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Introduction

Chemical equations are the stock-in-trade for much of what is expected of students in the first semester of college chemistry. Knowing how to read them, balance them, and interpret them in terms of molar ratios or stoichiometry is one thing; being able to write them from chemical knowledge or from experimental observations is another. The latter ability is developed only when students have acquired observational skills and a fairly good background in chemical principles. At some point each student must be able to bring whatever principle is needed to bear on the writing of a proper chemical equation. In the realm of inorganic chemistry it is possible to summarize those principles in a straight-forward manner to show that it is not usually necessary to memorize individual chemical equations.

Most of these reactions fall into the following categories, which are, of course, not mutually exclusive:

A. Reactions involving binary compounds
   1. Formation from the elements
   2. Decomposition
   3. Reaction with water
      a. Oxides
      b. Hydrides
      c. Binary salts
B. Reactions involving ternary compounds
   1. Formation
   2. Ionization - acids, bases, salts
C. Metathetical reactions
D. Oxidation-reduction

General Principles

The general principles that one must know to write valid chemical equations include the following:

1. The periodic classification of the elements.

The position of an element in the periodic chart is due to its electron configuration which in turn gives rise to the valences or oxidation states that characterize the element. Ionization potential, atomic size, electronegativity, and relative metallic and non-metallic behavior can all be inferred from the location of the element in the chart, and all of these properties determine the chemical behavior of the element.

2. The activity series of the metals and non-metals predicting the course of displacement reactions such as

$$\text{Element A+ Compound UC} \longrightarrow \text{Element U+ Compound AC.}$$

In some cases, particularly those involving the halogens, the position of the element in the periodic chart is the simplest clue.

3. The recognition of acids, bases, and salts is necessary to predict the course of reactions in which they occur in combination, such as neutralization reactions. As a starter a student must be able to identify a compound as an acid, base, salt, or oxide simply by looking at its formula, using the simple Arrhenius concept.

4. Going beyond the simple recognition of acids and bases from their formulas, a knowledge of their degree of ionization is necessary to predict the course of several different kinds of reactions. This information is available in tables of ionization constants that should be made available to students, or, on a more qualitative level, the location of the acid- or base-forming element in the periodic chart can be related to the degree of ionization.

5. In metathetical reactions (sometimes referred to as double displacement or partner-exchange reactions) the solubility rules of salts must be known and, in some cases, the degree of ionization of a possible product.

6. Especially in the case of oxidation-reduction reactions that take place in aqueous solution, one must know the most stable oxidation states of the non-metals (particularly groups VA, VIA, and VIIA) and the transition elements (B groups), since these are the elements most generally involved in such reactions. The most stable oxidation states are related directly to the electron configurations, which in turn determine the location of the element in the periodic chart. (Consider N, S, Cl, and Cr, Mn, and Fe.)

7. Particularly in the case of oxidation-reduction reactions that take place in aqueous solution one should be familiar with the ion-electron method of completing and balancing an ionic partial half-reaction, such as

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

(The procedure is described in virtually all beginning chemistry textbooks.)

8. Finally, if one is to predict with confidence that a possible reaction will in fact occur, simple thermodynamic principles are needed such as relative heats of formation or the meaning of standard reduction potentials. The level to which this is carried depends on the instructor's philosophy and the level of the course. If one is to write a chemical equation to describe what he has observed experimentally, he doesn't so much need a thermodynamic basis for interpreting observations, but rather a background that permits an intelligent choice among possible alternatives. At some levels it is gratifying to know that the student can at least make an educated guess.

Simple illustrative reaction types can be summarized in terms of the categories and principles listed above:

Illustrative Reaction Types

A. Reactions Involving Binary Compounds

1. Formation of binary compounds by direct union of the elements:

Examples

Element A + Element B ⟶ Compound AB

or

Non-metal + Non-metal ⟶ Covalent compound

or

Metal + Non-metal ⟶ Ionic compound

2. Decomposition of binary compounds:

Compound AB ⟶ Element A + Element B

3. Reaction of binary compounds with water:

a. Oxides:

Examples

Metal oxide + Water ⟶ a Base

or

Non-metal oxide + Water ⟶ an Acid

b. Hydrides:

Examples

Non-metal Hydrides + Water $\rightarrow$ A positive ion + A negative ion

or

Metal Hydrides + Water $\rightarrow$ Metal hydroxide + Hydrogen gas

c. Salts:

If a binary salt is the salt of a weak acid or a weak base it will react with water (hydrolyze) to form a small excess of H^+ or OH^- making the solution of the salt either acid or alkaline (basic). This principle applies to all salts, whether they be binary or ternary.

Example

$$\underset{}{Na_2S} + 2HOH \rightarrow \underset{\text{weak}}{H_2S} + \underset{\text{strong}}{2\,NaOH}$$

B. Reactions Involving Ternary Compounds:

The ternary compounds considered here consist of two elements besides oxygen. They are acids, bases, or salts.

1. Formation of ternary compounds:

Examples

Metal oxide + water A base (metal hydroxide)

or

Non-metal oxide + Water $\rightarrow$ an Acid

2. Ionization of ternary compounds:

a. Acids. All are covalent (molecular) solutes and ionize according to the general equation

$$HA + HOH \rightarrow H_3O^+ + A^-$$

(A^- = the oxy-anion)

Some ionize slightly and are called weak acids. Others ionize to a larger extent and are relatively stronger.

b. Bases. All group IA and IIA hydroxides are strong electrolytes, consisting of ions even in the solid state. Ionization occurs by simple dissolving and hydration of the ions.

c. Salts. Most salts are completely ionized both in the solid state and in solution. Solubilities vary greatly. Ionization is by simple dissolving and hydration of the ions.

C. Metathesis

Metathetical reactions occur between ionic compounds in which there is an exchange of ionic partners. They proceed only if an insoluble product is formed or if a non-ionic product is possible.

1. Neutralization (acid-base). Proceeds because of the formation of the almost completely non-ionic product HOH.

Example

$$\text{Acid} + \text{Base} \longrightarrow \text{a Salt} + \text{Water}$$

2. Double decomposition:

Example

$$\text{Salt MA} + \text{Salt NB} \longrightarrow \text{Salt MB} + \text{Salt NA}$$

Either salt MB or salt NA or both must be relatively insoluble compared to the reactants. To predict the course of such reactions one must know the rules of solubility of salts (easily found in most introductory chemistry texts) and also the relative degree of ionization of certain covalent compounds if double decomposition is generalized as

$$\text{Cpd AB} + \text{Cpd CD} \longrightarrow \text{Cpd AD} + \text{Cpd CB}.$$

Such a generalization would include neutralization reactions and the following.

3. Displacement of a "low-boiling" acid from one of its salts by reaction with a high-boiling acid such as concentrated sulfuric acid.
Examples of "low-boiling" acids: hydrochloric, nitric, acetic, any of the hydrogen halides, carbonic, sulfurous.

Examples

$$\underset{\text{(solid)}}{NaCl} + \underset{\text{(conc.)}}{H_2SO_4} \longrightarrow HCl + NaHSO_4$$

or

$$CaF_2 + H_2SO_4 \longrightarrow 2\,HF + Ca(HSO_4)_2$$

D. Oxidation-Reduction:

Any reaction in which the oxidation state of one or more elements changes is an oxidation-reduction reaction. Therefore some reactions outlined above are included here as well:

1. Direct union of the elements:
2. Decompostion of a compound into its elements:
3. Displacement:

Examples

$$\text{Elt. A} + \text{Cpd UC} \longrightarrow \text{Elt U} + \text{Cpd AC}$$

or

$$Zn + CuCl_2 \longrightarrow Cu + ZnCl_2$$

or

$$Cl_2 + KI \longrightarrow KCl + I_2$$

There is no direct or simple device for predicting the products of these reactions since the concentration of the reactants and the pH of the solution influence the extent of oxidation and reduction in many cases. The possible reduction products of various oxidizing agents must be memorized or a table of reduction half reactions can be used to determine what the possibilities are. Students are most likely to be making direct observations of these reactions and in attempting to write proper equations they must be knowledgeable about the characteristic colors of the oxidation states of various elements, particularly the transition elements, chemical tests for certain colorless ions such as sulfate and chloride, and above all they must develop a keen sense of observation and awareness of the difference between observation and inference.

Summary

Bridging the gap between the operational and the conceptual (as expressed by chemical equations) is the greatest challenge to both the chemistry student and the chemistry teacher. The successful student will, when confronted with a problem, ask himself or herself what information must be mastered to bear on the solution. The teacher can help by giving some thought to this when planning instruction and then deliberately sequencing experiences to develop the conceptual aspect of chemistry without leaving gaps here and there. These simple notes on learning to write chemical equations were written with that in mind.