**Calculating with Kw**

**We can calculate the extent of the dissociation of water molecules starting from experimental measurements that show H3O+ concentration in pure water to be 1.0 x 10-7 M @ 25 C.**

**[H3O+] = 1.0 x 10-7 M @ 25 C.**

**Since the dissociation reaction of water produces equal concentrations of H3O+ and OH- ions, the OH- concentration in pure water is also 1.0 x 10-7 M @ 25 C.**

**[H3O+] = [OH-] = 1.0 x 10-7 M @ 25 C.**

**The molar concentration of pure water can be calculated using it’s density (.997 g/ml) and its Molar mass (18.0 g).**

**Molar Concentration of water [H2O] =**

**(997 g /L) x (1 mole / 18 g )=**

**55.4 mol / L @ 25 C.**

**The ratio of dissociated to undissociated molecules of pure water can be found using these facts (55.4 mol / L @ 25 C and 1.0 x 10-7 M @ 25 C).**

**[H2O]dissociated = 1.0 x 10-7**

**[H2O]undissociated 55.4 M =**

**1.8 x 10-9 about 2 in 109 (2 : 1000000000).**

**In addition, we can calculate Kw (ion – product constant for water).**

**Kw = [H3O+][OH-]**

1. **x 10-7) (1.0 x 10-7) = 1.0 x 10-14 @ 25C**

**In very dilute solutions, the water is almost a pure liquid and the product of the H3O+ and OH- concentrations are unaffected by the presences of a solute (dissolved solids).**

**This is not true in more concentrated solutions, but we’ll ignore that complication and assume that the product of the H3O+ and OH- concentrations is always 1.0 x 10-14 @25C in any aqueous solution.**

**You can determine if a solution is acidic, basic or neutral by comparing [H3O+] to [OH-] (brackets mean concentration).**

**Acidic [H3O+] > [OH-]**

**Neutral [H3O] = [OH-]**

**Basic [H3O] < [OH-]**

**Since Kw = 1.0 x 10-14**

**[H3O+] = 1.0 x 10-14 and [OH-] = 1.0 x 10-14**

**[OH-] [H3O+].**