WRITING NET IONIC EQUATIONS: A PREVIEW **Chemistry BC2001x**

Writing properly balanced net ionic equations among acids, bases, and salts is an important part of this course. Such equations convey in the **most concise way possible** the **predominant chemical process** that occurs. Some textbooks include a lot of so-called molecular equations for ionic reactions. One example is

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

As a description of what happens, this is useless, and would be graded as incorrect. Aqueous HCl molecules and NaOH units simply do not exist, and the salt NaCl(aq) is not formed. While HCl(aq) is a suitable label for a bottle, it is not acceptable in a properly written chemical equation. Why not? HCl is a *strong* acid: fully dissociated into H^+(aq) and Cl^-(aq) ions in water. NaOH is a very soluble ionic salt, so a solution labeled NaOH(aq) contains water and ions: Na^+(aq) and OH^-(aq). Sodium chloride, NaCl(s), is also very soluble, so the ions Na^+(aq) and Cl^-(aq) remain separate when the solutions are mixed: they do not combine to make NaCl(aq), as the equation above seems to imply. The one process that actually does occur is the combination of H^+ with OH^- to make water. So we could write the ionic equation

$$H^+(aq) + Cl^-(aq) + Na^+(aq) + OH^-(aq) \rightarrow H_2O(\ell) + Cl^-(aq) + Na^+(aq)$$

However to focus on what happens, omit **spectator** species, like Na^+ and Cl^-, that remain unchanged. This the correct **net ionic equation** is:

$$H^+(aq) + OH^-(aq) \rightarrow H_2O(\ell)$$

In this lab this term, you work almost exclusively with aqueous solutions. Water, the solvent, may be written $H_2O(\ell)$ or just H_2O. Dissolved (aqueous) species should be designated (aq). Other phase information is indicated as (s) for solids, (ℓ) for liquids, and (g) for gases.

The equation should convey symbolically what actually happens. If a reaction starts with *solid* sodium hydroxide, then NaOH(s) should appear on the left. But if the reaction starts with an aqueous solution of a soluble salt, so the reagent bottle contains ions, then only the ions may appear. If a precipitate, for example silver hydroxide, forms, then AgOH(s) appears on the right.

All this means that *you must know some chemistry* in order to write proper equations.

Acids. Strong acids are fully ionized, so they are written as ions. Weak acids are only slightly ionized, most of the acid molecules remain in the molecular form, so the molecule appears in the equation. Compare nitric acid (HNO_3), which is strong, with nitrous acid (NHO_2), which is weak. In net ionic equations, the former would be written $H^+(aq) + NO_3^-(aq)$ while the latter would be written $HNO_2(aq)$. $HNO_3(aq)$ may appear on a label, but never in a net ionic equation.

The acids you will encounter in this course are essentially all soluble. Chemical formulas of inorganic acids are written in the form H_nX, where n indicates the number of hydrogens that can dissociate as H^+(aq). The anion is X^{n-}. There are only a few strong acids, all others are weak. Memorize the following six strong acids: **HCl, HBr, HI, HNO_3, $HClO_4$, and H_2SO_4.**

Bases. Two classes of bases will appear in this course. The **soluble hydroxide salts** (see below) are strong bases, sources of large amounts of OH^-(aq). Common weak bases are **ammonia**, NH_3(aq), and related compounds called **amines**. For example methylamine, CH_3NH_2(aq), can be viewed as the result of substituting one –H unit with a $-CH_3$ (methyl) unit in ammonia. These weak bases appear in solution in molecular form. The reaction equations below show ammonia acting as a base, accepting protons either directly from solution or from a weak acid:

$$NH_3(aq) + H^+(aq) \rightarrow NH_4^+(aq)$$

$$NH_3(aq) + HNO_2(aq) \rightarrow NH_4^+(aq) + NO_2^-(aq)$$

Salts. Almost all salts are strong electrolytes, meaning that to the extent that they dissolve, they form ions. (Contrast this with acids, some of which ionize readily, and some do not.) However the solubilities of salts differ widely: some are very soluble, some moderately so, some quite insoluble. The nomenclature here may be a bit misleading: insoluble salts *do* dissolve, forming ions, they just do so to a very small extent.

The objective in writing **net ionic equations** is to describe the major process: what happens to *most* of the substance. If a salt is insoluble, then it should appear as a solid, whereas if it is soluble and is in water, it should appear as ions. Solubility depends on a delicate balance of factors, so there is no way to look at a formula of a salt and predict its solubility. Instead, we make certain useful empirical generalizations: **solubility rules**. Three are given below; a longer list will be distributed later. "Potassium salt" means a salt whose cation is potassium, K^+.

Preliminary Solubility Rules

1) **Potassium, sodium, and ammonium salts are *soluble*.**
2) **Nitrate and acetate salts are *soluble*.**
3) **Carbonate, phosphate, and hydroxide salts are generally *insoluble*, except with cations from rule 1) above.**

Hydronium ion. Free protons do not actually exist in solution as $H^+(aq)$, instead they attach to water (which is acting as a base), forming $H_3O^+(aq)$. Thus to be very explicit, one can write $H_3O^+(aq)$, as below. Observe that water is left when the proton is exchanged. However chemists routinely write $H^+(aq)$, understanding that it implies $H_3O^+(aq)$; see the equation above.

$$NH_3(aq) + H_3O^+(aq) \rightarrow NH_4^+(aq) + H_2O(\ell)$$

Other processes. Later, you will learn about some other things that can happen in ionic solutions, for example, formation of gases. These are special case that may be ignored for now. You will also learn how to deal with reactions of polyprotic acids, like H_2SO_4 and H_3PO_4.

Finally, always check that the overall charge is balanced and that each element appears the same number of times on each side. It is possible that nothing happens: everything is a spectator. In that case, write "no reaction".

Some Examples (translating words into a net ionic equation):

Solid nickel hydroxide dissolves in concentrated nitrous acid. No precipitate forms.

$$Ni(OH)_2(s) + 2\ HNO_2\ (aq) \rightarrow 2\ H_2O(\ell) + Ni^{2+}(aq) + 2\ NO_2^-(aq)$$

A solution of barium chloride reacts with a solution of potassium carbonate:

$$Ba^{2+}(aq) + CO_3^{2-}(aq) \rightarrow BaCO_3(s)$$

For Practice (answers on WWW):

Nitrous acid (HNO_2) reacts with methylamine, CH_3NH_2.

Nitric acid (HNO_3) reacts with methylamine, CH_3NH_2.

Solid aluminum hydroxide dissolves in HCl(aq).

Solid aluminum hydroxide dissolves in HOCl(aq).

Aqueous sodium hydroxide is mixed with aqueous zinc nitrate.

Copper (II) nitrate solution is mixed with potassium acetate solution.

Copper (II) nitrate solution is mixed with potassium hydroxide solution.

Copper (II) nitrate solution is mixed with potassium phosphate solution.