BLUFFER'S GUIDE TO PREDICTING REACTIONS

A REACTION WILL "GO" FOR ANY ONE OF THE FOLLOWING REASONS

1. For electron-transfer reactions, if the electrode potential (E) for the reducing half-reaction is more negative than that for the oxidizing half-reaction:
2. For non-electron-transfer reactions, if
   (a) an insoluble product is formed ($K_c > 1$ because $K_{sp} < 1$), or
   (b) a weak (or non-) electrolyte is formed ($K_c > 1$ because $K_i < 1$), or
   (c) a gas is formed in an open vessel (loss of gas product continually shifts equilibrium to right).

Guidelines

Some simple, general guides for using these principles are listed below.

1. *Electrode potentials* for half-reactions are most easily obtained from the Table of Standard Electrode Potentials; they may be modified for marked changes in concentration as needed.
2. *General solubility rules* for water as a solvent must be stated in two parts, one emphasizing the negative ions of the salts, and the other emphasizing the positive ions, as follows:
   (a) All alkali metals, ammonium (NH_4^+), nitrates (NO_3^-), chlorates (ClO_3^-), perchlorates (ClO_4^-), and acetates ($C_2H_3O_2^-$), halides (Cl^-, Br^-, and I^-), and sulfates (SO_4^{2-}) *are soluble*, with the following important exceptions:
      (1) the halides (Cl^-, Br^-, and I^-) of Ag^+, Hg_2^{2+}, and Pb^{2+} (AP/H) are insoluble;
      (2) the sulfates (SO_4^{2-}) of Ca^{2+}, Sr^{2+}, Ba^{2+}, Pb^{2+} (P/CBS) are insoluble.
   (b) All oxides (O^{2-}), hydroxides (OH^-), carbonates (CO_3^{2-}), phosphates (PO_4^{3-}), sulfides (S^{2-}), and sulfates (SO_3^{2-}) *are insoluble* except with alkali metals and NH_4^+ with the following exception:
      (1) the oxides (O^{2-}) and hydroxides (OH^-) of Ca^{2+}, Sr^{2+}, Ba^{2+} (CBS) are soluble.
   (c) In general, all other common types of inorganic compounds are insoluble.
3. *Strong electrolytes* (those that ionize in water) comprise, with minor exception:
   (a) all soluble salts (this does not include acids and bases);
   (b) all strong acids (there are eight common ones: HCl, HBr, HI, HNO_3, $HClO_3$, $HClO_4$, H_2SO_4, and HIO_4);
   (c) all strong bases (these are the alkali metal and Ba^{2+} hydroxides)
4. *Weak electrolytes* include all those substances that are not classed as strong electrolytes. A particularly important weak electrolyte is water.
5. *Common gases* at room temperature are: H_2, O_2, N_2, CO_2, CO, NO, NO_2, N_2O, Cl_2, NH_3, HCl, HCN, H_2S, SO_2. Other substances, particularly H_2O, may be driven off as a gas at elevated temperatures.

Balancing Ionic Equations

A properly balanced chemical equation shows all the information we have just discussed. Soluble salts and strong electrolytes in aqueous solutions are always written in ionic form, for example, $Na^+ + Cl^-$, not NaCl, or H^+ (or H_3O^+) + ClO_4^-, not $HClO_4$. Insoluble salts and weak electrolytes are not written in ionic form, even though a minor fraction of each may actually exist in solution. For example, we indicate the formula for calcium carbonate as $CaCO_3$, not $Ca^{2+} + CO_3^{2-}$. Similarly, for slightly dissociated acetic acid, we use the molecular formula $HC_2H_3O_2$ instead of $H^+ + C_2H_3O_2^-$.

Insoluble products are usually indicated by $_{(s)}$ and evolved gases by $_{(g)}$. The following equations illustrate correct procedures:

$$BaCO_3 + 2HC_2H_3O_2 \rightarrow Ba^{2+} + 2C_2H_3O_2^- + H_2O + CO_{2(g)}$$
$$BaCO_3 + 2H^+ + SO_4^{2-} \rightarrow BaSO_{4(s)} + H_2O + CO_{2(g)}$$
$$BaCO_3 + 2H^+ + 2Cl^- \rightarrow Ba^{2+} + 2Cl^- + H_2O + CO_{2(g)}$$

Whenever there are chemical species in solution that are *identical* on both sides of the equation, like Cl^- in the last example, this species does not participate in any way in the reaction, and it may be omitted from the final balanced equation:

$$BaCO_3 + 2H^+ \rightarrow Ba^{2+} + H_2O + CO_{2(g)}$$

Such equations often emphasize the generality of a reaction. This last one makes it obvious that *any* strong acid will dissolve $BaCO_3$ to produce CO_2 and H_2O.

There are times when the conditions of an experiment determine whether a reaction goes. In the reaction

$$CdS + 2H^+ = Cd^{2+} + H_2S_{(g)}$$

CdS dissolves in an excess of strong acid because the equilibrium is shifted to the right and the H_2S is evolved to the atmosphere. However, if we should saturate an aqueous solution of Cd^{2+} (with no added acid) with H_2S, we should actually obtain a precipitate of CdS. The reduced acidity shifts the equilibrium to the left, as does maintaining a constant high pressure of H_2S.

Remember that a weak electrolyte does not have to be a molecule; it may also be a weakly dissociated ion. For example, $Ca_3(PO_4)_2$ dissolves in an excess of strong acid, because the weak acid ion HPO_4^{2-} (as well as some $H_2PO_4^-$) is formed by the reaction

$$Ca_3(PO_4)_2 + 2H^+ \rightarrow 3Ca^{2+} + 2HPO_4^{2-}$$

In general, you could expect the insoluble salt of a weak acid to dissolve in an excess of strong acid for the reasons just given (extremely insoluble substances, such as some of the metal sulfides, are exceptions).

The last two examples illustrate how precipitation or solution of an insoluble compound may be controlled by controlling the *p*H.

SUMMARY OF PREDICTION PROCEDURE

1. First, decide whether an electron-transfer reaction is possible, using approximate half-reaction potentials. A reducing agent and an oxidizing agent must be present for electron transfer; one alone is not enough. If one of the reactants is a metal, it *must* be a reducing agent if anything.
2. If the reaction is non-electron-transfer, then treat it as a metathesis reaction.
   (a) Write a preliminary balanced molecular equation in which the reaction partners have been interchanged:
   $$A_2B + 2CD \rightarrow 2AC + BD_2$$
   (b) Study each compound in the molecular equation and decide whether it is an insoluble solid, a weak (or non-) electrolyte, or a gas. Rewrite the equation as is appropriate. If A_2B is a strong electrolyte, CD a strong electrolyte, AC a gas, and BD_2 an insoluble solid, then
   $$2A^+ + B^- + 2CD = 2AC_{(g)} + BD_{2(s)}$$
   Eliminate all entities that are identical in form on both sides to get the final balanced ionic equation.
   (c) The reaction will go if any *one* of the products is a gas, a solid, or a weak electrolyte. If there is competition (insoluble solids or weak electrolytes on *both* sides), you must consider (1) whether one of the reactants is used in excess or high concentration to favor a shift to the right, and (2) the relative insolubility of solids or weakness of electrolytes.