CHEMICAL REACTIONS

Some Common Types of Chemical Reactions

1. When two elements react, a combination reaction occurs (think: could any other type of reaction occur?), producing a binary compound (that is, one consisting of only two types of atoms). If a metal and a nonmetal react, the product is ionic with a formula determined by the charges on the ions the elements form. If two nonmetals react, the product is a molecule with polar covalent bonds, with a formula consistent with the normal valences of the atoms involved. Some pairs of elements may react only slowly and require heating for significant reaction to occur.

   Examples

   $K + S_8 \rightarrow K_2S$ (ionic)

   $Ca + O_2 \rightarrow CaO$ (ionic)

   $Al + I_2 \rightarrow AlI_3$ (ionic)

   $H_2 + 0_2 \rightarrow H_2O$ (covalent)

   $I_2 + Cl_2 \rightarrow ICl, ICl_3$, or ICl_5 (covalent)

   [exact product depends on relative amounts of I_2 and Cl_2]

(NOTE: The above reactions are not balanced, nor were they intended to be. They, like the others in this handout, are meant only to show the correct formulae for the reactants and products. You may wish to balance the reactions in the handout as an exercise.)

2. Reaction of a metal oxide with water produces a metal hydroxide; that is, a strong base. Reaction of a nonmetal oxide with water produces and oxyacid in which the nonmetal is in the same oxidation state as in the oxide you started with. Both of these are combination reactions, and both can be reversed by heating the products. Metal hydroxides decompose on heating to give the metal oxide and water, and oxyacids decompose on heating to give water and the nonmetal oxide in the appropriate oxidation state.

   Examples

$Na_2O + H_2O \rightarrow NaOH$	$MgO + H_2O \rightarrow Mg(OH)_2$
$SO_2 + H_2O \rightarrow H_2SO_3$	$Cl_2O_5 + H_2O \rightarrow HClO_3$
$HNO_3 \xrightarrow{\Delta} N_2O_5 + H_2O$	$Fe(OH)_3 \xrightarrow{\Delta} Fe_2O_3 + H_2O$

3. Reaction of a metal oxide with a nonmetal oxide gives an oxysalt; reaction of a metal hydroxide with a nonmetal oxide produces a "hydrogen" oxysalt. This is essentially a reaction of the O^{2-} or OH^- in the metal compound with the molecular nonmetal oxide. This combination reaction occurs only if no water is present; in the presence of water, the nonmetal and metal oxides react with the water to produce acid and hydroxide, respectively (as shown in (2) above), then these react as in (4) below.

   Examples

   $CaO(s) + SO_3(g) \rightarrow CaSO_4(s)$

   $NaOH(s) + CO_2(g) \rightarrow NaHCO_3(s)$

CHEMICAL REACTIONS

4. Reaction of an acid with a base gives a salt plus water. The cation in the salt comes from the base; the anion comes from the acid. The base may be a metal hydroxide, a metal oxide, or a weak base such as NH3. The acid and/or base may be pure solids, liquids, or gases, or in aqueous solution. The oxidation states of the anion of the acid and cation of the base normally remain unchanged.

   Examples

   $HCl(aq) + Ca(OH)_2(aq) \rightarrow CaCl_2(aq) + H_2O(l)$

   $H_2SO_4(aq) + Fe(OH)_3(s) \rightarrow Fe_2(SO_4)_3(aq) + H_2O(l)$

   $NH_3(g) + HC_2H_3O_2(l) \rightarrow NH_4C_2H_3O_2(s)$

   $Al_2O_3(s) + HClO_4(aq) \rightarrow Al(ClO_4)_3(aq) + H_2O(l)$

5. Ammonium salts react with metal hydroxides and oxides in an acid-base reaction to produce ammonia. This is essentially the reverse of one of the reaction types mentioned in (4) above. Either one or both of the reactants may be a pure material or in aqueous solution.

   Examples

   $NH_4Cl(aq) + KOH(aq) \rightarrow NH_3(g) + H_2O(l) + KCl(aq)$

   $NH_4NO_3(s) + CaO(s) \rightarrow NH_3(g) + H_2O(l) + Ca(NO_3)_2(s)$

6. Reaction of the salt of a weak acid (that is, a compound containing the anion of a weak acid) with a strong acid produces the weak acid and a salt. This is another example of an acid-base reaction, in addition to the ones give in (4) and (5) above. The original salt of the weak acid may be either a pure solid or in aqueous solution. The cation in the salt formed as the product comes from the weak acid salt; the anion in the product salt comes from the strong acid. In many cases, the weak acid produced is unstable and decomposes to give the oxide of a nonmetal and water (see (2) above). This is especially true if the nonmetal oxide is a compound of limited solubility in water such as SO_2, CO_2, or the nitrogen oxides. The best-known examples of this type of reaction involve carbonates, bicarbonates, sulfides, and sulfites, but many other examples are known as well. Normally, these reactions do not involve oxidation or reduction.

   Examples

   $BaCo_3(s) + HBr(aq) \rightarrow BaBr_2(aq) + H_2O(l) + Co_2(g)$

   $NaHCO_3(aq) + H_2SO_4(aq) \rightarrow Na_2SO_4(aq) + CO_2(g) + H_2O(l)$

   $MgS(s) + HCl(aq) \rightarrow H_2S(g) + MgCl_2(aq)$

   $K_2SO_3(aq) + HNO_3(aq) \rightarrow KNO_3(aq) + SO_2(g) + H_2O(l)$

   $Ca_3(PO_4)_2(s) + HCl(aq) \rightarrow CaCl_2(aq) + H_3PO_4(aq)$

   $Zn(C_2H_3O_2)_2(aq) + HBr(aq) \rightarrow ZnBr_2(aq) + HC_2H_3O_2(aq)$

<u>CHEMICAL REACTIONS</u>

7. Reaction of solutions of <u>two soluble salts</u> with one another can give a <u>precipitate</u> of an insoluble salt formed by a double replacement reaction (also called a <u>metathesis</u>). Whether or not a precipitate forms depends on the exact combination of salts used. To make a prediction as to whether a reaction will take place or not, <u>you must know the solubility rules</u> for common salts. Some combinations of salts may give oxidation-reduction reactions (see (11) below), but most do not.

   Examples

   $CaCl_2(aq) + K_2CO_3(aq) \rightarrow CaCO_3(s) + KCl(aq)$

   $AgNO_3(aq) + FeCl_3(aq) \rightarrow AgCl(s) + Fe(NO_3)_3(aq)$

   But: $NiSO_4(aq) + MgI_2(aq) \rightarrow$ no reaction

   [NiI_2 and $MgSO_4$ are both soluble]

   $Al(NO_3)_3(aq) + Pb(C_2H_3O_2)_2(aq) \rightarrow$ no reaction

   [$Al(C_2H_3O_2)_3$ and $Pb(NO_3)_2$ are both soluble]

8. <u>Heating an oxysalt</u> produces a <u>metal oxide plus a nonmetal oxide</u> or a <u>metal salt plus oxygen</u>, or some combination of these two decomposition reactions.

   Examples

   $KClO_3(s) \xrightarrow{\Delta} KCl(s) + O_2(g)$

   $CaCO_3(s) \xrightarrow{\Delta} CaO(s) + CO_2(g)$

   $Pb(NO_3)_2(s) \xrightarrow{\Delta} PbO(s) + NO_2(g) + O_2(g)$

9. <u>Heating a hydrated material</u> initially causes a decomposition reaction to produce the anhydrous compound and water. Further heating may yield further decomposition, depending on the material. (See (2) and (8) above.) Most binary compounds are stable to heat.

   Examples

   $H_2C_2O_4 \cdot 2H_2O(s) \xrightarrow{\Delta} H_2O(g) + H_2C_2O_4(s)$; followed by

   $H_2C_2O_4(s) \xrightarrow{\Delta} H_2O(g) + CO(g) + CO_2(g)$

   $CaCl_2 \cdot 6H_2O(s) \xrightarrow{\Delta} H_2O(g) + CaCl_2(s)$; followed by

   $CaCl_2(s) \xrightarrow{\Delta}$ no reaction

   $CuSO_4 \cdot 5H_2O(s) \xrightarrow{\Delta} H_2O(g) + CuSO_4(s)$; followed by

   $CuSO_4(s) \xrightarrow{\Delta} CuO(s) + SO_3(g)$ (requires strong heating)

10. Reaction of an <u>element with a compound</u> often gives a <u>single replacement</u> reaction in which a nonmetallic element <u>can</u> replace a combined nonmetal, and a metallic element <u>can</u> replace a combined metal, or hydrogen from an acid. As a general rule, a more active (reactive) element will replace a less active (reactive) element from its compounds. In general (but with many exceptions), the most reactive nonmetals are found to the upper right in the periodic table, and the

most reactive metals are found to the lower left. The order of reactivity of the halogens is $F_2>Cl_2>Br_2>I_2$. For hydrogen and the more common metals, the order of reactivity (the activity series) is
$Li>K>Ca>Na>Mg>Al>Zn>Cr>Fe>Ni>Sn>Pb>H_2>Cu>Hg>Ag>Pt>Au$

In these two series, one element can replace another one to its right in the series. Metals to the left of H_2 can replace H^+ from acids. The very reactive metals (Li, K, Na, Ca) can replace H^+ from cold water; metals of intermediate reactivity (Mg, Al) can replace H^+ from hot water or steam. Any single replacement reaction can also be categorized as an oxidation-reduction (redox) reaction.

Examples

$Al(s) + NiSO_4(aq) \rightarrow Al_2(SO_4)_3(aq) + Ni(s)$

$Fe(s) + HBr(aq) \rightarrow FeBr_3(aq) + H_2(g)$

$Cl_2(g) + KI(aq) \rightarrow KCl(aq) + I_2(s)$

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

$Zn(s) + Cu(NO_3)_2(aq) \rightarrow Cu(s) + Zn(No_3)_2(aq)$

But: $Ag(s) + HClO_4(aq) \rightarrow$ no reaction

$Br_2(l) + ZnCl_2(aq) \rightarrow$ no reaction

$Sn(s) + H_2O(l) \rightarrow$ no reaction

$Pb(s) + CrF_3(aq) \rightarrow$ no reaction

11. Compounds containing one or more atoms in <u>high oxidation states</u> often act as <u>oxidizing agents</u>; compounds containing atoms in <u>low oxidation states</u> often act as <u>reducing agents</u>. For most elements, the (old) group number of the atom in the periodic table gives the <u>highest</u> oxidation state possible for that element. For nonmetals, the <u>lowest</u> oxidation state possible is given by the (old) <u>group number minus eight</u>. Elemental metals most often act as reducing agents (they are oxidized); nonmetals frequently act as oxidizing agents (they are reduced).

    For the <u>representative elements</u> (i.e., those in the first two and last six columns of the periodic table), oxidation states most often are two units apart. For example, Sn forms Sn(II) and Sn(IV); Br forms Br^{1-}, Br(I), Br(III), Br(V), and Br(VII). For the <u>transition elements</u>, (i.e., those in the "center" ten columns of the periodic table), oxidation states are often one unit apart, but can be in almost any relationship to one another. For the transition elements, the common oxidation states (charges on their ions) must be memorized. For example, Fe forms Fe^{2+} and Fe^{3+}; Cu forms Cu^+ and Cu^{2+}, etc. Some of the transition elements form oxyanions as well as cations. For example, Mn forms Mn^{2+}, Mn^{3+}, MnO_4^{2-}, and MnO_4^-; Cr forms Cr^{2+}, Cr^{3+}, CrO_4^{2-}, and $Cr_2O_7^{2-}$.

CHEMICAL REACTIONS

Any atom in its highest possible oxidation state can only act as an oxidizing agent; any atom in its lowest possible oxidation state can only act as a reducing agent. Atoms in intermediate oxidation states can be either oxidized or reduced; that is, they can act as either reducing or oxidizing agents. Some of the oxidizing agents most commonly encountered are MnO_4^-, CrO_4^{2-}, $Cr_2O_7^{2-}$, HNO_3, H_2O_2, and the halogens. Some of the more common reducing agents are elemental H_2, metals, carbon, and I^-.

In predicting products of oxidation-reduction reactions, don't forget their name – oxidation and reduction must occur simultaneously! It is impossible for oxidation to occur without reduction or vice versa.

Examples

$Sn^{2+}(aq) + F_2(g) \rightarrow Sn^{4+}(aq) + F^-(aq)$

$Mn^{2+}(aq) + BiO_3^-(aq) \rightarrow Bi^{3+}(aq) + MnO_4^-(aq)$

[note that the Bi is in its highest possible oxidation state in BiO_3^-]

$K(s) + P_4O_{10}(s) \rightarrow K_3PO_3(s)$

[note that P is reduced from P(V) to P(III)]

$MnO_4^-(aq) + I^-(aq) \rightarrow Mn^{2+}(aq) + I_2(aq)$

$CuS(s) + HNO_3(aq) \rightarrow Cu(NO_3)_2(aq) + S_8(s) + NO_2(g)$

[note $S^{2-} \rightarrow S^0$ and N(V) $\rightarrow$ N(IV)]

$Fe_2O_3(s) + C(s) \rightarrow CO_2(g) + Fe(s)$