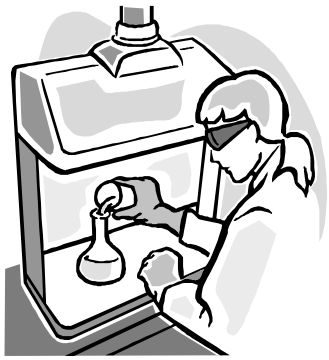

3.16 Oxidizing and Reducing Agents

- **Oxidizing agent** - causes oxidation (removes electrons from the species being oxidized)
  - is the species **being reduced**
  - contains an element which undergoes a decrease in Ox# (reduction)
- **Reducing agent** - causes reduction (supplies electrons to the species being reduced)
  - is the species **being oxidized**
  - contains an element which undergoes an increase in Ox# (oxidation)

- Identification of Ox/Red agents - need to examine the Ox# of all elements in the reaction

Example: Identify the Ox. and Red. agents in the reaction of **Cu** with hot, concentrated H_2SO_4.

$$Cu(s) + 2H_2SO_4(aq,conc.) \rightarrow$$
$$\rightarrow Cu^{2+} + SO_4^{2-} + SO_2(g) + 2H_2O(l)$$

$$\overset{0}{Cu}(s) + 2\overset{+1\ +6\ -2}{H_2SO_4}(aq,conc.) \rightarrow$$
$$\rightarrow \overset{+2}{Cu^{2+}} + \overset{+6\ -2}{SO_4^{2-}} + \overset{+4\ -2}{SO_2}(g) + 2\overset{+1\ -2}{H_2O}(l)$$

Cu $\Rightarrow$ 0 (in Cu) $\rightarrow$ +2 (in Cu^{2+})

S $\Rightarrow$ +6 (in H_2SO_4) $\rightarrow$ +4 (in SO_2)

$\Rightarrow$Cu is oxidized $\Rightarrow$ Cu is the reducing agent

$\Rightarrow$S in H_2SO_4 is reduced $\Rightarrow$ H_2SO_4 is the oxidizing agent

$\Rightarrow$$H_2SO_4$ oxidizes Cu; Cu reduces H_2SO_4

3.17 Balancing Simple Redox Equations

- Charge balance in redox equations
  - electrons lost in oxidation must be gained in reduction
  - total charge of reactants must equal total charge of products

$Zn(s) + Ag^+ \rightarrow Zn^{2+} + Ag(s)$

the mass is balanced, but the charge is not balanced

$Zn(s) + 2Ag^+ \rightarrow Zn^{2+} + 2Ag(s)$

Example:

- Write the net ionic equation for the oxidation of iodide ions to iodine by cerium(IV) ions which are reduced to cerium(III) ions.

$Ce^{4+} + I^- \rightarrow Ce^{3+} + I_2(s)$

$\Rightarrow$mass balance:

$Ce^{4+} + 2I^- \rightarrow Ce^{3+} + I_2(s)$

$\Rightarrow$charge balance:

$2Ce^{4+} + 2I^- \rightarrow 2Ce^{3+} + I_2(s)$

3.18 Classifying Reactions

- Redox
  - change in Ox# of elements
  - products
    - carbon dioxide and water for combustion
    - ionic compound for metal + nonmetal
- Neutralization
  - transfer of protons from an acid to a base
  - products are salt (cation from base, anion from acid) and/or water (sometimes gas formation)

- Precipitation
  - exchange of ions between two soluble salts
  - formation of an insoluble product
  - solubility rules

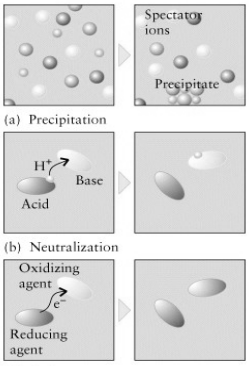

Fig. 3.33

- **Example:** Balance the equation of the reaction between ammonia gas and and copper(II) oxide with products nitrogen, copper and liquid water. Identify the reaction type.

⇒**skeletal equation:**

$NH_3(g) + CuO(s) \rightarrow N_2(g) + Cu(s) + H_2O(l)$

⇒**balanced equation:**

$2NH_3(g) + CuO(s) \rightarrow N_2(g) + Cu(s) + H_2O(l)$

$2NH_3(g) + CuO(s) \rightarrow N_2(g) + Cu(s) + 3H_2O(l)$

$2NH_3(g) + 3CuO(s) \rightarrow N_2(g) + Cu(s) + 3H_2O(l)$

$2NH_3(g) + 3CuO(s) \rightarrow N_2(g) + 3Cu(s) + 3H_2O(l)$

⇒**formation of elements (check Ox#):**

$2NH_3(g) + 3CuO(s) \rightarrow N_2(g) + 3Cu(s) + 3H_2O(l)$

-3 +2 0 0

⇒**change in (Ox#) → Redox reaction**

Example: Predict the products of the reaction between sodium carbonate and hydrobromic acid.

⇒**HBr is a strong acid, Na_2CO_3 is a salt → probably a proton transfer reaction**

⇒ **Na_2CO_3 is a salt of a weak acid → gas formation**

$2HBr(aq) + Na_2CO_3(aq) \rightarrow NaBr(aq) + H_2CO_3(aq)$

$H_2CO_3(aq) \rightarrow CO_2(g) + H_2O(l)$

⇒**gas formation reaction**

$2HBr(aq) + Na_2CO_3(aq) \rightarrow 2NaBr(aq) + CO_2(g) + H_2O(l)$

$2H^+ + 2Br^- + 2Na^+ + CO_3^{2-} \rightarrow 2Na^+ + 2Br^- + CO_2(g) + H_2O(l)$

⇒**net ionic equation:**

$2H^+ + CO_3^{2-} \rightarrow CO_2(g) + H_2O(l)$

⇒**consider the hydronium ions:**

$2H_3O^+ + CO_3^{2-} \rightarrow CO_2(g) + 3H_2O(l)$

Assignments:

- **Homework:** Chpt. 3/3, 5, 7, 11, 15, 19, 23, 27, 31, 33, 35, 37, 39, **45, 49, 51, 53, 59, 61, 73**
- **Student companion:** 3.1, 3.4, 3.5, 3.6