SCH4U-C

UNIT 3

Oxidation Reduction Reactions and Energy

Sept. 1, 2016

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Introduction

So far in this course, you've learned that the tiny electron is responsible for all of the chemical reactions that happen in the substances around us. The electron arrangement of the atom of an element determines the element's reactivity. Furthermore, knowing how many outermost or valence electrons an atom has allows chemists to predict trends in the reactivity of elements.

Similarly, the reactivity of compounds is also determined largely by the electrons within the bonds of the compound. For example, in Unit 2, you saw that electron-rich portions of organic molecules, like the double bonds of alkenes, are often the most reactive parts of the molecule. Many chemical reactions occur as a result of electron transfers from one reactant to another. Chemical reactions in which one entity loses electrons, while another gains electrons, are called oxidation reduction reactions (redox reactions).

In the first half of this unit, you will study the nature of redox reactions, as well as their applications. The second half of the unit focuses on the energy changes associated with these reactions.

Overall Expectations

After completing this unit, you will be able to

- assess health, safety, and environmental issues related to industrial applications of electrochemistry
- recognize and explain redox reactions, in terms of a loss and gain of electrons
- explain the energy changes in exothermic and endothermic reactions, and describe their applications
- solve problems using $q = mc\Delta T$, Hess's law, and tables of standard enthalpies of formation
- evaluate and report on research into the feasibility of replacing fossil-fuel energy sources with fuel-cell technology

You can access this online Periodic Table and this Scientific Calculator any time you need them. You will find links to them on the Course Materials page.

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9

Oxidation Reduction Reactions

Introduction

Lead contamination in drinking water is a problem facing many Ontario municipalities. This is a major concern because prolonged exposure to lead can result in brain and nervous system impairments, especially in children. One major source of lead is the solder used to join lengths of copper pipe together. Solder is a soft, pliable metal that can be melted into the joint between two copper pipes. Once the solder hardens, it forms a watertight seal, fusing the pipes together permanently.

Figure 9.1: Two pipes can be soldered together with lead or other soft metals.

Source: http://commons.wikimedia.org/wiki/File:US_Navy_060528-N-1332Y-022_Hull_Maintenance_Technician_Fireman_Bob_Chambers_uses_a_brazing_torch_to_weld_a_one-and-a-half-inch_copper-nickel_pipe_in_the_pipe_shop_aboard_USS_Kitty_Hawk_%28CV_63%29.jpg

If your home was built before 1980, lead solder was likely used in the plumbing. Most tap water is slightly acidic. Lead ions get into the water by slowly reacting with any acidity (hydrogen ions) that may be in the water. The net ionic equation for this reaction is:

$$Pb_{(s)} + 2H^{+}_{(aq)} \rightarrow Pb^{2+}_{(aq)} + H_{2(g)}$$

Some of the hydrogen ions in this reaction come from the water itself. Most, however, come from any acids that may be in the water.

The conversion of metals to positive ions or corrosion of metals like lead and steel is one common example of an oxidation reduction reaction, often called a redox reaction. This lesson introduces you to redox reactions and some of their applications. You will learn how to recognize redox reactions, and how to write the chemical equations used to describe them.

Planning Your Study

You may find this time grid helpful in planning when and how you will work through this lesson.

Suggested Timing for This Lesson (Hours)	
Redox Reactions	1
Oxidizing and Reducing Agents	¾
Activity: Observing Oxidizing and Reducing Agents	¾
Balancing Redox Reactions	1
Key Questions	1

What You Will Learn

After completing this lesson, you will be able to

- identify redox reactions
- assign oxidation numbers to elements in a redox reaction
- distinguish between oxidizing and reducing agents
- balance redox equations using half-reactions
- assess the impact of the corrosion of lead on human health

Redox Reactions

Most of the chemical reactions you encounter each day are oxidation reduction (or redox) reactions. For example, the formation of rust on the bolts of a car, the browning of bread as it toasts, and the combustion reactions that you learned about in the previous unit, are all examples of redox reactions. Before "oxidation" and "reduction" are defined, take a historical look at how these terms developed.

Figure 9.2: Smelting metal with a primitive forge.

Source: http://commons.wikimedia.org/wiki/File:Escalade_2010_forge.JPG

Thousands of years ago, people developed the skills they needed to build primitive smelters that could extract metals from the rocks that they found or mined from the earth. Our ancestors used these skills to craft metal tools, weapons, and jewellery. By the eighteenth and nineteenth centuries, chemists realized that metals were bound within compounds (two or more different elements bonded together), within the rock. The process of producing a purer, simpler substance (the metal) from something more complex (the minerals) became known as reduction. Early chemists also knew that some metals readily react with oxygen. These reactions became known as oxidation reactions.

Historically, oxidation and reduction were seen as unrelated processes. This changed with the discovery of the electron in the late nineteenth century. Consider, for example, the reaction between zinc and oxygen, which forms zinc oxide:

$$2Zn_{(s)} + O_{2(g)} \rightarrow 2ZnO_{(s)}$$

During this reaction, zinc and oxygen atoms react to form the ionic compound, zinc oxide. Zinc oxide is composed of two ions: Zn^{2+} and O^{2-}. To produce these ions, two electrons from zinc metal must be transferred to each oxygen atom in the reaction. In other words, zinc loses two electrons. In fact, the oxygen atom takes the electrons from the zinc atom.

Conversely, when oxides like zinc oxide are heated, they can be decomposed into the elements they contain. For example:

$2ZnO_{(s)} \rightarrow 2Zn_{(s)} + O_{2(g)}$

Scientists working in the early twentieth century were able to analyze the reaction and explain what happens at the atomic level. During this reaction, two electrons from the oxide ion are transferred to the zinc ion to form neutral atoms. In other words, zinc gains two electrons. To do this kind of analysis, it is essential to have the periodic table (which is located at the back of this unit) for reference. In general, when metals and non-metals bond, metals lose electrons and non-metals gain electrons.

As a result of the analysis of reactions like these, the traditional definitions of oxidation and reduction changed. Oxidation became known as the loss of electrons, while reduction became known as the gain of electrons. For example, during the formation of zinc oxide, zinc loses two electrons to form an ion:

$Zn \rightarrow Zn^{2+} + 2e^-$ (oxidation) When oxidation of an atom or ion occurs, the atom or ion in question becomes more positive; the oxidation value has increased.

$O + 2e^- \rightarrow O^{2-}$ (reduction) When reduction of an atom or ion occurs, the atom or ion becomes less positive; the oxidation value has decreased. The term “reduction” becomes more logical when considered this way.

This example illustrates that reduction and oxidation are related terms. Oxidation cannot occur without reduction, because electrons lost by one reactant must be gained by another.

Have a look at how the definitions of oxidation and reduction apply to reactions that don’t involve oxygen. For example, magnesium reacts with sulfur to form magnesium sulfide, MgS, an ionic solid:

$Mg_{(s)} + S_{(s)} \rightarrow MgS_{(s)}$

This reaction can also be broken down into two steps:

$Mg \rightarrow Mg^{2+} + 2e^-$ (loss of electrons: oxidation) The Mg atom became more positive.

$S + 2e^- \rightarrow S^{2-}$ (gain of electrons: reduction) The S atom became less positive. The charge decreased from 0 to –2.

The definitions of oxidation and reduction also apply to reactions involving only non-metals. For example, hydrogen burns vigorously in chlorine to form hydrogen chloride:

$H_{2(g)} + Cl_{2(g)} \rightarrow 2HCl_{(g)}$

The hydrogen and chlorine atoms undergo the following changes, as they form each molecule of HCl:

$H \rightarrow H^+ + e^-$ (loss of electrons: oxidation) H became more positive. The charge increased from 0 to +1.

$Cl + e^- \rightarrow Cl^-$ (gain of electrons: reduction) The Cl became less positive. The charge decreased from 0 to –1.

Oxidation Numbers

In order to recognize which reactant is oxidized and which is reduced, you must be able to follow the movements of electrons in a reaction. To facilitate this, chemists invented the concept of oxidation numbers. An oxidation number is the real or apparent charge of an ion or an atom, if the atom's bonds are considered to be ionic.

This idea can be illustrated with a couple of examples. Consider any element that is not part of a compound (or an ion). The oxidation number of any element is 0, because elements with equal numbers of protons and electrons are electrically neutral. For example, magnesium, Mg, and diatomic chlorine, Cl_2, have an oxidation number of 0. Conversely, ions or charged atoms never have an oxidation number of 0. The oxidation number of simple ions, such as oxide, O^{2-}, is indicated by the charge of the ion. Therefore, the oxidation number of oxygen in the oxide ion is –2.

In ionic compounds, such as zinc oxide, ZnO, the oxide ion still has a charge of –2. Therefore, the oxidation number of the oxygen ion in ZnO is –2 as well, although the compound itself is neutral. In order for the compound to be neutral, you can deduce that zinc must have an oxidation number of +2. In molecular compounds, such as carbon monoxide, CO, oxygen is covalently bonded to carbon. Therefore, carbon monoxide does not contain any ions. However, for the purposes of defining oxidation numbers, the oxygen atom in this compound still has an oxidation number of –2, even though an O^{2-} ion does not exist in this compound. In order for carbon monoxide to be electrically neutral, the carbon atom in this compound must have an oxidation number of +2. Note, however, that a C^{2+} ion does not exist in this compound.

Rules for Assigning Oxidation Numbers

The following rules summarize how to assign oxidation numbers. (Think of these rules as being like those an accountant would follow when doing simple accounting.)

1. The oxidation number of an element is always 0, regardless of its subscript. For example, Zn, C, H_2, O_3, and S_8 all have an oxidation number of 0 because they are all naturally occurring elements. This is always true, as long as they are not part of a compound.
2. The oxidation number of a simple ion is the charge of the ion. For example, the oxidation number of chlorine in Cl^- is –1. The oxidation number of nitrogen in N^{3-} is –3.
3. The oxidation number of oxygen in compounds is –2. For example, oxygen has the oxidation number of –2 in each of the following: MgO, PbO_2, ClO_3^-, and $Cr_2O_7^{2-}$.
4. The oxidation number of hydrogen in compounds is +1. For example, the hydrogen in these compounds has an oxidation number of +1: HCl, H_2O, NaOH, and HCO_3^-.
5. Group 1 (Li family) and Group 2 (Be family) metals always have a +1 and +2 oxidation value, respectively, in compounds.
6. All other oxidation numbers are assigned so that the net sum of the oxidation numbers equals the net charge of the molecule or complex ion.

Example

Determine the oxidation number of sulfur in the following molecules:

a) H_2S

b) SO_2

c) SO_3^{2-}

Solution

a) H_2S (This molecule has an overall charge of zero.)

$(+1)2 + S = 0$ (Rule 1: The value for H in all compounds is +1.)
(Rule 6: The sum of all oxidation values in the formula must add to zero.)

$2 + S = 0$

$S = -2$

For the compound to have an overall charge of 0, sulfur must have an oxidation number of −2.

b) SO_2 (This molecule has no net charge.)

$S + (-2)2 = 0$ (Rule 3: The oxidation value for oxygen in compounds is −2.
Rule 6: The sum of all values must equal zero then.)

$S - 4 = 0$

$S = +4$

For the compound to have an overall charge of 0, sulfur must have an oxidation number of +4. Some periodic tables will list all of the common oxidation values that elements exhibit in compounds. These are shown on your periodic table at the end of this unit, although the +4 value for S is not shown on that table. In the Grade 11 chemistry course, you probably would not have encountered this +4 value for sulfur. In Grade 11, we would have predicted that sulfur would have an oxidation value of -2 because of its proximity to the noble gas, Argon. When dealing with redox reactions, you will encounter oxidation values that you will not be able to rationalize using the "nearest noble gas concept."

c) SO_3^{2-} (This complex ion has an overall charge of –2. Therefore, the sum of all oxidation values in this compound must add to –2, Rule 6 again.)

$S + (-2)3 = -2$

$S - 6 = -2$

$S = +4$

For the ion to have an overall charge of –2, sulfur must have an oxidation number of +4.

The above examples illustrate that elements can have two or more different oxidation numbers, depending on the other elements they bond with. The periodic table at the end of this unit includes common oxidation values (states) that elements will exhibit in various compounds.

Use of Oxidation Numbers

In a redox reaction, one element gains electrons, while another loses them. Oxidation numbers are a useful way of keeping track of these electron transfers. They are also a tool for identifying redox reactions. Consider the following sequence of chemical reactions:

1. $S + O_2 \rightarrow SO_2$
2. $2SO_2 + O_2 \rightarrow 2SO_3$
3. $SO_3 + H_2O \rightarrow H_2SO_4$

These three reactions are key steps in the formation of acid rain. Sulfur is often found as an impurity in fossil fuels like coal. Combustion of coal containing sulfur impurities results in the formation of sulfur dioxide (reaction 1). This compound then undergoes a further reaction with oxygen in the atmosphere to form sulfur trioxide (reaction 2). Then, water droplets in the atmosphere combine with sulfur trioxide to form sulfuric acid, which falls as acid rain. Acid rain harms animals, insects, and plants and can wreak havoc on aquatic ecosystems found in our streams, rivers, and lakes as these ecosystems continue to grow more acidic with each passing acid rainfall.

Now, take a closer look at the three reactions that form acid rain. Here is how to assign the oxidation numbers to the elements in these compounds, using the rules established previously.

1. $S + O_2 \rightarrow SO_2$

   $0 \quad 0 \rightarrow +4 \; -2$

2. $2SO_2 + O_2 \rightarrow 2SO_3$

   $+4 \; -2 \quad 0 \rightarrow +6 \; -2$

3. $SO_3 + H_2O \rightarrow H_2SO_4$

$+6 \;\; -2 \;\; +1 \;\; -2 \rightarrow +1 \;\; +6 \;\; -2$

Can all three reactions be classified as redox reactions? In reactions 1 and 2, the oxidation numbers of sulfur and oxygen changed. Therefore, these are redox reactions. Sulfur was oxidized in both reactions because its oxidation number became more positive (as if it had lost electrons). Oxygen was reduced in both cases, because its oxidation number decreased, becoming more negative (it had gained electrons). However, no change in oxidation numbers occurred in reaction 3. Therefore, this reaction is not a redox reaction.

Support Questions

Be sure to try the Support Questions on your own before looking at the suggested answers provided.

1. Describe why oxidation can't occur without reduction.

Suggested answer

2. Determine the oxidation number of the underlined elements, in each of the following. (**Hint:** The rules remind us that certain elements are special, so their known values are assigned based on the rules. Review rules 1, 3, 4 , and 5.)

a) $\underline{S}_8$

b) $H_3\underline{P}$

c) $Na_2\underline{C}O_3$

d) $K_2\underline{Cr}O_4$

e) $\underline{Mn}O_4^-$

f) $\underline{Cr}_2O_7^{2-}$

Suggested answer

3. Determine whether the following reactions are redox reactions (answer "yes" or "no"). (**Hint:** Has one element increased in value, meaning that oxidation has occurred to it? Has another decreased in value, meaning that reduction has occurred to it? Place the values above or below each atom in the formulas, using our rules, so that you can decide.)

a) $Co + 2Fe^{3+} \rightarrow Co^{2+} + 2Fe^{2+}$

b) $HNO_3 + NaOH \rightarrow NaNO_3 + H_2O$

c) $2Al + 3Br_2 \rightarrow 2AlBr_3$

d) $2HClO_2 + 3I^- + 3H^+ \rightarrow Cl_2 + 3HIO + H_2O$

Suggested answer

Oxidizing and Reducing Agents

The previous examples illustrate that during a redox reaction, one reactant is oxidized, while the other is reduced. Both reactants must be present; otherwise the reaction does not occur. Therefore, one reactant causes the other's oxidation. The reactant that causes the other's oxidation is called the oxidizing agent. An oxidizing agent would be something that takes electrons from another. Similarly, the reducing agent is a reactant that causes another reactant's reduction. These concepts will be clarified below.

In the reaction of magnesium with sulfur, you must assign oxidation numbers to decide which element is the oxidizing agent.

$Mg_{(s)} + S_{(s)} \rightarrow MgS_{(s)}$

0 0 → +2 −2

In the reaction of magnesium with sulfur, we can see that the magnesium increased from an oxidation value of 0 to a value of +2. This increase in oxidation value is what is termed "oxidation." It was the sulfur atom that caused this oxidation of magnesium to occur, as the sulfur took electrons from the magnesium atom. Because the sulfur atom caused the oxidation, we call sulfur the oxidizing agent. Because sulfur was reduced by contact with the magnesium atom, we say that magnesium was the reducing agent.

In summary:

- Sulfur was reduced by taking electrons, so acted as the oxidizing agent.
- Magnesium was oxidized as it lost electrons, so acted as the reducing agent.

The terminology *oxidation*, *reduction*, *reducing agent*, and *oxidizing agent* can be confusing. Consider, very carefully, the meaning of oxidation and reduction in terms of electron loss and gain. When something has been oxidized, it has had electrons taken from it (by the oxidizing agent). It is really the oxidizing agent that makes everything happen, since that element takes the electrons from the other.

In the hydrogen and chlorine example mentioned earlier and shown again below, chlorine is the oxidizing agent because it causes the oxidation of hydrogen. Hydrogen is a reducing agent because it causes the reduction of chlorine.

$H_{2(g)} + Cl_{2(g)} \rightarrow 2HCl_{(g)}$

0 0 → +1 −1

Consider lead reacting with acid—another reaction that you read about earlier in this lesson. This reaction is shown below.

$Pb_{(s)} + 2H^{+}_{(aq)} \rightarrow Pb^{2+} + H_{2(g)}$

0 +1 → +2 0

In this case, hydrogen ions from the acidity of the water oxidize the lead, while the lead reduces the hydrogen ions. The hydrogen ions are the oxidizing agent here. The hydrogen ions took electrons from the lead metal.

Complex ions or neutral compounds may also be oxidizing and reducing agents. Consider the reaction that occurs when nitrogen dioxide gas bubbles in a solution containing the silver ions (Ag^+):

$Ag^+_{(aq)} + NO_{2(g)} + H_2O \rightarrow Ag_{(aq)} + NO_3^-{}_{(aq)} + 2H^+_{(aq)}$

$+1 \quad +4\ -2 \quad +1\ -2 \rightarrow 0 \quad +5\ -2 \quad +1$ (oxidation values assigned by rule; N was calculated)

Note that silver is reduced, while the nitrogen in nitrogen dioxide is oxidized. However, the oxidation numbers of hydrogen and oxygen do not change. Silver ions were reduced from +1 in value to 0 in value. Therefore the silver ions acted as the oxidizing agent. The silver ions took electrons from something.

The N, in NO_2, increased in value from +4 to +5, so was oxidized (by the silver ion, as we said). The NO_2 is said to be the reducing agent.

Support Questions

4. Use oxidation numbers to identify the oxidizing and reducing agents in the following reactions. (**Hint:** Assign oxidation values by rule. Determine what was oxidized—that was the reducing agent. Decide what was reduced—that was the oxidizing agent.)

a) $CH_4 + O_2 \rightarrow CO_2 + H_2O$

b) $Pb + H_2SO_4 \rightarrow H_2 + PbSO_4$

c) $2Mg + CO_2 \rightarrow C + 2MgO$

Suggested answer

Activity: Observing Oxidizing and Reducing Agents

This experiment is an online activity.

Metals vary in their reactivity with acids. In this experiment, you will compare the reactivity of copper, nickel, zinc, and iron with acetic acid (vinegar). The general equation for the reaction of a metal with acid is:

$M_{(s)} + 2HA_{(aq)} \rightarrow H_{2(g)} + MA_{(aq)}$, where HA is the acid.

Now, oxidation numbers will be assigned to this reaction:

0 +1 –1 → 0 +1 –1

$M_{(s)} + 2HA_{(aq)} \rightarrow H_{2(g)} + MA_{(aq)}$

You should note that the charge on the acid anion, A^-, is –1, in this case, because the acid has only one hydrogen atom. Since the metal's oxidation number changes from 0 to +1, the metal is oxidized. Hydrogen's oxidation number changes from +1 to 0. Therefore, hydrogen is reduced. In conclusion, the metal (which was oxidized) is the reducing agent, while the hydrogen ion (which was reduced) is the oxidizing agent.

In this experiment, you will compare the relative strength of four reducing agents. The metals are the potential reducing agents since they could be oxidized. So when we say that we will compare the relative strengths of the these metals as reducing agents, we are saying we will compare their tendencies to be oxidized to positive ions.

Purpose: To rank copper, nickel, zinc, and iron in order of decreasing strength as reducing agents

Procedure:

Step 1: Check out Developing an Activity Series. In this activity, you will be asked to study the observations and decide which metal was most readily oxidized and so is the strongest reducing agent.

The acid used in this experiment is acetic acid. The chemical formula of acetic acid can be written in two forms: CH_3COOH or $HC_2H_3O_2$. The second form of the formula fits the pattern in the general equation, $HA = HC_2H_3O_2$, where "A" is the anion released by the acid (which is the acetate ion, $C_2H_3O_2^-$). "M" is any metal that will react with the acid. Therefore, the reaction with acetic acid becomes:

$M_{(s)} + 2HC_2H_3O_{2(aq)} \rightarrow H_{2(g)} + MC_2H_3O_{2(aq)}$

Step 2: Do the online activity. You will observe or read about a number of chemical changes occurring. Focus only on reactions that produce hydrogen gas.

Step 3: Study the observation table provided and answer the Support Questions related to the experiment.

Support Questions

These questions are based on the experiment you have just done.

5. What evidence in the experiment shows that hydrogen ions are being reduced?

Suggested answer

6. Which metal produced hydrogen gas most vigorously?

Suggested answer

7. Which metals showed no evidence of being able to reduce hydrogen ions to produce hydrogen gas?

Suggested answer

8. Rank the metals in increasing order of their strengths as reducing agents (tendency to be oxidized to metal ions). (**Hint:** The strongest reducing agent is the metal that was the most aggressively oxidized.)

Suggested answer

9. In a related experiment, strips of the metals A, B, C, and D were placed in their corresponding solutions, A^{2+}, B^{2+}, C^{2+}, and D^{2+}. The following data were observed, showing which metal–metal ion combinations resulted in a chemical reaction. For example, the chemical equation for the reaction of metal A with a B^{2+} solution is:

$$A_{(s)} + B^{2+}_{(aq)} \rightarrow B_{(s)} + A^{2+}_{(aq)}$$

If this type of reaction does occur, then we would conclude that A has a greater tendency to be oxidized than B. This would make A a stronger reducing agent than B.

	A	B	C	D
A^{2+}	No reaction	No reaction	No reaction	No reaction
B^{2+}	Reaction	No reaction	Reaction	Reaction
C^{2+}	Reaction	No reaction	No reaction	Reaction
D^{2+}	Reaction	No reaction	No reaction	No reaction

a) Rank the metals in order of increasing strength as reducing agents (tendency to be oxidized). Write them in a column with the strongest reducing agent (SRA) on the bottom of the column. Label it "SRA" to the right of the metal letter. Beside "SRA" write "greatest tendency to be oxidized."

b) Rank the metal ions in increasing strength as oxidizing agents, but this time place them in a column with the strongest oxidizing agent (SOA) on the top. Place the letters "SOA" to the left of the ion and write "greatest tendency to be reduced." If an ion took electrons from a certain metal, as we see in the example reaction above the table, then that ion is an oxidizing agent. Notice in the table that when metal $A_{(s)}$ met the ion B^{2+}, a reaction did occur: the ion was able to take electrons from the metal A. However, in another section of the table grid we see that when the $B_{(s)}$ metal meets the A^{2+} ion, no reaction occurred. The A^{2+} was unable to take electrons form B. Conclusion: The B^{2+} ion is a better oxidizing agent than the A^{2+} ion.

c) Place the two columns side by side, with the ion column to the left of the metal column. Place the acronyms SRA and SOA in the proper locations. Place the phrases beside SRA and SOA as earlier.

Compare the ion and metal pairs, M+ and M, at the top of the columns and at the bottom of the columns. Comment on the pairs in terms of strengths. Suggested answer

Balancing Redox Reactions

Many redox equations can be balanced through trial and error. Using this approach, you alter the coefficients of each chemical in the chemical equation, and then check if your trial equation balances. You continue doing this until the total numbers of atoms of each type on both sides are equal. The following example is a balanced chemical equation for a redox reaction:

$2Al + 6HCl \rightarrow 2AlCl_3 + 3H_2$

Note that there are two aluminums, six hydrogens, and six chlorines on both sides of the equation.

Redox reactions occurring in aqueous solutions are often far more complicated than this example. Therefore, balancing them through trial and error can be very time-consuming. This section introduces you to a systematic and efficient way of balancing redox reactions, using your understanding of these reactions.

Example

Write a balanced chemical equation for the reaction of the permanganate ion, MnO_4^-, with the iron(II) ion, Fe^{2+}. The net ionic equation for this reaction is:

$MnO_4^- + Fe^{2+} \rightarrow Mn^{2+} + Fe^{3+}$

Solution

Here is a summary of the steps and sub-steps involved in balancing a redox equation, using half-reactions:

Step 1: Identify and write the two half-reactions.

Step 2: Balance all of the elements except hydrogen and oxygen.

Step 3: Balance the oxygen atoms by adding H_2O to the appropriate side.

Step 4: Balance the hydrogen by adding H^+.

Step 5: Balance the charges by adding electrons.

Step 6: Multiply one or both of the half-reactions by a whole number so that the number of electrons gained and lost is equal.

Step 7: Add the half-reactions. Subtract any chemicals that are common to both sides.

Step 1: Identify and write the two half-reactions. Write the oxidation and reduction half-reactions and assign oxidation numbers to elements other than hydrogen and oxygen.

$Fe^{2+} \rightarrow Fe^{3+}$

$+2 \rightarrow +3$

The Fe^{2+} lost an electron so it is oxidized.

$MnO_4^- \rightarrow Mn^{2+}$

$+7 \rightarrow +2$

The Mn gained the equivalent of five electrons, so it is reduced.

Therefore, the overall reaction can be broken down into two half-reactions:

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

and

$Fe^{2+} \rightarrow Fe^{3+}$ (oxidation)

Step 2: Balance all of the elements except hydrogen and oxygen.

$MnO_4^- \rightarrow Mn^{2+}$

Manganese is already balanced.

Step 3: Balance the oxygen atoms by adding H_2O to the appropriate side. To balance oxygen atoms, water molecules (which contain oxygen) are added to the right-hand side (RHS) of the equation. Since you know that the reaction occurs in an aqueous (water-based) solution, it makes sense that water could be involved in the reaction.

$MnO_4^- \rightarrow Mn^{2+} + 4H_2O$ (There are no oxygen atoms to balance in the iron half-reaction.)

Step 4: Balance the hydrogen by adding H^+. Now hydrogen ions, H^+, are added to the left-hand side (LHS) to balance the addition of water. H^+ is used rather than H_2, because it has the same oxidation number as the hydrogen in water (+1).

$8H^+ + MnO_4^- \rightarrow Mn^{2+} + 4H_2O$ (There are no H atoms to balance in the iron half-reaction.)

Step 5: Balance the total charges by adding electrons. The total number of atoms of each type have already been balanced. However, the total charges of the LHS and RHS are not balanced. You find the total charge of each side by multiplying the charge of each ion by its coefficient and then adding all of the ions on one side. First examine the manganese half-reaction for charges.

The sum of charges of the LHS is 8(+1) + (–1) = +7. (8 hydrogen ions plus one permanganate ion)

The sum of charges of the RHS is +2. The right-hand side has one manganese ion with a +2 charge and four water molecules that have no charge.

Electrons are added to balance the charges. In this case, $5e^-$ are added to the LHS to bring the sum of its charges to the same total value of +2.

$8H^+ + MnO_4^- + 5e^- \rightarrow Mn^{2+} + 4H_2O$

$8(+1) + (-1) + 5(-1) = +2$

Now, the same is done for the other half-reaction:

$Fe^{2+} \rightarrow Fe^{3+} + e^-$ (This half-reaction needs just one electron on the RHS to make the charge on both sides be equal. It is coincidental that both half-reactions became equal charges of +2.)

$(+2) = (+3) + (-1)$

$+2 = +2$

Step 6: Multiply one or both of the half-reactions by a whole number so that the number of electrons gained and lost is equal. Remember that redox is the transfer of a specific number of electrons from one entity to another. Now, consider the two half-reactions, which have been numbered (i) and (ii).

i) $8H^+ + MnO_4^- + 5e^- \rightarrow Mn^{2+} + 4H_2O$ (reduction)

ii) $Fe^{2+} \rightarrow Fe^{3+} + e^-$ (oxidation)

Notice that electrons are on the LHS of (i). This shows that electrons are gained by one or more reactants (LHS) when they form the products (RHS). This agrees with the earlier prediction, using oxidation numbers, that Mn is reduced. Notice also that in (ii), an electron is lost. This is oxidation and agrees with the earlier prediction. However, in a redox reaction, the electrons gained by one element have to come from the element that loses electrons. Therefore, the number of electrons gained and lost must be equal.

To accommodate this, both sides of (ii) are multiplied by 5. Since the same operation is applied to both sides of the equation, the equation remains valid.

i) $8H^{+} + MnO_4^{-} + 5e^{-} \rightarrow Mn^{2+} + 4H_2O$

ii) $[Fe^{2+} \rightarrow Fe^{3+} + e^{-}] \times 5 = 5Fe^{2+} \rightarrow 5Fe^{3+} + 5e^{-}$

Now the half-equations can be added.

Step 7: Add the half-reactions. Subtract any chemicals that are common to both sides.

$8H^{+} + MnO_4^{-} + 5e^{-} + 5Fe^{2+} \rightarrow 5Fe^{3+} + 5e^{-} + Mn^{2+} + 4H_2O$

Since there are five e^{-} on both sides, these can be subtracted (removed) from the equation:

$8H^{+} + MnO_4^{-} + 5Fe^{2+} \rightarrow 5Fe^{3+} + Mn^{2+} + 4H_2O$ (final balanced chemical equation)

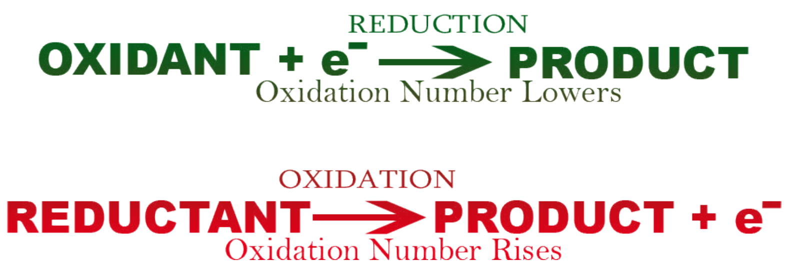

Figure 9.3: Remember that gaining electrons is reduction, while losing electrons is is oxidation.

Source: http://commons.wikimedia.org/wiki/File:Redox_Reminder.png

Support Questions

10. Write a balanced chemical equation for the following net ionic equations. (Follow the steps illustrated earlier. Be extra careful with the step involving balancing total charges—this is the step where most errors are made.)

a) $MnO_4^{-} + SO_3^{2-} \rightarrow Mn^{2+} + SO_4^{2-}$ Suggested answer

b) $I_2 + HNO_3 \rightarrow HIO_3 + NO + H_2O$ (**Hint:** if the water part of this equation does not have any elements that are oxidized or reduced, ignore it for now. The balancing procedure will take care of water.) Suggested answer

Key Questions

Now work on your Key Questions in the online submission tool. You may continue to work at this task over several sessions, but be sure to save your work each time. When you have answered all the unit's Key Questions, submit your work to the ILC.

33. Determine the oxidation number of the underlined element in each of the following chemical formulas. Your answers must include any calculations and/or reasoning in how you derived the oxidation numbers. **(4 marks total)**

a) $\underline{O}_3$

b) $H_3\underline{P}O_4$

c) $\underline{Mn}O_3^{-}$

d) $\underline{C}_2O_4^{2-}$

34. **(5 marks total)**

a) Identify the element oxidized and the element reduced, in this chemical equation: $H_3AsO_4 + H_2S \rightarrow H_3AsO_3 + S + H_2O$ **(2 marks)**

b) Identify the oxidizing agent and the reducing agent. **(2 marks)**

c) Justify why this is a redox reaction. **(1 mark)**

35. Balance the following redox equation, using half-reactions. All steps must be shown in your solutions for full marks to be earned. Assume that the reaction occurs in an aqueous solution. **(6 marks)**

$Cr_2O_7^{2-} + NO \rightarrow Cr^{3+} + NO_3^{-}$

36. Small pieces of silver, copper, and magnesium are placed in solutions that contain one of the following ions: Ag^{+}, Cu^{2+}, and Mg^{2+}. The metal/solution combinations tested are summarized in the following table. Any reaction that occurred followed the pattern

$M_{(s)} + B^{+}_{(aq)} \rightarrow B_{(s)} + M^{+}_{(aq)}$, where M is a general metal and B^{+} is a general metal ion.

Here is the summary of the observations:

	Ag^{+}	Cu^{2+}	Mg^{2+}
Ag	No reaction	No reaction	No reaction
Cu	Reaction	No reaction	No reaction
Mg	Reaction	Reaction	No reaction

In a column, list the oxidizing agents (those being reduced), with the SOA (strongest oxidizing agent) at the top. In a second column, list the reducing agents (those being oxidized), with the SRA (strongest reducing agent) at the bottom. Label the SOA and the SRA. **(4 marks)**

37. A water-testing report showed that the water in the drinking fountains in a 50-year-old elementary school contained slightly more than the acceptable level of dissolved lead. Lead levels were highest in the morning and then decreased during the school day. **(8 marks total)**

a) Identify a possible source of the lead contamination.

b) Why was the presence of lead a concern?

c) Why was the amount of lead in the water highest in the sample taken early each morning?

d) Closing or renovating the school was not an option. The principal considered two options for supplying the school with fresh, safe drinking water. Identify one advantage and one disadvantage of each of these two options, which follow. In your opinion, which option should she choose? Why?

i) Supply bottled water to 500 students and staff each day.

ii) Have custodians open the taps each morning for about 30 minutes.

Now go on to Lesson 10. Send your answers to the Key Questions to the ILC when you have completed Unit 3 (Lessons 9 to 12).