Appendix N

CONCENTRATION UNITS

Chemists and engineers often encounter technical situations that require the description of the relative quantity of some substance in terms of its concentration in a solution or in air. For these purposes a variety of concentration units have been developed that can be used according to the needs and requirements of the job.

Concentration units can be divided into two broad categories. There are those units that are used in research and scientific applications and those units that are used in industrial applications. These categories are not firm and fast, but they are useful in understanding when one unit is used in preference to another. Several common *scientific* and *industrial* units are listed below:

SCIENTIFIC	INDUSTRIAL
molarity (*M*)	mass percent (%)
molality (*m*)	parts per million (ppm)
mole fraction (*X*)	parts per billion (ppb)

The primary difference between scientific and industrial units is that in scientific units the identities of the solvents and solutes are known and are usually pure substances; in industrial units, the solutes and solvents may be mixtures of materials whose composition is not well defined. The use of scientific units requires the calculation of a number of moles, therefore the molar mass must be defined. However, this requirement is not present with the industrial units.

A. SCIENTIFIC UNITS

1. Molarity (*M*)

Molarity is defined as:

$$M = \frac{(\text{mol solute})}{(\text{L solution})}$$

and is reported in units of mol/L or molar. Molar concentrations are the most common of the scientific concentration units because molar solutions are easy to prepare. In general, a molar solution is prepared by placing a known quantity of solute in a volumetric flask and adding solvent until the flask is filled to the calibration mark.

Molar concentration is used in many common laboratory applications and is the concentration unit most frequently encountered in the general chemistry laboratory. One limitation in the use of molar concentrations is that the volume of the solution changes with temperature. This results in a change in the molar concentration. For work at constant temperature, however, this limitation does not exist.

2. Molality (*m*)

Molality is defined as:

$$m = \frac{(\text{mol solute})}{(\text{kg solvent})}$$

and is reported in units of mol/1000 g, mol/kg, or molal. Molal concentration is less convenient to use than molarity because the mass of the solvent must be determined rather than its volume. In practice, generally only small quantities of these solutions are prepared and therefore this inconvenience is minor.

Molal concentrations are used in experiments determining freezing point depressions, boiling point elevations, and in other cases where the temperature of the solution changes over the course of the experiment. Molal concentrations eliminate the problem of the dependence of volume on temperature seen with molar solutions because masses are not temperature dependent.

3. Mole Fraction (*X*)

Mole fraction is defined as:

$$\overline{X} = \frac{(\text{mol solute})}{(\text{mol solute } + \text{ mol solvent})}$$

As with molality, preparing solutions using mole fraction units requires the weighing of solute and solvent so that the number of moles can be calculated. For many solutions, the quantity of solute is small compared to that of solvent, so the numerical values of mole fraction are frequently very small numbers. Because moles are the units of both numerator and denominator, mole fraction is a unit-less quantity.

B. Industrial Units

1. Mass Percent (%) (or percent mass)

Mass % is defined as:

$$\% = \left[\frac{(\text{mass of solute})}{(\text{total mass of solution})}\right] \times 100$$

A similar relationship holds for volume percent (or percent by volume).

$$\%_{\text{volume}} = \left[\frac{(\text{volume of solute})}{(\text{total volume of solution})}\right] \times 100$$

Mass percent is used to express the concentration of substances that are not pure, for example, the content of butterfat in milk. Butterfat is not a pure substance, but its mass percent in milk determines the legal difference between skim milk, whole milk, and table cream.

Mass percent is also used in the metallurgical industry to describe the quantities of different components in alloys. Certain generic names such as "stainless steel" imply the presence of certain minimum levels of different components in the steel. In this case, the definitions are not legal ones, but are standard compositions accepted by industry. Examples are:

COMMON NAME	MASS-PERCENT ELEMENTAL COMPOSITION
brass	67-90% Cu, 10-33% Zn
dental amalgam	50% Hg, 35% Ag, 15% Sn
18-carat gold	75% Au, 10-20% Ag, 5-15% Cu
stainless steel	73-79% Fe, 14-18% Cr, 7-9% Ni

2. Parts per Million

Parts per million (ppm) is defined as:

$$\text{ppm} = \left[\frac{(\text{mass of solute})}{\text{total mass of solution}}\right] \times 10^6$$

The unit "parts per million" is usually used for very dilute aqueous solutions. For very dilute aqueous solutions the density of the solution is assumed to be the same as the density of pure water (1.00 g/mL), so that one liter of solution has a mass of 1000 g and 1000 L of solution has a mass of 10^6 g. This information gives two additional definitions of parts per million.

$$\text{ppm} = \frac{(\text{g solute})}{(1000\ \text{L solution})}$$

$$\text{ppm} = \frac{(\text{mg solute})}{(\text{L solution})}$$

Parts per million is used to describe concentrations in solutions containing poorly defined or unidentified solutes or mixtures of solutes. Parts per million is also used to describe concentration levels to those who are unfamiliar with the concept of moles.

Tap water contains a variety of dissolved impurities, mostly minerals. A simple way of describing the levels of these impurities, often called dissolved solids, is to measure out a sample, 10 mL for example, into a weighed aluminum cup and evaporate the water. The mass of the remaining solids is determined by difference. Dividing the mass in mg of remaining solids by 0.010 L gives the concentration of the dissolved solids in ppm. Notice that the identity of the solids is not specified, but a measure of their concentration is known.

The Navy uses the level of chloride ions in boiler water as an early warning indicator for seawater leaks into condensers. The titration test that is used precipitates the chloride ions as a mercury compound. The chloride ions themselves are present in water along with many cations including sodium, potassium, magnesium, and calcium ions. Because it is not appropriate to talk about the amount of sodium chloride in the boiler water, the results from the mercury titration are used to determine the mass in mg of chloride ions in the sample and are in turn expressed as ppm of chloride.

3. Parts per Billion

Parts per billions is defined as:

$$\text{ppb} = \left[\frac{(\text{mass of solute})}{(\text{total mass of solution})}\right] \times 10^9$$

For very dilute aqueous solutions the approximation is:

$$\text{ppb} = \frac{(\mu g \text{ solute})}{(\text{L solution})}$$

Parts per billion is used for reporting trace contaminants of substances in water supplies and for substances present in extremely small quantities in other systems.

Workshop 2
Stoichiometry

Introduction

Chemical reactions have been known and applied for many years. For example, ancient people knew that if carbon (charcoal) was mixed with iron oxides and heated, elemental iron and carbon dioxide would be formed. This practical knowledge was attained without any concept of atoms, molecules and reactions. By the nineteenth century, the study of stoichiometry allowed chemists to determine masses of reactants and products during the reaction. We will use stoichiometry to allow us to predict the mass of a reactant or product using a balanced equation.

Balanced Equations

Chemical equations may take many forms. The statement above that iron oxide reacts with charcoal to produce elemental iron and carbon dioxide is one form of describing a chemical reaction. However, this type of statement does not allow any quantitative information to be determined. Using the information above, and a little knowledge of chemical formulas, we can produce a balanced chemical equation to use in quantitative determinations. The term "iron oxide" is ambiguous because iron and oxygen may react with several different ratios. For our purposes, we will assume that the iron oxide in question is iron(III) oxide. The formula of iron(III) oxide is Fe_2O_3. Carbon is a monatomic element with the symbol C. Elemental iron is Fe. Carbon dioxide has the formula CO_2. We can write an unbalanced equation as:

$$Fe_2O_3 + C \rightarrow Fe + CO_2$$

In order to get quantitative information from this equation it must be balanced. Balancing the equation would result in the following equation:

$$2Fe_2O_3 + 3C \rightarrow 4Fe + 3CO_2$$

This equation tells us that 2 moles of Fe_2O_3 will react with 3 moles of C to produce 4 moles of Fe and 3 moles of CO_2. We will use this information to form conversion factors to allow us to calculate the amount of one substance from the amount of another substance.

Molar Masses

An additional conversion factor necessary in stoichiometric calculations is molar masses. As we learned in the last workshop, the atomic masses from the periodic table expressed in grams represent Avagadro's number of molecules. One mole is defined to be Avagadro's number of particles. If the atomic mass of an element is expressed in grams, this represents one mole of atoms.

$$1 \text{ mol} = 6.022\text{x}10^{23} \text{ atoms} = \text{atomic mass in grams}$$

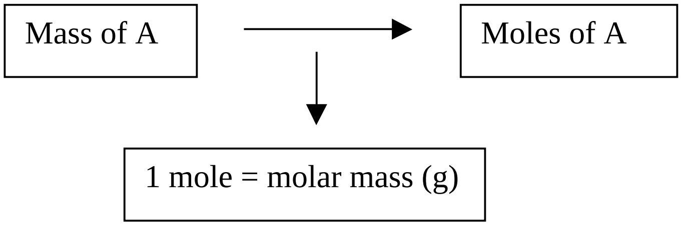
Similarly, the molar mass of substances is defined to be the mass of one mole of substance. For compounds, calculation of the molar mass is accomplished by adding the atomic masses. If the molar mass is expressed in grams, this represents one mole of the compound.

1 mol = 6.022x10^{23} formula units = molar mass in grams

We use the more general term “formula units” instead of molecules because ionic compounds do not exist as molecules.

Calculation of Moles

We can use the conversion factors above to calculate the number of moles of a substance from the mass of the substance. Using a flow chart method, we can diagram this process:



Name: ___________________________

Self Test

Complete these problems before coming to Workshop. Be sure to show your work.

1. Complete the following table

Compound	Molar Mass	Number of moles in 100.0 g
C		
Fe_2O_3		
CO_2		

2. A survey requires responses from 3 men for every 2 women. If you have received responses from 3241 men, how many women do you need?

Workshop

For the first two questions, use paired problem solving. After each problem is completed, review the listener's notes to discover the path of problem solving that was taken. After the two questions have been completed, work as a large group to develop a universal flow chart for stoichiometry problems.

1. Consider the equation below describing the decomposition of hydrogen peroxide:

$$2H_2O_2(aq) \rightarrow 2H_2O(l) + O_2(g)$$

Using pennies to represent hydrogen atoms and nickels to represent oxygen atoms, build models to represent the following scenarios:

a. How many molecules of H_2O_2 are required to produce 4 H_2O molecules?
b. How many molecules of H_2O_2 are required to produce 4 O_2 molecules?

2. Using the balanced equation in the introduction for the production of iron from Fe_2O_3, calculate

a. the number of moles of iron produced by the reaction of 15.0 moles of Fe_2O_3 with an unlimited supply of carbon.

b. the number of grams of iron produced by the reaction of 15.0 moles of Fe_2O_3 with an unlimited supply of carbon.

c. the number of grams of iron produced by the reaction of 100.0 grams of Fe_2O_3 with an unlimited supply of carbon.

Use group brainstorming to solve parts (a) through (d) of question 3. Use a round robin approach to solve the remaining questions.

3. Vitamin C is a compound of carbon, hydrogen and oxygen. A 35.5 mg sample of vitamin C was completely burned in a combustion apparatus, and 53.3 mg of CO_2 and 14.4 mg of H_2O were recovered.

a. Write an unbalanced chemical equation, identifying the reactants and products. (Hint: Combustion involves a reaction with oxygen.) The equation cannot yet be balanced because vitamin C can only be represented as $C_xH_yO_z$.

b. What is the reactant source of the carbon in the carbon dioxide product?

c. What is the reactant source of the hydrogen in the water product?

d. What is the reactant source of the oxygen in the carbon dioxide and water products?

e. How many moles of carbon were in the original vitamin C sample?

f. How many moles of hydrogen were in the original vitamin C sample?

g. How many moles of oxygen were in the sample?

h. What is the ratio of moles of hydrogen to moles of carbon to moles of oxygen in vitamin C? Reduce to whole numbers.

i. What is the empirical formula of vitamin C?

CHAPTER 2. INTRODUCTION TO CHEMICAL REACTIONS

Throughout your study of chemistry you will be confronted with the nomenclature, formulas and reactions of chemical compounds. Therefore, it is important to develop a familiarity with the general principles that pertain to the naming of chemical compounds, and to the prediction of their formulas and reactions. Because over 100 elements exist in nature and because many elements exhibit the capacity to combine with others in a variety of proportions, the subject of chemical reactions and nomenclature is complicated by the sheer number of possible compounds. Nevertheless, relatively simple rules may be applied to the naming of compounds and to the prediction of their formulas.

A. Oxidation Numbers

The first problem one faces in writing the formula for a chemical compound is predicting the relative numbers of atoms of the different elements involved. This task is facilitated by using what is termed the oxidation numbers of the elements to be combined. The oxidation number is related to the number of electrons an atom must gain or lose to form the most stable electron configuration possible in combination with the other elements that make up a compound.

In general, metals tend to lose electrons in chemical interactions to become positively charged cations;

$$Li \rightarrow Li^{+} + e-$$

$$Mg \rightarrow Mg^{2+} + 2\ e^{-}$$

$$Al \rightarrow Al^{3+} + 3\ e^{-}.$$

The loss of electrons in a chemical process is called **oxidation**. Nonmetals, on the other hand, tend to gain electrons in chemical interactions to become negatively charged anions;

$$\tfrac{1}{2}\ F_2 + e^{-} \rightarrow F$$

$$S + 2\ e^{-} \rightarrow S^{2-}.$$

The gain of electrons in a chemical process is called **reduction**. The charge that an element appears to adopt in a chemical compound is the **oxidation number**. Most elements have multiple oxidation numbers, such as iron, which commonly exhibits an oxidation number of 2+ (the ferrous ion) and 3+ (the ferric ion). The oxidation numbers of many of the common elements may be assigned according to the following rules.

<u>Rules for assigning oxidation numbers</u>

1. The oxidation number of atoms of a free element or of a compound composed of only one element is zero. (For example, Na, H_2, Ca, Cl_2).
2. The oxidation number of a monatomic ion is equal to its ionic charge. (For example, Na^+ = 1, $C1^-$ = −1, Fe^{3+} = 3).
3. The oxidation number of hydrogen is 1 in all compounds except when it is combined with the group I and group II metals, in which case it is -1. (For example, 1 in H_2S, 1 in $HClO_4$, but −l in NaH).
4. The oxidation number of oxygen is −2 in all compounds except those which involve O-O bonds (i.e., peroxides such as H_2O_2) where it is −1, and in superoxides (such as KO_2) where it is −1/2.
5. The oxidation numbers of some other elements in all compounds are:

   group I = 1
   group II = 2
   group VII = −l (but not when combined with oxygen).
6. The oxidation numbers of other elements in polyatomic compounds or ions may usually be determined by requiring the sum of the oxidation numbers to equal zero in the case of a neutral molecule or the ionic charge in the case of an ion.

<u>Example</u>: Determine the oxidation number of Cr in $Na_2Cr_2O_7$.

Na = 1 (rule 5), O = −2 (rule 4)

Since the compound is neutral, 2(Na) + 2(Cr) + 7(O) = 0, hence Cr = (−2 + 14)/2 = 6.

<u>Exercises</u>:

a). Determine the oxidation number of S in $H_2S_2O_3$.

b). Determine the oxidation number of C in $C_{12}H_{22}O_{11}$.

B. Chemical Formulas and Names

Your first task in mastering chemical formulas and nomenclature will be to learn the names and symbols of the most common elements. At this stage of the game, you should concentrate on learning the names and symbols of the elements in the first three rows of the periodic table. Most of the symbols are the first two letters of the name. However, some symbols are derived from Latin names, which make them a little harder to recognize. Specifically, Na is sodium, K is potassium, Fe is iron, Ag is silver, Sn is tin, Sb is antimony, W is tungsten, Au is gold, Hg is mercury and Pb is lead.

1. Rules for naming inorganic compounds

Cations: A cation is an ion with a positive charge. Use the name of the element followed by a roman numeral indicating its oxidation number followed by the word ion. For example, Cu^{2+} is copper(II) ion, Cr^{3+} is chromium(III) ion, Sn^{4+} is tin(IV) ion.

An older convention which is still in use assigns the suffix *ous* to the ending of the (Latin) name in the case of the lowest ionic charge and *ic* in the case of the highest. For example, Cu^+ is cuprous ion and Cu^{2+} is cupric ion; Sn^{2+} is stannous ion and Sn^{4+} is stannic ion.

Anions: An anion is an ion with a negative charge. Replace the ending of the element name with *ide*. For example, $C1^-$ is chloride, O^{2-} is oxide, S^{2-} is sulfide, N^{3-} is nitride.

Binary Compounds:

For the purpose of naming, the rules for cations apply to the element with the positive oxidation number and the rules for anions apply to the element with the negative oxidation number. The name of the compound is the name of the cation followed by the name of the anion. For example, NaCl is sodium chloride, $CuBr_2$ is copper(II) bromide (or cupric bromide), Fe_2O_3 is iron(III) oxide (or ferric oxide).

Polyatomic compounds:

A polyatomic compound is usually composed of one or more complex ions which have special names. The names of the common complex ions and their charges are tabulated below and <u>should be committed to memory</u>. The name of a polyatomic compound is obtained by giving the cation or complex cation name followed by the name of the anion or complex anion. For example, $Na_2Cr_2O_7$ is sodium dichromate, $KMnO_4$ is potassium permanganate, $(NH_4)_2SO_4$ is ammonium sulfate.

Table 2.1. Names and formulas of complex ions.

Cations				Anions					
+1		+2		−1		−2		−3	
Ammonium	NH_4^+	Vanadyl	VO^{2+}	Hydroxide	OH^-	Carbonate	CO_3^{2-}	Phosphate	PO_4^{3-}
Nitryl	NO_2^+	Uranyl	UO_2^{2+}	Nitrite[a]	NO_2^-	Sulfite[a]	SO_3^{2-}	Arsenite[a]	AsO_3^{3-}
Nitrosyl	NO^+			Nitrate[a]	NO_3^-	Sulfate[a]	SO_4^{2-}	Arsenate[a]	AsO_4^{3-}
				Hypochlorite[b]	ClO^-	Thiosulfate	$S_2O_3^{2-}$	Borate	BO_3^{3-}
				Chlorite[b]	ClO_2^-	Chromate	CrO_4^{2-}		
				Chlorate	ClO_3^-	Dichromate	$Cr_2O_7^{2-}$		
				Perchlorate[b]	ClO_4^-	Oxalate	$C_2O_4^{2-}$		
				Cyanide	CN^-	Peroxide	O_2^{2-}		
				Thiocyanate	SCN^-	Hydrogen - phosphate[c]	HPO_4^{2-}		
				Acetate	$CH_3CO_2^-$				
				Permanganate	MnO_4^-				
				Hydrogen - carbonate[c]	HCO_3^-				
				Hydrogen - sulfite[c]	HSO_3^-				
				Hydrogen - sulfate[c]	HSO_4^-				
				Dihydrogen - phosphate[c]	$H_2PO_4^-$				

[a] When an element forms two oxoanions, the ion with the smaller number of oxygens takes the suffix -*ite* and the ion with the larger number of oxygens the suffix -*ate*.

[b] When an element forms more than two oxoanions the prefixes *hypo*- and *per*- are used to designate the ions with the smallest and largest number of oxygen atoms, respectively.

[C] When two or more anions differ in the number of hydrogen atoms they contain, they are distinguished by preceding the name of the anion by *hydrogen* (one H atom) or *dihydrogen* (two H atoms). The prefix *bi*- to denote the presence of hydrogen is part of an older but still commonly used nomenclature.

2. <u>Writing formulas for inorganic compounds</u>

Now that you know all about oxidation numbers and complex (or polyatomic) ions, writing formulas for chemical compounds is simply a matter of making the oxidation numbers add up to zero.

Example: Write the chemical formula for sodium carbonate.

Na has ox. no. = 1, carbonate is CO_3^{2-} (see the Table 2.1) and hence has ox. no. = −2. Thus, to form the neutral compound sodium carbonate requires 2 Na^+ and 1 CO_3^{2-}. Therefore, the formula is Na_2CO_3.

Exercise: Write chemical formulas for the following compounds:

potassium hydrogen phosphate
ammonium dichromate
iron(III) sulfate

C. Classification of Inorganic Reactions

Chemistry has traditionally been divided into two main branches: organic chemistry, which involves the chemistry of carbon compounds, and inorganic chemistry, which is the chemistry of everything else. In organic chemistry, the large variety of carbon bearing compounds is attributable to the ability of carbon to form strong bonds with itself. Inorganic chemistry, on the other hand, derives its complexity from the variety of oxidation states available to many of the elements. A general classification scheme for inorganic reactions is shown on the following page.

1. Oxidation-reduction reactions

Oxidation-reduction reactions involve the transfer of electrons from one atom to another. Since an exchange of electrons requires that one atom gain electrons while another atom lose electrons, oxidation and reduction must always occur together. To determine whether an element has been oxidized or reduced in a chemical reaction, we must calculate the change in its oxidation number in going from the reactant side to the product side.

Example: Determine which element has been oxidized and which element has been reduced in the following reaction:

$$KMnO_4 + KCl + H_2SO_4 \rightarrow MnSO_4 + K_2SO_4 + H_2O + Cl_2$$

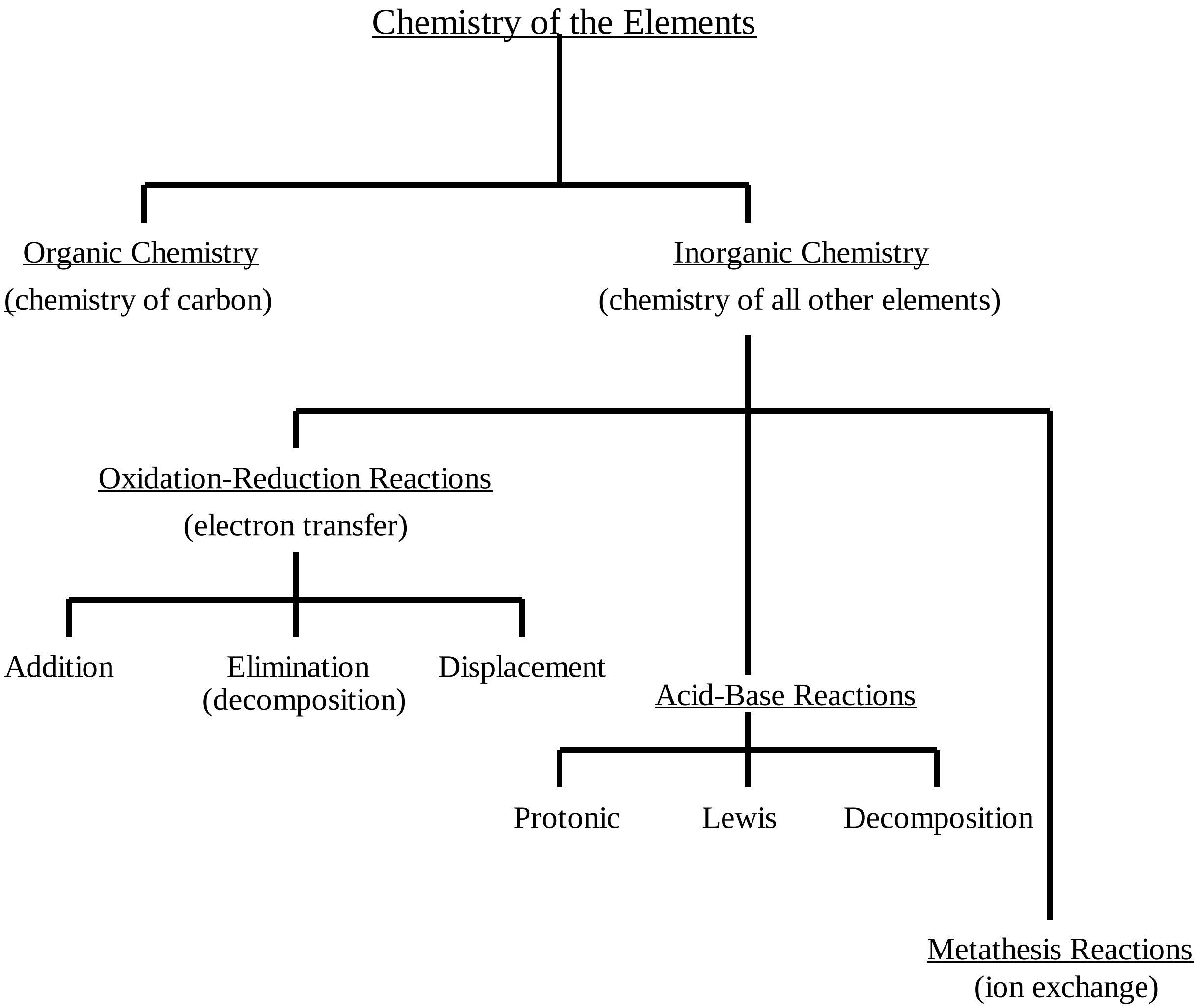


Figure 2.1. Classification scheme for chemical reactions.

First calculate the oxidation number per atom for each element:

Left side	Right side
K = +l	K = +1
Mn = +7	Mn = +2
O = −2	O = −2
K = +1	

Cl = −1	Cl = 0
H = +1	H = +1
S = +6	S = +6
O = −2	O = −2

It is evident that the element Mn changed from +7 to +2 and hence gained 5 electrons per atom, while the element Cl changed from −1 to 0 and hence lost 1 electron per atom. Therefore Mn was reduced and Cl was oxidized.

There are three types of oxidation-reduction reactions.

a). *Addition reactions*:

The combination of a reactive metal with an active nonmetal to produce an ionic salt constitutes a redox addition reaction;

Metal + Nonmetal → Ionic Salt

Examples:

$$4\ Mg_{(s)} + O_{2(g)} \rightarrow 2\ MgO_{(s)} \qquad \text{(flare)}$$

$$4\ Fe_{(s)} + 3\ O_{2(g)} \rightarrow 2\ Fe_2O_3 \qquad \text{(rust)}$$

Exercise: Determine the change in oxidation number (per atom) for each element in the above reactions.

The different oxidation states of oxygen are formed only in reactions with the most electropositive metals (i.e., the alkali metals);

$$4\ Li + O_2 \rightarrow 2\ Li_2O \qquad (-2)\ \text{oxide}$$

$$2\ Na + O_2 \rightarrow Na_2O_2 \qquad (-1)\ \text{peroxide}$$

$$K + O_2 \rightarrow KO_2 \qquad (-\tfrac{1}{2})\ \text{superoxide}$$

Fluorine is so reactive that under conditions of high temperature and pressure it will react with the noble gas Xe;

$$Xe_{(g)} + n\ F_{2(g)} \rightarrow \begin{array}{l} XeF_{2(g)} \\ XeF_{4(g)} \\ XeF_{6(g)} \end{array}$$

Other examples of redox addition reactions which do not involve metals are

$$2\ NO_{(g)} + O_{2(g)} \rightarrow 2\ NO_{2(g)}$$

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

$$PF_{3(g)} + F_{2(g)} \rightarrow PF_{5(g)}$$

$$PtCl_4^{2-}{}_{(aq)} + Cl_{2(g)} \rightarrow PtCl_6^{2-}{}_{(aq)}$$

$$Fe(CN)_6^{4-}{}_{(aq)} + IrCl_6^{2-}{}_{(aq)} \rightarrow Fe(CN)_6^{3-}{}_{(aq)} + IrCl_6^{3-}.$$

b). *Elimination Reactions*

In redox elimination (or decomposition) reactions, a complex molecule splits up into smaller components. Many reactions of this type are quite violent and can produce explosions. Some rather spectacular examples of elimination reactions are:

$NH_4NO_{3(s)} \rightarrow N_2O_{(g)} + 2\ H_2O_{(g)}$ (The infamous fertilizer bomb)

$$2\ KC1O_{3(s)} \rightarrow 2\ KC1_{(s)} + 3\ O_{2(g)}$$

$$(NH_4)_2Cr_2O_{7(s)} \rightarrow N_{2(g)} + Cr_2O_{3(s)} + 4\ H_2O_{(g)}$$

$$14\ KMnO_{4(s)} + 4\ C_3H_5(OH)_{3(l)} \rightarrow 7\ K_2CO_{3(s)} + 7\ Mn_2O_{3(s)} + 5\ CO_{2(g)} + 6\ H_2O_{(g).}$$

c). *Displacement Reactions*

In redox displacement reactions, one ion displaces or replaces another. An example is:

$$2\ M_{(s)} + H_2O_{(l)} \rightarrow H_{2(g)} + 2\ M^+{}_{(aq)} + 2\ OH^-{}_{(aq)},$$

where M is an alkali metal. The reactivity of the metals decreases down the column of the periodic table and less active metals require acids to dissolve. The more active halogens displace the less active halogens;

$$Cl_{2(g)} + 2\ NaBr_{(aq)} \rightarrow Br_{2(g)} + 2\ NaCl_{(aq)}.$$

d). *Balancing Oxidation-Reduction Reactions*

Most simple chemical reactions are easily balanced using the trial and error method whereby one attempts to choose stoichiometric coefficients that will result in equal numbers of the same elements and charges on both sides of the equation, one at a time, with the help of just a little intuition. This method works fine for relatively simple reactions involving up to two reactants and two products. Most oxidation-reduction reactions are more complicated than this and so a more systematic method is required for this class of reactions.

A general procedure which works well for most oxidation-reduction reactions is outlined below.

A procedure for balancing oxidation-reduction reactions

1). Determine oxidation numbers for the reactants and products and identify the elements for which the oxidation numbers (per atom) change in going from reactants to products. Usually H^+, OH^- and H_2O can be ignored in this step (but not when peroxides are involved).

2). Balance the elements whose oxidation numbers change and compute the total oxidation numbers. The total oxidation number = oxidation number per atom x number of atoms (including the stoichiometric coefficient).

3). Take the differences in total oxidation numbers to determine the numbers of electrons transferred. (Note: at least one element must gain electrons and at least one element must lose electrons).

4). Reduce these numbers by dividing by a common denominator, if any, and cross multiply (see the examples to follow).

5). Balance the charge by adding the required number of H^+'s (if the reaction is in acid solution) or OH^-'s (if the reaction is in basic solution) to whichever side is necessary.

6). Balance the H's and O's by adding the required number of H_2O's to whichever side is necessary.

Example: Balance the following reaction which occurs in acidic solution:

$$Cr_2O_7^{2-} + I^- \rightarrow Cr^{3+} + I_2.$$

Step 1. Cr goes from +6 to +3; I goes from −1 to 0.

Step 2. Balance the Cr and I atoms;

$$Cr_2O_7^{2-} + 2\ I^- \rightarrow 2\ Cr^{3+} + I_2.$$

Step 3. To get the total oxidation number (TON), multiply the oxidation number per atom by the number of atoms (including the stoichiometric coefficients inserted in step 2).

TON(Cr) on left = 2 x 6 = 12
TON(Cr) on right = $\underline{2 \text{ x } 3 = \ 6}$
$\Delta e = +6$ (gain)

TON(I) on left = 2 x −1 = −2
TON(I) on right = $\underline{2 \text{ x } 0 = \ 0}$
$\Delta e = -2$ (loss)

Step 4. The above TON may be reduced to 3 and 1 by dividing by 2. Now cross multiply (i.e., multiply the stoichiometric coefficients in front of the substances containing Cr by 1 and those in front of the substances containing I by 3):

$$1\ Cr_2O_7^{2-} + 6\ I^- \rightarrow 2\ Cr^{3+} + 3\ I_2.$$

(Note: at this point, these coefficients are fixed and should not be changed in any subsequent operations).

Step 5. The total charge on the left side = −2 + −6 = −8. The total charge on the right side = +6. Therefore, add 14 H^+ on the left side (since the reaction takes place in acid solution). The charges are now balanced.

Step 6. The number of O's on the left side = 7. The number of O's on the right side = 0. Therefore, add 7 H_2O's on the right side to balance the O's. Check the H's: 14 on the left side and 14 on the right side. Hence, the reaction is now fully balanced!

$$14\ H^+ + Cr_2O_7^{2-} + 6\ I^- \rightarrow 2\ Cr^{3+} + 3\ I_2 + 7\ H_2O.$$

Example: Balance the following reaction which takes place in basic solution:

$$MnO_4^- + NO_2^- \rightarrow MnO_2 + NO_3^-.$$

Steps 1 - 4.

$$\begin{array}{lcccc}
 & \multicolumn{3}{c}{\overbrace{\hspace{10em}}^{+3e}} & \\
\text{TON's} \rightarrow & +7-8 & +3-4 & +4-4 & +5-6 \\
 & 2\,MnO_4^- + & 3\,NO_2^- \rightarrow & 2\,MnO_2 + & 3\,NO_3^- \\
\text{Ox.No./atom} \rightarrow & +7-2 & +3-2 & +4-2 & +5-2 \\
 & & \multicolumn{3}{c}{\underbrace{\hspace{10em}}_{-2e}}
\end{array}$$

Steps 5& 6.

$$H_2O + 2\,MnO_4^- + 3\,NO_2^- \rightarrow 2\,MnO_2 + 3\,NO_3^- + 2\,OH^-.$$

Exercises: Balance the following:

a). $CuS + NO_3^- \rightarrow Cu^{2+} + SO_4^{2-} + NO$ (in acid solution)

b). $C_6H_5CHO + Cr_2O_7^{2-} \rightarrow C_6H_5COOH + Cr^{3+}$ (in acid solution)

c). $Zn + NO_3^- \rightarrow NH_3 + Zn(OH)_4^{2-}$ (in basic solution)

A complication arises when the same element or compound is being both oxidized and reduced. The easiest way to attack this problem is to write the formula of the compound twice.

Exercises:

a). Balance the reaction in acid solution

$$HS_2O_3^- \rightarrow S + HSO_4^-$$

by rewriting the reaction as

$$HS_2O_3^- + HS_2O_3^- \rightarrow S + HSO_4^-.$$

b). Balance the reaction:

$$Hg_2Cl_2 + NH_3 \rightarrow Hg + HgNH_2Cl + NH_4^+ + C1^-.$$

Some reactions are easier to balance using the method of **half reactions**. One first identifies the species being oxidized and reduced and then writes a separate reaction for each process. The half reactions are then balanced individually by applying essentially the same procedure as outlined above and including the number of electrons involved explicitly as a reactant or product. In the final step, each half reaction is multiplied by the smallest integer that will make the electrons in one half reaction exactly cancel the electrons in the other.

Example: Balance the following reaction in basic solution using the method of half reactions:

$$MnO_4^- + NO_2^- \rightarrow MnO_2 + NO_3^- .$$

(Note: this is the same reaction that was worked in the example on the preceding page).

Mn is reduced and N is oxidized, so write these processes as separate (half) reactions;

$$\underset{+7}{MnO_4^-} \rightarrow \underset{+4}{MnO_2}$$

$$\underset{+3}{NO_2^-} \rightarrow \underset{+5}{NO_3^-} .$$

Now include the number of electrons involved by inserting them as a reactant (for the reduction) and a product (for the oxidation);

$$3e^- + MnO_4^- \rightarrow MnO_2$$

$$NO_2^- \rightarrow NO_3^- + 2e^- .$$

Next, balance the charge in each half reaction by adding the required number of hydroxide ions to the appropriate side;

$$3\ e^- + MnO_4^- \rightarrow MnO_2 + 4\ OH^-$$

$$2\ OH^- + NO_2^- \rightarrow NO_3^- + 2\ e^- .$$

Now, balance the H's and O's by adding H_2O's;

$$2\ H_2O + 3\ e^- + MnO_4^- \rightarrow MnO_2 + 4\ OH^-$$

$$2\ OH^- + NO_2^- \rightarrow NO_3^- + 2\ e^- + H_2O .$$

To make the electrons cancel, multiply the first half reaction by 2 and the second half reaction by 3 (i.e., cross multiply);

$$4\ H_2O + 6\ e^- + 2\ MnO_4^- \rightarrow 2\ MnO_2 + 8\ OH^-$$

$$6\ OH^- + 3\ NO_2^- \rightarrow 3\ NO_3^- + 6\ e^- + 3\ H_2O.$$

Finally, add the two half reactions together and cancel the e's, OH^-'s, and H_2O's on opposite sides to get

$$H_2O + 2\ MnO_4^- + 3\ NO_2^- \rightarrow 2\ MnO_2 + 3\ NO_3^- + 2\ OH^-.$$

2. Acid-Base Reactions

According to the **Arrhenius definition**, an acid is a substance that increases the H^+ concentration in a solution, while a base is a substance that increases the OH^- concentration in a solution. The following are examples of Arrhenius acids and bases:

$HCl + H_2O \rightarrow H_3O^+ + C1^-$ (A strong acid - dissociates completely)

$CH_3COOH + H_2O \rightarrow H_3O^+ + CH_3COO^-$ (A weak acid - dissociates very little)

$NaOH_{(aq)} \rightarrow Na^+ + OH^-$ (A strong base - dissociates completely)

There are two other definitions of an acid/base which are more general than the Arrhenius definition. One is due to Bronsted and Lowry, who defined an acid as a **proton donor** and a base as a **proton acceptor**. The following is an example of a Bronsted-Lowry acid/base pair:

$$\begin{array}{ccccccc} HCl & + & NH_3 & \rightarrow & NH_4^+ & + & C1^- \\ \text{acid} & + & \text{base} & = & \text{conjugate} & + & \text{conjugate} \\ & & & & \text{acid} & & \text{base} \end{array}$$

Arrhenius and Bronsted-Lowry acid/base reactions are classified as **protonic reactions** because they involve H^+.

The third definition of an acid/base was first proposed by G. N. Lewis and states that an acid is an **electron pair acceptor** and a base is an **electron pair donor**. This definition is much more general and includes reactions that do not involve either H^+ or OH^-. Examples of Lewis acid/base reactions are:

$$BF_3 + NH_3 \rightarrow BF_3NH_3$$

$$B_2O_3 + 3\ H_2O \rightarrow \underset{\text{(Boric acid)}}{2\ B(OH)_3}\ .$$

Acid/base **decomposition** reactions are much less spectacular than redox decomposition reactions and, of course, do not involve changes in oxidation states. A typical example of an acid/base decomposition reaction is

$$CaCO_{3(s)} \rightarrow CaO_{(s)} + CO_{2(g)}.$$

3. Metathesis Reactions

Metathesis reactions involve simple ion exchange. Generally, the driving force behind this kind of reaction is the favorable energy change associated with the formation of an insoluble product. Some examples are:

$$AgNO_{3(aq)} + NaCl_{(aq)} \rightarrow AgCl_{(s)} + NaNO_{3(aq)}$$

$$Na_2SO_{4(aq)} + Pb(NO_3)_{2(aq)} \rightarrow PbSO_{4(s)} + 2\ NaNO_{3(aq)}.$$

Solubilities form the basis for predicting the products of many metathesis reactions. You will find the following solubility rules useful for this purpose, and for devising laboratory tests to identify ions by qualitative analysis.

Solubility rules

1. All salts of Group IA and of NH_3 are soluble.
2. All salts containing nitrate (NO_3^-), acetate (CH_3COO^-) and chlorate (ClO_3^-) are soluble.
3. All chlorides, bromides and iodides are soluble except those of Ag^+, Pb^{2+}, and Hg_2^{2+}.
4. All sulfates are soluble except those of Ca^{2+}, Sr^{2+}, Ba^{2+}, Hg_2^{2+}, and Ag^+.
5. All metal oxides are insoluble except those of Group IA and of Ca^{2+}, Sr^{2+}, and Ba^{2+}.
6. All hydroxides are insoluble except those of Group IA, and of Sr^{2+} and Ba^{2+}.
7. All carbonates, phosphates, sulfides, and sulfites are insoluble except those of Group IA and of NH_4^+.

D. Net ionic reactions

Most soluble compounds containing the group 1A and IIA metals and/or the complex ions listed in table 2.1 exist in solution as ions. Because ionic solutions conduct electricity, such compounds are sometimes referred to as **electrolytes**. The reactions that occur between two electrolyte compounds in aqueous solution usually do not involve all of the ionic species present and the equations for the chemical reactions can be simplified by leaving out those species that merely act as spectators (i.e., species that do not directly participate in the reaction). As an example, consider the reaction between silver nitrate and hydrochloric acid;

$$AgNO_{3(aq)} + HCl_{(aq)} \rightarrow Ag^{+} + NO_3^{-} + H^{+} + Cl^{-} \rightarrow AgCl_{(s)} + H^{+} + NO_3^{-} .$$

In this reaction, the H^+ and NO_3^- ions do not participate in the reaction and therefore, they may be left out of the *net ionic reaction*:

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

The advantage of using net ionic reactions is that it greatly simplifies the problem of understanding exactly what chemical change takes place.

Exercise: Write net ionic reactions for each of the following:

a). Sodium sulfate + barium chloride

b). Silver chloride and ammonium sulfide

Review Questions

1. Give chemical formulas for each of the following compounds:

   a). Ammonium dichromate

   b). Aluminum bromide

   c). Sodium hypochlorite

   d). Silicon tetrafluoride

   e). Iron(II) hydrogen sulfite

2. Name the following compounds:

   a). $Ca(H_2PO_4)_2$

   b). $K_2C_2O_4$

   c). LiN_3

   d). $Sr(IO_3)_2$

   e). $HClO_4$

3. Balance the following oxidation-reduction reactions:

   a). $MnO_4^- + Cl^- \rightarrow Mn^{2+} + Cl_2$ (in acid solution)

   b). $V^{2+} + H_2O_2$ (hydrogen peroxide) $\rightarrow H_4VO_4^+ + H_2O$ (in acid solution)

   c). $Cr(OH)_3 + ClO^- \rightarrow CrO_4^{2-} + Cl^-$ (in basic solution)

   d). $C_2H_5OH + MnO_4^- \rightarrow Mn^{2+} + CH_3COOH$ (in acidic solution)

4. Write net ionic reactions for the following:

   a). Strontium carbonate + hydrochloric acid

   b). Acetic acid + water (acetic acid is a weak acid)

   c). acetic acid + potassium hydroxide

   d). Ammonia + water (ammonia is a weak base)

   e). Ammonia + sulfuric acid

   f). Aluminum hydroxide + nitric acid.

5. When 40.00 mL of 0.100 M $Pb(NO_3)_2$ are mixed with 60.00 mL of 0.300 M HCl, what net ionic reaction takes place? Calculate the molar concentrations of all ions remaining in solution after the reaction and the mass of any precipitate that has formed.

Answers:

5. $[NO_3^-] = 0.0800$ M, $[H^+] = 0.200$ M, $[Cl^-] = 0.100$ M, mass($PbCl_2$) = 1.11 g.

THINGS YOU SHOULD KNOW

How to determine oxidation numbers.

How to write chemical formulas for compounds,

How to name compounds.

Names and formulas of common complex ions.

How to balance oxidation-reduction reactions.

Solubility rules.

How to write net ionic reactions.

Definitions:

- oxidation and reduction
- Arrhenius acid/base
- Bronsted-Lowry acid/base
- Lewis acid/base

TIPOS DE REACCIONES

A) Reacciones de síntesis

$$A + B \longrightarrow C$$

B) Reacciones de descomposición

$$C \longrightarrow A + B$$

C) Reacciones de desplazamiento

$$A + BC \longrightarrow AC + B$$

C1 Desplazamiento de hidrógeno

$$2Na_{(s)} + 2H_2O_{(l)} \longrightarrow 2NaOH_{(ac)} + H_{2(g)}$$

$$Zn_{(s)} + 2HCl_{(ac)} \longrightarrow ZnCl_{2(ac)} + H_{2(g)}$$

C2 Desplazamiento de metal

$$V_2O_{5(s)} + 5Ca_{(l)} \longrightarrow 2V_{(l)} + 5CaO_{(s)}$$

C3 Desplazamiento de halógeno

$$Cl_{2(g)} + 2KBr_{(ac)} \longrightarrow 2KCl_{(ac)} + Br_{2(l)}$$

D) Reacciones de precipitación

$$Pb(NO_3)_{2(ac)} + NaI_{(ac)} \longrightarrow PbI_{2(s)} + NaNO_{3(ac)}$$

E) Reacciones de neutralización

$$H^+ + OH^- \longrightarrow H_2O$$

$$HCl + NaOH \longrightarrow NaCl + H_2O$$

Problems:
Ch. 17: 18, 20, 32ab, 40, 42, 44, 46, 48

Chemical Reactions: Part Two

Net Ionic Equations and Redox Reactions

Topics

- Total ionic and net ionic equations
- Oxidation numbers
- Redox reactions

Looking Back

- Consider the following reaction:

$NaCl_{(aq)} + KNO_{3(aq)} \rightarrow$

- Since both the products of this reaction are aqueous, and since no molecular compound is formed, we say that there is no reaction.
- However, since we did bring these two solutions together, why is it incorrect to say that the products are $KCl_{(aq)}$ and $NaNO_{3(aq)}$?
- Earlier, we defined a chemical change as one that involves transforming one or more chemical compounds into one or more different products.
- By combining two solutions which do not make an insoluble product or a molecular compound, we fail to create a new product, as we shall now prove.

Total Ionic Equations

- Total molecular equations show all species in solution for a given reaction.
  - The term <u>species</u> refers to the specific ions and/or compounds in the solution.
- Examples of species:

For $NaCl_{(aq)}$, $Na^{+}_{(aq)}$ and $Cl^{-}_{(aq)}$.
For $HNO_{3(aq)}$, $H^{+}_{(aq)}$ and $NO_3^{-}{}_{(aq)}$.
For $H_3PO_{4(aq)}$, it remains $H_3PO_{4(aq)}$.

- Some simple guidelines for writing down the species.
  - <u>Strong</u> electrolytes, including strong acids and bases, dissociate into their ions.
  - <u>Weak</u> electrolytes, including weak acids and bases, generally do not dissociate, so the species is just the original compound.
  - <u>Nonelectrolytes</u> do not dissociate, so the species is again the same as the original compound.

Example:

$NaCl_{(aq)} + AgNO_{3(aq)} \rightarrow AgCl_{(s)} + NaNO_{3(aq)}$

Total Molecular Equation:

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

Spectator Ions

- A *spectator* of sports is someone who watches the game from the sidelines, but does not participate.
- Similarly, in chemical reactions, spectator ions "hang out" in a solution but do not actively participate in the reaction itself.
  - In other words, any ion which is both on the reactants and products side of a reaction is a spectator ion, for it has not undergone a chemical change.
- The ions' main purpose is to maintain constant charge in the solution.

Net Ionic Equations

- Net ionic equations only show those chemicals which participate in the reaction. Spectator ions are not included.
- To write a net ionic equation, first write down the total ionic equation.
- Then, cancel anything which appears *identically* on both sides of the reaction.

Net Ionic Equations

- Consider the total ionic equation

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

- Now, factor out the spectator ions

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

- The net ionic equation is left over.

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

Examples

Write chemical equations, total ionic equations, and net ionic equations for the following:

a. a solution of barium chloride reacts with a sodium sulfate solution.

b. sodium hydroxide is neutralized by hydrochloric acid.

c. phosphoric acid is reacted with potassium hydroxide.

d. solutions of sodium nitrate and ammonium chloride are brought together.

e. a piece of sodium metal is dropped into cold water.

Oxidation Numbers

- Oxidation numbers are assigned to atoms to keep track of electrons in redox reactions.
- While an oxidation number is not always the same as charge, it gives us an idea as to where the electrons tend to accumulate in a molecule.
- An element's oxidation number changes in a redox reaction; generally one element is oxidized and one reduced.

General Rules for Assigning Oxidation Numbers

- The oxidation number of an atom in an element (i.e. not a compound) is 0.
  Examples:
- The oxidation number of a monatomic ion is the same as its charge.
  Examples:
- The oxidation number of fluorine in any compound is always -1.
- The oxidation number of oxygen in most compounds is -2; it is -1 in peroxides.
  - Note that other, less common oxidation numbers are possible for oxygen.
- The oxidation number of hydrogen is +1 in most compounds. It is -1 in hydrides, where hydrogen is combined with one metal (examples: NaH, LiH, CaH_2, etc.)

- The sum of all oxidation numbers must equal the total charge on the compound.
- Oxidation numbers for atoms not described in detail by these rules are usually deduced from the oxidation numbers of other atoms.

Examples

Calculate the oxidation numbers for each atom in the following compounds/ions.

O_2 H_2O

NaCl HSO_4^-

$KHCO_3$ ClO_4^-

Redox Reactions:
More Complicated Cases

- We have already seen redox reactions as they occur in single displacement reactions.
- With the concept of oxidation numbers, we can study redox reactions involving a diverse range of compounds, including molecular compounds (instead of just ionic ones).

Balancing Redox Reactions ***(Note: This topic may __not__ be covered)***

- Most redox reactions are far too complex to balance in the fashion we have studied so far.
- We must use a new method in order to balance these, involving a series of detailed steps to take us toward our goal.
- The steps for balancing these reactions in an acidic environment are slightly different from those for a basic one.

Balancing Redox Reactions

1. Identify which atoms are being oxidized and reduced in the reaction.

2. Separate the equation into two “half reactions”, one containing only the oxidation reaction, the other containing the reduction (do the steps 3-8 for each half reaction).

3. Balance the atoms of the element being oxidized or reduced for each equation.

4. Balance any other atoms *except* oxygen and hydrogen.

5. Balance oxygen by adding enough water molecules to the side with the least number of oxygen atoms to make the two sides balance.

6. Balance hydrogen by adding enough H^+ ions to the side with the least number of hydrogen atoms to make the two sides balance.

7. <u>If the reaction is under basic conditions:</u> You are not allowed to have the acidic H^+ ions in your final answer! Add as many OH^- ions to *both* sides as are necessary to react out the H^+ ions.
   Remember that $H^+ + OH^- \rightarrow H_2O$*!!!*

8. Add up the total charge for each side of the reaction. Then, add enough electrons (e^-) to the side with the more positive charge to make it equal the charge of the opposite side.

9. In order to add your two half-reactions together, each must contain the same number of electrons. Multiply each *entire* reaction by the lowest possible factor to make each reaction have the same number of electrons.

10. Combine the two half-reactions into a single reaction.

11. Make sure that the reaction is balanced with the lowest common multiples of the coefficients.

Examples

Balance each of the following redox reactions:

$Fe^{2+} + NO_3^- \rightarrow Fe^{3+} + NO$ (in acid)

$Fe^{2+} + CrO_4^{2-} \rightarrow Fe^{3+} + Cr^{3+}$ (in base)

$Cr^{3+} + MnO_4^{-} \rightarrow Cr_2O_7^{2-} + Mn^{2+}$ (in acid)

The Value of Redox Reactions

- Electric currents produced by redox reactions can be harnessed in many practical reactions.
- The driving force behind a battery is a redox reaction.
- Redox reactions can sometimes be driven backwards, such as in the electroplating of metals. This is how chrome is plated onto a car's bumper, or silver onto a tarnished piece of silverware.

V. SOLUCIONES.

5.1 Naturaleza de soluciones (Tipos de soluciones).

SOLUTO (NaCl, $NaNO_3$, H_2SO_4): Sustancia disuelta en un SOLVENTE (H_2O, NH_3, HF, etc.)

SOLVENTE: Sustancia en la cual se disuelve el SOLUTO

SOLUCION: ***Solvente + Soluto***

¿CUÁNDO SE CONSIDERA UNA SUSTANCIA SOLUTO O SOLVENTE?: *Cuando la cantidad Relativa de una sustancia en una Solución es mucho más GRANDE, la sustancia presente en MAYOR cantidad se considera el SOLVENTE.*

Disolución	
SATURADA	*Contiene la máxima cantidad de un soluto que se disuelve en un disolvente en particular a una temperatura especifica*
NO SATURADA	*Contiene menos cantidad de soluto que la que puede disolver.*
SOBRESATURADA	*Contiene más soluto que el que puede haber en una disolución sobresaturada.*

Soluciones sobresaturadas....
NO MUY ESTABLES

→

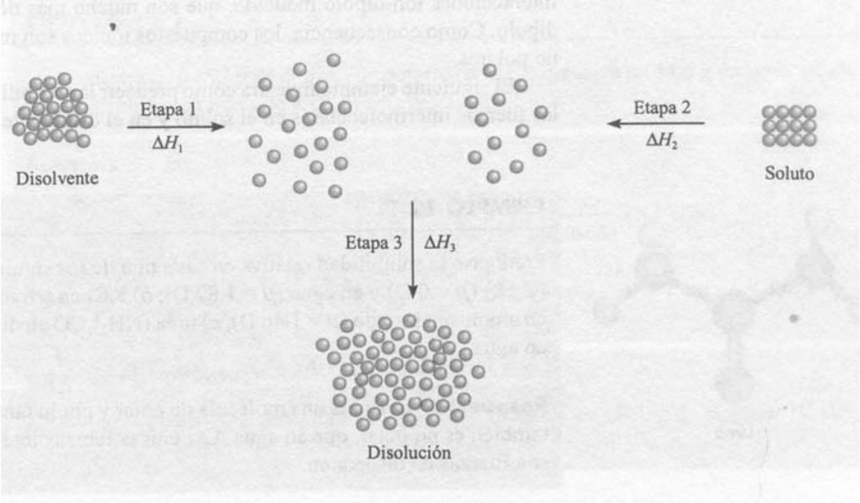
Una parte del soluto...
SE CRISTALIZA

CRISTALIZACIÓN: *Proceso en el cual un SOLUTO disuelto se separa de la disolución y forma cristales.*

5.2 Procesos de disolución en las soluciones.

- La facilidad con que una partícula de solutos sustituye a una molécula de disolvente depende de la fuerza relativa de 3 tipos de interacciones:
  - Interacción disolvente – disolvente
  - Interacción soluto – soluto
  - Interacción disolvente – soluto

El proceso de disolución puede simplificarse en 3 etapas:



$$\Delta H_{\text{disolución}} = \Delta H_1 + \Delta H_2 + \Delta H_3$$

- El proceso de disolución puede ser controlado por 2 factores:

1) Factor Energético: Determina si el proceso será exotérmico o Endotérmico.
2) Se refiere a la tendencia hacia el desorden inherente a todos los procesos naturales.

⇩

Al mezclarse el soluto y el solvente, la entropía (***S***) aumenta.

Solubilidad: *Es una medida de la cantidad de soluto que se disolverá en un cierto disolvente a una T especifica.*

- 2 sustancias cuyas *Fuerzas Intermoleculares* son del mismo TIPO y MAGNITUD son solubles entre sí.
  - Ejemplo: CCl_4 y BENCENO: sustancias NO Polares:

  - Las únicas fuerzas Presentes son de dispersión
  - Cuando se mezclan rápido se disuelven porque las FUERZAS DE ATRACCIÓN entre moléculas son similares.

- Los compuestos que son:

$$\text{Iónicos} \xrightarrow{\text{DISUELVEN}} \text{Solventes Polares}$$
$$\text{Polares} \xrightarrow{\text{DISUELVEN}} \text{Solventes Polares}$$
$$\text{No Polares} \xrightarrow{\text{DISUELVEN}} \text{Solventes No Polares}$$

Compuestos Solubles	Excepciones
Compuestos que contengan iones de metales alcalinos (Li^+, Na^+, K^+, Rb^+, Cs^+) y el ion amonio (NH_4^+). Nitratos (NO_3^-), bicarbonatos (HCO_3^-) y cloratos (ClO_3^-)	
Haluros (Cl^-, Br^-, I^-)	Haluros de Ag^+, Hg_2^{2+} y Pb^{2+}
Sulfatos (SO_4^{2-})	Sulfatos de Ag^+, Ca^{2+}, Sr^{2+}, Ba^{2+} y Pb^{2+}.
Compuestos Insolubles	**Excepciones**
Carbonatos (CO_3^{2-}), fosfatos (PO_4^{3-}), cromatos (CrO_4^{2-}), sulfuros (S^{2-})	Compuestos que contengan iones de metales alcalinos y el ion amonio
Hidróxidos (OH^-)	Compuestos que contengan iones de metales alcalinos y el ion Ba^{2+}.

- Cuando se disuelve un compuesto iónico (NaCl) en un solvente Polar (H_2O), éste se estabiliza por la formación de INTERACCIONES Ion –Dipolo. (Estabilización de iones por Solvatación).
  - **Solvatación**: *Proceso mediante el cual un ion o una molécula es rodeada por moléculas de disolvente, distribuidas en forma especifica.*

- Ya que los solventes No Polares no presentan momento dipolo, estos no pueden SOLVATAR a los iones

- Los compuestos iónicos serán más solubles en disolventes Polares: H_2O, $NH_{3(l)}$, $HF_{(l)}$, que en disolventes No Polares: Benceno, CCl_4, debido a que carecen de momento dipolo.

5.3 Soluciones entre líquido y sólido

- La Habilidad de un sólido para disolverse depende de su Energía de Red Cristalina.

- Energía de Red Cristalina: cambio energético que acompaña a la formación de 1 mol de unidades fórmula en el estado cristalino a partir de sus moléculas constituyentes en el estado gaseoso (Proceso de corte Exotérmico).

Depende de la Fuerza de atracción entre los iones en el sólido.

$$M^{+}_{(g)} + X^{-}_{(g)} \rightarrow MX_{(s)} + \text{Energía}$$

- Para el proceso contrario:

$$MX_{(s)} + \text{Energía} \rightarrow M^{+}_{(g)} + X^{-}_{(g)}$$

Este es el primer paso en la disolución de un sólido (hipotético): Proceso endotérmico

- Cuando menores sean las interacciones soluto-soluto (Energía de red cristalina), menor será la Energía de disolución.

- Si el disolvente es agua, la energía necesaria para expandir el disolvente incluye la requerida para romper algunos de los enlaces de hidrógeno entre las moléculas de agua.

- El tercer factor que contribuye al calor de disolución es la extensión en que las moléculas del disolvente interaccionan con las partículas del sólido (Solvatación: Cambio Energético implicado en la hidratación de un mol de iones gaseosos):

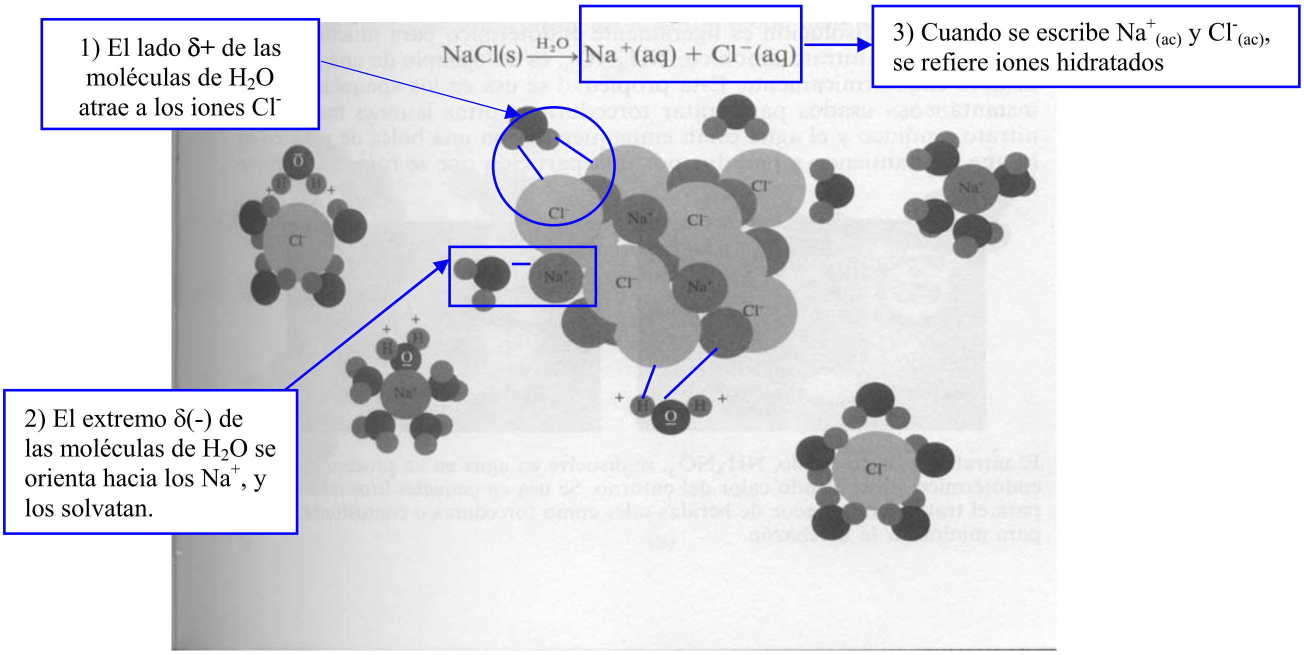
$M^{n+}_{(g)} + \boldsymbol{x}\, H_2O \rightarrow M(H_2O)^{n+}_{\boldsymbol{x}} + \text{Energía}$ (para el catión)

$X^{y-}_{(g)} + \boldsymbol{r}\, H_2O \rightarrow X(H_2O)^{y-}_{\boldsymbol{r}} + \text{Energía}$ (para el anión)

- El calor de disolución total para un sólido que se disuelve en un líquido, es:

$$\Delta H_{\text{disolución}} = (\text{calor de Solvatación}) - (\text{energía de red cristalina})$$

- Ejemplo de la disolución de un sólido: Cuando un trozo de NaCl (sólido iónico) se introduce en Agua.



1) El lado δ+ de las moléculas de H_2O atrae a los iones Cl^{-}

2) El extremo δ(-) de las moléculas de H_2O se orienta hacia los Na^{+}, y los solvatan.

3) Cuando se escribe $Na^{+}_{(ac)}$ y $Cl^{-}_{(ac)}$, se refiere iones hidratados

- La ***Hidratación*** y los ***efectos de las atracciones*** en el cristal se oponen entre sí en el proceso de ***DISOLUCIÓN***. Las Energías de hidratación y las energías de red son normalmente de la misma magnitud para especies de carga baja, de modo que a menudo se anulan unas a otras. Como resultado, el proceso de disolución es ligeramente endotérmico

- Algunos sólidos iónicos se disuelven con desprendimiento de calor: Na_2SO_4 anhidro, $Ca(CH_3COO)_2$, $CaCl_2$, $LiSO_4{\cdot}H_2O$.

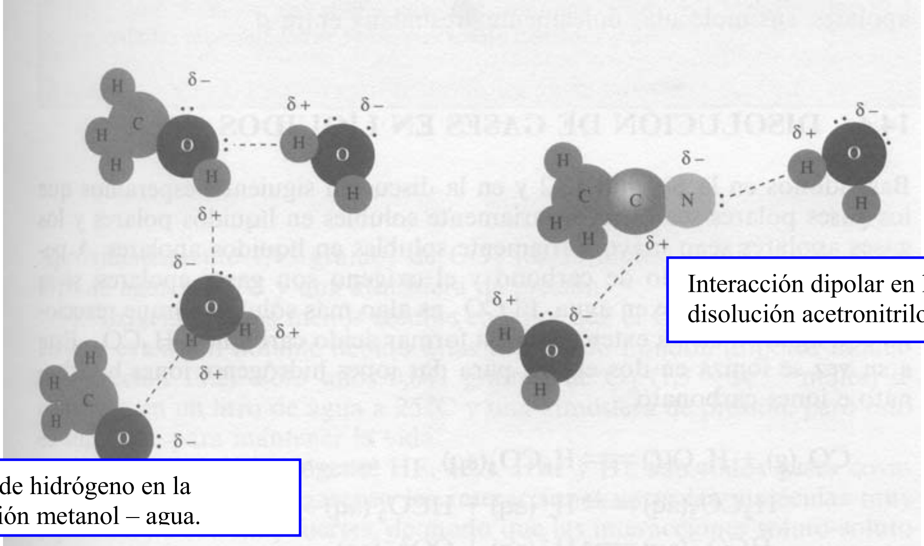
5.4 Soluciones de Líquidos en Líquidos (Miscibilidad).

- La Miscibilidad: es la habilidad de un líquido para disolverse en otro. Los 3 tipos de interacciones son:
  - Disolvente – Soluto
  - Disolvente – Disolvente
  - Soluto – Soluto (interacciones mucho menores)

- Los Líquidos Polares tienden a interaccionar y a disolverse en otros líquidos polares, ejemplo:
  - *Metanol* (CH_3OH)
  - *Etanol* (CH_3CH_2OH)
  - *Acetronitrilo* (CH_3CN)
  - *Ácido Sulfúrico* (H_2SO_4)

→ Agua (H_2O)

- Ejemplo: Interacciones entre:



Interacción dipolar en la disolución acetronitrilo-agua.

Enlace de hidrógeno en la disolución metanol – agua.

¿Porqué se desprende calor cuando se mezclan H_2SO_4 y H_2O?:

5.5 Concentración de las soluciones

5.5.1 Concentraciones expresadas en Unidades Físicas.

a) Masa del soluto por unidad de volumen (g/L): 20 g de KCl por litro de solución: 20 g/L.
b) Concentración en porcentaje: El % en peso de B,

$$\% \text{ en peso de B} = \frac{m_B}{m_A + m_B} x100 = \frac{\text{gramos de soluto}}{\text{gramos de solucion}} x100$$

GRAMOS DE SOLUCIÓN = gramos de **soluto** + gramos de **solvente**

- Si la concentración de una solución se da simplemente en porcentaje, se considera Porcentaje en Peso (% w). También puede expresarse en Porcentaje en Volumen (% *v*).

- Ejemplo: 100 g de una solución de NaCl al 20 %w: $\left\{\begin{array}{l} 20 \text{ g de NaCl y} \\ 80 \text{ g de } H_2O \end{array}\right.$

5.5.2 Concentraciones expresadas en Unidades Químicas.

a) Fracción molar: La fracción molar de B (**X_B**) es:

$$X_B = \frac{n_B}{n_A + n_B} = \frac{\textit{mol de soluto}}{\textit{mol totales}}$$

$$\mathbf{X_A + X_B = 1.0}$$

b) Molaridad (***M***):

$$M = \frac{g_{\text{soluto}}}{PM_{\text{soluto}} V(L)_{\text{solución}}}$$

c) Molalidad (***m***):

$$m = \frac{\text{mol}_{\text{soluto}}}{\text{kg}_{\text{solvente}}}$$

d) Normalidad (***N***):

$$N = \frac{\text{equivalente} - \text{gramo de soluto}}{L_{\text{solucion}}}$$

- Equivalente-gramo de ácidos y bases:

$$\text{Peso} - \text{equivalente}_{\text{acido}} = \frac{PM_{\text{acido}}}{\#\ \text{de}\ H^+\ \text{producidos}}$$

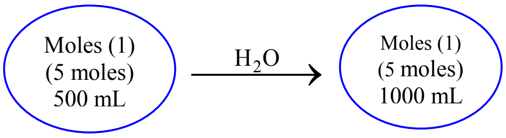
$$\text{Peso} - \text{equivalente}_{\text{base}} = \frac{PM_{\text{base}}}{\#\ \text{de}\ OH^-\ \text{producidos}}$$

5.5.3 Dilución.

- En la práctica de laboratorio con frecuencia se usan soluciones concentradas de concentración conocida para preparar a partir de ellas soluciones diluidas.
- Cuando la concentración se expresa en una escala volumétrica, la cantidad de soluto contenido en un volumen determinado de la solución es igual al producto del volumen por la concentración.

$$V_1C_1 = V_2C_2$$

V = Volumen
C = Concentración



Cambia la concentración, pero el # de moles contenidos es el mismo (Proceso Físico)

5.6 Problemas de de unidades de concentración.

1. Cuando se evaporan 50 g de una solución de Sulfato de sodio ($Na_2SO_{4(ac)}$) hasta completar sequedad, se producen 20 g de sal, ¿cuál es el porcentaje de sal en solución?

2. ¿Cuántos gramos de agua deberán usarse para disolver 150 g de NaCl (cloruro de sodio), para producir una solución al 20% en peso).

3. Calcular la fracción molar de ácido sulfúrico (H_2SO_4) en 100 g de solución al 20 % en peso.

4. Se disuelven 25 g de metanol (CH_3OH) en 50 g de agua, calcular la fracción molar del metanol del agua en la solución.

5. ¿Cuál es la molalidad (*m*) de una solución que se prepara disolviendo 29.22 g de NaCl en 100 mL de H_2O.

6. ¿Cuál es la molalidad de una disolución que contiene 40 g de azúcar ($C_{12}H_{22}O_{11}$) disueltos en 150 g de agua?

7. Una disolución de alcohol etílico (C_2H_5OH) en agua es 1.6 *m*, ¿Cuántos gramos de alcohol están disueltos en 2000 g de H_2O?.

8. ¿Cuántos gramos de NaCl se necesitan para preparar 2000 mL de solución =.2 M?

9. ¿Cuál es la molaridad de una solución al 40% de H_2SO_4 si la densidad es de 1.19 g/mL.?

10. Calcular (a) la molaridad y (b) la molalidad de una solución de H_2SO_4 de peso específico 1.1 que contiene 25% de H_2SO_4 en peso.

11. ¿Cuántos equivalentes-gramo de HCl están contenidos en: (a) 2 L de solución 1 *N*, (b) 2 L de solución 0.5 *N*, (c) 0.2 L de solución 0.5 *N* .

12. Un frasco de laboratorio tiene escrito un rótulo con 10 *M* NaOH. ¿Cuántos mL de esta solución se necesitan para preparar 50 mL de una solución 2 *M* de NaOH?

4.1 EQUILIBRIO QUÍMICO

4.4.1 Conceptos Básicos.

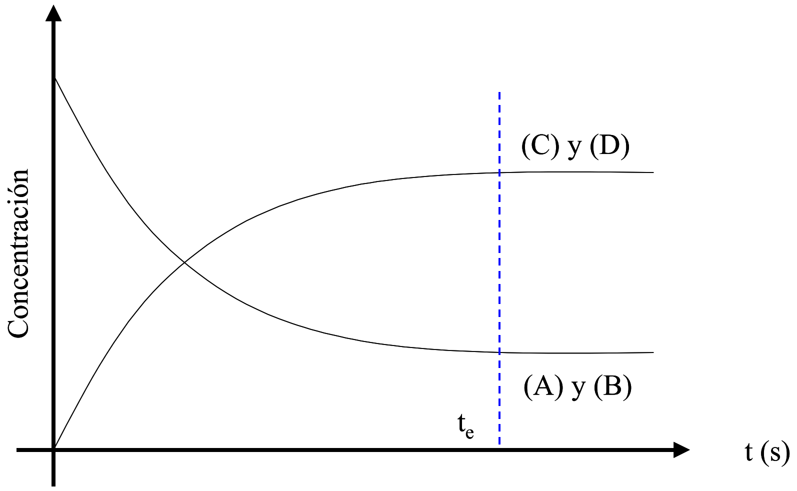
1) La mayoría de las Reacciones no se llevan a cabo completamente.
2) Las reacciones que no se completan del todo y pueden producirse en ambas direcciones se denominan REACCIONES REVERSIBLES.
3) Una reacción reversible puede ser:

$$aA + bB \boxed{\leftrightarrows} cC + dD$$

Indica Reversibilidad

[A], **[B]**, **[C]**, **[D]:** Concentraciones molares [mol/L] de los reaccionantes **A**, **B** y así como de los productos **C** y **D.**.
a, ***b***, ***c***, ***d***, : Representan los coeficientes estequiométricos, resultado del balanceo de la reacción.

4) El Equilibrio Químico se logra cuando 2 reacciones opuestas ocurren simultáneamente con la misma velocidad (Equilibrios Dinámicos).



Ejemplo: Para la reacción reversible entre **SO_2** (dióxido de azufre) y **O_2** (oxígeno) para formar **SO_3** (trióxido de azufre).

$$2\,SO_{2(g)} + O_{2(g)} \leftrightarrows 2\,SO_{3(g)}$$

Nota: Siempre hay que iniciar balanceando la ecuación.

- Se introducen 0.4 mol de $SO_{2(g)}$ y 0.2 mol de $O_{2(g)}$ en un recipiente de 1 L.

	$2\,SO_{2(g)}$	+	$O_{2(g)}$	$\leftrightarrows$	$2\,SO_{3(g)}$
Concentración Inicial	0.4 *M*		0.2 *M*		0

- Cuando se alcanza el EQUILIBRIO: se han formado 0.056 mol de $SO_{3(g)}$ y <u>han reaccionado</u>: 0.056 mol de $SO_{2(g)}$ y 0.028 mol de $O_{2(g)}$.

	$2SO_{2(g)}$	+	$O_{2(g)}$	$\leftrightarrows$	$2SO_{3(g)}$
Concentración Inicial	0.4 *M*		0.2 *M*		0
Cambios en la concentraciones debido a Reacción	-0.056 *M*		-0.028 *M*		+0.056 *M*
Concentración de Equilibrio	0.344 *M*		0.172 *M*		0.056 *M*

Ejemplo 2: Otro experimento se introdujo sólo 0.4 mol de $SO_{3(g)}$ (producto), en otro recipiente de 1 L.

	$2\ SO_{2(g)}$	+	$O_{2(g)}$	⇆	$2\ SO_{3(g)}$
Concentración Inicial	0		0		0.4 *M*

- Cuando se alcanza el EQUILIBRIO: se han formado 0.344 mol de $SO_{2(g)}$ y 0.172 mol de $O_{2(g)}$, quedando sin reaccionar 0.056 mol de $SO_{3(g)}$.

	$2\ SO_{2(g)}$	+	$O_{2(g)}$	⇆	$2\ SO_{3(g)}$
Concentración Inicial	0		0		0.4 *M*
Cambios en las concentraciones debido a Reacción	+0.344 *M*		+0.172 *M*		-0.344 *M*
Concentración de Equilibrio	0.344 *M*		0.172 *M*		0.056 *M*

4.4.2 Constante de Equilibrio.

Para la reacción: $2\ A + B \leftrightarrows A_2B$ **(1 SOLO PASO)**

→

- La Velocidad de reacción directa (de IZQUIERDA a DERECHA), es:

  $$\text{Velocidad}_D = \boldsymbol{k_d}\,[A]^2\,[B]$$

←

- Y la Velocidad de Reacción Inversa (de DERECHA a IZQUIERDA), es:

  $$\text{Velocidad}_I = \boldsymbol{k_i}\,[A_2B]$$

- Las dos velocidades son idénticas en el equilibrio: $\text{Velocidad}_D = \text{Velocidad}_I$

  $$\boldsymbol{k_d}\,[A]^2\,[B] = \boldsymbol{k_i}\,[A_2B]$$

- Dividiendo entre: $\boldsymbol{k_i}\,[A]^2[B]$, los 2 miembros de la ecuación:

  $$\frac{k_d}{k_i} = \frac{[A_2B]}{[A]^2[B]}$$

- @ una T determinada $\boldsymbol{k_d}$ y $\boldsymbol{k_i}$ son constantes:

  $$K_C = \frac{[A_2B]}{[A]^2[B]} \qquad \text{(en el equilibrio)}$$

- Para una reacción que implica 2 mecanismos (2 PASOS)

Paso(1) $$2A \overset{k_{1,D}}{\underset{k_{1,I}}{\leftrightarrow}} A_2 \qquad K_1 = \frac{k_{1,D}}{k_{1,I}} = \frac{[A_2]}{[A]^2}$$

Paso (2) $$A_2 + B \overset{k_{2,D}}{\underset{k_{2,I}}{\leftrightarrow}} A_2B \qquad K_2 = \frac{k_{2,D}}{k_{2,I}} = \frac{[A_2B]}{[A_2][B]}$$

- Si se multiplica K_1 ***x*** K_2:

$$K_1 x K_2 = \frac{[A_2]}{[A]^2} x \frac{[A_2B]}{[A_2][B]} = \frac{[A_2B]}{[A]^2[B]} = K_C$$

Independientemente del mecanismo de las reacciones, en términos generales, para una reacción, la constante de equilibrio es:

$$aA + bB \leftrightarrows cC + dD$$

$$K_C = \frac{[D]^d [C]^c}{[A]^a [B]^b}$$

KK_C: constante de equilibrio de la reacción.

[A], **[B]**, **[C]**, **[D]**: Concentraciones molares **EN EQUILIBRIO** [mol/L] de los reaccionantes **A**, **B** así como de los productos **C** y **D.**.

a, ***b***, ***c***, ***d***: Representan los coeficientes estequiométricos, resultado del balanceo de la reacción.

a) K_C, no presenta unidades
b) Sólo varía con la temperatura
c) Es constante a una temperatura dada
d) Es independiente de las concentraciones iniciales

4.4.3 Cociente de Reacción (*Q*).

Para una reacción:

$$\boldsymbol{a}\,A + \boldsymbol{b}\,B \leftrightarrows \boldsymbol{c}\,C + \boldsymbol{d}\,D$$

El cociente de reacción será : $Q = \frac{[D]_0^d [C]_0^c}{[A]_0^a [B]_0^b}$

$[A]_0$, $[B]_0$, $[C]_0$, $[D]_0$: Concentraciones molares **INICIALES** [mol/L] de los reaccionantes **A**, **B** así como de los productos **C** y **D.**.

a, ***b***, ***c***, ***d***: representan los coeficientes estequiométricos.

- El Cociente de Reacción (***Q***) sirve para predecir la dirección en la que procederá una mezcla de reacción para lograr el equilibrio. Y se comparan los valores de ***Q*** y K_C, habiendo tres posibilidades:

  - $Q > K_C$: Para alcanzar el equilibrio, los productos deben transformarse en reactivos. La reacción INVERSA predomina hasta alcanzar el equilibrio (de *derecha* a *izquierda*).

  - $Q = K_C$: El sistema está en equilibrio.

  - $Q < K_C$: : Para alcanzar el equilibrio, los reactivos deben transformarse en productos. La reacción DIRECTA predomina hasta alcanzar el equilibrio (de *izquierda* a *derecha*).

4.4.4 Uso de la constante de equilibrio.

Cálculo de las concentraciones de equilibrio:

- La ecuación para la siguiente reacción y el valor de K_C a una determinada temperatura son las siguientes. Una mezcla de equilibrio en un recipiente de 1.0 L contiene 0.25 mol de PCl_5 y 0.16 mol de PCl_3. ¿Qué concentración de equilibrio de Cl_2 debe haber presente?

$$PCl_{3(g)} + Cl_{2(g)} \leftrightarrows PCl_{5(g)} \qquad K_C = 1.9$$

- Para la reacción descrita por la ecuación:

$$N_{2(g)} + C_2H_{2(g)} \leftrightarrows 2\,HCN_{(g)}$$

K_C= 2.3x10^{-4} @ 300°C. ¿Cuál es la concentración de equilibrio del cianuro de hidrógeno si las concentraciones iniciales de N_2 y acetileno (C_2H_5) eran de 2.5 mol/L y 1.0 mol/L, respectivamente.

- La constante de equilibrio para la reacción:

$$\mathrm{Br}_{2(g)} + \mathrm{F}_{2(g)} \leftrightarrows 2\ \mathrm{BrF}_{(g)}$$

Es 54.7. ¿Cuáles son las concentraciones de equilibrio de todos estos gases si las concentraciones iniciales de bromo y de flúor eran ambas de 0.250 mol/L.

- El pentacloruro de antimonio se descompone según una reacción en fase gaseosa @ 448°C

$$\mathrm{SbCl}_{5(g)} \leftrightarrows \mathrm{SbCl}_{3(g)} + \mathrm{Cl}_{2(g)}$$

Una mezcla en equilibrio en un recipiente de 5.0 L se sabe que contiene 3.84 g de $SbCl_5$, 9.14 g de $SbCl_3$ y 2.84 g de Cl_2. Evaluar K_C @ 448°C.

4.4.5 Factores que afectan el equilibrio.

- Una vez que el sistema ha alcanzado el equilibrio, permanece en equilibrio, HASTA que es perturbado por algún cambio en las condiciones (principio de Le Chatelier).
- El cociente de Reacción (***Q***), nos ayuda a predecir el sentido de esta respuesta.

a) CAMBIOS EN LA CONCENTRACIÓN.

Para una reacción:

$$a\,\mathrm{A} + b\,\mathrm{B} \leftrightarrow c\,\mathrm{C} + d\,\mathrm{D} \qquad\qquad K_C = \frac{[\mathrm{D}]^d[\mathrm{C}]^c}{[\mathrm{A}]^a[\mathrm{B}]^b}$$

- Si se ***añade*** más REACTIVO o PRODUCTO al sistema, los cambios se aminorizan desplazando el equilibrio en la dirección que consume algo de sustancia añadida.
  - Cuando se agrega más A o B: La reacción directa prosigue en mayor extensión que la reacción inversa: $Q < K_C$.
  - Si se añade más C o D: La reacción inversa ocurre en mayor extensión, hasta que se restablece el equilibrio: $Q > K_C$.
- Si un PRODUCTO o REACTIVO se ***retira*** de un sistema en equilibrio: La reacción que produce esa sustancia tiene lugar en mayor extensión que su inversa.
  - Si se retira algo de C o D: La reacción ***directa*** está favorecida
  - Si se retira algo de A o B, se favorece la reacción ***inversa***.

Cambio	***Q***	Dirección del cambio de **A + B ⇆ C + D**
↑ [A] y [B]	$Q < K$	→ derecha
↑ [C] o [D]	$Q > K$	izquierda ←
↓ [A] o [B]	$Q > K$	Izquierda ←
↓ [C] o [D]	$Q < K$	→ derecha

Ejemplo:

- Se mezclan algo de hidrógeno y de yodo @ 229°C en un recipiente de 1.0 L . Cuando se establece el equilibrio, hay presentes las siguientes concentraciones: [HI] = 0.49 *M*, $[H_2]$ = 0.08 *M* e $[I_2]$ = 0.06 *M*. Si se añaden 0.3 mol adicionales de HI, ¿qué concentraciones habrá presentes cuando se establezca el nuevo equilibrio?. K_C= 30

b) CAMBIOS EN VOLUMEN Y PRESIÓN.

- Los cambios en volumen y presión tienen poco efecto en las concentraciones de sólidos o líquidos porque son ligeramente compresibles.

- Para gases si tienen gran impacto. Para un gas ideal

$$PV = nRT \text{ o } P = \left(\frac{n}{V}\right)\left(\frac{R}{T}\right)$$

  - Si el Volumen del gas aumenta, tanto su presión parcial como su concentración disminuyen. Para la siguiente reacción @ temperatura constante:

$$A_{(g)} \leftrightarrows 2\, D_{(g)} \qquad K = \frac{[D]^2}{[A]}$$

- @ T= cte., un ↓ en el volumen (aumento de presión) , aumenta las concentraciones de A y D. Así, ***Q*** > ***K***, y este equilibrio se desplaza a la izquierda.

- Por el contrario, un ↑ del volumen (descenso de presión) desplaza esta reacción a la derecha hasta que se reestablece el equilibrio, porque ***Q*** < ***K***.

Cambio	***Q****	Dirección del cambio de **$A_{(g)} \leftrightarrow 2D_{(g)}$**
Descenso de volumen, aumento de presión	Q >K	Hacia un menor número de moles de gas (izquierda para esta reacción)
Aumento de volumen, descenso de presión	Q < K	Hacia un mayor número de moles de gas (derecha para esta reacción)

Ejemplo:

- @ 22°C la constante de equilibrio, *Kc*, para la siguiente reacción es $4.66x10^{-3}$. (a) Si 0.8 mol de N_2O_4 se inyectan en un recipiente cerrado de 1.0 L @ 22°C, ¿cuántos mol de cada gas estarán presentes en el equilibrio?, (b) Si el volumen disminuyese a la mitad (0.5 L) a temperatura constante, ¿cuántos mol de cada gas habría presentes después de que se establezca el nuevo equilibrio?

$$N_2O_{4(g)} \leftrightarrows 2\, NO_{2(g)} \qquad K_c = \frac{[NO_2]^2}{[N_2O_4]} = 4.66x10^{-3}$$

c) CAMBIOS DE TEMPERATURA.

Para la reacción exotérmica

$$A + B \leftrightarrows C + D + \textit{calor}$$

- Se produce calor en la reacción directa (exotérmica). Se ↑ la T @ presión cte. por adición de calor. Esto favorece el sistema hacia la izquierda

- ↓ la temperatura favorece la reacción hacia la derecha a medida que el sistema repone el calor que se eliminó

Para la reacción endotérmica

$$\textit{calor} + A + B \leftrightarrows C + D$$

- Un ↑ de temperatura a@ presión cte. favorece la reacción hacia la derecha. Un descenso de temperatura favorece la reacción hacia la izquierda.

- Un ↓ de temperatura favorece la reacción hacia la izquierda.

d) INTRODUCCIÓN DE UN CATALIZADOR.

Añadir un catalizador a un sistema cambia la velocidad de la reacción, pero esto no puede desplazar el equilibrio a favor ni de reactivos ni de productos. Debido a que un catalizador afecta a la energía de activación tanto de la reacción directa como de la inversa por igual, cambia ambas reacciones por igual. El equilibrio se establece más rápidamente en presencia de un catalizador.

4.4.6 Presiones parciales y constante de equilibrio

A partir de

$$\mathrm{P} = \frac{\mathrm{n}}{\mathrm{V}}(\mathrm{RT})$$ Presión de un gas $\propto$ a su concentración (n/V)

$$\mathrm{P} = M(\mathrm{RT})$$

Para la reacción:

$$a\,\mathrm{A} + b\,\mathrm{B} \leftrightarrow c\,\mathrm{C} + d\,\mathrm{D} \qquad K_{\mathrm{p}} = \frac{(P_{\mathrm{C}})^{c}(P_{\mathrm{D}})^{d}}{(P_{\mathrm{A}})^{a}(P_{\mathrm{B}})^{b}}$$

4.4.7 Relación entre *K*p y *K*c.

de

$$\mathrm{P} = M(\mathrm{RT})$$

$$\left(\frac{n}{V}\right) = M = \frac{\mathrm{P}}{\mathrm{RT}}$$

para la síntesis del amoniaco

$$N_{2(g)} + 3\,H_{2(g)} \leftrightarrows 2\,NH_{3(g)}$$

$$K_{\mathrm{C}} = \frac{[\mathrm{NH}_3]^2}{[\mathrm{N}_2][\mathrm{H}_2]^3} = \frac{\left(\frac{P_{\mathrm{NH}_3}}{\mathrm{RT}}\right)^2}{\left(\frac{P_{\mathrm{N}_2}}{\mathrm{RT}}\right)\left(\frac{P_{\mathrm{H}_2}}{\mathrm{RT}}\right)^3} = \frac{(P_{\mathrm{NH}_3})^2}{(P_{\mathrm{N}_2})(P_{\mathrm{NH}_3})^3} \mathrm{x} \frac{\left(\frac{1}{\mathrm{RT}}\right)^2}{\left(\frac{1}{\mathrm{RT}}\right)^4}$$

$$K_{\mathrm{C}} = \mathrm{K}_{\mathrm{P}}(\mathrm{RT})^2$$
$$K_{\mathrm{P}} = \mathrm{K}_{\mathrm{C}}(\mathrm{RT})^{-2}$$

$$K_{\mathrm{P}} = K_{\mathrm{C}}(\mathrm{RT})^{\Delta n}$$
$$K_{\mathrm{C}} = \mathrm{K}_{\mathrm{P}}(\mathrm{RT})^{-\Delta n}$$
$$\Delta n = (n_{\text{gas prod}} - n_{\text{gas react}})$$

4.4.8 Equilibrios Heterogéneos.

- Estos equilibrios implican especies en más de una fase

$$2\,HgO_{(s)} \leftrightarrows 2\,Hg_{(l)} + O_{2(g)}$$

- Cuando se alcanza el equilibrio, hay un sólido, un gas y un líquido.

Ni sólidos ni líquidos se ven afectados significativamente por cambios en la presión. La definición fundamental de la constante de equilibrio en termodinámica se hace en términos de las actividades implicadas.

Para sólidos y líquidos la actividad se toma como 1.0, de modo que *no aparecen* términos para sólidos y líquidos puros en las expresiones de K para equilibrios heterogéneos.

Así $K_C = [O_2]$ y $K_P = P_{O2}$

4.4.9 Relación entre $\Delta G°$ y la constante de equilibrio.

- $\Delta G°$ = Cambio de Energía que se produce al establecerse un Equilibrio. Cuando $[\text{Reactivos}]_0$ y $[\text{Productos}]_0$ están a 1 M en disolución acuosa como sólidos y líquidos puros y como gases a presión parcial de 1 atm.

- ΔG= cambio de Energía para condiciones diferentes de éstas:

$$\Delta G = \Delta G° + RT \ln Q$$
$$\Delta G = \Delta G° + 2.303RT \log Q$$

R = cte. de los gases: 8.314 J/K mol
T = temperatura en K
Q = Cociente de reacción.

- Si el sistema está en equilibrio ($\Delta G = 0$)

$$0 = \Delta G° + RT \ln Q$$
$$\Delta G° = -2.303RT \log Q$$

- Además en el equilibrio $Q = K$:

$$\Delta G° = -2.303RT \log K$$

$\Delta G°$ es -	K>1	La reacción inversa es espontánea con concentraciones o presiones paralelas unitarias
$\Delta G°$ es -	K =0	El sistema está en equilibrio con concentraciones o presiones parciales unitarias (caso raro)
$\Delta G°$ es +	$K < 1$	La reacción inversa es espontánea con concentraciones o presiones paralelas unitaria

4.4.10 Cálculo de constante de equilibrio a distintas temperaturas.

- Las constantes de equilibrio (K), se calculan a una temperatura.

- A partir de la constante obtenida a una temperatura, puede calcularse el valor de la constante a Otra temperatura, mediante el uso de **$\Delta H°$** y la ecuación de Clausius – Clapeyron.

$$\frac{\Delta H°(T_2 - T_1)}{2.303RT_2T_1} = \log\left(\frac{K_{T_2}}{K_{T_1}}\right)$$

4.4.11 Problemas de Equilibrio Químico

1. Se introduce algo de nitrógeno y de hidrógeno en un recipiente vacío de 5.0 L @ 500°C. Cuando se establece el equilibrio, hay presentes 3.01 mol de N_2, 2.1 mol de H_2 y 0.565 mol de NH_3. Calcular la K_C para la reacción @ 500°C.

2. Para la siguiente reacción:

$$2\ HBr_{(g)} + Cl_{2(g)} \leftrightarrows 2\ HCl_{(g)} + Br_{2(g)} \qquad K_C = 4x10^4$$

Escribir la expresión, y calcular el valor numérico de la constante de equilibrio para cada una de las siguientes reacciones a la misma temperatura.

(a) $4\ HBr_{(g)} + 2\ Cl_{2(g)} \leftrightarrows 4\ HCl_{(g)} + 2\ Br_{2(g)}$
(b) $HBr_{(g)} + ½\ Cl_{2(g)} \leftrightarrows HCl_{(g)} + ½\ Br_{2(g)}$

3. A temperatura muy elevada, $K_C = 1.0x10^{-13}$ para la siguiente reacción:

$$2\ HF_{(g)} \leftrightarrows H_{2(g)} + F_{2(g)}$$

A un tiempo determinado se detectaron las siguientes concentraciones ¿ El sistema está en equilibrio? Si no es así, ¿qué debe suceder para que se establezca el equilibrio?

$[HF] = 0.5\ M$, $[H_2] = 1x10^{-3}\ M$, y $[F_2] = 4x10^{-3}\ M$.

4. La constante de equilibrio (KC) para la formación de cloruro de nitrosilo, un compuesto de color amarillo – naranja, a partir de óxido nítrico y cloro molecular:

$$2\ NO_{(g)} + Cl_{2(g)} \leftrightarrows 2\ NOCl_{(g)}$$

tiene el valor de $6.5x10^4$ @ 35°C. En un experimento se mezclan $2.0x10^{-2}$ mol de NO, $8.3x10^{-3}$ mol de Cl_2 y 6.8 mol de NOCl en un matraz de 2.0 L. ¿En qué dirección procederá el sistema para alcanzar el equilibrio?

5. Sobre los valores de las constantes de equilibrio, escoger las reacciones en las que se ven favorecidos los reactivos:

(a) $[AlF_6]^{3-}_{(ac)} \leftrightarrows Al^{3+}_{(ac)} + 6\ F^-_{(ac)}$ $K = 2x10^{-24}$
(b) $Ca_3(PO_4)_{2(s)} \leftrightarrows 3\ Ca^{2+}_{(ac)} + 2\ PO_4^{3-}{}_{(ac)}$ $K = 1x10^{-25}$

(c) $2\ Fe^{3+}_{(ac)} + 3\ S^{2-}_{(ac)} \leftrightarrows Fe_2S_{3(s)}$ $K = 1x10^{88}$

6. La reacción entre el nitrógeno y el oxígeno para formar $NO_{(g)}$ se representa mediante la ecuación química:

$$N_{2(g)} + O_{2(g)} \leftrightarrows 2\ NO_{(g)}$$

Las concentraciones de los gases en equilibrio @ 1500 K son $1.7x10^{-3}$ mol/L para O_2, $6.4x10^{-3}$ mol/L para N_2 y $1.1x10^{-5}$ mol/L para NO. Calcular el valor de K_C @ 1500 K a partir de estos datos.

7. Calcular la constante de equilibrio (***K***) @ 50°C para la reacción:

$$C_2H_{2(g)} + O_{2(g)} \leftrightarrows CO_{2(g)} + H_2O_{(g)}.$$

8. Para la reacción: $N_{2(g)} + O_{2(g)} \leftrightarrows 2\ NO_{(g)}$, la $\Delta G°_{reacción} = + 41.38$ Kcal. Calcular la K de la reacción a 25°C.

9. Para la reacción: $C_2H_{4(g)} + H_{2\ (g)} \leftrightarrows C_2H_{6(g)}$, la ***K*** de la reacción es $2.415x10^4$ a 25°C. Calcular la $\Delta G°$.

10. Para la reacción: $N_{2\ (g)} + O_{2\ (g)} \leftrightarrows 2\ NO_{(g)}$, la ***Kp*** $= 4.6x10^{-31}$ a 25°C y $\Delta H° = 43.14$ kcal. Calcular la constante de equilibrio a 2400 K y comparar con la K a 298 K.

11. Se coloca 10 g de $SbCl_3$ en un recipiente de 5 L @ 448°C y dejamos que la reacción llegue al equilibrio. ¿Cuántos gramos de $SbCl_5$ hay presentes en el equilibrio?. Resolver este problema (a) usando K_C y las concentraciones molares y (b) usando K_p y las presiones parciales.

12. Escribir las expresiones de K_C y K_P para las siguientes reacciones:

- $2\ SO_{2(g)} + O_{2(g)} \leftrightarrows 2\ SO_{3(g)}$
- $2\ NH_{3(g)} + H_2SO_{4(l)} \leftrightarrows (NH_4)_2SO_{4(s)}$
- $S_{(s)} + H_2SO_{3(ac)} \leftrightarrows H_2S_2O_{3(ac)}$

Equilibrio Ácido-Base

Los apuntes de este tema están basados en el capítulo 15 de la referencia 1.

Referencia:

1. Química; Chang Raymond; 6ª. Edición, Mc Graw Hill, México 1999.

WS 5.7 - Review

Balance these following chemical reactions:

1. ___ CO + ___ O_2 ---> CO_2
2. ___ Al + ___ HNO_3 ---> ___ $Al(NO_3)_3$ + ___H_2
3. ___ $Ca(OH)_2$ + ___ HBr ---> ___ $CaBr_2$ + ___ H_2O
4. ___ C_4H_9OH + ___ O_2 ---> ___ CO_2 + ___H_2O

<u>Use dimensional analysis to determine the following:</u>

5. How many moles are in 3.98 g of $CuSO_4$?

Ans ________

6. How many molecules are in 0.1029 moles of He?

Ans ________

7. 8.4 x 10^{24} boron atoms weigh how many grams?

Ans ________

8. **2 $KClO_3$ ---> 2 KCl + 3 O_2**

How many grams of O_2 will be produced from 55.4 g of $KClO_3$?

Ans ________

9. **2 Na + Cl_2 ---> 2 NaCl**

a. Starting with 30.1 g of Na and 22.4 g of Cl_2, how many grams of NaCl can be made?

Ans ________

b. Afterwards, 17.1 grams of NaCl are produced by the reaction. What is the % yield?

Ans ________

10. A compound is 87.5% N and is 12.5% H. What is its emperical formula?

Ans ________

11. The compound in #10 has a formula mass of 32 amu. What is its molecular formula?

Ans ________

Use the activity series (at right) to predict whether the following reactions will occur...

If **YES**, then write the products -- If **NO**, then write 'N. R.' (no reaction)

12. Al + $FeCl_2$ ---> ____________________
13. Pb + $Cu(NO_3)_2$ ---> ______________________
14. Zn + NaCl ---> _______________________
15. K + Ag_2S ---> ___________________

<u>Predict the products:</u>

16. $CuSO_4$ + AgCl ---> ___________________
17. C_2H_6 + O_2 ---> ___________________

Li
K
Ca
Na
Mg
Al
Mn
Zn
Cr
Fe
Cd
Co
Ni
Sn
Pb
H
Cu
Hg
Ag
Pt
Au

Se resolvieron la mayoría de los ejercicios complementarios del Capítulo 4 de la referencia 1.

Los ejercicios complementarios comprenden del ejercicio 4.91 al 4.140 (Pág. 148-151)

Referencia:

1. Química; Chang Raymond; 6ª. Edición, Mc Graw Hill, México 1999.

Chemical Reactions Worksheet

Directions: Complete and balance each of the reactions below. Identify each reaction type.

1. chlorine gas + sodium metal →

2. aluminum sulfate + barium chloride →

3. sodium hydrogen carbonate → water + ?

4. beryllium fluoride + magnesium →

5. pentane (C_5H_{12}) + oxygen →

6. hydrogen gas + oxygen gas →

7. sodium iodide + chlorine gas →

8. magnesium chloride + silver nitrate →

9. aluminum sulfate + calcium phosphate →

Worksheet #20 Predicting Products of Chemical reactions **Name** ________________________
Text Reference: pgs 256-264

A) For the following **synthesis reactions**, complete the chemical equation and balance the equation.

1. $Na(s) + O_2(g) \rightarrow$

2. $Mg(s) + F_2(g) \rightarrow$

3. $Al(s) + S(s) \rightarrow$

4. $Ca(s) + P(s) \rightarrow$

B) For the following **decomposition reactions**, complete the chemical equation and balance the equation.

1. $HgO(s) \rightarrow$

2. $CuCl_2(s) \rightarrow$

3. $Al_2O_3(s) \rightarrow$

C) For the following **single displacement reactions**, complete the chemical equation and balance the equation.

1. $Zn(s) + Pb(NO_3)_2(aq) \rightarrow$

2. $Al(s) + NiSO_4(aq) \rightarrow$

3. $Na(s) + H_2O(l) \rightarrow$

4. $F_2(g) + NaCl(aq) \rightarrow$

D) For the following **double displacement reactions**, complete the chemical equation and balance the equation.

1. $AgNO_3(aq) + CaCl_2(aq) \rightarrow$

2. $KOH(aq) + ZnCl_2(aq) \rightarrow$

3. $FeS(s) + HCl(aq) \rightarrow$

4. $CoCl_2(aq) + Na_3PO_4(aq) \rightarrow$

E) For the following **combustion reactions**, complete the chemical equation and balance the equation.

1. $C_7H_{16}(g) + O_2(g) \rightarrow$

2. $C_2H_2(g) + O_2(g) \rightarrow$

3. $C_3H_8O(g) + O_2(g) \rightarrow$

F) For each of the following reactions, you will need to identify the reaction type from the reactants, predict the products and then balance the equation.

1. $HCl(aq) + NaOH(aq) \rightarrow$

2. $CH_3OH + O_2(g) \rightarrow$

3. $NaI(aq) + Cl_2(g) \rightarrow$

4. $Li(s) + N_2(g) \rightarrow$

5. $C_3H_6(g) + O_2(g) \rightarrow$

6. $AlCl_3(s) \rightarrow$

7. $Ca(s) + HCl(aq) \rightarrow$

8. $LiOH(aq) + Fe(NO_3)_3(aq) \rightarrow$

Unit 2 Practice Problems (with answers at end)

> The gods too are fond of a joke. --Aristotle

Balancing chemical reactions

1. Balance each equation below and tell which of the three types of reactions is involved.

$AgNO_3 + H_2S \rightarrow Ag_2S + HNO_3$
$Cl_2 + CuBr_2 \rightarrow CuCl_2 + Br_2$
$H_2 + N_2 \rightarrow NH_3$
$NaHCO_3 \rightarrow Na_2CO_3 + CO_2 + H_2O$
$C_6H_{14} + O_2 \rightarrow CO_2 + H_2O$
$H_2SO_4 + KOH \rightarrow K_2SO_4 + H_2O$

2. The following are the beginnings of *combination* reactions. Write balanced reactions for them.

sodium + bromine →
potassium + oxygen →

3. The following are *decomposition* reactions. Write balanced reactions for them.

aluminum hydroxide → aluminum oxide + water
ammonium nitrite → nitrogen + water

4. The following are the beginnings of *redox displacement* reactions. Write the balanced equations for them.

potassium iodide + bromine →
calcium + hydrochloric acid →

5. The following are the beginnings of *precipitation* reactions. Write balanced equations for them.

magnesium chloride + sodium hydroxide →
iron(II) sulfate + ammonium sulfide →

6. Where the word equation is complete, write and balance the chemical equation. Where the word equation is incomplete, complete it and write and balance a chemical equation.

calcium chloride + sodium hydroxide →
iron(II) sulfide + hydrochloric acid → dihydrogen sulfide + ?
nickel(II) chloride + ammonium sulfide →
ammonia + oxygen → nitric acid + water
magnesium + nitric acid →
potassium + water → (hint, think of water as HOH)
sodium bromide + chlorine →
iron(III) oxide + carbon monoxide → iron + carbon dioxide
copper + sulfuric acid → copper(II) sulfate + water + sulfur dioxide

7. If the reactions above are known to occur in water solution, write net-ionic equations that represent the reactions.

[insoluble substances are: $Ca(OH)_2$, FeS, NiS, Mg, K, Fe_2O_3, Fe, Cu]
[gases include: NH_3, H_2S, H_2, Cl_2, CO_2, SO_2]

> Innovators are inevitably controversial.
> --Eva LeGallienne

Balancing redox reactions

8. Balance the following redox reactions.

$Cr + Sn^{4+} \rightarrow Cr^{3+} + Sn^{2+}$ (neutral)

$Br_2 + SO_3^{2-} \rightarrow Br^- + SO_4^{2-}$ (H^+)

$CuS + NO_3^- \rightarrow Cu^{2+} + NO_2 + S$ (H^+)

$NH_3 + OCl^- \rightarrow N_2H_4 + Cl^-$ (OH^-)

$Cl_2 + I^- \rightarrow Cl^- + IO_3^-$ (OH^-)

Oxidation numbers and definitions

9. Return to the previous problem and for each reaction, assign oxidation numbers to each atom in the reaction, label the oxidizing agent (OA) and the reducing agent (RA).

Predicting when a reaction will occur

10. In each case below use solubility rules or the activity series to predict if a reaction will occur or not.

a. $NaCl + AgNO_3 \rightarrow ?$
b. $K_2SO_4 + NH_4Cl \rightarrow ?$
c. $Cu + AgNO_3 \rightarrow ?$
d. $Zn + MgSO_4 \rightarrow ?$
e. $Br_2 + FeCl_3 \rightarrow ?$
f. $Ag + SnCl_2 \rightarrow ?$
g. $Ba(NO_3)_2 + Li_2SO_4 \rightarrow ?$

Answers

1. 2,1,1,2 precipitation
1,1,1,1 halogen displacement (redox)
3,1,2 combination
2,1,1,1 decomposition
2,19,12,14 combustion
1,2,1,2 acid/base
2. $2\,Na + Br_2 \rightarrow 2\,NaBr$
$4\,K + O_2 \rightarrow 2\,K_2O$
3. $2\,Al(OH)_3 \rightarrow Al_2O_3 + 3\,H_2O$
$NH_4NO_2 \rightarrow N_2 + 2\,H_2O$
4. $2\,KI + Br_2 \rightarrow I_2 + 2\,KBr$
$Ca + 2\,HCl \rightarrow CaCl_2 + H_2$
5. $MgCl_2 + 2\,NaOH \rightarrow Mg(OH)_2 + 2\,NaCl$
$FeSO_4 + (NH_4)_2S \rightarrow FeS + (NH_4)_2SO_4$
6. $CaCl_2 + 2\,NaOH \rightarrow Ca(OH)_2 + 2\,NaCl$
$FeS + 2\,HCl \rightarrow H_2S + FeCl_2$
$NiCl_2 + (NH_4)_2S \rightarrow NiS + 2\,NH_4Cl$
$NH_3 + 2\,O_2 \rightarrow HNO_3 + H_2O$
$Mg + 2\,HNO_3 \rightarrow Mg(NO_3)_2 + H_2$
$2\,K + 2\,H_2O \rightarrow 2\,KOH + H_2$
$2\,NaBr + Cl_2 \rightarrow 2\,NaCl + Br_2$
$Fe_2O_3 + 3\,CO \rightarrow 2\,Fe + 3\,CO_2$
$Cu + 2\,H_2SO_4 \rightarrow CuSO_4 + 2\,H_2O + SO_2$
7. $Ca^{2+}_{(aq)} + 2\,OH^-_{(aq)} \rightarrow Ca(OH)_{2(s)}$
$FeS_{(s)} + 2\,H^+_{(aq)} \rightarrow H_2S_{(g)} + Fe^{2+}_{(aq)}$
$Ni^{2+}_{(aq)} + S^{2-}_{(aq)} \rightarrow NiS_{(s)}$
$NH_{3(g)} + 2\,O_{2(g)} \rightarrow H^+_{(aq)} + NO_3^-{}_{(aq)} + H_2O_{(l)}$
$Mg_{(s)} + 2\,H^+_{(aq)} \rightarrow Mg^{2+}_{(aq)} + H_{2(g)}$
$2\,K_{(s)} + 2\,H_2O_{(l)} \rightarrow 2\,K^+_{(aq)} + 2\,OH^-_{(aq)} + H_{2(g)}$
$2\,Br^-_{(aq)} + Cl_{2(g)} \rightarrow 2\,Cl^-_{(aq)} + Br_{2(aq)}$
$Fe_2O_{3(s)} + 3\,CO_{(g)} \rightarrow 2\,Fe_{(s)} + 3\,CO_{2(g)}$
$Cu_{(s)} + 4\,H^+_{(aq)} + SO_4^{2-}{}_{(aq)} \rightarrow Cu^{2+}_{(aq)} + SO_{2(g)} + 2\,H_2O_{(l)}$
8. $2\,Cr + 3\,Sn^{4+} \rightarrow 2\,Cr^{3+} + 3\,Sn^{2+}$
$H_2O + Br_2 + SO_3^{2-} \rightarrow SO_4^{2-} + 2\,Br^- + 2\,H^+$
$4\,H^+ + CuS + 2\,NO_3^- \rightarrow Cu^{2+} + S + 2\,NO_2 + 2\,H_2O$
$2\,NH_3 + OCl^- \rightarrow N_2H_4 + Cl^- + H_2O$
$3\,Cl_2 + 6\,OH^- + I^- \rightarrow 6\,Cl^- + IO_3^- + 3\,H_2O$

9. 0 +4 +3 +2
$2\,Cr + 3\,Sn^{4+} \rightarrow 2\,Cr^{3+} + 3\,Sn^{2+}$
RA OA

+1 -2 0 +4 -2 +6 -2 -1 +1
$H_2O + Br_2 + SO_3^{2-} \rightarrow SO_4^{2-} + 2\,Br^- + 2\,H^+$
OA RA

+1 +2 -2 +5 -2 +2 0 +4 -2 +1 -2
$4\,H^+ + CuS + 2\,NO_3^- \rightarrow Cu^{2+} + S + 2\,NO_2 + 2\,H_2O$
RA OA

-3 +1 -2 +1 -2 +1 -1 +1 -2
$2\,NH_3 + OCl^- \rightarrow N_2H_4 + Cl^- + H_2O$
RA OA

0 -2 +1 -1 -1 +5 -2 +1 -2
$3\,Cl_2 + 6\,OH^- + I^- \rightarrow 6\,Cl^- + IO_3^- + 3\,H_2O$
OA RA

10. a. yes, ppt b. no c. yes d. no e. no f. no g. yes, ppt

Name ______________________________ Date ________ Class ____________________

MATH SKILLS

Balancing Chemical Equations

Aluminum reacts with copper(II) chloride, $CuCl_2$, to form copper metal and aluminum chloride, $AlCl_3$. Write the balanced equation for this reaction.

1. Identify the reactants and products.
   Aluminum and copper(II) chloride are the reactants, and aluminum chloride and copper are the products.
2. Write a word equation for the reaction.

   aluminum + copper(II) chloride → aluminum chloride + copper

3. Write the equation using formulas for the reactants and products.

   $$Al + CuCl_2 \rightarrow AlCl_3 + Cu$$

4. Balance the equation one element at a time. The same number of each type of atom must appear on both sides. So far, there are three chlorine atoms on the right and only two on the left. To balance the number of chlorine atoms, you must multiply the amount of copper(II) chloride by 3 and multiply the amount of aluminum chloride by 2.

   $$Al + 3CuCl_2 \rightarrow 2AlCl_3 + Cu$$

Atom	Reactants	Products	Balance
Al	1	2	×
Cu	3	1	×
Cl	6	6	√

This equation gives you two aluminum atoms on the right and only one on the left. So you need to multiply the amount of aluminum by 2.

$$2Al + 3CuCl_2 \rightarrow 2AlCl_3 + Cu$$

Atom	Reactants	Products	Balance
Al	2	2	√
Cu	3	1	×
Cl	6	6	√

Balancing Chemical Equations *continued*

To balance the equation, multiply the amount of copper produced by 3.

$$2Al + 3CuCl_2 \rightarrow 2AlCl_3 + 3Cu$$

Atom	Reactants	Products	Balance
Al	2	2	√
Cu	3	3	√
Cl	6	6	√

Your Turn to Think

1. Combustion in automobile engines takes place when fuel and oxygen are combined and ignited in the cylinders of the engine. However, the air that provides the oxygen for combustion also introduces nitrogen into the engine. The nitrogen reacts with oxygen at the high temperatures present in the engine, producing nitrogen oxide compounds, which are a major component of smog. In one of these reactions, nitric oxide, NO, reacts with oxygen, O_2, to form nitrogen dioxide, NO_2. Write the balanced equation for this reaction.
2. During the centuries following the collapse of the western Roman Empire, marble (calcium carbonate, $CaCO_3$) was taken from the monuments of Rome and heated to form quicklime (calcium oxide, CaO), which was used to make plaster. Carbon dioxide, CO_2, was also produced in this decomposition reaction. Write the balanced equation for this reaction.
3. When a match is lit, sulfur (S_8) reacts with oxygen to release energy and form sulfur dioxide, SO_2. Write the balanced equation for this reaction.
4. Zinc reacts with water to produce zinc hydroxide, $Zn(OH)_2$, and molecular hydrogen gas, H_2. Write the balanced equation for this reaction.
5. Barium, Ba, reacts with sulfur, S_8, to form barium sulfide, BaS. Write the balanced equation for this synthesis reaction.
6. Automobile airbags rely on the decomposition of the compound sodium azide (NaN_3) to produce the nitrogen gas, N_2, needed to rapidly inflate the bag. Sodium is also produced in the reaction. Write the balanced equation for this decomposition reaction.
7. A useful single-displacement reaction involves thermite, which is a mixture of aluminum and iron oxide, Fe_2O_3. When the thermite reaches a high temperature, the components react to produce molten iron, aluminum oxide (Al_2O_3), and a great deal of energy. Write the balanced equation for the thermite reaction.

Balancing Chemical Equations *continued*

8. Acid indigestion can occur when the stomach produces too much hydrochloric acid, HCl. An old and effective remedy for this involves drinking a solution of baking soda (sodium hydrogen carbonate, $NaHCO_3$), which reacts with the hydrochloric acid to produce sodium chloride (NaCl), water, and carbon dioxide. Write the balanced equation for this reaction.
9. A problem with the remedy given in problem 8 for acid indigestion is that the carbon dioxide produced can cause discomfort. In many modern antacids, the active ingredient is magnesium hydroxide, $Mg(OH)_2$. When this compound reacts with the hydrochloric acid, a double-displacement reaction occurs that produces only water and magnesium chloride, $MgCl_2$. Write the balanced equation for this reaction.

Sample Problem

Write the equation that describes the formation of glucose and oxygen, by means of photosynthesis, from carbon dioxide and water.

1. Identify the reactants and products. Carbon dioxide and water, the reactants, form glucose and oxygen, the products.
2. Write a word equation for the reaction.

carbon dioxide + water → glucose + oxygen

3. Write the equation using formulas for the reactants and products. Some gaseous elements are molecules, not atoms. Oxygen in air is O_2, not O.

$$CO_2 + H_2O \rightarrow C_6H_{12}O_6 + O_2$$

4. Balance the equation one element at a time. The same number of each type of atom must appear on both sides. So far, there are six carbon atoms on the right and only one on the left. To balance the number of carbon atoms, multiply the amount of carbon dioxide by 6.

$$6CO_2 + H_2O \rightarrow C_6H_{12}O_6 + O_2$$

Atom	Reactants	Products	Balance
C	6	6	√
H	2	12	×
O	13	8	×

Balancing Chemical Equations *continued*

This equation gives you twelve hydrogen atoms on the right and only two on the left. So you need to multiply the amount of water by 6.

$$6CO_2 + 6H_2O \rightarrow C_6H_{12}O_6 + O_2$$

Atom	Reactants	Products	Balance
C	6	6	√
H	12	12	√
O	18	8	×

To balance the equation, multiply the amount of oxygen produced by 6.

$$6CO_2 + 6H_2O \rightarrow C_6H_{12}O_6 + 6O_2$$

Atom	Reactants	Products	Balance
C	6	6	√
H	12	12	√
O	18	18	√

Your Turn to Think

10. Uranium reacts with fluorine gas, F_2, to form uranium(VI) fluoride, UF_6. Write the balanced equation for this synthesis reaction.
11. Iron reacts with chlorine gas, Cl_2, to form iron(III) chloride, $FeCl_3$. Write the balanced equation for this synthesis reaction.
12. Aluminum sulfate, $Al_2(SO_4)_3$, decomposes to form aluminum oxide, Al_2O_3, and sulfur trioxide, SO_3. Write the balanced equation for this reaction.
13. Water is decomposed by electrolysis to form the gaseous products hydrogen, H_2, and oxygen, O_2. Write the balanced equation for this reaction.
14. Potassium chlorate, $KClO_3$, decomposes to form potassium chloride, KCl, and oxygen gas. Write the balanced equation for this decomposition reaction.
15. Chlorine gas, Cl_2, reacts with potassium bromide, KBr, to form potassium chloride and bromine, Br_2. Write the balanced equation for this single-displacement reaction.
16. Aluminum reacts with lead nitrate, $Pb(NO_3)_2$, to form lead and aluminum nitrate, $Al(NO_3)_3$. Write the balanced equation for this single-displacement reaction.

Name ______________________________ Date _______ Class ____________________

MATH SKILLS

Balancing Chemical Equations *continued*

17. Ammonium chloride, NH_4Cl, reacts with calcium hydroxide, $Ca(OH)_2$, to form calcium chloride ($CaCl_2$), ammonia (NH_3), and water. Write the balanced equation for this reaction.

18. Zinc reacts with hydrochloric acid, HCl, to form zinc chloride, $ZnCl_2$, and hydrogen gas. Write the balanced equation for this reaction.

19. Fluorine reacts with sodium chloride, NaCl, to form sodium fluoride, NaF, and chlorine. Write the balanced equation for this single-displacement reaction. (**Hint:** Remember that both fluorine and chlorine have two atoms per molecule when they are in elemental form.)

20. Calcium oxide, CaO, reacts with sulfur dioxide to form calcium sulfite, $CaSO_3$. Write the balanced equation for this synthesis reaction.

21. In air, calcium sulfite, $CaSO_3$, reacts slowly with oxygen to form calcium sulfate, $CaSO_4$. Write the balanced equation for this reaction.

22. When heated, mercury(II) oxide, HgO, decomposes to form mercury and oxygen. Through this reaction, Joseph Priestley demonstrated the existence of oxygen in 1774. Write the balanced equation for the decomposition of mercury(II) oxide.

23. Methanol, CH_3OH, decomposes to form carbon monoxide, CO, and hydrogen gas, H_2. Write the balanced equation for this decomposition reaction.

Mixed Review

24. Potassium nitrate, KNO_3, decomposes to form potassium nitrite, KNO_2, and oxygen gas, O_2. Write the balanced equation for this decomposition reaction.

25. Antimony(V) chloride, $SbCl_5$, reacts with potassium iodide, KI, to form the products potassium chloride, KCl, iodine, I_2, and antimony(III) chloride, $SbCl_3$. Write the balanced equation for this reaction.

26. Nitric acid, HNO_3, reacts with hydrogen sulfide, H_2S, to form nitrogen dioxide (NO_2), water, and sulfur (S). Write the balanced equation for this reaction.

27. Chromium(III) oxide, Cr_2O_3, reacts with silicon to form chromium metal and silicon dioxide, SiO_2. Write the balanced equation for this reaction.

28. Ammonium dichromate, $(NH_4)_2Cr_2O_7$, decomposes to form chromium(III) oxide, Cr_2O_3, nitrogen gas, N_2, and water. Write the balanced equation for this decomposition reaction.

29. Iron(II) chloride, $FeCl_2$, reacts with water at high temperatures to form triiron tetraoxide (Fe_3O_4), hydrochloric acid, and hydrogen gas. Write the balanced equation for this reaction.

Name ________________________________ Date _______ Class ______________________

MATH SKILLS

Balancing Chemical Equations *continued*

30. Aluminum sulfate, $Al_2(SO_4)_3$, reacts with calcium hydroxide, $Ca(OH)_2$, to form aluminum hydroxide, $Al(OH)_3$, and calcium sulfate, $CaSO_4$. Write the balanced equation for this double-displacement reaction.

31. Lead chloride, $PbCl_2$, reacts with sodium chromate, Na_2CrO_4, to form lead chromate, $PbCrO_4$, and sodium chloride, NaCl. Write the balanced equation for this double-displacement reaction.

32. Vanadium pentoxide, V_2O_5, reacts with calcium to form vanadium metal and calcium oxide, CaO. Write the balanced equation for this reaction.

33. When heated, tantalum metal reacts with chlorine gas to form tantalum pentachloride, $TaCl_5$. Write the balanced equation for this reaction.

34. Potassium manganate(VI), K_2MnO_4, can be made to react with water, forming potassium permanganate, $KMnO_4$, potassium hydroxide, KOH, and hydrogen gas. This reaction is used in industry to produce potassium permanganate, which is widely used in analytical chemistry. Write the balanced equation for this reaction.

35. Phosphorus pentoxide, P_4O_{10}, can be made to react with water (as steam) to form phosphoric acid, H_3PO_4. This reaction is an important industrial source of phosphoric acid. Write the balanced equation for this reaction.

36. Iron sulfide, FeS, reacts with hydrochloric acid, HCl, to form hydrogen sulfide, H_2S, and iron(II) chloride, $FeCl_2$. Write the balanced equation for this reaction.

Se resolvieron la mayoría de los ejercicios complementarios del Capítulo 14 de la referencia 1.
Los ejercicios complementarios comprenden del ejercicio 14.61 al 14.102 (Pág. 592-595)

Referencia:

1. Química; Chang Raymond; 6ª. Edición, Mc Graw Hill, México 1999.

Solubility Equilibrium

1. For equilibrium to exist, two opposing processes must be occurring at the same rate. Use this definition to explain how a "saturated solution" is an example of equilibrium.

2. Write balanced equations for the dissociation of the following salts. Include the solubility product expressions for each salt, also.

   a) silver phosphate, Ag_3PO_4
   b) magnesium fluoride, MgF_2
   c) calcium carbonate, $CaCO_3$
   d) lead iodide, PbI_2
   e) cerium(III) hydroxide, $Ce(OH)_3$
   f) aluminum chromate, $Al_2(CrO_4)_3$

3. What mass of sodium hydroxide, NaOH, must be dissolved to prepare 2.00 L of 0.0500-M solution?

4. What mass of iron(III) nitrate, $Fe(NO_3)_3$, must be dissolved to prepare 500.0 mL of 0.120-M solution?

5. A student used a pipette to transfer 25.00 mL of 0.50-M sulfuric acid to a clean 500-mL volumetric flask. She then filled the flask with distilled water and shook to mix. What is the concentration of the new solution?

6. A chemistry student mixed 20.0 mL of 0.010-M NaOH with 25.0 mL of 0.025-M KNO_3. What is the concentration of each ion in the mixed solution? No reaction occurs during the mixing.

7. Indicate whether the following statements are true or false.

   a) Sodium salts are generally very soluble in water.
   b) The solubility of salts is generally higher in cold water than in hot water.
   c) The solubility of gases is generally higher in cold water than in hot water.
   d) Salts containing the chloride ion are soluble, except silver chloride and lead chloride.
   e) Salts containing the nitrate ion are insoluble.
   f) "Sparingly soluble" is a more accurate description for salts considered to be "insoluble" in water.
   g) A salt with a small Ksp is generally very soluble in water.
   h) To compare solubilities of two salts in water, one can simply compare the sizes of their solubility product constants.

8. A student mixed 1.0-M solutions of the following salts. In which case(s) will a precipitate likely form? Identify any precipitates that form.

   a) NaI is mixed with KNO_3
   b) $Pb(NO_3)_2$ is mixed with KNO_3
   c) $FeCl_3$ is mixed with KOH
   d) $CuSO_4$ is mixed with $CaCl_2$
   e) KF is mixed with $AgNO_3$
   f) $ZnCl_2$ is mixed with $Pb(NO_3)_2$
   g) $CaCl_2$ is mixed with NaOH
   h) KOH is mixed with NaCl

9. At 25°C, the solubility product constant for silver bromide, AgBr, is 5.0×10^{-13}. Calculate the molar solubility of silver bromide at this temperature.

10. Calculate the mass of AgBr that can be dissolved in 2.00 L of water at 25°C.

11. As temperature increases, does Ksp for a salt increase or does it decrease? Justify your answer.

12. At 25°C, the Ksp of calcium fluoride (CaF_2) is 3.9×10^{-11}. What is the molar solubility of calcium fluoride in water at 25°C?

13. The Ksp of magnesium hydroxide, $Mg(OH)_2$, is 7.1×10^{-12} at 25°C. Calculate the solubility of this salt, expressed in "grams per liter".

14. Lead sulfate, $PbSO_4$, has a solubility product constant, Ksp, of 6.3×10^{-7} at 25°C. The Ksp of lead fluoride, PbF_2, is smaller, 3.6×10^{-8} at 25°C. Calculate the molar solubility of each salt. Which salt is more soluble in water?

15. The Ksp of aluminum hydroxide, $Al(OH)_3$, is very small (3.0×10^{-34}). How much water is required to dissolve 1.0 g of this salt?

16. Calculate the volume of water needed to dissolve 0.50g of silver chromate. The Ksp of Ag_2CrO_4 is 1.2×10^{-12} at 25°C.

17. What mass of lead iodide, PbI_2, is needed to make 1.00 L of saturated solution at 25°C? The Ksp of lead iodide is 7.9×10^{-9} at 25°C.

18. What mass of iron(II) hydroxide, $Fe(OH)_2$, is needed to prepare 500.0 mL of saturated solution? The Ksp of $Fe(OH)_2$ is 7.9×10^{-16} at 25°C.

19. A student added a teaspoon of barium sulfate, $BaSO_4$, to a glass of water and stirred for several minutes. Use calculations to predict whether the salt will appear to dissolve in the water. The Ksp of barium sulfate is 1.1×10^{-10}.

20. Lead iodide, PbI_2, is a bright yellow precipitate that makes a colorful demonstration. The Ksp of lead iodide is 7.9×10^{-9}. A chemist added 1.00 g of sodium iodide, NaI, to 2.00 L of 0.100-M lead nitrate, $Pb(NO_3)_2$. Will the chemist observe a precipitate?

21. What mass of potassium hydroxide, KOH, can be dissolved in 500.0 mL of 0.20-M magnesium nitrate, $Mg(NO_3)_2$, without causing a precipitate to form? The Ksp of $Mg(OH)_2$ is 7.1×10^{-12}.

22. Calculate the mass of Na_2CO_3 that must be dissolved in 2.00 L of 0.050-M $AgNO_3$ to make a saturated solution of silver carbonate, Ag_2CO_3. The Ksp of silver carbonate is 8.1×10^{-12}.

23. The solubility of strontium carbonate, $SrCO_3$, is 0.0059 g per 250 mL of water. What is the Ksp for strontium carbonate?

24. The molar solubility of barium phosphate, $Ba_3(PO_4)_2$, is 0.0041 mol/L. Calculate the Ksp for this salt.

25. The solubility of calcium hydroxide, $Ca(OH)_2$, is measured to be 0.93 g/L at some temperature. What is the Ksp of the salt at this temperature?

26. You place 2.75 g of BaF_2 in 1.00 L of pure water at 25°C. After equilibrium has been established, the fluoride ion concentration is 0.0150 M. What is the Ksp of the salt?

27. Describe the *common ion effect* on the solubility of a salt.

28. Compare the solubility of AgCl (Ksp = 1.0×10^{-10}) in water to that in 0.10-M NaCl. T

29. The Ksp of silver iodide is 8.3×10^{-17}. What is the iodide ion concentration of a 1.00-L saturated solution of AgI, to which 0.020 mol of $AgNO_3$ is added?

30. Calculate the molar solubility of silver thiocyanate, AgSCN, in pure water and in a solution containing 0.010-M NaSCN. The Ksp of silver thiocyanate is 1.0×10^{-12}.

31. What is the solubility, in "grams per milliliter", of barium fluoride in a solution containing 1.5 g/L potassium fluoride, KF. The Ksp of BaF_2 is 1.7×10^{-6}.

32. Calculate the molar solubility of calcium hydroxide, $Ca(OH)_2$, in 0.100-M NaOH. The Ksp of $Ca(OH)_2$ is 6.5×10^{-6} at 25°C.

33. Rank the following compounds in order of increasing solubility in pure water: Na_2CO_3, BaF_2, $PbCl_2$, and Ag_2CrO_4. Refer to the Ksp table in ch. 17 of your text.

34. Rank the following compounds in order of increasing solubility in pure water: $CaSO_4$, $NiCO_3$, AgOH, and ZnS. Refer to the Ksp table in ch. 17 of your text.

35. Which has a greater solubility in pure water – Ag_2SO_4 or $CaSO_4$? Refer to the Ksp table in ch. 17 of your text.

36. Will a precipitate form if 100.0 mL of 0.0030-M NaCl is added to 200.0 mL of 0.0.012-M $AgNO_3$? The Ksp of silver chloride, AgCl, is 1.0×10^{-10}.

37. Will a precipitate form if 10.0 mL of 0.050-M NaOH is added to 500 mL of 0.050-M $Pb(NO_3)_2$? The Ksp of $Pb(OH)_2$ is 1.2×10^{-15}.

38. Will a precipitate form if 2.0 g of solid $Pb(NO)_3$ is added to 250 mL of 0.020-M NaI? The Ksp of PbI_2 is 1.4×10^{-8}. Assume that no volume change occurs.

39. Will a precipitate form if 50.0 mL of 0.10-M $AgNO_3$ is added to 1.0-L of 0.10-M KCl? The Ksp of AgCl is 1.0×10^{-10}.

40. Will a precipitate form if 25.0 mL of 0.0040-M NaF is added to 75.0 mL of 0.0160-M $Mg(NO_3)_2$? The Ksp of MgF_2 is 6.4×10^{-9}.

41. Define a "supersaturated solution" in terms of the reaction quotient.

42. A solution is prepared by dissolving 1.40 g of Ag_2SO_4 in 100.0 mL of hot water. Will a precipitate form if the solution is cooled to 25°C? At 25°C, the Ksp of Ag_2SO_4 is 1.2×10^{-5}.

PROBLEM SET 3: SOLUBILITY AND COMPLEX ION EQUILIBRIA
CHEM 102/105 Sept. 23, 2002

References: 4.5, 15.6–15.8 Solubility equilibria, qualitative analysis, complex ion equilibria
Similar problems in Zumdahl: Chap. 4: 35; Chap. 15: 79–111 odd

1. State whether the following salts are soluble or insoluble in water:

   Barium nitrate
   Barium chloride
   Barium carbonate
   Ammonium sulphate
   Potassium chromate
   Sodium sulphide
   Silver(I) bromide
   Lead(II) nitrate
   Chromium(III) hydroxide
   Lead(II) sulphide
   Lead(II) chromate
   Lead(II) carbonate
   Sodium phosphate
   Lead(II) iodide
   Sodium bromide

2. At 25°C, the K_{sp} of silver(I) phosphate is 1×10^{-19}. (What is the formula for this salt?)
   (a) Calculate the solubility of silver(I) phosphate in mol/L and in g/L at 25°C.
   (b) At 19°C, the solubility of silver(I) phosphate is 2.7×10^{-6} mol/L. Calculate K_{sp} at 19°C, and comment on how solubility depends on temperature.
   (c) Determine the solubility of silver(I) phosphate in a 0.20 M solution of $AgNO_3$ at 25°C.
3. Will a precipitate form when the solutions below are mixed? If so, determine the moles of precipitate formed and the equilibrium concentrations of the dissolved species.
   (a) 500 mL of 3.0×10^{-4} M thallium(I) nitrate and 500 mL of 3.0×10^{-4} M potassium iodide (K_{sp} of TlI = 4.0×10^{-8}).
   (b) 200 mL of 9.0×10^{-6} M $Hg_2(NO_3)_2$ and 100 mL of 1.5×10^{-6} M $BaCl_2$ (K_{sp} for $Hg_2Cl_2 = 1.1 \times 10^{-18}$; in aqueous solutions, mercury(I) exists as the dimeric Hg_2^{2+} ion.)
4. The K_{sp} values for some insoluble hydroxides are:

   $Al(OH)_3$ 1.3×10^{-33}; $Co(OH)_3$ 1.6×10^{-44}; $Mn(OH)_2$ 2×10^{-13}.

   (a) Calculate the solubility of $Al(OH)_3$ in pure water and in a solution buffered at *pH* 9.31.
   (b) Calculate the solubility of $Co(OH)_3$ in pure water and in a solution buffered at *pH* 9.00.
   (c) Will a precipitate of $Mn(OH)_2$ form when a solution of 1.0 M $MnCl_2$ is prepared in plain water? Using le Châtelier's principle, explain how you could enhance or diminish the solubility of $Mn(OH)_2$ by adding an acid or base. Determine the range of *pH* in which $Mn(OH)_2$ will form from a 1.0 M Mn^{2+} solution.
5. A solution contains 1.0×10^{-5} M Cl^- and 1.0×10^{-5} M I^- ions. To this solution, Ag^+ ions are slowly added.
   (a) Which ion, Cl^- or I^-, precipitates first? ($K_{sp} = 1.8 \times 10^{-10}$ for AgCl and $K_{sp} = 8.3 \times 10^{-17}$ for AgI.) Determine the range of $[Ag^+]$ that will permit a clean separation of these ions.
   (b) Calculate the concentration of the first ion when the second ion just starts to precipitate.
6. A solution contains 1.0×10^{-4} M of both Pb^{2+} and Zn^{2+} ions. To this solution, S^{2-} ions are slowly added, with the intention of precipitating either PbS ($K_{sp} = 8.0 \times 10^{-28}$) or ZnS ($K_{sp} = 1.6 \times 10^{-24}$).
   (a) Which ion, Pb^{2+} or Zn^{2+}, precipitates first? Determine the range of $[S^{2-}]$ that will permit a clean separation of Pb^{2+} and Zn^{2+} ions.
   (b) Calculate the concentration of the first ion when the second ion just starts to precipitate.
   (c) In practice, the amount of S^{2-} is regulated via the equilibrium:

   $$H_2S(aq) \rightleftharpoons 2\,H^+(aq) + S^{2-}(aq) \qquad K_a = 1 \times 10^{-26}$$

   Suppose the concentration of $H_2S(aq)$ is fixed at 0.10 M. What *pH* range is necessary for the clean separation described in (a)?

7. The overall formation constant for $Zn(NH_3)_4^{2+}$ is 2.9×10^9. Calculate the equilibrium concentration of Zn^{2+} in a solution prepared by dissolving 0.220 mol $ZnCl_2$ in 500 mL of a 2.0 M ammonia solution.
8. Calculate the solubility (in mol/L) of AgI in a 4.0 M CN^- solution (K_{sp} for AgI $= 8.3 \times 10^{-17}$ and overall K_f for $Ag(CN)_2^- = 1 \times 10^{21}$).
9. Calculate the solubility (in mol/L) of AgCl in a 1.0 M NH_3 solution (K_{sp} for AgCl $= 1 \times 10^{-10}$ and overall K_f for $Ag(NH_3)_2^+ = 1.6 \times 10^7$).
10. Play detective and use solubility rules to deduce which salts in the samples below <u>must</u> be present, <u>must</u> be absent, or cannot be determined from the data. (Acetate salts are soluble.)
    (a) A sample consists of one or more of these salts: Na_2CO_3, KNO_3, $CaCl_2$. The sample dissolves completely in water, but when dilute HCl is added, effervescence is observed.
    (b) A sample consists of one or more of these salts: $NaC_2H_3O_2$, KCl, $Hg_2(NO_3)_2$. The sample dissolves completely in water, but when aqueous KI is added, a precipitate forms.
    (c) A sample consists of one or more of these salts: NH_4NO_3, Na_2S, K_2CO_3, $Pb(C_2H_3O_2)_2$. The sample dissolves completely in water, but when aqueous Na_2SO_4 is added, a precipitate forms.

Answers:

1. soluble: $Ba(NO_3)_2$, $BaCl_2$, $(NH_4)_2SO_4$, K_2CrO_4, Na_2S, $Pb(NO_3)_2$, Na_3PO_4, NaBr
   insoluble: $BaCO_3$, AgBr, $Cr(OH)_3$, PbS, $PbCrO_4$, $PbCO_3$, PbI_2
2. (a) Solubility of Ag_3PO_4 is 7.8×10^{-6} mol/L or 3.3×10^{-3} g/L (b) $K_{sp} = 1.4 \times 10^{-21}$; solubility generally increases with temperature (c) 1.2×10^{-17} mol/L
3. (a) no ppt (b) 8.4×10^{-8} mol Hg_2Cl_2 formed; $[Hg_2^{2+}] = 5.5 \times 10^{-6}$ M, $[Cl^-] = 4.4 \times 10^{-7}$ M
4. (a) 1.3×10^{-12} mol/L in water *vs* 1.6×10^{-19} mol/L at *pH* 9.31 (b) 1.6×10^{-23} mol/L in water *vs* 1.6×10^{-29} mol/L at *pH* 9.00 (c) ppt does not form at *pH* 7; only under basic conditions (when *pH* > 7.7) will $Mn(OH)_2$ be present.
5. (a) I^-; $8.3 \times 10^{-12}\text{ M} < [Ag^+] < 1.8 \times 10^{-5}$ M (b) $[I^-] = 4.6 \times 10^{-12}$ M
6. (a) Pb^{2+}; $8.0 \times 10^{-24}\text{ M} < [S^{2-}] < 1.6 \times 10^{-20}$ M (b) $[Pb^{2+}] = 5.0 \times 10^{-8}$ M (c) $2.0 < pH < 3.6$
7. $[Zn^{2+}] = 4.6 \times 10^{-8}$ M
8. $s = 2.0$ mol/L
9. $s = 0.037$ mol/L
10. (a) Na_2CO_3 present, KNO_3 can't say, $CaCl_2$ absent (b) $NaC_2H_3O_2$ can't say, KCl absent, $Hg_2(NO_3)_2$ present (c) NH_4NO_3 can't say, Na_2S absent, K_2CO_3 absent, $Pb(C_2H_3O_2)_2$ present