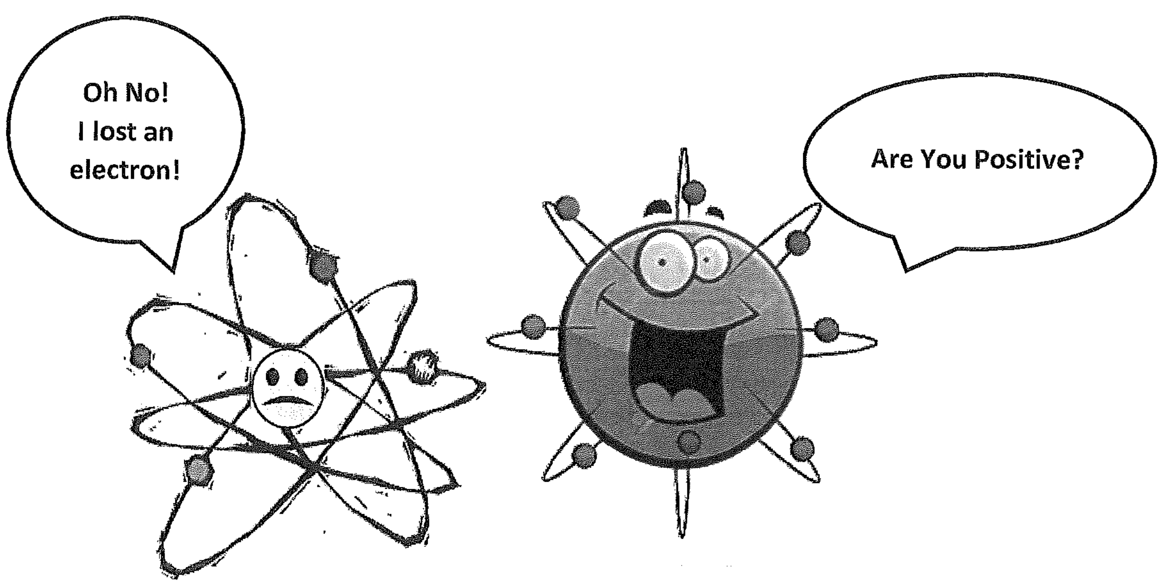
40S Chemistry

Preview

Welcome to Grade 12 Chemistry!

Before starting this course, take some time to review some of the fundamentals of the chemistry you will be studying. Your understanding of these topics will be necessary for your success in this course.

Once you are comfortable with the material presented in this review, you will complete a short assessment. This will provide you with opportunity to discuss these topics with your instructor.



Lesson 1
The Atom

Just as a cell is considered the building block of life, atoms are considered the building blocks of matter. And just as a cell is made of smaller components, atoms can be further broken down into subatomic particles. The three main subatomic particles are **protons, neutrons**, and **electrons**. Today, scientists know subatomic particles in turn can be broken down into still smaller particles.

The Bohr model of the atom suggested that electrons travel in a circular orbit or energy level in a well-defined path. Scientists have done experiments to show that this is not the path electrons actually travel. Electron location and movement is more complicated than Bohr thought; however, the planetary model of the atom is used because it is easier to understand and is adequate for a beginning course in chemistry.

The modern atomic theory states that an atom consists of the following:

- Protons are located in the nucleus of the atom. Protons have a positive charge and have a mass of one atomic mass unit (u).
- Neutrons are also located in the nucleus of the atom. Neutrons have no electrical charge and have a mass of approximately 1 u.
- Electrons are located around the nucleus in what is referred to as orbits. Electrons have a single negative electrical charge but their mass is considered zero since it is so small.

Atoms

An atom consists of a dense nucleus surrounded by electrons moving in space. A nucleus is made of protons and neutrons. The total mass of an atom is equal to the sum of the masses of its protons and neutrons. (remember: electrons have negligible mass!)

Number protons + number neutrons = atomic mass

The atomic number of an atom is equal to the number of protons in the nucleus. This is a unique number that can be used to identify any one specific atom.

Electrons

A neutral atom will have the same number of electrons as it has protons. Electrons can be found in orbits surrounding a nucleus. These orbits are actually energy levels. It takes less energy to keep an electron in orbit near the nucleus, than it does to have it orbit further away from the nucleus. Elections, therefore, want to be as close to the nucleus as possible (it requires less energy). However, there is not always enough room for all an atom's electrons near the nucleus.

There are "rules" for how electrons are arranged around a nucleus. An atom can have a maximum of only two electrons occupying the first energy level, or orbit, nearest the nucleus, but it can hold up to eight electrons in each of the next energy levels. This is the **Octet Rule** (rule of 8). Each lower energy level must be filled before electrons will occupy the next level.

Electrons, called **valence electrons,** occupy the outermost levels and are of importance because they are the ones available for reaction. Valence electrons determine the chemical properties of an atom; how an atom interacts with other atoms. Some atoms can gain extra valence electrons, becoming more negative, while others can lose valence electrons, becoming more positive.

Atomic Number

All atoms are made of protons, neutrons, and electrons. A proton in a carbon atom is the same as a proton in an oxygen atom. The same is true for electrons and neutrons. So, if all atoms are made of the same subatomic materials, what determines the difference between atoms of different elements? The answer is that atoms of different elements have different numbers of protons in the nucleus and electrons around the nucleus.

The atomic number is equal to the number of protons in an atom's nucleus. The atomic number is special because it can be used to identify any known element. It is easy to identify copper because it has an atomic number of 29. Carbon has an atomic number of six, oxygen has an atomic number of eight, and so on.

Mass Number

All atoms are given a mass number. The mass number is equal to the number of protons plus the number of neutrons; for example, carbon has six protons and six neutrons.

The mass number for carbon would be 6 (protons)+ 6 (neutrons) = 12.

The atomic mass of an atom is expressed in atomic mass units (u). This means the atomic mass of carbon is expressed as 12 u.

The Periodic Table

The properties of elements are a periodic function of their atomic numbers, and therefore, the elements can be organized in a periodic table. It is an important tool as it allows chemists to quickly determine some key facts about an element and allows us to predict how that element should interact with another element.

Columns in the Periodic Table

A single column in the periodic table is called a **family**. A family contains elements that have similar but not identical properties. The alkali metals, alkaline-earth metals, chalcogens, halogens, and noble gases are all examples of families.

Hydrogen is a special case because it is a family of one. Sometimes hydrogen behaves as a metal and sometimes as a non-metal. Hydrogen has one electron in its outermost energy level, so it is reactive. Almost all the hydrogen on Earth is combined with other materials or with itself.

Alkali Metals

The alkali metal family occupies the first column in the periodic table and includes lithium (Li), sodium (Na), and potassium (K). Each element has one valence electron in its outer energy level. These metals are the most reactive metals in the periodic table because of the single electron in the outer energy level.

In their natural state, alkali metals are always found combined with other substances because of their reactivity. The most common element in the family is sodium, which is found all over Earth in compounds like salt (sodium chloride).

Alkaline-Earth Metals

The alkaline-earth metals family is located in the second column of the periodic table. Alkaline-earth metals are less reactive than the alkali metal family because they have two valence electrons in their outer energy level. These atoms need to lose two electrons to become stable. Beryllium is the first member of the family, followed by magnesium, calcium, strontium, barium, and radium.

Halogens

The halogen (fluorine) family is the seventeenth family in the periodic table and includes fluorine (F), chlorine (Cl), bromine (Br), iodine (I), and astatine (At). The halogens have seven valence electrons, making them one electron short of filling their outermost energy level. The halogens are the most reactive non-metals in the periodic table. In their natural state they are found combined with another element.

Chalcogens

The chalcogen (oxygen) family is located in the sixteenth column of the periodic table. The chalcogen family is slightly less reactive than the halogen family. They have six valence electrons in their outer energy level (that is, they are two electrons short of having a completely filled outer energy level). The first member of the chalcogen family is oxygen, followed by sulphur, selenium, tellurium, and polonium.

Noble Gases

The noble gases (helium) family is the eighteenth family in the periodic table. It includes helium (He), neon (Ne), argon (Ar), krypton (Kr), xenon (Xe), and radon (Rn). They are called noble gases because they do not generally form compounds with other elements. They are unreactive because their outer energy levels are completely filled with electrons. No natural compounds formed from these gases exist.

Rows in the Periodic Table

Rows in the periodic table are called periods. Elements found in the same period do not demonstrate similar properties as they do in families. Periods, however, show trends. As you look from the left side to the right side of the table, the elements change from metals (Li) to non-metals (C) and to gases (Ne). There are other trends across periods that chemists use, like atomic radius and electronegativity.

Lesson 2
Compounds

Atoms have a tendency to combine and form new materials. Only about 100 different atoms combine to form the millions of materials that are found on Earth. In some ways, atoms behave like letters of the alphabet in that 26 letters combine to form the huge number of words in the English language.

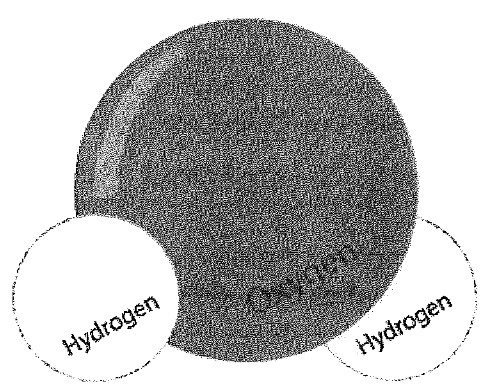
A bond is a kind of glue that holds atoms together to form compounds. There are two types of bonds: ionic bonds and covalent bonds.

Naming and Writing Compounds

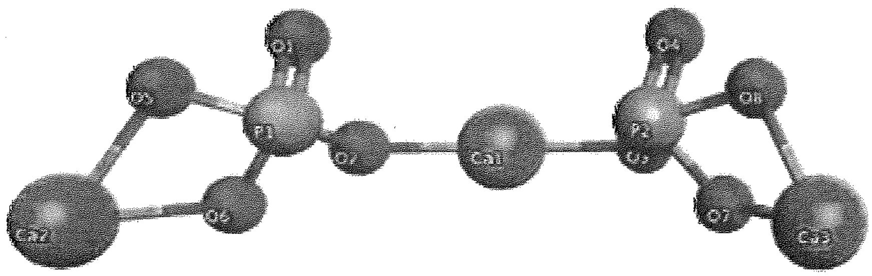
Chemistry has its own language. Chemists all over the world communicate in this language to describe the millions of known compounds. This communication depends on a standard system of naming and writing of formulas for compounds, called the IUPAC system (the acronym stands the International Union of Pure and Applied Chemistry). The ability to name compounds is an essential skill that will allow you to write chemical formulas and fully understand chemical reactions.

A chemical formula is a shorthand method to represent compounds. It uses the elements' symbols and subscripts, and gives the following information:

- The different elements in the compound represented by their symbols
- The number of atoms of each element in the compound represented by subscripts.



In the figure above, the compound water (H_2O) contains hydrogen (H) and oxygen (0) atoms. The subscripts that follow each element indicate the number of atoms of that element in the compound. The subscript following the symbol for hydrogen is 2, indicating that there are two hydrogen atoms in each water molecule. Notice that there is no subscript following the symbol for oxygen; this shows that there is only one atom of oxygen in a water molecule. Chemists do not write the number one as a subscript when only one atom of that kind of element exists in the compound.



The above figure shows a more complex chemical formula. The compound $Ca_3(PO_4)_2$ contains calcium (Ca), phosphorous (P), and oxygen (O). The subscript after the symbol for calcium indicates that there are three atoms of calcium in one formula unit. The subscript "2" outside of the bracket indicates that every subscript inside the bracket is to be multiplied by two. This means that there are 2 x 1 = 2 atoms of phosphorous and 2 x 4 = 8 atoms of oxygen in each formula unit of $Ca_3(PO_4)_2$

Ionic Compounds

An ion is a charged particle that forms when a neutral atom gains or loses electrons. Positively charged ions, called **cations,** are formed when an atom loses one or more electrons. A negative ion, called an **anion**, is formed when an atom gains one or more electrons.

Ionic compounds are formed when two or more oppositely charged ions (usually a metal and a non-metal) are attracted to each other. This chemical attraction is called a chemical bond. An ionic bond is formed when a negatively charged ion is attracted to a positively charged ion. Ions combine together so that their charges add to zero.

Examples of ionic compounds:

$NaCl$ - sodium chloride

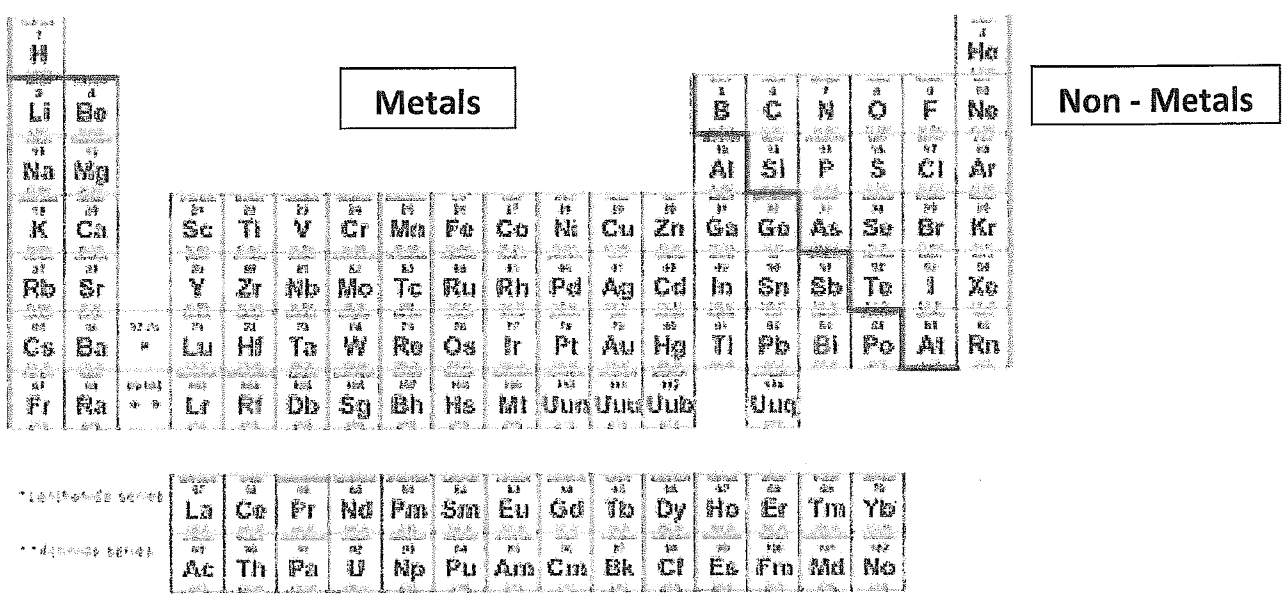
Fe_2O_3 - iron (III) oxide

$CuSO_4$ -copper(II)sulfate

$Ca_3(PO_4)_2$ - calcium phosphate

Predicting the Charge of an Ion

The periodic table groups atoms according to their properties. The periodic table below shows the names of several groups, or families. Remember that metals are found to the left of the dividing line (the "staircase"), while non-metals are found to the right of the line.



If you do not know the charge of an ion, you can use the periodic table to predict the charge. For example, the **alkali metals** are found in the first group or family. This means they will each tend to lose 1 electron to produce ions with a 1+ charge. This loss leaves the outermost electron shell empty, but the shell "underneath" is full. The alkali metal now resembles the closest noble gas and is considered to be stable. Remember that an outermost shell of electrons is stable when it contains eight electrons (or two electrons, in the case of the first shell only). The **alkaline earth metals**, group 2, tend to form ions with a 2+ charge. Metals on the left-hand side of the periodic table tend to lose their valence electrons, leaving a complete octet of electrons in their next-lowest energy level.

Interestingly, many of the **transition metals**, because of their electron arrangement, tend to form more than one ion charge. The reason for this is beyond the scope of this course.

On the non-metal side of the table, the **chalcogens** occupy group 6. This indicates that they have six electrons in their outermost shell and will need to gain two to form a stable group of eight electrons. Therefore, the chalcogens tend to form 2- ions, while the **halogens**, group 7, tend to form 1 - ions. The noble gases (group 8) are already stable. They do not need to gain or lose electrons, and therefore do not carry a charge.

Naming Binary Ionic Compounds

A binary compound contains two different kinds of elements, although there can be more than one atom of each of those elements. Binary ionic compounds usually contain one positively charged metal ion combined with one negatively charged non-metal ion.

When naming an ionic compound from its formula, follow the rules below.

Example 1

Write the name for NaCl.

Step 1: *Name the first element.*

Na = sodium

Step 2: *Name the second element and change the ending to " - ide."*

Cl = chlorine → chloride

The name of the compound is sodium chloride.

Example 2

Write the name for Mg_3P_2

Step 1: *Name the first element.*

Mg = magnesium

Step 2: *Name the second element and add "-ide."*

P = phosphorous → phosphide

The name of the compound is magnesium phosphide.

Writing Ionic Formulas

Remember that formulas contain the symbols of the elements involved as well as subscripts indicating the number of atoms of each element. Here are the key points to remember when writing the formulas for binary ionic compounds:

- The formula must have the cation listed first, followed by the anion.
- The sum of the charges of the ions must be zero. That is, the number of positive charges must equal the number of negative charges.
- You cannot change the charge of the ions to make the ion charges equal zero. You can only modify the subscripts.

Steps for Writing Formulas

Step 1: Write the symbols and charges for the ions involved.

Step 2: Determine the lowest common multiple between the charges, and add whatever subscripts are needed to balance the charges. That is, the number of positive charges must equal the number of negative charges.

Example 3

Write the formula for aluminum oxide.

Step 1: *Write the symbols and charges for the ions involved.*

Aluminum is Al^{3+} and oxide (oxygen) is 0^{2-}

Step 2: *Add whatever subscripts are needed to balance the charges.*

How can you balance a 3+ charge and a 2- charge? The lowest common multiplier of 3 and 2 is 6. Therefore, you must have 2 Al (3+ x 2 = 6+) and 3 O (2- x 3 = 6-)

The formula for aluminum oxide is Al_2O_3

Metal Ions with Multiple Charges

Most of the transition metals have more than one possible ion charge. Here are some examples:

Ion	Possible Ion Charges
Copper	Cu^{1+}, Cu^{2+}
Iron	Fe^{2+}, Fe^{3+}
Cobalt	Co^{2+}, Co^{3+}
Chromium	Cr^{2+}, Cr^{3+}
Lead	Pb^{2+}, Pb^{4+}
Tin	Sn^{2+}, Sn^{4+}

The Stock naming system uses Roman numerais following the metal ion's name to indicate an ion's charge. For example,

Cu^{+} = Copper (I)

Cu^{2+} = Copper (II)

Fe^{2+} = Iron (II)

Fe^{3+} = Iron (III)

<u>Example 4</u>

Write the formula for iron (III) chloride.

Step 1: *Write the symbols and charges for the ions involved in the usual order, the cation first followed by the anion.*

iron (III) is Fe^{3+} and chloride is Cl^{-1}

Step 2: *Add whatever subscripts are needed to balance the charges.*

How can you balance a 3+ charge and a 1- charge?
The lowest common multiplier of 3 and 1 is 3.
Therefore, you must have 1 Fe (3+ x 1 = 3+) and 3 Cl (1- x 3 = 3-)

The formula for iron (III) chloride is $FeCl_3$.

Example5

Write the formula for lead (IV) sulfide.

Step 1: *Write the symbols and charges for the ions involved in the usual order, the cation first followed by the anion.*

lead (IV) is Pb^{4+} and sulfide is S^{2-}

Step 2: *Add whatever subscripts are needed to balance the charges.*

How can you balance a 4+ charge and a 2- charge?
The lowest common multiplier of 4 and 2 is 4.
Therefore, you must have 1 Pb (4+ x 1 = 4+) and 2 S (2- x 2 = 4-)

Therefore, the formula for lead (IV) sulfide is PbS_2.

Naming Compounds Having Metal Ions with Multiple Charges

You will be naming compounds with multiple charges using the Stock system. The same rules (which you already know) will apply, but you must also determine the charge on the metal ion to correctly name the compound.

Example 6

Write the name for Fe_2O_3

Step 1: *Determine the original charges of the individual ions.*

You already know the charge for O is (2-). If there are three oxygen atoms, then the total charge would be 6- (3 x 2- = 6-). This negative charge must be cancelled by iron. Since there are two atoms of iron and they must have a total charge of 6+, the charge of each individual ion must be 3+ (3+ x 2 = 6+).

Therefore, we have charges of, Fe^{3+} and O^{2-}

Step 2: *Write the name of the compound using Roman numerals following the metal ion's name to indicate its charge.*

Therefore, the name of Fe_2O_3 is iron (III) oxide.

Writing Names for Compounds Containing Polyatomic Ions

Some ions are composed of several atoms joined covalently. These are called polyatomic ions (poly meaning many). Examples of common polyatomic ions are in the table below:

Some Common Polyatomic Ions and Their Names		
Ion	**Name**	**Charge**
NH^{4+}	ammonium	1+
SO_4^{2-}	sulfate	2-
PO_4^{3-}	phosphate	3-
$N0_3^-$	nitrate	1-
OH^-	hydroxide	1-
CO_3^{2-}	carbonate	2-
$C_2H_30_2^-$	acetate	1-

Although polyatomic ions have more than one atom, we will name polyatomic ionic compounds like binary compounds. In other words, you will treat polyatomic ions as though they were one single ion. This means that the charge for polyatomic ions is for the whole group of atoms, not just for the atom that is written last. Do not change the subscripts of polyatomic ions; if you do, you change the identity of these ions.

When indicating the presence of more than one polyatomic ion in a compound, we use parentheses around the polyatomic ion, followed by its subscript. For example, the compound $Al(C_2H_3O_2)_3$ has one aluminum ion and three acetate ions. Placing the acetate ion in parentheses and following it with the subscript 3 identifies that 3 acetate ions are required to balance the charge of one aluminum ion. This also helps you avoid confusion as to which atoms belong to which polyatomic ion.

Example 7

Write the name for KNO_3.

Step 1: *Identify the cation and the anion.*

K^+ is the cation. It only has one possible charge, so we don't need to use a Roman numeral in the formula. The name of the K^+ ion is potassium. NO^{-3} will be the anion. It is called the nitrate ion.

Step 2: *Write the name of the cation first, followed by the anion.*

The name of the compound is potassium nitrate.

Example 8

Write the name of $Cu_3(PO_4)_2$

Step 1: *Identify the anion and the cation.*

Cu will be the cation, and PO_4 will be the anion. Because copper is an ion with more than one possible charge, we must look to the anion to determine its value.

The previous chart tells us that the anion is the phosphate ion (PO_4). The parentheses in the formula is followed by the number 2 indicating there are 2 phosphate ions in this compound. If phosphate has a charge of (3^-), the total charge will be 2 x 3- = 6 - The total charge of the anions is 6-, so the charges of all the cations must add up to 6+ for there to be a net charge of zero. There are three copper ions in the formula, and $6^+ \div 3 = 2^+$, so the charge on the copper ion is 2^+.

Step 2: *Write the name of the cation first, followed by the name of the anion.*

Copper is one of the ions with more than one possible charge, so you must use a Roman numeral to indicate the 2^+ charge.

The name of the compound is then copper (II) phosphate.

Example 9

Write the name of NH_4SCN.

Step 1: *Identify the cation and the anion.*

The cation is NH^{4+}, or ammonium, while the anion is SCN^-, thiocyanate.

Step 2: *Write the name of the cation first, followed by the name of the anion.*

The name of the compound is ammonium thiocyanate.

Writing Formulas for Compounds Containing Polyatomic Ions

Writing formulas for compounds containing polyatomic ions is similar to writing formulas for binary compounds.

Step 1: *Write out the symbols and charges for the ions involved.*

Step 2: *Determine the lowest common multiple between the charges and add whatever subscripts are needed to balance the charges. Be sure to include brackets around polyatomic ions when they have a subscript.*

Example 10

Write the formula for sodium perchlorate.

Step 1: *Write out the symbols and charges for the ions involved.*

Sodium is Na^+ and perchlorate is ClO^{-4}

Step 2: *Determine the lowest common multiple between the charges and add whatever subscripts are needed to balance the charges.*

The sodium is a 1^+ and the perchlorate is 1^-. The lowest common multiple here is 1, so the charges balance with 1 ion each.

The formula for sodium perchlorate is then $Na(ClO_4)$ or $NaClO_4$.

Example 11

Write the formula for iron (III) cyanide.

Step 1: *Write out the symbols and charges for the ions involved.*

Iron (III) is Fe^{3+} and cyanide is CN-.

Step 2: *Determine the lowest common multiple between the charges and add whatever subscripts are needed to balance the charges. You may find the following chart method to be helpful:*

Fe^{3+}	CN- CN- CN-
3+	3-

Three CN- atoms were needed to balance the charge of the iron atom. The formula for iron (III) cyanide is then $Fe(CN)_3$.

Example 12

Write the formula for potassium hydrogen phosphate.

Step 1: *Write out the symbols and charges for the ions involved.*

Potassium is K^+ and hydrogen phosphate is $HPO_4{}^{2-}$

Step 2: *Determine the lowest common multiple between the charges and add whatever subscripts are needed to balance the charges.*

K^+ K^+	$HPO_4{}^{2-}$
2+	2-

The formula for potassium hydrogen phosphate is $K_2(HPO_4)$ or K_2HPO_4

Polyatomic Ions of Different Form

For those polyatomic ions that are made of different amounts of the same elements, we need a system of naming to help us differentiate their compounds. For example, PO_4^{3-} and PO_3^{3-}. These ions contain the same elements (phosphate and oxygen) but differ in the number of oxygen atoms they contain. In order to name these ions and the compounds they make, we need to know the common form of the ion. This common form of the ion is designated with an "*ate*" ending. For example, PO_4^{3-} is the common form and is named phosphate.

Once we know the common, or "*ate*" form of the ion, we continue to name similar ions based on the number of oxygen atoms they contain in comparison to the common form. This is illustrated in the table below.

	NO_3^{1-} nitrate	ClO_3^{1-} chlorate	SO_4^{2-} sulfate
1 more oxygen		ClO_4^{1-} perchlorate	
1 less oxygen	NO_2^{1-} nitrite	ClO_2^{1-} chlorite	SO_3^{2-} sulfite
2 less oxygen		ClO^{1-} hypochlorite	

We can also use this system of naming to help us identify compounds that contain these polyatomic ions and hydrogen. We start by indicating the number of hydrogen followed by the name of the correct polyatomic ion.

PO_4^{3-} phosphate	SO_4^{2-} sulfate
HPO_4^{2-} Hydrogen phosphate	HSO_4^{1-} Hydrogen sulfate
$H_2PO_4^{1-}$ Dihydrogen phosphate	HSO_3^{1-} Hydrogen sulfite

Covalent Compounds

Non-metals tend to combine chemically by sharing electron pairs. These bonds are known as covalent bonds. Neutral compounds made of atoms joined covalently are called molecular or covalent compounds.

We name covalent compounds differently than ionic compounds. We must indicate the number of each element by adding a Greek prefix in front of the element's name.

The prefixes are:

One = mono	Two = di
Three = tri	Four = tetra
Five = penta	Six = hexa
Seven = hepta	Eight = octa
Nine = nona	Ten = deca

Naming Covalent Compounds

Step 1: *Name the first element in full, using a prefix only when there are two or more of that element. You can omit "mono-" if there is only one of that element in the compound. For example, NO is nitrogen monoxide, but N_2O is dinitrogen monoxide.*

Step 2: *Name the second element, but use an "-ide" ending. Use prefixes to indicate the number of that element (including mono).*

Step 3: *Write the name of the compound.*

There are two exceptions to the naming rules. The common names for the following compounds are used instead of their IUPAC names:

H_2O =water

NH_3 = ammonia

Example 13

Write the name for CO_2

Note: This is a covalent compound since it is made of two non-metal atoms.

Step 1: *Name the first element in full, using a prefix only when there are two or more of that element.*

There is only one carbon atom. You can omit the "mono-" for the first element, so the first part of the name is carbon.

Step 2: *Name the second element. Use prefixes and end the name with " -ide".*

The second element is oxygen. There are two oxygen atoms, so the second part of the name is dioxide.

Step 3: *Write the name of the compound.*

The name of CO_2 is carbon dioxide.

Example 14

Write the name for N_2O_4.

This is another covalent compound, as both nitrogen and oxygen are non-metals.

Step 1: *Name the first element in full, using a prefix only when there are two or more of that element.*

There are two nitrogen atoms, so the first part of the name is dinitrogen.

Step 2: *Name the second element. Use prefixes and end the name with "-ide,"*

The second element is oxygen. There are four oxygen atoms, so the second part of the name is tetraoxide.

Step 3: *Write the name of the compound.*

The name of N_2O_4 is then dinitrogen tetraoxide.

Writing Formulas for Covalent Compounds

Step 1: *Write the symbol for the first element, followed by the subscript indicated by the prefix.*

Step 2: *Write the symbol for the second element, followed by the subscript indicated by its prefix.* ***Do not reduce the subscripts for covalent compounds!***

Step 3: *Write the name of the compound.*

Example 15

Write the formula for dinitrogen monoxide.

Step 1: *Write the symbol for the first element, followed by the subscript indicated by the prefix.*

The symbol for nitrogen is N and the prefix "di" is a subscript of 2.

Step 2: *Write the symbol for the second element, followed by the subscript indicated by its prefix.*

The symbol for oxide (oxygen) is O and the prefix "mono" stands for a subscript of 1.

Step 3: *Write the name of the compound.*

The formula for dinitrogen monoxide is then N_2O

Example 16

Write the formula for sulfur hexafluoride.

Step 1: *Write the symbol for the first element, followed by the subscript indicated by the prefix.*

The symbol for sulfur is S and the lack of a prefix identifies that its subscript is 1.

Step 2: *Write the symbol for the second element, followed by the subscript indicated by its prefix.*

The symbol for fluoride is F and the prefix "hexa" stands for a subscript of 6.

Step 3: *Write the name of the compound.*

The formula for sulfur hexafluoride is then SF_6.

Diatomic Molecules

Some elements do not exist as single atoms. These elements exist as pairs of atoms joined covalently, and are called **diatomic molecules**. The elements that exist as diatomic molecules are:

hydrogen gas (H_2)

oxygen gas (0_2)

nitrogen gas (N_2)

fluorine gas (F_2)

chlorine gas (Cl_2)

bromine gas or liquid bromine (Br_2)

solid iodine (I_2)

solid astatine (At_2)

You can also use the "7+1" method to help you remember the diatomic elements. On a periodic table, form the number 7, starting with nitrogen (N) and working towards fluorine (F). Now trace the backbone of the number seven going all the way down to iodine (I). All the elements that make up the number seven are diatomic elements. Since hydrogen is also a diatomic element, we say "+1" to help you remember to consider it as well.

Lesson 3
Atomic Mass Units

Formula Mass

The formula mass of a substance is the sum of the atomic masses of all the atoms in one molecule or particle of that substance written in atomic mass units (amu). This terminology applies only to ionic substances, such as NaCl. For covalent molecules, such as water (H_2O), the sum of the atomic masses of all of the atoms in a molecule is called the molecular mass.

Example 1:

Find the molecular mass of water (H_2O)

One molecule of water contains 2 hydrogen atoms and one oxygen atom.

The atomic mass of hydrogen is 1.0 amu

The atomic mass of oxygen is 16.0amu

H_2O = (2 X H) + (1 X O) = (2 X 1.0 amu) + (1 X 16.0 amu) = 18.0amu

The molecular mass for water is 18.0 amu

Example 2:

What is the formula mass of $Ca_3(PO_4)_2$

Remember that a subscript found outside of the parentheses means that you will have multiple ions present. A formula of $Ca_3(PO_4)_2$ means that two phosphate (PO_4) ions are present. This means that there are two Ps and eight Os!

$Ca_3(PO_4)_2$ = (3 x Ca) + (2 x P) + (8 X O)

(3 X 40.1 amu) + (2 X 31.0 amu) + (8 X 16.0 amu) = 310.3 amu

The formula mass for calcium phosphate is 310.3 amu

Lesson 4
Balancing Chemical Equations

A balanced chemical equation tells us about the quantities of reactants and products in a reaction. We balance chemical equations in order to obey the Law of Conservation of Mass.

Chemical Equations

This section provides a brief overview of what you need to know before you attempt to balance more complex chemical equations. For some, this may be a review, while for others this may be new material.

A chemical equation indicates the substances reacting and the substances produced in a chemical reaction. Generally, we call the substance or substances that react together the reactants, and we call the resulting substance or substances the products of a reaction. A chemical equation will usually be written with the reactants on the left side of an arrow and the products on the right side of the arrow.

$$\text{REACTANTS} \rightarrow \text{PRODUCTS}$$

A chemical equation also shows the ratio in which these substances react or are produced. For example:

$$2H_2\,(g) + O_2\,(g) \rightarrow 2H_2O\,(g)$$

You can also use words instead of symbols to describe a chemical reaction. Here is the word equation that describes the same chemical reaction you see above:

"Hydrogen gas and oxygen gas react to form (or yield) water vapour."

A balanced chemical reaction provides the same information that a recipe does. In addition to symbols, a chemical reaction uses numbers to indicate the quantity of reactants used and products created. For example, the "2"s in front of both the H_2 and the H_2O are called coefficients. Coefficients indicate the ratio in which the substances combine or are produced in a chemical reaction. The number "1" is not written as a coefficient, so the coefficient for O_2 is 1. The balanced word equation for the above example is:

"Two molecules of hydrogen gas react with one molecule of oxygen gas to yield two molecules of water."

Chemists show the states or phases of the reactants and products by using abbreviations in parentheses following each reactant and product. The abbreviations are as follows:

(s) or (c) means solid or crystalline

(l) means liquid

(g) means gas or vapour

(aq) means the substance is aqueous or dissolved in water

Balanced Chemical Equations

According to the Law of Conservation of Mass (or matter), mass cannot be created or destroyed. During a chemical reaction then, since atoms are a form of matter, atoms cannot be created or destroyed. This means the number of each kind of atom on the reactant side of the equation must equal the number of that same kind of atom on the product side of the equation.

When you looked at the example in the previous section, you might have wondered why it was not written as

$$H_2\,(g) + O_2\,(g) \rightarrow H_2O\,(g)$$

We call this equation *unbalanced* because the atoms are not conserved. In other words, there are not an equal number of atoms on each side of the chemical equation. While there may be two H atoms on the reactant side and two H-atoms on the product side, there are two O atoms on the reactant side compared to only one O atom on the product side. The coefficients in the reaction from the previous section ensure the equation is balanced. A **balanced chemical equation** has an equal number of atoms of each element on both the reactant and product sides of the equation. The balanced chemical equation for the burning of hydrogen in oxygen is the one we saw previously:

$$2H_2\,(g) + O_2\,(g) \rightarrow 2H_2O\,(g)$$

In this reaction, there are four H atoms on the reactant side and four H atoms on the product side. Furthermore, there are two O atoms on the reactant side and two O atoms on the product side. The Law of Conservation of Mass has been respected, and this chemical equation is balanced.

Balancing Chemical Equations

When balancing a chemical equation, you *cannot change the subscripts* of the substances in the reaction. Changing subscripts changes the identity of the compounds. To balance atoms, you can *insert coefficients* rather than change subscripts. Keep the following in mind:

- Sometimes an equation is already balanced, and you won't need to change anything.
- In a balanced equation, each side of the equation will have the same number of atoms of each element.
- It is not necessary for coefficients to be the same on both sides of the equation in order for the number of atoms of each type to balance.

You may already have a system of balancing equations that works best for you. If not, or if you need a review, then you can use the following steps:

1. If it has not already been done for you, *write out the formulas of each reactant and product.*
2. *Determine the number of atoms of each element.* Write each element in a table and keep a tally of the number of atoms of each element.
3. *Use coefficients to balance the equation* so that it obeys the Law of Conservation of Mass.
4. If necessary, *reduce the coefficients to the lowest whole number ratio.* Multiply fractions, if present, by the denominator to make all coefficients whole numbers.

Example 1:

Balance the equation: $C_3H_3 + O_2 \rightarrow CO_2 + H_2O$

Step 1: *Write out the formulas of each reactant and product.*

This has already been done, so we can move directly to Step 2.

Step 2: *Determine the number of atoms of each element.*

C	H	O	C	H	O
3	3	2	1	2	3

Step 3: *Use coefficients to balance the equation.*

Balance the carbon's and hydrogen's first, and leave the oxygen molecule until last because it will be the easiest to balance.

$$\underline{2}\ C_3H_3 + \underline{7.5}\ O_2 \rightarrow \underline{6}\ CO_2 + \underline{3}\ H_2O$$

C	H	O	C	H	O
6	6	15	6	6	15

Step 4: *Reduce the coefficients to the lowest whole number ratio.*

Multiply the entire equation by two in order to make 7.5 a whole number.

The balanced equation is: $\underline{4}\ C_3H_3 + \underline{15}\ O_2 \rightarrow \underline{12}\ CO_2 + \underline{6}\ H_2O$

Example 2:

Balance the equation: $Al_2(SO_4)_3 + CaCl_2 \rightarrow AlCl_3 + CaSO_4$

Step 1: *Write out the formulas of each reactant and product.*

Step 2: *Determine the number of atoms of each element.*

Note: Since the sulfate ion remains unchanged from reactant to product, we can balance the polyatomic ion as if it were a single element, or we could balance S and O atoms separately.

Al	SO_4	Ca	Cl	Al	SO_4	Ca	Cl
2	3	1	2	1	1	1	3

Step 3: *Use coefficients to balance the equation.*

Balance the metals first. Be sure to recount other elements that are affected by the change.

$$Al_2(SO_4)_3 + \underline{3}\ CaCl_2 \rightarrow \underline{2}\ AlCl_3 + \underline{3}\ CaSO_4$$

Step 4: *Reduce the coefficients to the lowest whole number ratio.*

The coefficients are in lowest terms.

The balanced reaction is $Al_2(SO_4)_3 + \underline{3}\ CaCl_2 \rightarrow \underline{2}\ AlCl_3 + \underline{3}\ CaSO_4$

Classifying Reactions

By knowing some simple rules of chemistry, we can actually predict how different compounds will interact with each other. Classifying reactions can allow you to predict the products of a reaction There are five general types of chemical reactions.

1. synthesis (also called combination)
2. decomposition
3. single-replacement
4. double-replacement
5. combustion.

Synthesis Reactions

As the name suggests, synthesis reactions synthesize a product. Generally, synthesis reactions involve the reaction of two simple substances to produce a single, more complex substance.

The following are examples of synthesis reactions:

$8\ Fe(s) + 8\ S(s) \rightarrow 8FeS(s)$

$2Na(s) + Cl_2(g) \rightarrow 2NaCl(s)$

$4Na(s) + O_2(g) \rightarrow 2Na_2O(s)$

$CO_2(g) + H_2O(1) \rightarrow H_2CO_3(aq)$ (carbonic acid)

Remember that when you predict the products of a reaction, their charges must balance. You must always remember the ion charges when writing formulas for ionic compounds. In addition, you must finish your work by balancing the chemical equation.

Decomposition Reactions

Decomposition reactions are the opposite of synthesis reactions. These reactions have only one reactant producing two or more simpler substances. The decomposition of a compound usually requires the input of energy in the form of heat, electricity, or light.

Here are some examples of decomposition reactions:

$2Na_20\ (s) \rightarrow 4Na\ (s) + 0_2(g)$

$CaC0_3(s) \rightarrow CaO\ (s) + C0_2(g)$

Single Replacement Reactions

Single replacement (also called single displacement) reactions are reactions between a compound and an element. The element replaces (or displaces) one element in the compound to produce a new element and a new compound. In these reactions, metals replace metals (cations replace cations) and non-metals replace non-metals (anions replace anions).

Here are more examples of single replacement reactions:

$$Cu(s) + 2AgNO_3(aq) \rightarrow Cu(NO_3)_2(aq) + 2Ag(s)$$

$$2Na(s) + 2H_2O(l) \rightarrow 2NaOH\ (aq) + H_2(g)$$

$$Ca(s) + 2H_2O(l) \rightarrow Ca(OH)_2\ (aq) + H_2(g)$$

Remember that when ions combine, the charges must balance. Don't forget that the equation then needs to be balanced!

Double Replacement Reactions

Double replacement (or double displacement) reactions occur when the positive ions of two compounds change positions. Basically, the reactants switch partners. These types of reactions generally produce a precipitate, a gas, or a molecular compound such as water.

Here are examples of double replacement reactions:

$$Ca(NO_3)_2(aq) + Na_2CO_3(aq) \rightarrow 2NaNO_3(aq) + CaCO_3(s)$$

$$HCl(aq) + NaOH(aq) \rightarrow NaCl(aq) + H_2O\ (l)$$

Combustion Reactions

A combustion reaction is a chemical change where oxygen reacts with another element or compound, often resulting in the production of heat or light energy. You can usually identify this type of reaction by looking for oxygen on the reactant side of the chemical equation. Often, the other reactant is a hydrocarbon (a compound composed of hydrogen and carbon) such as methane (CH_4). However, the other reactant can sometimes be another element like sulfur or magnesium, which burn in the presence of oxygen.

Here are some examples of combustion reactions:

$$2Mg(s) + O_2(g) \rightarrow 2MgO\ (s)$$

$$2C_8H_{18}\ (l) + 25O_2(g) \rightarrow 16CO_2(g) + 18H_2O(l)$$

$$C_7H_{16}(l) + 11O_2(g) \rightarrow 7CO_2(g) + 8H_2O(l)$$

LESSON 5
Working with Units

You will be using number values to solve problems, not only in this course, but also possibly in future courses. It is important to pay attention to units. A number means nothing without its unit.

Knowing the units of measurement that correspond with a number can give you so much more information than a digit sitting there by itself. Units can:

- Help to show another person the exact amount you have
- Assist in solving a mathematical problem, especially in chemistry, where you can follow the units to get to the answer
- Show which measurement system the person is using (i.e. metric or standard)

Making Unit Conversions

It is often necessary, especially in chemistry or physics courses, to make conversions between units. You may need to convert units because a value provided to you in a question is not given in a useful unit. Or, you are being asked to express your answer in a unit different from what you needed to use for a calculation.

Converting units is a very useful skill that you should spend some time developing. Although there is more than one way to complete a unit conversion, you will be introduced to a ratio method that allows you to visualize both the units you wish to cancel or eliminate, and those units you wish to obtain. It is a method that you will be expected to use throughout your chemistry course. Practice converting these values of measurement.

Example 1:

325m = _______cm

There are 100 cm in one metre. We can show this relationship with the ratios,

$$\frac{100\text{ cm}}{1\text{ m}} \quad \textbf{OR} \quad \frac{1\text{ m}}{100\text{ cm}}$$

Now, we can set up a system of conversion to solve the question. We start with the value given, or known. We use the ratio that allows us to cancel the unit we don't want, and be left with the unit we do want. (remember algebra!)

$$325\text{ m} \quad \text{x} \quad \frac{100\text{ cm}}{1\text{ m}}$$

$$325\ \cancel{\text{m}} \quad \text{x} \quad \frac{100\text{ cm}}{1\ \cancel{\text{m}}}$$

The math tells us: $\frac{325 \text{ x } 100}{1} = 32{,}500$

So the answer is: 325 m is equal to 32,500 cm

Example 2:

100 km/hr = _______m/s

We can use the same approach for this question, there are simply two units to convert, rather than one!

There are 1000 m in one kilometer and there are 3600 seconds in one hour.

We can show this relationship with the ratios,

$$\frac{1000\text{ m}}{1\text{ km}} \quad \text{AND} \quad \frac{3600\text{ s}}{1\text{ hr}}$$

We set up a system of conversion similar to the previous example. We can even convert both units in one math calculation, as long as we set up each ratio first, one ratio to convert km to m and another ratio to convert hours to seconds.

$$100\ \frac{km}{hr} \times \frac{1000\ m}{1\ km} \times \frac{1\ hr}{3600\ s}$$

$$100\ \frac{\cancel{km}}{\cancel{hr}} \times \frac{1000\ m}{1\ \cancel{km}} \times \frac{1\ \cancel{hr}}{3600\ s}$$

The math tells us: $\frac{100 \times 1000 \times 1}{1 \times 3600} = 27.777....$

So the answer is: 100 km/hr is equal to 27.7 m/s

Try these other conversions and check your answers below.

a) 7,527 cm = ______km

b) 32 °C = ________°F

c) 12 ft = ________m

d) 150 pounds = ____kg

e) 15 days = _________s

f) 348 K = _______°C

Answers: a) 0.07527 km b) 89.6 ° F c) 3.7 m d) 68 kg e) 1296000 s f) 75 ° C

LESSON 6 Calculations

Scientific Notation

Scientific notation is a way of expressing numbers that are too big or too small to be conveniently written in decimal form. It is a form commonly used by scientists, and you will use it throughout any chemistry or physics course. It will be important for you to practice how to enter values of scientific notation into your calculator.

In scientific notation, all numbers are written in the form, **m × 10n**

- "m" is a real number, for example, 4 or 3.5 or 5.67
- "n" is an exponent of 10
- If n is positive then the overall value is a larger number
- If n is negative then the overall value is a smaller number

Number	Scientific Notation
60000	6×10^{4}
12000000	1.2×10^{7}
0.000436	4.36×10^{-4}
6030000000000	6.03×10^{12}
0.000000000000000000045	4.5×10^{-20}

You can see how the use of scientific notation eliminates the need to record or enter excessive zeros.

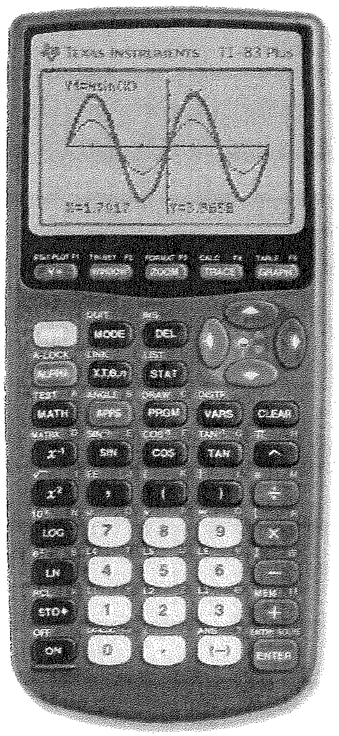
Scientific Notation and Your Calculator

It is important to learn how your calculator accepts and how it displays scientific notation entries.

You will want to know how to carry out a calculation using the value, 3 x 10^{22} without having to input 22 zeros!

You may run into trouble if you simply try to enter scientific notation as "3*10^22"

Instead, look for the buttons, "EE" or "EXP". This will allow you to enter the value as "3$_{E}$22", which lets the calculator know to use scientific notation. Below are two examples of scientific calculators. The one on the left uses an "EE" button that is activated with the "2nd" or second function button. The calculator on the right uses a "EXP" button.



Be sure you know how to enter scientific notation in YOUR calculator!

Example 1:

Practice entering scientific notation into your calculator. Make sure you are getting the correct answers!

$$(4.6 \times 10^{12}) \times (7.99 \times 10^{3}) = 3.6754 \times 10^{16}$$

$$\frac{6.0 \times 10^{23}}{45.6} = 1.315789474 \times 10^{22}$$

Significant Figures

There is variation in all measurements. Depending on the measuring device used, we can minimize uncertainty, but we can never eliminate it. For example, in the kitchen, you might use a measuring cup when you need 1 cup of milk. But if you are in the chemistry lab and need 236 mL of distilled water, you would not use a 1-cup measuring cup but rather a device with more accuracy, like a graduated cylinder for example.

In order to accurately and consistently report the uncertainty in measurements, scientist use an agreed upon system of rounding. Significant figures are those digits in a measurement (or a calculation using measurements) that includes all certain digits plus a final one that is uncertain, or the least significant digit.

Counting Significant Figures

In order to record your answers with the correct significant figures, you first need to know how to determine, or count these digits. These are the rules for counting significant figures:

1. All digits are significant except zeros at the beginning of a number and possibly zeros at the end of a number.

Value	# of significant figures
3.56	3
3.06	3
0.00036	2

2. End zeros are significant if the number contains a decimal point.

Value	# of significant figures
3600	**2**
3600.00	**6**

3. If unclear, either rewrite the number in scientific notation to clarify or if possible, add a decimal point.

   As written, 360, 000 has only two sig figs. If it should be written with 3 significant figures, then rewrite it as, 3.60×10^5

Recording Your Answer with Proper Significant Figures

These are the rules for using significant figures in calculations:

1. When multiplying or dividing, report answers with as many significant figures as there are in the quantity with the fewest number of significant figures.
2. When adding or subtracting, report answers with the same number of decimal places as there are in the quantity with the fewest number of decimal places.

Calculation	Complete Answer	Rounded Answer
$(4.6 \times 10^{12}) \times (7.99 \times 10^{3})$	$= 3.6754 \times 10^{16}$	$= 3.8 \times 10^{16}$
456 x 0.0010	= 0.456	= 0.46
63.4 + 783.01	= 846.41	= 846.4
125 - 2.5	= 122.5	=123

Solving Equations

Many calculations in introductory chemistry require some simple algebra. You will also be expected to be able to convert from one unit to another, which we have studied in a previous lesson.

Make sure you are able to perform the following calculations using mathematical formulas. Always remember proper units and significant figures!

Example 1:

If force (kg · m/s) is calculated by multiplying the mass and acceleration of an object, with what force will a 1500 kg car hit a brick wall if it was travelling 32 m/s at the moment of impact?

$$\boxed{\textbf{force} = \textbf{mass} \times \textbf{acceleration}}$$

force = (1500 kg) x (32 m/s)

force = 48,000 kg · m/s

Example 2:

If water has a density of 1.0 g/mL, how many grams of water are in 240 mL?

$$\boxed{\textbf{density} = \frac{\textbf{mass}}{\textbf{volume}}}$$

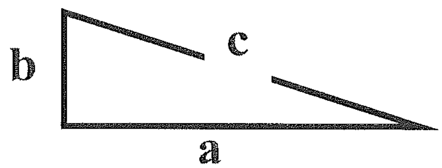
mass = density x volume

mass = 1.0 g/mL x 240 mL = 240 g

Example 3:

The Pythagorean theorem can be used to solve the sides of right angle triangles.

$$\boxed{a^2 + b^2 = c^2}$$



If a = 75.2 cm and c = 87.6 cm what does b equal?

$$\boxed{a^2 + b^2 = c^2} \Rightarrow \boxed{b^2 = c^2 - a^2} \Rightarrow \boxed{b = 44.93016804}$$

$$\boxed{b = 44.9 \text{ cm}}$$

Example 4:

Can you solve this algebraic expression? The answer is 35!

$$68.75 = \frac{X^2 - (20)^2}{(2)(6)}$$

LESSON 7
THE MOLE

Eggs have a grouping unit that you are familiar with - the dozen. An egg farmer, who deals with thousands of eggs, counts the number of eggs in groups of twelve, the same way you would buy them at the store.

In chemistry, there is an important unit, the **mole**, used to represent the number of particles in a substance. Since particles are extremely small, and a substance contains thousands or even millions of particles, it is easier to count particles in terms of moles, just like we use dozens to count eggs.

The Mole

Just like there are 12 eggs in one dozen, there are 6.02×10^{23} particles in one mole of any substance. That is a pretty large number; 602,000,000,000,000,000,000,000!!

For example:

6.02×10^{23} doughnuts is one mole of doughnuts

6.02×10^{23} pennies is one mole of pennies

6.02×10^{23} is a very large number that is used to measure very small things, like formula units, ions, molecules, and atoms.

The number that which represents the mole, is called **Avogadro's Number**.

Avogadro's Number

Formula and molecular mass deal with individual atoms, formula units, and molecules. Chemists do not work with amounts of individual atoms or molecules because these particles are far too small to see or mass.

Furthermore, balances or scales do not measure mass in terms of atomic mass units, so chemists need a practical unit that relates mass (in grams) to the number of particles present in a sample. Note that the word "particle" is a generic term meaning atoms, molecules, ions, or formula units. The mole (mol) is the unit that relates the number of particles in a sample to its mass.

6.02×10^{23} particles is called Avogadro's Number.

Molar Mass

Avogadro's Number relates the number of particles to mass. By definition, one mole of carbon atoms has a mass of 12.0000 g. If the mass of one mole of any atom is its atomic mass in grams, then

- one mole of aluminum atoms has a mass of 27.0 g
- one mole of silver atoms has a mass of 107.9 g
- one mole of sodium atoms has a mass of 23.0 g
- one mole of iron atom has a mass of 55.8 g

Similarly, one mole of any compound has a mass equal to its formula mass (you learned how to calculate formula mass!) in grams. For example, since the formula mass of water is 18.0 amu, the mass of one mole of water molecules will be18.0 g.

The mass of one mole of a substance is called molar mass. It is also referred to as molecular weight in some textbooks. The units for molar mass are grams per mole (g/mol), so the molar mass of water is 18.0 g/ mol.

Mole Calculations

The mole is an intermediate step allowing you to convert from one unit to another. It is possible to use the mole to convert amounts to grams (for weighing samples), or to number of particles in a sample. By first converting values into moles, we can then convert to the unit we are seeking.

Step 1: *Convert the given quantity to moles.*

Step 2: *Convert the number of moles to the desired quantity.*

Calculating Moles from Mass

The units for molar mass are g/mol (grams per mole). This represents the mass, in grams, of one mole of a substance.

To find the number of moles, given a sample's mass,

Step 1: *Find the molar mass.*

Step 2: *Use the molar mass to calculate the number of moles.*

Example 1:

How many moles of carbon atoms are in 24.0 g of carbon?

Step 1: *Find the molar mass.*

Looking at the periodic table, we find that carbon has a molar mass of 12 g/mol.

Step 2: *Use the molar mass to calculate the number of moles.*

Multiplying the mass by the reciprocal of the molar mass will mathematically cancel out the "grams" and leave "moles" as the only unit. In other words, you are dividing the mass of carbon by the molar mass of carbon.

$$24.0 \text{ g carbon} \times \frac{1 \text{ mole}}{12 \text{ g carbon}} \quad \textbf{OR} \quad \frac{24.0 \text{ g carbon} \times 1 \text{ mole}}{12 \text{ g carbon}}$$

$$= \frac{24.0 \text{ g } \cancel{\text{carbon}} \times 1 \text{ mole}}{12 \text{ g } \cancel{\text{carbon}}} = \frac{24.0 \text{ g}}{12 \text{ g}} = 2.0 \text{ mol carbon}$$

Example 2:

How many moles of water molecules are there in 81.0 g of water?

Step 1: *Find the molar mass.*

H_2O = 2 hydrogen + 1 oxygen

= (2 x 1.0 g/mol) + (1 x16.0 g/mol) = 18.0 g/mol

Step 2: *Use the molar mass to calculate the number of moles.*

$$81.0 \text{ g } H_2O \times \frac{1 \text{ mole}}{18 \text{ g } H_2O} \quad \textbf{OR} \quad \frac{81.0 \text{ g } H_2O \times 1 \text{ mole}}{18 \text{ g } H_2O}$$

$$= \frac{81.0 \text{ g } \cancel{H_2O} \times 1 \text{ mole}}{18 \text{ g } \cancel{H_2O}} = \frac{81.0 \text{ g}}{18 \text{ g}} = 4.5 \text{ mol water}$$

Example 3:

How many moles are there in 50.0 g of lead (II) chloride?

Step 1: *Find the molar mass.*

$PbCl_2$ = 1 lead + 2 chlorine

= (1 x 207.2 g/mol) + (2 x 35.5 g/mol) = 278.2 g/mol

Step 2: *Use the molar mass to calculate the number of moles.*

$$50.0 \text{ g } PbCl_2 \times \frac{1 \text{ mole}}{278.2 \text{ g } PbCl_2} \quad \textbf{OR} \quad \frac{50.0 \text{ g } PbCl_2 \times 1 \text{ mole}}{278.2 \text{ g } PbCl_2}$$

$$= \frac{50.0 \cancel{\text{ g } PbCl_2} \times 1 \text{ mole}}{278.2 \cancel{\text{ g } PbCl_2}} = \frac{50.0 \text{ g}}{278.2 \text{ g}} = 0.1797268\ldots \text{mol } PbCl_2 = 0.180 \text{ mol}$$

(3 figures because 50.0g is 3 figures!)

Calculating Mass Given Moles

If one mole of carbon has a mass of 12.0 g, then 2 moles of carbon will have a mass of 2 x 12.0 g = 24.0 g. If you need a certain number of moles of a substance, you can calculate the mass of the substance by multiplying the number of moles of that substance by its molar mass.

Step 1: *Find the molar mass of the substance.*

Step 2: *Convert the number of moles to mass.*

Multiply so that units cancel to yield grams

Example 4:

What is the mass of 2.50 moles of gold?

Step 1: *Find the molar mass of the substance.*

Au = 197.0 g/mol

Step 2: *Convert the number of moles to mass.*

$2.50 \text{ mol Au} \times \frac{197.0 \text{ g}}{1 \text{ mole}}$ **OR** $\frac{2.50 \text{ mol Au} \times 197.0 \text{ g}}{1 \text{ mole}}$

$= \frac{2.50 \text{ mol}\cancel{\text{Au}} \times 197.0 \text{ g}}{1 \cancel{\text{mole}}} = 2.50 \times 197.0 = 492.5 \text{ g gold}$

$= 493 \text{ g}$

Example 5:

What is the mass of 1.20×10^{-5} moles of carbon tetrachloride?

Step 1: *Find the molar mass of the substance.*

CCl_4 = 1 carbon + 4 chlorine

= (1 x 12 g/mol) + (4 x 35.5 g/mol) = 154 g/mol

Step 2: *Convert the number of moles to mass.*

$1.20 \times 10^{-5} \text{ mol } CCl_4 \times \frac{154.0 \text{ g}}{1 \text{ mole}}$ **OR** $\frac{1.20 \times 10^{-5} \text{ mol } CCl_4 \times 154.0 \text{ g}}{1 \text{ mole}}$

$= \frac{1.20 \times 10^{-5} \text{ mol}\cancel{CCl_4} \times 154.0 \text{ g}}{1 \cancel{\text{mole}}} = 1.20 \times 10^{-5} \times 154.0 = 0.001848 \text{ g } CCl_4$

$= 0.00185 \text{ g}$

Converting Mass to Particles

The mole allows for the conversion between the mass and number of particles in a sample. To convert from mass to a number of particles, follow the following steps:

Step 1: *Convert mass to moles.*

Step 2: *Convert the number of moles to number of particles.*

Multiply so that units cancel to yield particles.

<u>Example 6:</u>

How many molecules of water are in a 10.0 g sample of water?

Step 1: *Convert mass to moles.*

To convert from mass to moles, we must know the molar mass of water.

$H_2O = 18.0\ g/mol$

$$10.0\ g\ H_2O \times \frac{1\ mole}{18.0\ g\ H_2O} \quad \textbf{OR} \quad \frac{10.0\ g\ H_2O \times 1\ mol}{18.0\ g\ H_2O}$$

$$= \frac{\cancel{10}.0\ g\ H_2O \times 1\ mol}{18.0\ g\ \cancel{H_2}O} = \frac{10.0}{18.0} = 0.555555....mol\ H_2O$$

***Note: do not round at this step!**

Step 2: *Convert the number of moles to number of particles.*

$$0.55\ mol\ H_2O \times \frac{6.02 \times 10^{23}\ particles}{1\ mole}$$

$$= \frac{0.55\ mol\ \cancel{H_2}O \times 6.02 \times 10^{23}\ particles}{1\ \cancel{mo}le} = 3.344 \times 10^{23}\ particles\ H_2O$$

$$= 3.34 \times 10^{23}$$

Stoichiometry

Stoichiometry is the use of ratios to determine the quantities of reactants used and products produced in a chemical reaction. When we balance a chemical equation, the numbers we place in front of formulas are coefficients. The ratio of coefficients in a chemical equation is known as the molar ratio for those substances. This is similar to reading a recipe in the kitchen. If the balanced equation is the recipe, the chemical formulas are the ingredients and the coefficients are the quantities of those ingredients.

In stoichiometric calculations, the molar ratio will help you convert between moles of any reactant and moles of any product.

Solving Stoichiometry Problems

The coefficients in a balanced chemical equation can be used to create a ratio of moles that are being reacted or produced. It is therefore important that any quantity first be converted to moles before using the molar ratio.

When solving stoichiometry problems, follow these steps:

Step 1: If not already balanced, balance the chemical equation.

Step 2: If necessary, convert the given amount of mass or volume to moles.

Step 3: Use the molar ratio to calculate the moles of the unknown substance based on amount of known substance.

Step 4: If needed, convert the moles of the unknown substance to mass or volume as required by the problem.

Example 1:

Propane burns according to the reaction below. What mass of oxygen is formed by the reaction of 75.0 g of propane?

$$\mathbf{C_3H_8\,(g) + O_2\,(g) \rightarrow CO_2\,(g) + H_2O\,(g)}$$

Step 1: *Balance the chemical equation.*

$$\mathbf{C_3H_8\,(g) + 5\,O_2\,(g) \rightarrow 3\,CO_2\,(g) + 4\,H_2O\,(g)}$$

Step 2: *Convert the given amount of mass to moles.*

The given mass is 75.0 grams of propane. The mass of oxygen is unknown.

$$75.0 \text{ g } C_3H_8 \times \frac{1 \text{ mol}}{44.0 \text{ g}} = 1.70454545 \text{ mol } C_3H_8$$

Step 3: *Use the molar ratio to calculate the moles of the unknown substance based on amount of known substance.*

$$1.70454545 \text{ mol } C_3H_8 \times \frac{5 \text{ mol } O_2}{1 \text{ mol } C_3H_8} = 8.5227272 \text{ mol } O_2$$

Step 4: *Convert the moles to mass.*

$$8.5227272 \text{ mol } O_2 \times \frac{32.0 \text{ g}}{1 \text{ mol}} = 272.7272 \text{ g } O_2$$

$$= \mathbf{273 \text{ g } O_2}$$

Example 2:

You need to produce 10.0 grams of sodium sulfate. How many grams of sodium chloride will you need to react according to the reaction below?

$$(NH_4)_2SO_4 \text{ (aq)} + 2\ NaCl \text{ (s)} \rightarrow 2\ NH_4Cl \text{ (aq)} + Na_2SO_4 \text{ (aq)}$$

Step 1: *Balance the chemical equation.*

This has been done for you. The molar ratio is 1:2:2:1

Step 2: *Convert the given amount of mass to moles.*

The given mass is 10.0 grams of sodium sulfate. The mass of sodium chloride is unknown.

$$10.0 \text{ g } Na_2SO_4 \times \frac{1 \text{ mol}}{142.1 \text{ g}} = 0.070372977 \text{ mol } Na_2SO_4$$

Step 3: *Use the molar ratio to calculate the moles of the unknown substance based on amount of known substance.*

$$0.070372977 \text{ mol } Na_2SO_4 \times \frac{2 \text{ mol NaCl}}{1 \text{ mol } Na_2SO_4} = 0.140745954 \text{ mol NaCl}$$

Step 4: *Convert the moles to mass.*

$$0.140745954 \text{ mol NaCl} \times \frac{58.5 \text{ g}}{1 \text{ mol}} = 8.233638283 \text{ g NaCl}$$
$$= \mathbf{8.23 \text{ g NaCl}}$$

Lesson 8
Acids and Bases

Characteristics of Adds and Bases

You have probably experienced the sour taste of acids. Carbonated beverages or, "pop" are made with carbonic acid. Pickles need vinegar (acetic acid) to prevent them from spoiling. These are what we consider weak acid. Strong acids, such as, sulfuric acid and hydrochloric acid are much more reactive, and can even corrode certain metals. Strong acids are used in many industrial processes like paper and steel production.

Bases include important household chemicals used for cleaning and disinfecting. Bases feel slippery (ex. soap) because they dissolve oils and fatty acids in your skin, reducing the friction between your fingers as you rub them together. Bases destroy protein. Most animal material contains protein, making bases dangerous to handle. Extreme care should be taken whenever you handle a base.

Acids	Bases
The term acid comes from the Latin word *acere,* meaning sour. Some common sour substances like lemon juice (citric acid), and vinegar (acetic acid), are all acids.	Bases all taste bitter. An example of a base is soap. Do not taste bases!
Turn litmus paper red.	Turn litmus paper blue.
Tend to have H (hydrogen) in their formula.	Tend to have OH (hydroxide) in their formula.
Dissociate and contribute H^+ to solution	Dissociate and remove H+ from solution. (usually by producing OH^- that combine to the H^+)
Have pH values less than 7.	Have pH values higher than 7.

Neutralization Reaction

An acid and a base, when combined, will neutralize each other; that is, acids will lose their acidic properties and bases will lose their basic properties. A neutral solution has a pH of 7, although adding an acid and a base together does not necessarily result in a neutral solution. For example, if a strong acid is added to a weak base, the resulting solution may have a more acidic (less than 7) pH.

When acids and bases are added together, the products always include a salt and water. This is a double displacement reaction. The reactants (acid and base) exchange partners to form a salt from the acid's anion and the base's cation. The leftovers are hydrogen (H) and hydroxide (OH) which join to form water.

To write the equation for a neutralization reaction, you will have to identify the acid and the base, and then predict the products of the reaction. Remember, acids and bases mix to form a salt and water.

Step 1: *Write the formulas for the reactants.*

Step 2: *Write the balanced chemical formula for the double displacement reaction, including water as a product.*

Example 1:

Write the neutralization reaction for sodium hydroxide and hydrochloric acid.

Step 1: *Write the formulas for the reactants.*

sodium hydroxide = $NaOH$

hydrochloric acid = HCl

Step 2: *Write the balanced chemical formula for the double displacement reaction, including water as a product.*

$$NaOH + HCl \rightarrow \underset{\textbf{Salt}}{NaCl} + \underset{\textbf{Water}}{H_2O}$$

Appendix A: Periodic Table of Elements

Atomic Number → 19
K ← Symbol
Name → Potassium
39.1 ← Relative Atomic Mass

Period	Group 1	2	3	4	5	6	7	8	9	10	11	12	13	14	15	16	17	18
1	1 **H** Hydrogen 1.0																	2 **He** Helium 4.0
2	3 **Li** Lithium 6.9	4 **Be** Beryllium 9.0											5 **B** Boron 10.8	6 **C** Carbon 12.0	7 **N** Nitrogen 14.0	8 **O** Oxygen 16.0	9 **F** Fluorine 19.0	10 **Ne** Neon 20.2
3	11 **Na** Sodium 23.0	12 **Mg** Magnesium 24.3											13 **Al** Aluminum 27.0	14 **Si** Silicon 28.1	15 **P** Phosphorus 31.0	16 **S** Sulphur 32.1	17 **Cl** Chlorine 35.5	18 **Ar** Argon 39.9
4	19 **K** Potassium 39.1	20 **Ca** Calcium 40.1	21 **Sc** Scandium 45.0	22 **Ti** Titanium 47.9	23 **V** Vanadium 50.9	24 **Cr** Chromium 52.0	25 **Mn** Manganese 54.9	26 **Fe** Iron 55.8	27 **Co** Cobalt 58.9	28 **Ni** Nickel 58.7	29 **Cu** Copper 63.5	30 **Zn** Zinc 65.4	31 **Ga** Gallium 69.7	32 **Ge** Germanium 72.6	33 **As** Arsenic 74.9	34 **Se** Selenium 79.0	35 **Br** Bromine 79.9	36 **Kr** Krpton 83.8
5	37 **Rb** Rubidium 85.5	38 **Sr** Strontium 87.6	39 **Y** Yttrium 88.9	40 **Zr** Zirconium 91.2	41 **Nb** Niobium 92.9	42 **Mo** Molybdenum 96.0	43 **Tc** Technetium (98)	44 **Ru** Ruthenium 101.1	45 **Rh** Rhodium 102.9	46 **Pd** Palladium 106.4	47 **Ag** Silver 107.9	48 **Cd** Cadmium 112.4	49 **In** Indium 114.8	50 **Sn** Tin 118.7	51 **Sb** Antimony 121.8	52 **Te** Tellurium 127.6	53 **I** Iodine 126.9	54 **Xe** Xenon 131.3
6	55 **Cs** Cesium 132.9	56 **Ba** Barium 137.3	57–71 **Lanthanide Series**	72 **Hf** Hafnium 178.5	73 **Ta** Tantalum 180.9	74 **W** Tungsten 183.8	75 **Re** Rhenium 186.2	76 **Os** Osmium 190.2	77 **Ir** Iridium 192.2	78 **Pt** Platinum 195.1	79 **Au** Gold 197.0	80 **Hg** Mercury 200.6	81 **Tl** Thallium 204.4	82 **Pb** Lead 207.2	83 **Bi** Bismuth 209.0	84 **Po** Polonium (209)	85 **At** Astatine (210)	86 **Rn** Radon (222)
7	87 **Fr** Francium (223)	88 **Ra** Radium (226)	89–103 **Actinide Series**	104 **Rf** Rutherfordium (261)	105 **Db** Dubnium (268)	106 **Sg** Seaborgium (271)	107 **Bh** Bohrium (272)	108 **Hs** Hassium (270)	109 **Mt** Meitnerium (276)	110 **Ds** Darmstadium (281)	111 **Rg** Roentgenium (280)	112 **Cn** Copernicium (285)	113 **Uut** Ununtrium (284)	114 **Uuq** Ununquadium (289)	115 **Uup** Ununpentium (288)	116 **Uuh** Ununhexium (293)		118 **Uuo** Ununoctium (294)

Inner Transition Elements

Lanthanide Series	57 **La** Lanthanum 138.9	58 **Ce** Cerium 140.1	59 **Pr** Praseodymium 140.9	60 **Nd** Neodymium 144.2	61 **Pm** Promethium (145)	62 **Sm** Samarium 150.4	63 **Eu** Europium 152.0	64 **Gd** Gadolinium 157.2	65 **Tb** Terbium 158.9	66 **Dy** Dysprosium 162.5	67 **Ho** Holmium 164.9	68 **Er** Erbium 167.3	69 **Tm** Thulium 168.9	70 **Yb** Ytterbium 173.0	71 **Lu** Lutetium 174.9
Actinide Series	89 **Ac** Actinium (227)	90 **Th** Thorium 232.0	91 **Pa** Protactinium 231.0	92 **U** Uranium 238.0	93 **Np** Neptunium (237)	94 **Pu** Plutonium (244)	95 **Am** Americium (243)	96 **Cm** Curium (247)	97 **Bk** Berkelium (247)	98 **Cf** Californium (251)	99 **Es** Einsteinium (252)	100 **Fm** Fermium (257)	101 **Md** Mendelevium (258)	102 **No** Nobelium (259)	103 **Lr** Lawrencium (262)

Appendix C: Table of Electronegativities

Period	Group 1	2	3	4	5	6	7	8	9	10	11	12	13	14	15	16	17	18
1	1 **H** 2.20																	2 **He** —
2	3 **Li** 0.97	4 **Be** 1.47											5 **B** 2.01	6 **C** 2.50	7 **N** 3.07	8 **O** 3.50	9 **F** 4.10	10 **Ne** —
3	11 **Na** 1.01	12 **Mg** 1.23											13 **Al** 1.47	14 **Si** 1.74	15 **P** 2.06	16 **S** 2.44	17 **Cl** 2.83	18 **Ar** —
4	19 **K** 0.91	20 **Ca** 1.04	21 **Sc** 1.20	22 **Ti** 1.32	23 **V** 1.45	24 **Cr** 1.56	25 **Mn** 1.60	26 **Fe** 1.64	27 **Co** 1.70	28 **Ni** 1.75	29 **Cu** 1.75	30 **Zn** 1.66	31 **Ga** 1.82	32 **Ge** 2.02	33 **As** 2.20	34 **Se** 2.48	35 **Br** 2.74	36 **Kr** —
5	37 **Rb** 0.89	38 **Sr** 0.99	39 **Y** 1.11	40 **Zr** 1.22	41 **Nb** 1.23	42 **Mo** 1.30	43 **Tc** 1.36	44 **Ru** 1.42	45 **Rh** 1.45	46 **Pd** 1.35	47 **Ag** 1.42	48 **Cd** 1.46	49 **In** 1.49	50 **Sn** 1.72	51 **Sb** 1.82	52 **Te** 2.01	53 **I** 2.21	54 **Xe** —
6	55 **Cs** 0.86	56 **Ba** 0.97	57–71 Lanthanide Series	72 **Hf** 1.23	73 **Ta** 1.33	74 **W** 1.40	75 **Re** 1.46	76 **Os** 1.52	77 **Ir** 1.55	78 **Pt** 1.44	79 **Au** 1.42	80 **Hg** 1.44	81 **Tl** 1.44	82 **Pb** 1.55	83 **Bi** 1.67	84 **Po** 1.76	85 **At** 1.90	86 **Rn** —
7	87 **Fr** 0.86	88 **Ra** 0.97	89–103 Actinide Series	104 **Rf** —	105 **Db** —	106 **Sg** —	107 **Bh** —	108 **Hs** —	109 **Mt** —	110 **Ds** —	111 **Rg** —	112 **Cn** —	113 **Uut** —	114 **Uuq** —	115 **Uup** —	116 **Uuh** —		118 **Uuo** —

Inner Transition Elements

Lanthanide Series	57 **La** 1.08	58 **Ce** 1.08	59 **Pr** 1.07	60 **Nd** 1.07	61 **Pm** 1.07	62 **Sm** 1.07	63 **Eu** 1.01	64 **Gd** 1.11	65 **Tb** 1.10	66 **Dy** 1.10	67 **Ho** 1.10	68 **Er** 1.11	69 **Tm** 1.11	70 **Yb** 1.06	71 **Lu** 1.14
Actinide Series	89 **Ac** 1.00	90 **Th** 1.11	91 **Pa** 1.14	92 **U** 1.30	93 **Np** 1.29	94 **Pu** 1.25	95 **Am** —	96 **Cm** —	97 **Bk** —	98 **Cf** —	99 **Es** —	100 **Fm** —	101 **Md** —	102 **No** —	103 **Lr** —

Appendix D: Common Ions

Cations (Positive Ions)

1$^+$ charge		2$^+$ charge		3$^+$ charge	
NH_4^+	Ammonium	Ba^{2+}	Barium	Al^{3+}	Aluminum
Cs^+	Cesium	Be^{2+}	Beryllium	Cr^{3+}	Chromium(III)
Cu^+	Copper(I)	Cd^{2+}	Cadmium	Co^{3+}	Cobalt(III)
Au^+	Gold(I)	Ca^{2+}	Calcium	Ga^{3+}	Gallium
H^+	Hydrogen	Cr^{2+}	Chromium(II)	Au^{3+}	Gold(III)
Li^+	Lithium	Co^{2+}	Cobalt(II)	Fe^{3+}	Iron(III)
K^+	Potassium	Cu^{2+}	Copper(II)	Mn^{3+}	Manganese
Rb^+	Rubidium	Fe^{2+}	Iron(II)	Ni^{3+}	Nickel(III)
Ag^+	Silver	Pb^{2+}	Lead(II)		
Na^+	Sodium	Mg^{2+}	Magnesium	**4$^+$ charge**	
		Mn^{2+}	Manganese(II)	Pb^{4+}	Lead(IV)
		Hg_2^{2+}	Mercury(I)	Mn^{4+}	Manganese(IV)
		Hg^{2+}	Mercury(II)	Sn^{4+}	Tin(IV)
		Ni^{2+}	Nickel(II)		
		Sr^{2+}	Strontium		
		Sn^{2+}	Tin(II)		
		Zn^{2+}	Zinc		

(continued)

Anions (Negative Ions)

1^- charge		1^- charge		2^- charge	
CH_3COO^- ($C_2H_3O_2^-$)	Acetate (or ethanoate)	HS^-	Hydrogen sulfide	CO_3^{2-}	Carbonate
				CrO_4^{2-}	Chromate
BrO_3^-	Bromate	OH^-	Hydroxide	$Cr_2O_7^{2-}$	Dichromate
Br^-	Bromide	IO_3^-	Iodate	O^{2-}	Oxide
ClO_3^-	Chlorate	I^-	Iodide	O_2^{2-}	Peroxide
Cl^-	Chloride	NO_3^-	Nitrate	SO_4^{2-}	Sulfate
ClO_2^-	Chlorite	NO_2^-	Nitrite	S^{2-}	Sulfide
CN^-	Cyanide	ClO_4^-	Perchlorate	SO_3^{2-}	Sulfite
F^-	Fluoride	IO_4^-	Periodate	$S_2O_3^{2-}$	Thiosulfate
H^-	Hydride	MnO_4^-	Permanganate		
HCO_3^-	Hydrogen carbonate (or bicarbonate)	SCN^-	Thiocynate	**3^- charge**	
				N^{3-}	Nitride
ClO^-	Hypochlorite			PO_4^{3-}	Phosphate
HSO_4^-	Hydrogen sulfate			P^{3-}	Phosphide
				PO_3^{3-}	Phosphite

Appendix E: Solubility Chart

Negative Ions	Positive Ions	Solubility
essentially all	alkali ions (Li^+, Na^+, K^+, Rb^+, Cs^+)	soluble
essentially all	hydrogen ion $H^+_{(aq)}$	soluble
essentially all	ammonium ion (NH_4^+)	soluble
nitrate, NO_3^-	essentially all	soluble
acetate, CH_3COO^-	essentially all (**except** Ag^+)	soluble
chloride, Cl^- bromide, Br^- iodide, I^-	Ag^+, Pb^{2+}, Hg^{2+}, Cu^+, Tl^+, Hg^+	low solubility
	all others	soluble
sulfate, SO_4^{2-}	Ca^{2+}, Sr^{2+}, Ba^{2+}, Pb^{2+}, Ra^{2+}	low solubility
	all others	soluble
sulfide, S^{2-}	alkali ions, $H^+_{(aq)}$, NH_4^+, Be^{2+}, Mg^{2+}, Ca^{2+}, Sr^{2+}, Ba^{2+}, Ra^{2+}	soluble
	all others	low solubility
hydroxide, OH^-	alkali ions, $H^+_{(aq)}$, NH_4^+, Sr^{2+}, Ba^{2+}, Ra^{2+}, Tl^+	soluble
	all others	low solubility
phosphate, PO_4^{3-} carbonate, CO_3^{2-} sulfite, SO_3^{2-}	alkali ions, $H^+_{(aq)}$, NH_4^+	soluble
	all others	low solubility
chromate, CrO_4^{2-}	Ba^{2+}, Sr^{2+}, Pb^{2+}, Ag^+	low solubility
	all others	soluble

APPENDIX D: RELATIVE STRENGTHS OF ACIDS TABLE

Acid	Reaction	Ka
Perchloric acid	$HClO_4 + H_2O \rightarrow H_3O^+ + ClO_4^-$	very large
Hydriodic acid	$HI + H_2O \rightarrow H_3O^+ + I^-$	very large
Hydrobromic acid	$HBr + H_2O \rightarrow H_3O^+ + Br^-$	very large
Hydrochloric acid	$HCl + H_2O \rightarrow H_3O^+ + Cl^-$	very large
Nitric acid	$HNO_3 + H_2O \rightarrow H_3O^+ + NO_3^-$	very large
Sulfuric acid	$H_2SO_4 + H_2O \rightarrow H_3O^+ + HSO_4^-$	very large
Oxalic acid	$H_2C_2O_4 + H_2O \rightarrow H_3O^+ + HC_2O_4^-$	5.4×10^{-2}
Sulfurous acid	$H_2SO_3 + H_2O \rightarrow H_3O^+ + HSO_3^-$	1.7×10^{-2}
Hydrogen sulfate ion	$HSO_4^- + H_2O \rightarrow H_3O^+ + SO_4^{2-}$	1.3×10^{-2}
Phosphoric acid	$H_3PO_4 + H_2O \rightarrow H_3O^+ + H_2PO_4^-$	7.1×10^{-3}
Ferric ion	$Fe(H_2O)_6^{3+} + H_2O \rightarrow H_3O^+ + Fe(H_2O)_5(OH)^{2+}$	6.0×10^{-3}
Hydrogen telluride	$H_2Te + H_2O \rightarrow H_3O^+ + HTe^-$	2.3×10^{-3}
Hydrofluoric acid	$HF + H_2O \rightarrow H_3O^+ + F^-$	6.7×10^{-4}
Nitrous acid	$HNO_2 + H_2O \rightarrow H_3O^+ + NO_2^-$	5.1×10^{-4}
Hydrogen selenide	$H_2Se + H_2O \rightarrow H_3O^+ + HSe^-$	1.7×10^{-4}
Chromic ion	$Cr(H_2O)_6^{3+} + H_2O \rightarrow H_3O^+ + Cr(H_2O)_5(OH)^{2+}$	1.5×10^{-4}
Benzoic acid	$C_6H_5COOH + H_2O \rightarrow H_3O^+ + C_6H_5COO^-$	6.6×10^{-5}
Hydrogen oxalate ion	$HC_2O_4^- + H_2O \rightarrow H_3O^+ + C_2O_4^{2-}$	5.4×10^{-5}
Acetic acid	$HC_2H_3O_2 + H_2O \rightarrow H_3O^+ + C_2H_3O_2^-$	1.8×10^{-5}
Aluminum ion	$Al(H_2O)_6^{3+} + H_2O \rightarrow H_3O^+ + Al(H_2O)_5(OH)^{2+}$	1.4×10^{-5}
Carbonic acid	$H_2CO_3 + H_2O \rightarrow H_3O^+ + HCO_3^-$	4.4×10^{-7}
Hydrogen sulfide	$H_2S + H_2O \rightarrow H_3O^+ + HS^-$	1.0×10^{-7}
Dihydrogen phosphate ion	$H_2PO_4^- + H_2O \rightarrow H_3O^+ + HPO_4^{2-}$	6.3×10^{-8}
Hydrogen sulfite ion	$HSO_3^- + H_2O \rightarrow H_3O^+ + SO_3^{2-}$	6.2×10^{-8}
Ammonium ion	$NH_4^+ + H_2O \rightarrow H_3O^+ + NH_3$	5.7×10^{-10}
Hydrogen carbonate ion	$HCO_3^- + H_2O \rightarrow H_3O^+ + CO_3^{2-}$	4.7×10^{-11}
Hydrogen telluride ion	$HTe^- + H_2O \rightarrow H_3O^+ + Te^{2-}$	1.0×10^{-11}
Hydrogen peroxide	$H_2O_2 + H_2O \rightarrow H_3O^+ + HO_2^-$	2.4×10^{-12}
Monohydrogen phosphate	$HPO_4^{2-} + H_2O \rightarrow H_3O^+ + PO_4^{3-}$	4.4×10^{-13}
Hydrogen sulfide ion	$HS^- + H_2O \rightarrow H_3O^+ + S^{2-}$	1.2×10^{-15}
Water	$H_2O + H_2O \rightarrow H_3O^+ + OH^-$	1.8×10^{-16}
Hydroxide ion	$OH^- + H_2O \rightarrow H_3O^+ + O^{2-}$	$< 10^{-36}$
Ammonia	$NH_3 + H_2O \rightarrow H_3O^+ + NH_2^-$	very small

Appendix F: Table of Standard Reduction Potentials with Values

Oxidized species	↔	Reduced Species	E°/V
$Li^+{}_{(aq)} + e^-$	↔	$Li_{(s)}$	−3.04
$K^+{}_{(aq)} + e^-$	↔	$K_{(s)}$	−2.93
$Ca^{2+}{}_{(aq)} + 2e^-$	↔	$Ca_{(s)}$	−2.87
$Na^+{}_{(aq)} + e^-$	↔	$Na_{(s)}$	−2.71
$Mg^{2+}{}_{(aq)} + 2e^-$	↔	$Mg_{(s)}$	−2.37
$Al^{3+}{}_{(aq)} + 3e^-$	↔	$Al_{(s)}$	−1.66
$Mn^{2+}{}_{(aq)} + 2e^-$	↔	$Mn_{(s)}$	−1.19
$H_2O_{(l)} + e^-$	↔	$\frac{1}{2}H_{2(g)} + OH^-{}_{(aq)}$	−0.83
$Zn^{2+}{}_{(aq)} + 2e^-$	↔	$Zn_{(s)}$	−0.76
$Fe^{2+}{}_{(aq)} + 2e^-$	↔	$Fe_{(s)}$	−0.45
$Ni^{2+}{}_{(aq)} + 2e^-$	↔	$Ni_{(s)}$	−0.26
$Sn^{2+}{}_{(aq)} + 2e^-$	↔	$Sn_{(s)}$	−0.14
$Pb^{2+}{}_{(aq)} + 2e^-$	↔	$Pb_{(s)}$	−0.13
$H^+{}_{(aq)} + e^-$	↔	$\frac{1}{2}H_{2(g)}$	0.00
$Cu^{2+}{}_{(aq)} + e^-$	↔	$Cu^+{}_{(aq)}$	+0.15
$SO_4{}^{2-}(aq) + 4\ H^+{}_{(aq)} + 2e^-$	↔	$H_2SO_3(aq) + H_2O_{(l)}$	+0.17
$Cu^{2+}{}_{(aq)} + 2e^-$	↔	$Cu_{(s)}$	+0.34
$\frac{1}{2}O_{2(g)} + H_2O_{(l)} + 2e^-$		$2OH^-{}_{(aq)}$	+0.40
$Cu^+{}_{(aq)} + e^-$		$Cu_{(s)}$	+0.52
$\frac{1}{2}I_{2(s)} + e^-$	↔	$I^-{}_{(aq)}$	+0.54
$Fe^{3+}{}_{(aq)} + e^-$	↔	$Fe^{2+}{}_{(aq)}$	+0.77
$Ag^+{}_{(aq)} + e^-$	↔	$Ag_{(s)}$	+0.80
$\frac{1}{2}Br_{2(l)} + e^-$	↔	$Br^-{}_{(aq)}$	+1.07
$\frac{1}{2}O_{2(g)} + 2H^+{}_{(aq)} + 2e^-$	↔	$H_2O_{(l)}$	+1.23
$Cr_2O_7{}^{2-}{}_{(aq)} + 14H^+{}_{(aq)} + 6e^-$	↔	$2Cr^{3+}{}_{(aq)} + 7H_2O_{(l)}$	+1.33
$\frac{1}{2}Cl_{2(g)} + e^-$	↔	$Cl^-{}_{(aq)}$	+1.36
$MnO_4{}^-{}_{(aq)} + 8H^+{}_{(aq)} + 5e^-$	↔	$Mn^{2+}{}_{(aq)} + 4H_2O_{(l)}$	+1.51
$\frac{1}{2}F_{2(g)} + e^-$	↔	$F^-{}_{(aq)}$	+2.87

An activity series can also list other species in order of reactivity. Most North American activity series are listed as Standard Reduction Potentials at 25°C and ionic concentrations of 1 mol/L solution.

Many activity series will be ranked **in terms of reduction reactions**, like the one that follows. The better **oxidizing agents**, like permanganate and peroxide, are found at the **bottom left of the table**. They will draw electrons away from almost any substance to become reduced.

↑ Decreasing Ability as an Oxidizing Agent	$Li^+ + e^- \rightarrow Li_{(s)}$	Strongest Reducing Agents Easiest to Oxidize
	$Rb^+ + e^- \rightarrow Rb_{(s)}$	
	$K^+ + e^- \rightarrow K_{(s)}$	
	$2\ H_2O + 2e^- \rightarrow 2\ OH^- + H_{2(g)}$	
	$Zn^{2+} + 2e^- \rightarrow Zn_{(s)}$	
	$Cr^{3+} + 3e^- \rightarrow Cr_{(s)}$	↓ Decreasing Ability as a Reducing Agent
	$Sn^{2+} + 2e^- \rightarrow Sn_{(s)}$	
	$Pb^{2+} + 2e^- \rightarrow Pb_{(s)}$	
	$Fe^{3+} + 3e^- \rightarrow Fe_{(s)}$	
	$2\ H^+ + 2e^- \rightarrow H_{2(g)}$	
	$Cu^{2+} + 2e^- \rightarrow Cu_{(s)}$	
	$Cu^+ + e^- \rightarrow Cu_{(s)}$	
	$I_{2(s)} + 2e^- \rightarrow 2\ I^-$	
	$O_{2(g)} + 2H^+ + e^- \rightarrow H_2O_2$	
	$Li^+ + e^- \rightarrow Li_{(s)}$	
	$Fe^{3+} + e^- \rightarrow Fe^{2+}$	
	$NO_3^- + 2\ H^+ + e^- \rightarrow NO_{2(g)} + H_2O$	
	$Ag^+ + e^- \rightarrow Ag_{(s)}$	
	$NO_3^- + 4\ H^+ + 3e^- \rightarrow NO_{(g)} + 2\ H_2O$	
	$Br_{2(g)} + 2e^- \rightarrow 2\ Br^-$	
Strongest Oxidizing Agents Easiest to Reduce	$½O_{2(g)} + 2\ H^+ + 2e^- \rightarrow H_2O$	
	$Au^{3+} + 3e^- \rightarrow Au_{(s)}$	
	$MnO_4^- + 8\ H^+ + 5e^- \rightarrow Mn^{2+} + 4\ H_2O$	
	$H_2O_2 + 2\ H^+ + 2e^- \rightarrow 2\ H_2O$	
	$F_{2(g)} + 2e^- \rightarrow 2\ F^-$	