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411_Chapter 4 OVERVIEW NOTES

Metathesis

$AX + BY \rightarrow AY + BX$ For RXN to occur AY or BX must = (s) (l) (g)

Soluble	Insoluble
NO_3^-	CO_3 (except ï IA, NH_4)
$C_2H_3O_2^-$	CrO_4 (except ï IA, NH_4, Ca, Sr)
ClO_3^-	OH (except ï IA, NH_4, Ca, Sr, Ba)
Cl (except Hg, Pb, Ag)	PO_4 (except ï IA, NH_4)
Br (except Hg, Pb, Ag)	SO_3 (except ï IA, NH_4)
I (except Hg, Pb, Ag)	S (except ï IA, IIA, NH_4)
SO_4 (except Hg, Pb, Ag, Sr, Ba)	

TABLE 4.2 ■ Common Strong Acids and Bases

Strong Acids	Strong Bases
Hydrochloric, HCl	Group 1A metal hydroxides (LiOH, NaOH, KOH, RbOH, CsOH)
Hydrobromic, HBr	Heavy group 2A metal hydroxides [$Ca(OH)_2$, $Sr(OH)_2$, $Ba(OH)_2$]
Hydroiodic, HI	
Chloric, $HClO_3$	
Perchloric, $HClO_4$	
Nitric, HNO_3	
Sulfuric, H_2SO_4	

Neutralization RXN: Acid + Base $\rightarrow$ Salt + H_2O

Acid-Base forms gas: Acid + Base $\rightarrow$ gas

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Acids-
Bases-
Difference between a strong and weak acid and base-
Difference between strong and weak electrolytes-

TABLE 4.3 ■ Summary of the Electrolytic Behavior of Common Soluble Ionic and Molecular Compounds

	Strong Electrolyte	Weak Electrolyte	Nonelectrolyte
Ionic	All	None	None
Molecular	Strong acids (see Table 4.2)	Weak acids Weak bases	All other compounds

$$AS + HY \rightarrow AY + H_2S(g)$$

$$ACO3 + HY \rightarrow AY + CO_2(g) + H_2O(l)$$

$$ASO3 + HY \rightarrow AY + SO_2(g) + H_2O(l)$$

$$NH4X + BOH \rightarrow BX + NH_3(g) + H_2O(l)$$

$$AX + B \rightarrow BX + A$$

Most - Li - K ï Ba ï Ca ï Na ï Mg ï Al ï Mn ï Zn ï Cr ï Fe ï Co ï Ni ï Sn ï Pb ï **H** ï Cu ï Ag ï Hg ï Pt - Au ï **Least**

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TABLE 4.1 ■ Solubility Guidelines for Common Ionic Compounds in Water

Soluble Ionic Compounds		Important Exceptions
Compounds containing	NO_3^-	None
	CH_3COO^-	None
	Cl^-	Compounds of Ag^+, Hg_2^{2+}, and Pb^{2+}
	Br^-	Compounds of Ag^+, Hg_2^{2+}, and Pb^{2+}
	I^-	Compounds of Ag^+, Hg_2^{2+}, and Pb^{2+}
	SO_4^{2-}	Compounds of Sr^{2+}, Ba^{2+}, Hg_2^{2+}, and Pb^{2+}

Insoluble Ionic Compounds		Important Exceptions
Compounds containing	S^{2-}	Compounds of NH_4^+, the alkali metal cations, and Ca^{2+}, Sr^{2+}, and Ba^{2+}
	CO_3^{2-}	Compounds of NH_4^+ and the alkali metal cations
	PO_4^{3-}	Compounds of NH_4^+ and the alkali metal cations
	OH^-	Compounds of the alkali metal cations, and NH_4^+, Ca^{2+}, Sr^{2+}, and Ba^{2+}

Solubility cutoff- 0.01mol a liter or 0.01M

Metathesis reaction-

Ionic equation- example each:

- Molecular-
- Complete ionic-
- Net ionic-
- Spectator ions-

Oxidation Reduction RXN-

Oxidation-

Reduction-

Oxidation Numbers-

We use the following rules for assigning oxidation numbers:

1. *For an atom in its **elemental form**, the oxidation number is always zero.* Thus, each H atom in the H_2 molecule has an oxidation number of 0, and each P atom in the P_4 molecule has an oxidation number of 0.
2. *For any **monatomic ion** the oxidation number equals the charge on the ion.* Thus, K^+ has an oxidation number of +1, S^{2-} has an oxidation number of −2, and so forth. The alkali metal ions (group 1A) always have a 1+ charge, and therefore the alkali metals always have an oxidation number of +1 in their compounds. Similarly, the alkaline earth metals (group 2A) are always +2, and aluminum (group 3A) is always +3 in its compounds. (In writing oxidation numbers, we will write the sign before the number to distinguish them from the actual electronic charges, which we write with the number first.)
3. ***Nonmetals*** usually have negative oxidation numbers, although they can sometimes be positive:
   - **(a)** *The oxidation number of **oxygen** is usually* −2 in both ionic and molecular compounds. The major exception is in compounds called peroxides, which contain the O_2^{2-} ion, giving each oxygen an oxidation number of −1.
   - **(b)** *The oxidation number of **hydrogen** is usually +1 when bonded to nonmetals and −1 when bonded to metals.*
   - **(c)** *The oxidation number of **fluorine** is −1 in all compounds.* The other ***halogens*** have an oxidation number of −1 in most binary compounds. When combined with oxygen, as in oxyanions, however, they have positive oxidation states.
4. ***The sum of the oxidation numbers** of all atoms in a neutral compound is zero. The sum of the oxidation numbers in a polyatomic ion equals the charge of the ion.* For example, in the hydronium ion, H_3O^+, the oxidation number of each hydrogen is +1 and that of oxygen is −2. Thus, the sum of the oxidation numbers is 3(+1) + (−2) = +1, which equals the net charge of the ion. This rule is very useful in obtaining the oxidation number of one atom in a compound or ion if you know the oxidation numbers of the other atoms, as illustrated in Sample Exercise 4.8.

$$\underset{0}{Mg(s)} + 2\,\underset{+1\;-1}{HCl(aq)} \longrightarrow \underset{+2\;-1}{MgCl_2(aq)} + \underset{0}{H_2(g)} \qquad [4.26]$$