Zn (s) } + \text{ CuSO}_4 \text{ (aq) } \rightarrow \text{ Cu (s) } + \text{ ZnSO}_4 \text{ (aq)}
\]

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

Note that in the first reaction, Zn goes from having zero charge on the left to having a +2 charge in the ionic zinc sulfate compound on the right, while Cu goes from a +2 charge to zero. Hence, Zn loses 2 electrons in this process (each electron has a -1 charge), while Cu gains 2 electrons, or in other words, Zn gets oxidized while Cu gets reduced. (Note that nothing happens to the \text{SO}_4^{2-} \text{ ion; it is ‘just along for the ride’—a spectator ion.})

Note that writing the reaction equation above in net ionic form makes the change in oxidation states (charges) for the elements more apparent:

Net Ionic Eqn: \[
\text{Zn (s) } + \text{ Cu}^{2+} \text{ (aq) } \rightarrow \text{ Cu (s) } + \text{ Zn}^{2+} \text{ (aq)}
\]

The reaction above represents one important subclass of redox reactions, namely electron transfer among metals in aqueous solutions. Other important subclasses of redox reactions are indicated below with examples. Note that the first two follow the single displacement reaction pattern shown earlier, while the third follows the synthesis pattern.

Electron Transfer Among Metals

\[
\text{Mg (s) } + \text{ 2 AgNO}_3 \text{ (aq) } \rightarrow \text{ 2 Ag (s) } + \text{ Mg(NO}_3)_2 \text{ (aq)}
\]

Reactions of Metals with Acids

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

Reactions of Metals with Nonmetals (Ionic Compound Formation)

\[
\text{16 Cu (s) } + \text{ S}_8 \text{ (s) } \rightarrow \text{ 8 Cu}_2\text{S (s)}
\]

Combustion Reactions

\[
\text{C}_3\text{H}_8 \text{ (g) } + \text{ 5 O}_2 \text{ (g) } \rightarrow \text{ 4 H}_2\text{O (g) } + \text{ 3 CO}_2 \text{ (g)}
\]
Practice Predicting Redox Reactions

- Write molecular and net ionic equations for each of the following.
- Also indicate which substance is oxidized and which reduced in each reaction.
- Finally, indicate the oxidation states of the elements before and after reaction.

1) aluminum metal + aqueous silver nitrate →
   \[ \text{Al}_\text{(s)} + 3 \text{AgNO}_3\text{(aq)} \rightarrow \text{Al(NO}_3\text{)}_3 + 3 \text{Ag}_{\text{(s)}} \]
   \[ \text{Al}_\text{(s)} + 3 \text{Ag}^+ \rightarrow \text{Al}^{3+} + 3 \text{Ag}_{\text{(s)}} \]
   Al is oxidized, going from 0 to 3+        Ag is reduced, going from 1+ to 0

2) magnesium metal + hydrobromic acid →
   \[ \text{Mg}_\text{(s)} + 2 \text{HCl}_\text{(aq)} \rightarrow \text{MgCl}_2\text{(aq)} + \text{H}_2\text{(g)} \]
   \[ \text{Mg}_\text{(s)} + 2 \text{H}^+ \rightarrow \text{Mg}^{2+} + \text{H}_2\text{(g)} \]
   Mg is oxidized, going from 0 to 2+        H+ is reduced, going from 1+ to 0

3) lithium metal + oxygen gas →
   \[ 4 \text{Li}_\text{(s)} + \text{O}_2\text{(g)} \rightarrow 2 \text{Li}_2\text{O}_\text{(s)} \]
   Li is oxidized, going from 0 to 1+        O is reduced, going from 0 to 2-

4) combustion of pentane (C5H12) →
   \[ \text{C}_5\text{H}_{12}\text{(l)} + 8 \text{O}_2\text{(g)} \rightarrow 5 \text{CO}_2\text{(g)} + 6 \text{H}_2\text{O}_\text{(g)} \]
   C5H12 is oxidized, with C going from −12/5 to +4       O2 is reduced going from 0 to −2

5) silver metal + aqueous calcium nitrate →
   \[ 2 \text{Ag}_\text{(s)} + \text{Ca(NO}_3\text{)}_2\text{(aq)} \rightarrow \text{Ca}_\text{(s)} + 2 \text{AgNO}_3\text{(aq)} \]
   \[ 2 \text{Ag}_\text{(s)} + \text{Ca}^{2+} \rightarrow \text{Ca}_\text{(s)} + 2 \text{Ag}^+ \]
   This would indicate that Ag is oxidized and Ca reduced, but this won’t happen (Ag is less “active” than Ca), so there will be NO REACTION