3.10 Neutralization

- Neutralization reactions

acid + base → salt + water(or other products)

– salt - an ionic compound with a cation from the base and an anion from the acid

$H_2SO_4(aq) + 2KOH(aq) \rightarrow K_2SO_4(aq) + 2H_2O(l)$

Example: Predict the products of the reaction between carbonic acid and calcium hydroxide.

$H_2CO_3(aq) + Ca(OH)_2(aq) \rightarrow CaCO_3(s) + 2H_2O(l)$

3.11 Proton Transfer

- Net ionic equations for reactions between strong acids and bases

$HCl(aq) + KOH(aq) \rightarrow KCl(aq) + H_2O(l)$

$H^+ + Cl^- + K^+ + OH^- \rightarrow K^+ + Cl^- + H_2O(l)$

$\Rightarrow H^+ + OH^- \rightarrow H_2O(l)$

– H^+ is present in the form of H_3O^+

$\Rightarrow H_3O^+ + OH^- \rightarrow 2H_2O(l)$

⇒**net ionic equation for all strong acid/strong base reactions (transfer of a proton from H_3O^+ to OH^-)**

- Net ionic equations for reactions between weak acids and strong bases

Example:

$HF(aq) + NaOH(aq) \rightarrow NaF(aq) + H_2O(l)$

HF(aq) → weak acid (only partially ionized)

$HF(aq) + Na^+ + OH^- \rightarrow Na^+ + F^- + H_2O(l)$

$\Rightarrow HF(aq) + OH^- \rightarrow F^- + H_2O(l)$

⇒**transfer of a proton from HF to OH^-**

- Net ionic equations for reactions between strong acids and weak bases

Example:

$HCl(aq) + NH_3(aq) \rightarrow NH_4Cl\ (aq)$

NH_3(aq) → weak base (only partially ionized)

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

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

– H^+ is present in the form of H_3O^+

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

⇒**transfer of a proton from H_3O^+ to NH_3**

3.12 Acidic and Basic Character in the Periodic Table

- Basic oxides - react with water to form bases

– most soluble metal oxides

$BaO(s) + H_2O(l) \rightarrow Ba(OH)_2(aq)$

– react with acids to form salts and water

$CaO(s) + 2HNO_3(aq) \rightarrow Ca(NO_3)_2(aq) + H_2O(l)$

– elements located in the lower left corner of the table

- Acidic oxides - react with water to form acids
  - most nonmetal oxides

  $SO_3(g) + H_2O(l) \rightarrow H_2SO_4(aq)$

  - react with bases

  $CO_2(g) + 2KOH(aq) \rightarrow K_2CO_3(aq) + H_2O(l)$

  - elements located in the upper right corner of the table
- Amphoteric oxides - have both acidic and basic properties
  - some metalloid oxides + oxides of Be, Al, Ga, Sn, Pb, Bi
  - react with both acids and bases
  - elements form a diagonal band in the table

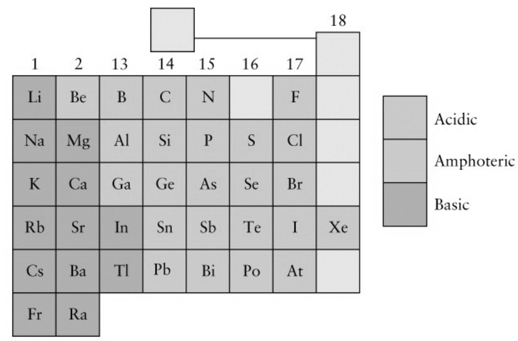

Fig. 3.23

3.13 Gas Formation Reactions

- Reactions of salts of weak or volatile acids with strong acids

Example:

$ZnS(s) + 2HCl(aq) \rightarrow ZnCl_2(aq) + H_2S(g)$

$ZnS(s) + 2H^+ + 2Cl^- \rightarrow Zn^{2+} + 2Cl^- + H_2S(g)$

$\Rightarrow ZnS(s) + 2H^+ \rightarrow Zn^{2+} + H_2S(g)$

- **H^+ is present in the form of H_3O^+**

$\Rightarrow ZnS(s) + 2H_3O^+ \rightarrow Zn^{2+} + H_2S(g) + 2H_2O(l)$

$\Rightarrow$ **transfer of a proton from H_3O^+ to S^{2-}**

Redox Reactions

3.14 Oxidation and reduction

- Transfer of electrons from one species to another

$2Na(s) + Cl_2(g) \rightarrow 2NaCl(s)$

NaCl(s) consists of ions:

$2Na(s) + Cl_2(g) \rightarrow 2Na^+(s) + 2Cl^-(s)$

$Na(s) \rightarrow Na^+(s)$ $\Rightarrow$ **loss of 1e$^-$ by Na**

$Cl_2(g) \rightarrow 2Cl^-(s)$ $\Rightarrow$ **gain of 2e$^-$ by Cl_2**

Result: transfer of electrons from Na to Cl_2

- **Oxidation** - loss of electrons (Na is oxidized)
  - term originates from reactions of substances with oxygen
- **Reduction** - gain of electrons (Cl_2 is reduced)
  - term originates from reactions of metal oxides with C, CO, H_2, etc. to extract (reduce) the pure metal
- Oxidation and reduction can not occur independently
  - electrons gained by one species must be lost by another (e$^-$ gained by Cl_2 are lost by Na)

3.15 Oxidation Numbers

- Oxidation number (Ox#) is assigned to each element in a substance
- Oxidation numbers can help determine whether substances are oxidized or reduced
  - **oxidation** - increase in Ox#
  - **reduction** - decrease in Ox#

$Na(s) \rightarrow Na^+(s)$ $\Rightarrow$ **Ox# increases (0** $\rightarrow$ +1)

$Cl_2(g) \rightarrow 2Cl^-(s)$ $\Rightarrow$ **Ox# decreases (0** $\rightarrow$ -1)

– for monoatomic ions → **Ox# = charge of ion**
– for free elements → **Ox# = 0**
– for **F** → **Ox# = -1**
– for **O** → **Ox# = -2** (except in combination with F and in peroxides)
– for **H** → **Ox# = +1** (in combination with nonmetals)
→ **Ox# = -1** (in combination with metals)
– for **halogens** → **Ox# = -1** (except in comb. with O or other halogen higher in the group)
– the **sum of Ox#** of all elements in a species equals the **charge** of the species

Example:

- Assign the oxidation numbers of all elements in NO_3^- and $HClO_3$.

NO_3^- ⇒ **O (-2) by rule**

$3\times(-2) + 1\times(X) = -1 \Rightarrow X = +5 \Rightarrow$ **N (+5)**

$HClO_3$ ⇒ **O (-2) by rule**

⇒ **H (+1) by rule**

$3\times(-2) + 1\times(+1) + 1\times(X) = 0 \Rightarrow X = +5 \Rightarrow$ **Cl (+5)**