Chapter 2: Introduction to Chemical Reactions

Part 2.1: Balancing Chemical Reactions

Sections 2.1.1 - 2.1.3: Introduction

A **chemical reaction** is an actual transformation of substances called reactants into substances called products.

To represent a chemical reaction we use a **chemical equation**, a sort of recipe which shows in a symbolic form 1) who the participating substances are (reactants and products), 2) the state or phase these substances are in (solid, liquid, gas, aqueous solution) and 3) the amount in which they must be present (number of atoms, molecules (for covalent compounds) or formula units (for ionic compounds)). This will be discussed further in Chapter 3.

Example: Consider the reaction of aluminum metal ($Al_{(s)}$) with solid iron oxide ($Fe_2O_{3\,(s)}$) forming solid aluminum oxide ($Al_2O_{3\,(s)}$) and solid iron ($Fe_{(s)}$). This reaction is represented by the following chemical equation:

$$\underbrace{2\ Al_{(s)} + Fe_2O_{3\,(s)}}_{\textbf{REACTANTS}} \rightarrow \underbrace{Al_2O_{3\,(s)} + 2\ Fe_{(s)}}_{\textbf{PRODUCTS}}$$

The arrow ($\rightarrow$) shows the direction in which the chemical transformation takes place. The **reactants** (shown on the left-hand side of the arrow) are the substances with which the reaction is started. The **products** (shown on the right-hand side of the arrow) are the substances resulting from the reaction.
In the above reaction all substances are in the solid state, as indicated by the subscript “s” in parentheses.

The **state** or **phase** of a substance is always indicated by a subscript in parentheses after the chemical formula of that substance. The following notations are used for the various phases encountered in chemical reactions:

(s)	for solids
(l)	for liquids
(g)	for gases
(aq)	for aqueous solutions

Finally, one of the most important pieces of information conveyed by a chemical equation is the number of atoms, ions, formula units or molecules associated with each substance. The number in front of each substance is called the **stoichiometric coefficients** or more simply the coefficient. The bulk of this information is often referred to as the **stoichiometry** of the chemical reaction (much more on that in Chapter 3!!).

For the above reaction, the stoichiometric coefficients are 2, 1, 1 and 2, respectively. Note that when a stoichiometric coefficient is 1, it is not shown (as is the case for Al_2O_3 and Fe_2O_3).

Stoichiometric coefficients play a very important role in chemical equations. Their presence insures that the number of atoms of each type is the same on the reactants and products sides. For instance, in the above reaction there are two aluminum atoms, two iron atoms and three oxygen atoms on each side of the chemical equation. This observation reflects Dalton's hypothesis that, in a chemical reaction, atoms are neither destroyed nor created.

To keep a chemical equation looking as simple as possible, we will generally ensure that stoichiometric coefficients are written using the smallest possible whole numbers (integers). **Balancing** a chemical equation consists in determining each of the stoichiometric coefficients.

Section 2.1.4: Balancing the Chemical Reaction: $CO_{2(g)} + H_2O_{(l)} \rightarrow C_6H_{12}O_{6(s)} + O_{2(g)}$

To balance a chemical reaction, we always start balancing the elements that are present in the least number of compounds. In the above equation, we can start with either carbon or hydrogen.

Balancing the element carbon.

There is one carbon atom on the left-hand side (in $CO_{2(g)}$) and six carbon atoms on the right-hand side (in $C_6H_{12}O_{6(s)}$). Hence, we place a coefficient of 6 in front of $CO_{2(g)}$ and 1 in front of $C_6H_{12}O_{6(s)}$. Doing so leads to the balancing of carbon.

When one assigns stoichiometric coefficients to some substances (here, $CO_{2(g)}$ and $C_6H_{12}O_{6(s)}$), one cannot change these coefficients when balancing another element.

Balancing the element hydrogen.

$$6\ CO_{2(g)} + H_2O_{(l)} \rightarrow C_6H_{12}O_{6(s)} + O_{2(g)}$$

There are two hydrogen atoms on the left-hand side (in $H_2O_{(l)}$) and 12 hydrogen atoms on the right-hand side (in $C_6H_{12}O_{6(s)}$). Hence, we place a coefficient of 6 in front of $H_2O_{(l)}$. Doing so leads to the balancing of hydrogen.

Assigning a stoichiometric coefficient to $H_2O_{(l)}$, implies that this coefficient can no longer be changed when balancing the last element, oxygen.

Balancing the element oxygen.

$$6\ CO_{2(g)} + 6\ H_2O_{(l)} \rightarrow C_6H_{12}O_{6(s)} + O_{2(g)}$$

There are 12 + 6 = 18 oxygen atoms on the left-hand side. On the right-hand side here are 6 oxygen atoms in $C_6H_{12}O_{6(s)}$ and 2 oxygen atoms in $O_{2(g)}$. At this stage, we can no longer change the stoichiometric coefficients of $CO_{2(g)}$, $H_2O_{(l)}$, or $C_6H_{12}O_{6(s)}$ since these coefficients have already been assigned when balancing C and H. Hence, to obtain 18 oxygen atoms on the right-hand side, we must assign a stoichiometric coefficient of 6 to $O_{2(g)}$.

The balanced chemical equation is:

$$\mathbf{6}\ CO_{2(g)} + \mathbf{6}\ H_2O_{(l)} \rightarrow C_6H_{12}O_{6(s)} + \mathbf{6}\ O_{2(g)}$$

The sum of the stoichiometric coefficients for reactants and products is **19**.

Section 2.1.5: Balancing the Equation of a Combustion Reaction

A combustion reaction is a reaction in which a substance (element or compound) is burnt with oxygen gas (O_2). The combustion reactions of organic molecules (molecules based on carbon, hydrogen, oxygen, etc…) lead to the formation of carbon dioxide (CO_2) gas and water (H_2O).

Consider the combustion reaction of hexane, $C_6H_{14(l)}$. For the combustion of hexane, oxygen gas must be a reactant and CO_2 and H_2O must be products.

$$C_6H_{14(l)} + O_{2(g)} \rightarrow CO_{2(g)} + H_2O_{(l)}$$

First, we **balance carbon:**

There are 6 carbon atoms in C_6H_{14} and 1 carbon atom in CO_2. Hence, we will use stoichiometric coefficients of 1 (not shown) for C_6H_{14} and 6 for CO_2.

The partially balanced equation is: $C_6H_{14(l)} + O_{2(g)} \rightarrow 6\ CO_{2(g)} + H_2O_{(l)}$

Next, we **balance hydrogen:**

There are 14 hydrogen atoms in one molecule of C_6H_{14} and 2 hydrogen atoms in one molecule of H_2O. Hence, we will use a coefficient of 7 for H_2O.

The partially balanced equation becomes: $C_6H_{14(l)} + O_{2(g)} \rightarrow 6\ CO_{2(g)} + 7\ H_2O_{(l)}$

Finally, we **balance oxygen:**

On the left-hand side there are 2 oxygen atoms. On the right-hand side there are $6 \times 2 + 7 \times 1 = 19$ oxygen atoms. However, oxygen is present as O_2 on the reactant side. Hence, we should use 19/2 O_2 molecules on the right-hand side.

The balanced equation is: $C_6H_{14(l)} + 19/2\ O_{2(g)} \rightarrow 6\ CO_{2(g)} + 7\ H_2O_{(l)}$

If we decide to use only whole numbers as stoichiometric coefficients in a balanced equation, then, we must multiply all coefficients by 2. Hence, the final form for the balanced combustion reaction of hexane is:

$$\mathbf{2}\ C_6H_{14(l)} + \mathbf{19}\ O_{2(g)} \rightarrow \mathbf{12}\ CO_{2(g)} + \mathbf{14}\ H_2O_{(l)}$$

The sum of the stoichiometric coefficients for reactants and products is 47.

Section 2.1.6: Balancing the Chemical Reaction: $Fe_2(SO_4)_{3(aq)} + NH_{3(g)} + H_2O_{(l)} \rightarrow Fe(OH)_{3(s)} + (NH_4)_2SO_{4(s)}$

First, we **balance iron:**

There are 2 iron atoms in $Fe_2(SO_4)_3$ and 1 iron atom in $Fe(OH)_3$. Hence, we will use stoichiometric coefficients of 1 (not shown) for $Fe_2(SO_4)_3$ and 2 for $Fe(OH)_3$.

The partially balanced equation is:

$$Fe_2(SO_4)_{3(aq)} + NH_{3(g)} + H_2O_{(l)} \rightarrow \mathbf{2}\ Fe(OH)_{3(s)} + (NH_4)_2SO_{4(s)}$$

Next, we **balance sulfur:**

There are 3 sulfur atoms in one formula unit of $Fe_2(SO_4)_3$ and 1 sulfur atom in one formula unit of $(NH_4)_2SO_4$. Hence, we will use a stoichiometric coefficient of 3 for $(NH_4)_2SO_4$.

The partially balanced equation becomes:

$$Fe_2(SO_4)_{3(aq)} + NH_{3(g)} + H_2O_{(l)} \rightarrow \mathbf{2}\ Fe(OH)_{3(s)} + \mathbf{3}\ (NH_4)_2SO_{4(s)}$$

Next, we **balance nitrogen:**

On the left-hand side there is 1 nitrogen atom in NH_3. On the right-hand side there are 3 x 2 = 6 nitrogen atoms in 3 formula units of $(NH_4)_2SO_4$. Hence, we will use a stoichiometric coefficient of 6 for NH_3.

The partially balanced equation becomes:

$$Fe_2(SO_4)_{3(aq)} + \mathbf{6}\ NH_{3(g)} + H_2O_{(l)} \rightarrow \mathbf{2}\ Fe(OH)_{3(s)} + \mathbf{3}\ (NH_4)_2SO_{4(s)}$$

Next, we **balance oxygen:**

On the left-hand side there are 3 x 4 = 12 oxygen atoms in one formula unit of $Fe_2(SO_4)_3$ and 1 oxygen atom in one molecule of $H_2O_{(l)}$. On the right-hand side there are 18 oxygen atoms [2 x 3 = 6 oxygen atoms in 2 formula units of $Fe(OH)_3$ and 3 x 4 = 12 oxygen atoms in 3 formula units of $(NH_4)_2SO_4$]. The only coefficient that can be modified is the coefficient for H_2O since all other coefficients have already been assigned. Hence, we will use a stoichiometric coefficient of 6 for H_2O.

The partially balanced equation becomes:

$$Fe_2(SO_4)_{3(aq)} + \mathbf{6}\ NH_{3(g)} + \mathbf{6}\ H_2O_{(l)} \rightarrow \mathbf{2}\ Fe(OH)_{3(s)} + \mathbf{3}\ (NH_4)_2SO_{4(s)}$$

Finally, we **balance hydrogen:**

Note that we expect the chemical equation to be balanced for hydrogen as all stoichiometric coefficients have already been assigned.
On the left-hand side there are 30 hydrogen atoms [6 x 3 = 18 hydrogen atoms in three formula units of NH_3 and 6 x 2 = 12 atoms in two molecules of H_2O]. On the right-hand side there are also 30 hydrogen atoms [2x3 = 6 hydrogen atoms in 2 formula units of $Fe(OH)_3$ and 3x4x2 = 24 hydrogen atoms in 3 formula units of $(NH_4)_2SO_4$].

Hence, the fully balanced chemical equation is:

$$Fe_2(SO_4)_{3(aq)} + \mathbf{6}\ NH_{3(g)} + \mathbf{6}\ H_2O_{(l)} \rightarrow \mathbf{2}\ Fe(OH)_{3(s)} + \mathbf{3}\ (NH_4)_2SO_{4(s)}$$

The sum of the stoichiometric coefficients for products and reactants is **18**.

Sections 2.1.7 - 2.1.8: Balancing Chemical Reactions

These are interactive sessions where you are guided to practice what you have learned in the first 6 sections of part 2.1.

Section 2.1.9: Practice Problems on Balancing Reactions

Practice, practice, practice.
If you have problems, go back and study Sections 2.1.1 through 2.1.8!

Sections 2.1.10 - 2.1.12: Balancing Chemical Equations in Word Problems

In these sections, you are taught to analyze statements in a word problem and come up with a chemical reaction. Demonstration of the reaction is shown in the video. Following this, a step by step explanation of balancing the chemical reaction is offered.

Chemical reactions are represented by chemical equations. Chemical equations consist of reactants and products.

Consider the reaction: Solid sodium metal reacts with water giving a solution of sodium hydroxide and releasing hydrogen gas.

Write a chemical reaction using complete formulas with phases.

Sodium is Na. It exists as a solid. Hence, put "(s)" next to $Na_{(s)}$

The formula for water is H_2O. Add H_2O

$$Na_{(s)} + H_2O$$

H_2O is a liquid. Hence, put "(l)" next to it. $Na_{(s)} + H_2O_{(l)}$

Giving sodium hydroxide means, produces sodium hydroxide.

Draw the line with arrow. $Na_{(s)} + H_2O_{(l)} \rightarrow$

Chemical formula for sodium hydroxide is NaOH.

Solution of NaOH, put "(aq)" next to it.

$$Na_{(s)} + H_2O_{(l)} \rightarrow NaOH_{(aq)}$$

Remember hydrogen exists as H_2.

$$Na_{(s)} + H_2O_{(l)} \rightarrow NaOH_{(aq)} + H_2$$

Since H_2 exists as gas, put "(g)" next to it.

$$Na_{(s)} + H_2O_{(l)} \rightarrow NaOH_{(aq)} + H_{2(g)}$$

Reactants are on the left-hand side (LHS) of the equation. Products are on the right-hand side (RHS) of the equation.

The coefficients of a balanced equation are positive integers. Note: Integers will be used here only for the sake of convenience. In reality one can use rational numbers (fractions) to balance chemical reactions. To balance a chemical reaction only the coefficients can be changed. The balanced chemical reaction contains the smallest possible coefficients.

$$Na_{(s)} + H_2O_{(l)} \rightarrow NaOH_{(aq)} + H_{2(g)}$$

On the left-hand side of the equation there is one Na. On the right-hand side of the equation there is one Na. Hence, Na is balanced.

On the left-hand side of the equation there are two H. On the right-hand side of the equation there are three H. In this case there should be an even number of H on both sides. The smallest even number that could work would be four. Hence, a coefficient of 2 is placed for H_2O and a coefficient of 2 is placed for NaOH.

$$Na_{(s)} + 2\ H_2O_{(l)} \rightarrow 2\ NaOH_{(aq)} + H_{2(g)}$$

On the left-hand side of the equation there are two O. On the right-hand side of the equation there are two O. Hence, O is balanced.

Now, on the left-hand side of the equation there is one Na. On the right-hand side of the equation there are two Na. Hence, a coefficient of 2 should be placed for Na on the left-hand side.

$$2\ Na_{(s)} + 2\ H_2O_{(l)} \rightarrow 2\ NaOH_{(aq)} + H_{2(g)}$$

Note: When no coefficient appears, it is understood that the coefficient is 1. Hence, the sum of coefficients of reactants and products for the balanced chemical reaction is 7.

Solid magnesium phosphate reacts with an aqueous solution of sulfuric acid giving magnesium sulfate as a solid and phosphoric acid as a solution. Write a chemical reaction using complete formulas with phases.

Magnesium phosphate has a chemical formula $Mg_3(PO_4)_2$. It exists as a solid. Hence, put "(s)" next to it.

$$Mg_3(PO_4)_{2\ (s)}$$

The formula for sulfuric acid is H_2SO_4

$$Mg_3(PO_4)_{2\,(s)} + H_2SO_4$$

H_2SO_4 is an aqueous solution. Hence, put “(aq)” next to it.

$$Mg_3(PO_4)_{2\,(s)} + H_2SO_{4(aq)}$$

Giving magnesium sulfate means produces magnesium sulfate. Draw the line with the arrow.

$$Mg_3(PO_4)_{2\,(s)} + H_2SO_{4(aq)} \rightarrow$$

Chemical formula for magnesium sulfate is $MgSO_4$

$$Mg_3(PO_4)_{2\,(s)} + H_2SO_{4(aq)} \rightarrow MgSO_4$$

$MgSO_4$ is produced as a solid. Hence, put “(s)” next to it.

$$Mg_3(PO_4)_{2\,(s)} + H_2SO_{4(aq)} \rightarrow MgSO_{4(s)}$$

Chemical formula for phosphoric acid is H_3PO_4

$$Mg_3(PO_4)_{2\,(s)} + H_2SO_{4(aq)} \rightarrow MgSO_{4(s)} + H_3PO_4$$

H_3PO_4 as a solution, put “(aq)” next to it.

$$Mg_3(PO_4)_{2\,(s)} + H_2SO_{4(aq)} \rightarrow MgSO_{4(s)} + H_3PO_{4(aq)}$$

Reactants are on the left-hand side (LHS) of the equation. Products are on the right-hand side (RHS) of the equation.

On the left-hand side of the equation there are three Mg. On the right-hand side of the equation there is one Mg. Hence, a coefficient of 3 is placed for Mg on the RHS.

On the left-hand side of the equation there are two P. On the right-hand side of the equation there is one P. Hence, a coefficient of 2 is placed for P on the RHS.

On the left-hand side of the equation there are two H. On the right-hand side there are six H. Hence, a coefficient of 3 is placed for H on the LHS.

On the left-hand side of the equation there are three S. On the right-hand side of the equation there are three S. Hence, S is balanced.

On the left-hand side of the equation there are eight + twelve O = 20 O.
On the right-hand side of the equation there are twelve + eight O = 20 O.
Hence, O is balanced.

Note: When no coefficient appears, it is understood that the coefficient is 1. Hence, the sum of coefficients of reactants and products for the balanced chemical reaction is 9.

Part 2.2: Solubility Rules (For Ionic Compounds in Water)

Section 2.2.1: Introduction

Solubility is defined as the ability of a compound to dissolve in a solvent, here, in water. Different compounds have different solubility in water. Here, we will only deal with ionic compounds since a number of simple rules exist to decide whether specific ions impart solubility or not. For example, NaCl dissolves very easily in water while $CaCO_3$ dissolves only very sparingly. Molecular compounds can also exhibit a wide range of solubility in water (for example sugar is very soluble in water, while oil and gasoline are not). The discussion of solubility of molecular compounds is however postponed until Chapter 10 where we study "Intermolecular Forces and Liquid Properties".

Section 2.2.2: Solubility Rules for Chlorides, Bromides and Iodides

All chloride, bromide and iodide ionic compounds are soluble in water except those containing silver(I), mercury(I), lead(II) and copper(I).

Section 2.2.3: Dissociation of Chlorides, Bromides and Iodides

Practice making chloride, bromide and iodide compounds with main group and transition metals and learn whether these compounds are soluble and if so, what species form upon dissolution in water.

Section 2.2.4: Solubility Rules for Acetate, Chlorate, Perchlorate, Nitrate and Hydroxide Compounds

All acetate, chlorate, perchlorate and nitrate ionic compounds are soluble in water.

All group 1 hydroxide compounds are soluble in water. Magnesium, calcium, strontium and barium hydroxides are only slightly soluble. All other hydroxides are insoluble in water.

Section 2.2.5: Dissociation of Acetate, Chlorate, Perchlorate, Nitrate and Hydroxide Compounds

Practice making acetate, chlorate, perchlorate, nitrate and hydroxide compounds with main group and transition metals and learn whether these compounds are soluble and if so, what species form upon dissolution in water.

Section 2.2.6: Solubility Rules for Sulfides

Group 1 sulfide ionic compounds are water soluble while all other sulfide compounds are not.

Section 2.2.7: Dissociation of Sulfide Compounds

Practice making sulfide compounds with main group and transition metals and learn whether these compounds are soluble and if so, what species form upon dissolution in water.

Section 2.2.8: Solubility Rules for Carbonates, Chromates and Sulfates

Group 1 carbonate and chromate ionic compounds are water soluble while all other carbonate and chromate compounds are not.

Most sulfate ionic compounds are soluble except ionic compounds made with Mg, Ca, Sr, Ba, Ra, Hg(I) and Pb(II).

Section 2.2.9: Dissociation of Carbonate, Chromate and Sulfate Compounds

Practice making carbonate, chromate and sulfate compounds with main group and transition metals and learn whether these compounds are soluble and if so, what species form upon dissolution in water.

Section 2.2.10: Solubility Rules and Dissociation of Ammonium Compounds

All ammonium compounds are water soluble.

Part 2.3: Ionic Equations and Precipitation Reactions

Section 2.3.1: Writing Ionic Equations

Ionic equations represent chemical reactions between ions. Such reactions require at least two ionic compounds (4 ions). To illustrate precipitation reactions, we will consider three examples.

2.3.1.a Reaction of $NaCl_{(aq)}$ with $AgNO_{3\,(aq)}$

$NaCl_{(s)}$ is soluble in water. Hence, when mixed with water, this solid breaks down into $Na^{+}_{(aq)}$ and $Cl^{-}_{(aq)}$ ions, according to the ionic equation:

$$NaCl_{(s)} \rightarrow Na^{+}_{(aq)} + Cl^{-}_{(aq)}$$

$AgNO_{3\,(s)}$ is soluble in water. Hence, when mixed with water, this solid breaks down into $Ag^{+}_{(aq)}$ and $NO_{3}^{-}{}_{(aq)}$ ions, according to the ionic equation:

$$AgNO_{3\,(s)} \rightarrow Ag^{+}_{(aq)} + NO_{3}^{-}{}_{(aq)}$$

Let us mix these two aqueous solutions. Does a reaction occur and if so, what reaction is it?

Obviously, $Cl^{-}_{(aq)}$ will not react with $Na^{+}_{(aq)}$. Does $Cl^{-}_{(aq)}$ form a compound with $Ag^{+}_{(aq)}$? Whether these ions form a compound in the presence of water depends on whether the resulting compound is soluble in water or not. The solubility rules for chlorides, bromides and iodides state that silver chloride is not soluble in water. Hence, the silver and chloride ions form an ionic compound. We say that a **precipitate** of solid silver chloride (white creamy solid) has formed by reaction of sodium chloride with silver nitrate.

Now, we must ask whether another precipitation reaction takes place between the sodium and the nitrate ions. Solubility rules for nitrate ions tell us that all nitrate compounds are soluble in water. Hence, sodium and nitrate ions remain as free ions in solution and do not form a precipitate.

The sodium and nitrate ions do not participate in the reaction. They are present and actually watch the silver and chloride ions react to form a solid ionic compound. We say that the sodium and nitrate ions are **spectator ions**.

The **complete ionic equation** between these compounds is written as:

$$Na^{+}_{(aq)} + Cl^{-}_{(aq)} + Ag^{+}_{(aq)} + NO_3^{-}{}_{(aq)} \rightarrow AgCl_{(s)} + Na^{+}_{(aq)} + NO_3^{-}{}_{(aq)}$$

Since the $Na^{+}_{(aq)}$ and $NO_3^{-}{}_{(aq)}$ ions are spectator ions (i.e. they do not participate in the reaction), they can be cancelled on both sides of the equation. The resulting equation is called the **net ionic equation** and is written as:

$$Ag^{+}_{(aq)} + Cl^{-}_{(aq)} \rightarrow AgCl_{(s)}$$

2.3.1.b Reaction of $CoCl_{2\,(aq)}$ with $Na_2S_{(aq)}$

$CoCl_{2(s)}$ is soluble in water. Hence, when mixed with water, this solid breaks down into $Co^{+2}_{(aq)}$ and $Cl^{-}_{(aq)}$ ions according to the ionic equation:

$$CoCl_{2(s)} \rightarrow Co^{+2}_{(aq)} + 2\,Cl^{-}_{(aq)}$$

$Na_2S_{(s)}$ is soluble in water. Hence, when mixed with water, this solid breaks down into $Na^{+}_{(aq)}$ and $S^{-2}_{(aq)}$ ions according to the ionic equation:

$$Na_2S_{(s)} \rightarrow 2\,Na^{+}_{(aq)} + S^{-2}_{(aq)}$$

Mixing of these two aqueous solutions then makes it possible for a reaction between these ions to take place. Will a reaction occur and if so, how is it represented?

Obviously, $Na^{+}_{(aq)}$ will not react with $Cl^{-}_{(aq)}$ since NaCl is water soluble but $Co^{+2}_{(aq)}$ will form a compound with $S^{-2}_{(aq)}$ since CoS is not water soluble (see solubility rules for sulfides). Mixing the two solutions lead to a dark **precipitate** of solid cobalt(II) sulfide.

Since the sodium and chloride ions remain as free ions in solution and do not form a precipitate, they are **spectator ions** during the precipitation of CoS.

The **complete ionic equation** between these compounds is written as:

$$2\,Na^{+}_{(aq)} + S^{-2}_{(aq)} + Co^{+2}_{(aq)} + 2\,Cl^{-}_{(aq)} \rightarrow CoS_{(s)} + 2\,Na^{+}_{(aq)} + 2\,Cl^{-}_{(aq)}$$

Since the $Na^{+}_{(aq)}$ and $Cl^{-}_{(aq)}$ ions are spectator ions (i.e. they do not participate in the reaction), they can be cancelled on both side of the equation. The resulting equation is called the **net ionic equation** and is written as:

$$Co^{+2}_{(aq)} + S^{-2}_{(aq)} \rightarrow CoS_{(s)}$$

2.3.1.c Reaction of NaCl $_{(aq)}$ and $Ca(NO_3)_{2\,(aq)}$

$NaCl_{(s)}$ and $Ca(NO_3)_{2(s)}$ are both soluble in water. However, $NaNO_{3(s)}$ and $CaCl_{2(s)}$ are also soluble in water. Hence, mixing aqueous solutions of sodium chloride and calcium nitrate does not lead to precipitation. All ions remain free in solution and there is NO reaction between them.

Sections 2.3.2 - 2.3.3: Interactive Ionic Reactions

Practice, practice, practice

Part 2.4: Electrolytes and Nonelectrolytes

Section 2.4.1: Electrolytes and Nonelectrolytes

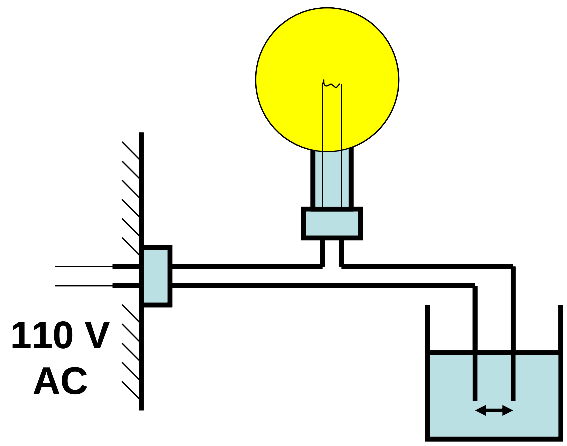
A substance which, when dissolved in water, makes an electrically conducting solution is called an **electrolyte**. Electrolytes are substances which dissociate into ions when dissolved in water.

Examples of electrolytes are ionic compounds which are water soluble, acids and bases. (Note that ionic compounds which are not soluble in water will not dissociate into ions when mixed with water, hence, they will not conduct electricity).

A substance which, when dissolved in water, makes an electrically non-conducting solution is called a **nonelectrolyte**. Nonelectrolytes do not dissociate into ions when dissolved in water. Examples of nonelectrolytes are molecular compounds other than acids and bases.

To decide whether a compound is an electrolyte or a nonelectrolyte, we can use the following experiment: (see video).

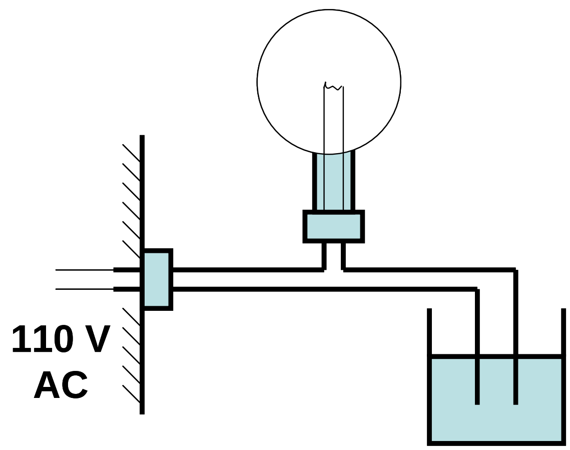
A compound is dissolved in water and the solution placed in a beaker. The two leads of the electrical circuit are dipped in the beaker. If dissolution of the compound is accompanied by dissociation of the compound into ions, the ions will be attracted to the oppositely charged leads. Motion of the ions between the leads (where the arrow is) closes the electrical circuit and current flows through the light bulb (see figure below).



Solution contains an Electrolyte
The bulb is lit !!!

Note that electrical current is characterized by the motion of electrons in the metallic wires and the motion of ions (not electrons) in the solution.

If the solution contains no ions (or a really low concentration of ions as in distilled water), then the circuit is not closed and electrical current does not flow through the light bulb.



Solution contains a
Non-electrolyte
The bulb is NOT lit!!!

Section 2.4.2: Strong and Weak Electrolytes

Electrolytes dissolved in water produce ions to various extents.

Some electrolytes when dissolved in water **dissociate completely** into cations and anions. Such compounds are called **strong electrolytes**.

Examples of strong electrolytes are soluble ionic compounds, strong acids and strong bases (see part 2.5). For example, when $NaCl_{(s)}$ is dissolved in water all NaCl dissociates into $Na^{+}_{(aq)}$ and $Cl^{-}_{(aq)}$. NaCl is a strong electrolyte. Similarly, when dissolving HCl gas (hydrogen chloride) in water, we form hydrochloric acid. Hydrochloric acid is written as $HCl_{(aq)}$ and exists as $H^{+}_{(aq)}$ and $Cl^{-}_{(aq)}$ ions. Using the above set-up, a strong electrolyte is characterized by a very brightly lit light bulb.

Some electrolytes, when dissolved in water **dissociate partially** into cations and anions. Partial dissociation means that some of the original compound exists in solution in the **undissociated** (molecular) form. Substances that dissociate partially in solution are called **weak electolytes**. Examples of weak electrolytes are weak acids and weak bases (see part 2.5). For instance, acetic acid, $HC_2H_3O_{2(l)}$, the main component of vinegar dissolves in water and dissociates partially into $H^{+}_{(aq)}$ and $C_2H_3O_2^{-}{}_{(aq)}$ ions. However, most of the molecules of $HC_2H_3O_{2(aq)}$ remain undissociated in solution. Using the above set-up, a weak electrolyte is characterized by a dimly lit light bulb.

Part 2.5: Acids and Bases

Section 2.5.1: Properties of Acids and Bases

Acids and bases taste different (sour vs. soapy) but do not try this at home unless it is vinegar, lemon juice (acids) or milk of magnesia (base).

Section 2.5.2: Arrhenius Acids and Bases

Svante Arrhenius proposed the first definition for acids and bases. An acid according to Arrhenius is a substance which, when dissolved in water, produces hydrogen ions (protons), $H^{+}_{(aq)}$.

Consider HCl gas. The dissolution reaction is written as:

$$HCl_{(g)} \rightarrow H^{+}_{(aq)} + Cl^{-}_{(aq)}$$

A base according to Arrhenius is a substance which, when dissolved in water, produces hydroxide ions, $OH^-_{(aq)}$.

Consider NaOH solid. The dissolution reaction is written as:

$$NaOH_{(s)} \rightarrow Na^+_{(aq)} + OH^-_{(aq)}$$

Section 2.5.3: Brønsted - Lowry Acids and Bases

Acids according to **Brønsted-Lowry** are substances that donate protons (H^+) to another species.

Consider the formation of hydrochloric acid:

$$HCl_{(g)} + H_2O_{(l)} \rightarrow H_3O^+_{(aq)} + Cl^-_{(aq)}$$

$H_3O^+_{(aq)}$ is called the hydronium ion. This reaction is also written as:

$$HCl_{(g)} \rightarrow H^+_{(aq)} + Cl^-_{(aq)}$$

as $H^+_{(aq)}$ is equivalent to $H_3O^+_{(aq)}$

Bases according to **Brønsted-Lowry** are substances that accept a proton (H^+) from another species.

Consider the dissolution of ammonia in water:

$$NH_{3(g)} + H_2O_{(l)} \rightarrow NH_4^+{}_{(aq)} + OH^-_{(aq)}$$

$NH_4^+{}_{(aq)}$ is the ammonium ion.

Section 2.5.4: Strong Acids and Weak Acids

Strong acids completely ionize when dissolved in water. The following list of strong acids is a list you need to memorize. All acids in this list are strong acids. All other acids are weak acids.

Strong Acids:			**Weak Acids**	
$HCl_{(aq)}$	hydrochloric acid		$HClO_{2\,(aq)}$	chlorous acid
$HBr_{(aq)}$	hydrobromic acid		$HClO_{(aq)}$	hypochlorous acid
$HI_{(aq)}$	hydroiodic acid		$HBrO_{4\,(aq)}$	perbromic acid
$HNO_{3\,(aq)}$	nitric acid		$HBrO_{3\,(aq)}$	bromic acid
$HClO_{4\,(aq)}$	perchloric acid		$HBrO_{2\,(aq)}$	bromous acid
$HClO_{3\,(aq)}$	chloric acid		$HBrO_{(aq)}$	hypobromous acid
$H_2SO_{4\,(aq)}$	sulfuric acid		$HF_{(aq)}$	hydrofluoric acid
			$HC_2H_3O_{2\,(aq)}$	acetic acid
MEMORIZE THE ABOVE LIST			$HNO_{2\,(aq)}$	nitrous acid
			$H_2SO_{3\,(aq)}$	sulfurous acid
			$H_3PO_{4\,(aq)}$	phosphoric acid
			$H_3PO_{3\,(aq)}$	phosphorous acid
			etc…	

The complete dissociation of a **strong acid** is represented by the following chemical equation:

$$HCl_{(aq)} \rightarrow H^+_{(aq)} + Cl^-_{(aq)}$$

The **single arrow** indicates that the ionization of HCl goes to completion. All HCl dissociates to $H^+_{(aq)}$ and $Cl^-_{(aq)}$ ions. <u>Strong acids are strong electrolytes</u>.

The partial dissociation of a **weak acid** is represented by the following chemical equation:

$$HF_{(aq)} \leftrightarrows H^+_{(aq)} + F^-_{(aq)}$$

The **double arrow** indicates that the ionization of HF does not go to completion. Some HF dissociates to $H^+_{(aq)}$ and $F^-_{(aq)}$ ions and some HF remains in molecular form. We say that the dissociation (or ionization) of HF reaches an equilibrium where HF molecules, $H^+_{(aq)}$ and $F^-_{(aq)}$ ions coexist in solution. <u>Weak acids are weak electrolytes.</u>

Section 2.5.5: Strong Bases and Weak Bases

Strong bases completely ionize when dissolved in water. The following list of strong bases is a list you need to memorize. All bases in this list are strong bases. All other bases are weak bases.

Strong Bases:		Weak Bases	
$LiOH_{(aq)}$	lithium hydroxide	$Mg(OH)_{2\ (aq)}$	magnesium hydroxide
$NaOH_{(aq)}$	sodium hydroxide	$NH_{3\ (aq)}$	ammonia
$KOH_{(aq)}$	potassium hydroxide	$F^{-}_{(aq)}$	fluoride ion
$CsOH_{(aq)}$	cesium hydroxide	$C_2H_3O_2^{-}{}_{(aq)}$	acetate ion
$Ca(OH)_{2\ (aq)}$	calcium hydroxide	$ClO_2^{-}{}_{(aq)}$	chlorite ion
$Sr(OH)_{2\ (aq)}$	strontium hydroxide		
$Ba(OH)_{2\ (aq)}$	barium hydroxide		

MEMORIZE THE ABOVE LIST

Note that strong bases are groups I and II hydroxides except beryllium and magnesium hydroxides. The complete dissociation of a **strong base** is represented by the following chemical equation (in the case of NaOH):

$$NaOH_{(aq)} \rightarrow Na^{+}{}_{(aq)} + OH^{-}{}_{(aq)}$$

The **single arrow** (→) indicates that the dissociation of NaOH goes to completion. All NaOH dissociate to $Na^{+}{}_{(aq)}$ and $OH^{-}{}_{(aq)}$ ions. <u>Strong bases are strong electrolytes.</u>

The partial dissociation of a **weak base** is represented by the following chemical equation (in the case of NH_3):

$$NH_{3\ (aq)} + H_2O_{(aq)} \leftrightarrows NH_4^{+}{}_{(aq)} + OH^{-}{}_{(aq)}$$

The **double arrow** (⇆) indicates that the ionization of $NH_{3\ (aq)}$ does not go to completion. Some $NH_{3\ (aq)}$ ionizes to $NH_4^{+}{}_{(aq)}$ and some NH_3 remains in molecular form. We say that the ionization of NH_3 reaches an equilibrium where NH_3 molecules, $NH_4^{+}{}_{(aq)}$ and $OH^{-}{}_{(aq)}$ ions coexist in solution. <u>Weak bases are weak electrolytes.</u>

Part 2.6: Oxidation-Reduction Reactions

Section 2.6.1: Introduction to Oxidation and Reduction Reactions

Oxidation-reduction reactions are chemical reactions involving the exchange of electrons between two substances.

During an **oxidation** reaction, there is a loss of electrons. For example, the oxidation of $Fe_{(s)}$ to $Fe^{+2}_{(aq)}$ is accompanied by the loss of two electrons. This reaction is represented by the net ionic equation:

$$Fe_{(s)} \rightarrow Fe^{+2}_{(aq)} + 2\ e^{-}$$

During a **reduction** reaction, there is a gain of electrons. For example, the reduction of $Cu^{+2}_{(aq)}$ to $Cu_{(s)}$ is accompanied by the gain of two electrons. This reaction is represented by the net ionic equation:

$$Cu^{+2}_{(aq)} + 2\ e^{-} \rightarrow Cu_{(s)}$$

To remember that **O**xidation **I**s the **L**oss of electrons (electrons on the right-hand side of the equation) and that **R**eduction **I**s the **G**ain of electrons (electrons on the left-hand side of the equation), remember **OIL-RIG**.

Oxidation and reduction occur together. Hence, they are called **redox** reactions. In a redox reaction, one substance gains electrons (undergoing a reduction), while the other substance loses electrons (undergoing an oxidation).

For instance, consider the following redox reaction:

$$Cu^{+2}_{(aq)} + Fe_{(s)} \rightarrow Fe^{+2}_{(aq)} + Cu_{(s)}$$

In this reaction, $Cu^{+2}_{(aq)}$ is reduced to $Cu_{(s)}$ and $Fe_{(s)}$ is simultaneously oxidized to $Fe^{+2}_{(aq)}$.

The substance undergoing reduction (substance being reduced) is called the **oxidizing** agent. The substance undergoing oxidation (substance being oxidized) is called the **reducing** agent.

A typical redox reaction is represented as:

OXIDIZING AGENT + REDUCING AGENT	→	REDUCED FORM OF OXIDIZING AGENT + OXIDIZED FORM OF REDUCING AGENT

Considering the same redox reaction as above:

$$Cu^{+2}_{(aq)} + Fe_{(s)} \rightarrow Fe^{+2}_{(aq)} + Cu_{(s)}$$

The oxidizing agent is $Cu^{+2}_{(aq)}$
The reducing agent is $Fe_{(s)}$
The reduced form of the oxidizing agent is: $Cu_{(s)}$
The oxidized form of the reducing agent is: $Fe^{+2}_{(aq)}$

How can we tell if a given chemical reaction is a redox reaction?

For some reactions, it is simple. Any reaction where a given element is combined with a different number of oxygen atoms on the reactants side and on the products side is a redox reaction.

For example, consider the reaction:

$$2\ Al_{(s)} + Fe_2O_{3\,(s)} \rightarrow Al_2O_{3\,(s)} + 2\ Fe_{(s)}$$

The ratio of oxygen to aluminum is higher on the product side (3 "O" for 2 "Al") than on the reactant side (0 "O" for 2 "Al"). Hence, aluminum is more oxidized in $Al_2O_{3\,(s)}$ than in the elemental state, $Al_{(s)}$. We say that aluminum undergoes an oxidation in this reaction. In contrast, you can see that iron is more oxidized in $Fe_2O_{3\,(s)}$ than in the elemental state, $Fe_{(s)}$. Hence, iron undergoes a reduction in this reaction.

However, there will be many reactions where it is more difficult to determine whether a given substance undergoes reduction or oxidation. In order to help us answer this question, we will learn to calculate **oxidation numbers**.

The oxidation number is a number which tells us how oxidized or reduced a given element of a given substance is. The higher the oxidation number is, the more oxidized the element is. Oxidation number will also help us greatly with the bookkeeping of electrons in reduction or oxidation reactions.

Section 2.6.2: Rules for Assigning Oxidation Numbers (11 rules)

Rule #1: The oxidation number for an element in the elemental state is 0.

For example, in the above reaction, the oxidation numbers for Al in $Al_{(s)}$ and for Fe in $Fe_{(s)}$ are both 0.

Rule #2: The oxidation number for any monoatomic ion is the charge of the ion.

For example, in the reaction: $Cu^{+2}_{(aq)} + Fe_{(s)} \rightarrow Fe^{+2}_{(aq)} + Cu_{(s)}$, the oxidation numbers for Cu in $Cu^{+2}_{(aq)}$ and for Fe in $Fe^{+2}_{(aq)}$ are both equal to **+2**.
The oxidation number of Fe in $Fe^{+3}_{(aq)}$ is **+3**.
The oxidation number of N in $N^{-3}_{(aq)}$ is **- 3**.

The oxidation number of S in $S^{-2}_{(aq)}$ is **-2**.
The oxidation number of Cl in $Cl^{-}_{(aq)}$ is **-1**.

Rule #3: The oxidation number for oxygen in most oxygen compounds (excluding peroxides and superoxides, a minor fraction of oxygen compounds) is equal to - 2.

Rule #4: In all group I compounds, the oxidation number of the metal element is +1. (does not apply to H, since it is not a metal)

Na in $NaCl_{(s)}$ or in $Na_2SO_{4(s)}$ has an oxidation number of +1

Rule #5: In all group II compounds, the oxidation number of the metal is +2.

Ca in $CaCO_3$ and in $Ca(NO_3)_2$ has an oxidation number of +2

Rule #6: In all fluorine compounds, the oxidation number of fluorine is -1.

The oxidation number of F in NaF, CaF_2 and AlF_3 is always -1.

Rule #7: In all Cl, Br, I compounds, the oxidation number of the halogen is -1, unless the halogen is combined with oxygen or with another halogen.

In NaCl, Cl has an oxidation number of -1.
In BrCl, Br has an oxidation number of +1 and Cl has an oxidation number of -1. Note the second halogen in the formula (Cl) plays the role of a nonmetal (oxidation number is -1 in this case, as for F another halogen), while the first halogen in the formula (Br) plays the role of a metal cation (positive oxidation number).

Rule #8: In a compound, the oxidation number of hydrogen is +1 if H is bonded to a nonmetal.

H in NH_3, in CH_4, in H_2O and in HCN has the same oxidation number of +1.

Rule #9: In a compound, the oxidation number of hydrogen is -1 if H is bonded to a metal. (Note: in this case, H behaves like an anion and is called hydride)

H in NaH (sodium hydride), in CaH_2 (calcium hydride) has the oxidation number of -1.

Rule #10: The sum of the oxidation numbers of all elements in a compound is equal to 0 (the charge of the compound).

For example in NO_2, oxidation #(nitrogen) + 2 x oxidation #(oxygen) = 0.
We will write this symbolically as: "N" + 2 "O" = 0

Rule #11: The sum of the oxidation numbers of all elements in a polyatomic ion is equal to the charge of the polyatomic ion.

For example in CO_3^{-2}, oxidation #(carbon) + 3 x oxidation #(oxygen) = -2.
We will write this symbolically as: "C" + 3 "O" = -2

Section 2.6.3: Assigning Oxidation Numbers

In this section we apply the above 11 rules to calculate the oxidation number of an element in a compound. Note that many elements can have very different values of the oxidation number.

Consider Na_2SO_4:

Oxidation number of Na = "Na" = +1
Oxidation number of O = "O" = -2
Oxidation number of S: = "S" to be calculated
2 "Na" + "S" + 4 "O" = 0
2x(+1) + "S" + 4x(-2) = 0
"S" – 6 = 0
"S" = +6

Consider SO_3^{-2}:

"S" + 3x(-2) = -2
"S" = +4

Consider S^{-2}:

"S" = -2

Consider S_8:

"S" = 0

Consider $H_2PO_4^-$:

Oxidation number of H = "H" = +1
Oxidation number of O = "O" = -2
Oxidation number of P: = "P" to be calculated
2 "H" + "P" + 4 "O" = -1
2x(+1) + "P" + 4x(-2) = -1
"P" – 6 = -1
"P" = +5

Section 2.6.4: Assigning Oxidation Numbers (Interactive)

Practice, practice, practice.

Section 2.6.5: Balancing Half-reactions (Oxidation and Reduction)

In this section we use our understanding of oxidation numbers to balance redox reactions. We need to emphasize at the onset that when balancing redox reactions, it is much easier to start by balancing the reduction half-reaction and the oxidation half-reaction (called half-reactions because any redox reaction has two parts: oxidation and reduction).

We will start with the simple case where the half-reaction involves only one element, as in the case:

$$Fe_{(s)} \rightarrow Fe^{+2}_{(aq)} + 2\ e^{-}$$

To balance such reaction equation, we only need to balance:

1) the element involved,

2) the oxidation number.

When we write: $Fe_{(s)} \rightarrow Fe^{+2}_{(aq)}$ we have actually already balanced the equation as far as the element iron is concerned (1 on each side of the equation). Next, we need to balance the equation as far oxidation number is concerned. Since Fe has an oxidation number equal to 0 on the left-hand side of the equation and an oxidation number of +2 on the right-hand side, we add two electrons on the right-hand side to get the fully balanced half-reaction:

$$Fe_{(s)} \rightarrow Fe^{+2}_{(aq)} + 2\ e^{-}$$

Similarly, consider the oxidation half-reaction of $Hg_{(l)}$ to $Hg_2{}^{+2}{}_{(aq)}$.
The oxidation half-reaction is written in balanced form as:

$$2\ Hg_{(l)} \rightarrow Hg_2{}^{+2}{}_{(aq)} + 2\ e^{-}$$

You can also consider the reduction of chlorine gas to chloride ion as:

$$Cl_{2\ (g)} + 2\ e^{-} \rightarrow 2\ Cl^{-}{}_{(aq)}$$

The above cases were very simple as only one element was involved in each half-reaction.

What should be done if more than one element is involved in a half-reaction? Consider for example the following half-reaction involving manganese and oxygen:

$$MnO_4{}^{-}{}_{(aq)} \rightarrow Mn^{+2}{}_{(aq)}$$

To balance this and similar half-reactions we need to follow these four steps.

STEP 1: Balance the element being oxidized or reduced.

Obviously, you need to be sure which element is being oxidized or reduced. In the above half-reaction, Mn is being reduced, since its oxidation number decreases from the value of +7 in $MnO_4^-{}_{(aq)}$ to the value of +2 in $Mn^{+2}{}_{(aq)}$.

The half-reaction: $MnO_4^-{}_{(aq)} \rightarrow Mn^{+2}{}_{(aq)}$ is already balanced for Mn.

STEP 2: Balance the oxidation number by adding the appropriate number of electrons on the appropriate side of the half-reaction.

Manganese has an oxidation number of +7 on the left-hand side and an oxidation number of +2 on the right hand side. Hence, this reaction is a reduction half-reaction and 5 electrons ((+7) – (+2) = 5) must be added on left-hand side of the half-reaction.

$$MnO_4^-{}_{(aq)} + 5\ e^- \rightarrow Mn^{+2}{}_{(aq)}$$
$$+7 \quad - \quad 5 \quad = \quad +2$$

STEP 3: Balance the charges by adding the appropriate number of H^+ on the appropriate side of the reaction when the reaction is run in acidic conditions, or, add the appropriate number of OH^- on the appropriate side of the reaction when the reaction is run in basic conditions.

Whether a reaction is run under acidic or basic conditions is always given in the problem of interest.

The reaction balanced up to step 2 is such that there are 6 negative charges on the LHS of the equation and 2 positive charges on the RHS of the equation.

Acidic Conditions:

We add H^+ to balance the charges. Since there is a differential of 8 charges between the LHS and the RHS, we must add 8 H^+ on the left-hand side of the equation to balance the charges.

The oxidation half-reaction becomes:

$$MnO_4^-{}_{(aq)} + 8\ H^+{}_{(aq)} + 5\ e^- \rightarrow Mn^{+2}{}_{(aq)}$$

Basic Conditions:

We add OH^- to balance the charges. Since there is a differential of 8 charges between the LHS and the RHS, we must add 8 OH^- on the right-hand side of the equation to balance the charges. The oxidation half-reaction becomes:

$$MnO_4^-{}_{(aq)} + 5\ e^- \rightarrow Mn^{+2}{}_{(aq)} + 8\ OH^-{}_{(aq)}$$

<u>STEP 4</u>: Balance the elements hydrogen and oxygen by adding the appropriate number of H_2O molecules on the appropriate side of the half-reaction.

If adding the appropriate number of H_2O molecules does not lead to a fully balanced reaction, then, you must have made a mistake along the way.

Acidic Conditions: You need to add 4 H_2O molecules on the RHS of the equation to get the fully balanced half-reaction:

$$MnO_4^-{}_{(aq)} + 8\ H^+{}_{(aq)} + 5\ e^- \rightarrow Mn^{+2}{}_{(aq)} + 4\ H_2O_{(l)}$$

Basic Conditions: You need to add 4 H_2O molecules on the LHS of the equation to get the fully balanced half-reaction:

$$MnO_4^-{}_{(aq)} + 4\ H_2O_{(l)} + 5\ e^- \rightarrow Mn^{+2}{}_{(aq)} + 8\ OH^-{}_{(aq)}$$

Once, you have balanced the oxidation half-reaction and the reduction half-reaction, you only need to learn how to combine these half-reactions into the full redox reaction (section 2.6.9).

Before moving to the “Interactive Sections on Balancing Half-Reactions”, you may also want to consider the following case, which involves two substances containing different numbers of the element undergoing oxidation or reduction.

Let us consider the oxidation half-reaction involving the following substances:

$$\mathbf{Al_2(Cr_2O_7)_{3\,(s)} \rightarrow CrCl_{2\,(s)} \quad (in\ acidic\ medium)}$$

When balancing half-reactions involving ionic compounds, we only concern ourselves with the monoatomic ions or polyatomic ions containing the element involved in the redox process.

In the above reaction, we see that the element chromium is involved in two compounds having different ratios of chromium to oxygen. Hence, chromium is the element involved in the redox process. We will therefore only focus on the ions, $Cr_2O_7^{-2}$ and Cr^{+2}. Hence, we start writing the half-reaction as:

$$Cr_2O_7^{-2}{}_{(aq)} \rightarrow Cr^{+2}{}_{(aq)}$$

STEP 1: Balance the element.

$$Cr_2O_7^{-2}{}_{(aq)} \rightarrow 2\ Cr^{+2}{}_{(aq)}.$$

STEP 2: Balance the oxidation numbers.

The oxidation number of Cr is +6 in $Cr_2O_7^{-2}$ and is +2 in Cr^{+2}. Difference of 4 ((+6) – (+2)). However, this difference is for one Cr atom. Here, we have two Cr atoms. Hence, 8 electrons must be added on the left-hand side.

$$Cr_2O_7^{-2}{}_{(aq)} + 8\ e^- \rightarrow 2\ Cr^{+2}{}_{(aq)}$$

STEP 3: Balance the charges assuming an acidic medium.

There are 10 negative charges on the left-hand side and 4 positive charges on the right-hand side. Hence, we must add 14 H^+ ions on the left-hand side to balance the charges (acidic medium).

$$Cr_2O_7^{-2}{}_{(aq)} + 8\ e^- + 14\ H^+{}_{(aq)} \rightarrow 2\ Cr^{+2}{}_{(aq)}$$

STEP 4: Balance the elements Hydrogen and Oxygen by adding water.

To balance hydrogen and oxygen, we only need to add 7 H_2O on the right-hand side. The fully balanced half-reaction is written as:

$$Cr_2O_7^{-2}{}_{(aq)} + 8\ e^- + 14\ H^+{}_{(aq)} \rightarrow 2\ Cr^{+2}{}_{(aq)} + 7\ H_2O_{(l)}$$

Sections 2.6.6 - 2.6.8: Balancing Half-Reactions (Interactive)

Practice balancing half-reactions under acidic and basic conditions using these guided Interactive Problems.

Sections 2.6.9 - 2.6.11: Balancing Overall Redox Reaction Equations

In this section, we use the results obtained in Section 2.6.5 (balancing half-reactions) to balance an overall redox reaction. The strategy for balancing an overall redox reaction is best understood through an example.

Let us consider the following redox reaction:

$$\mathbf{MnO_4^-{}_{(aq)} + Fe^{+2}{}_{(aq)} \rightarrow Mn^{+2}{}_{(aq)} + Fe^{+3}{}_{(aq)}}$$

Balancing the overall reaction is done in FOUR PARTS.

<u>PART I</u>: Split the redox equation into an OXIDATION HALF-REACTION and a REDUCTION HALF-REACTION.

Oxidation Half-Reaction: $Fe^{+2}{}_{(aq)} \rightarrow Fe^{+3}{}_{(aq)}$

Reduction Half-Reaction: $MnO_4^-{}_{(aq)} \rightarrow Mn^{+2}{}_{(aq)}$

<u>PART II</u>: Balance the OXIDATION HALF-REACTION.

Following the examples discussed in section 2.6.5, we write the balanced oxidation half-reaction as:

$$Fe^{+2}{}_{(aq)} \rightarrow Fe^{+3}{}_{(aq)} + e^-$$

<u>PART III</u>: Balance the REDUCTION HALF-REACTION.

Following the examples discussed in section 2.6.5, we write the balanced reduction half-reaction as:

$$MnO_4^-{}_{(aq)} + 8\,H^+{}_{(aq)} + 5\,e^- \rightarrow Mn^{+2}{}_{(aq)} + 4\,H_2O_{(l)}$$

Note that this reaction is fully balanced (Mn, H, and O elements, charges (+2 on each side)).

<u>PART IV</u>: Combine the OXIDATION and the REDUCTION HALF-REACTIONS.

Here, we combine the oxidation and the reduction half-reactions to obtain the equation for the full redox reaction. To do so, we need to ensure that <u>the same number of electrons is produced by the oxidation half-reaction and used up in the reduction half-reaction</u>. Since 5 electrons are used up by the reduction half-reaction, the oxidation half-reaction must be multiplied by 5 so that it produces 5 electrons.

$$5\,Fe^{+2}{}_{(aq)} \rightarrow 5\,Fe^{+3}{}_{(aq)} + 5\,e^-$$

5 e^- released by oxidation

$$MnO_4^-{}_{(aq)} + 8\,H^+{}_{(aq)} + 5\,e^- \rightarrow Mn^{+2}{}_{(aq)} + 4\,H_2O_{(l)}$$

5 e^- used up in reduction

We can now add the two half-reactions and obtain:

$$MnO_4^-{}_{(aq)} + 8\,H^+{}_{(aq)} + 5\,e^- + 5\,Fe^{+2}{}_{(aq)} \rightarrow 5\,Fe^{+3}{}_{(aq)} + 5\,e^- + Mn^{+2}{}_{(aq)} + 4\,H_2O_{(l)}$$

The redox equation is then simplified to yield the fully balanced equation (<u>for each element and for the charges</u>):

$$MnO_4^-{}_{(aq)} + 8\,H^+{}_{(aq)} + 5\,Fe^{+2}{}_{(aq)} \rightarrow 5\,Fe^{+3}{}_{(aq)} + Mn^{+2}{}_{(aq)} + 4\,H_2O_{(l)}$$

In Sections 2.6.10, 2.6.11, practice the Interactive Problems on balancing overall redox reaction equations under acidic or basic conditions.

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