The Brønsted theory encompasses any type of solvent that can donate and accept \( \text{(H}^+ \text{)} \) ions, not just aqueous solutions. The strength of an acid or a base varies depending on the solvent. Non-aqueous acid-base chemistry follows similar rules to those developed for acids and bases in water. For example in liquid ammonia, the solvent autodissociates in the reaction:

\[
\text{[2NH}_3\text{(l)} \rightleftharpoons \text{NH}_4^+ + \text{NH}_2^-]
\]

This equilibrium is analogous to the autodissociation of water, but has a smaller equilibrium constant (\( K \approx 10^{-30} \)). It follows by analogy to water that \( \text{NH}_4^+ \) is the strongest acid and \( \text{NH}_2^- \) is the strongest base that can exist in liquid ammonia. Because ammonia is a basic solvent, it enhances the acidity and suppresses the basicity of substances dissolved in it. For example, the ammonium ion (\( \text{NH}_4^+ \)) is a weak acid in water (\( K_a = 6 \times 10^{-10} \)), but it is a strong acid in ammonia. Similarly, acetic acid is weak in water but strong in ammonia. Solvent leveling in fact makes \( \text{CH}_3\text{COOH} \) and \( \text{NH}_4\text{Cl} \) both strong acids in ammonia, where they have equivalent acid strength.

Strong acids that are leveled in water have different acid strengths in acidic solvents such as HF or anhydrous acetic acid. For example, acid dissociation of \( \text{HX} \) in acetic acid (\( \text{CH}_3\text{COOH} \)) involves protonating the solvent to make its conjugate acid (\( \text{CH}_3\text{COOH}_2^+ \)) and the \( \text{X}^- \) anion. Because \( \text{CH}_3\text{COOH}_2^+ \) is a stronger acid than \( \text{H}_3\text{O}^+ \), the anion \( \text{X}^- \) (which is a spectator in water) can become a weak base in \( \text{CH}_3\text{COOH} \):

\[
\text{[HX} + \text{CH}_3\text{COOH} \rightleftharpoons \text{CH}_3\text{COOH}_2^+ + \text{X}^-\]
\]

It follows that acidic solvents magnify the Brønsted basicities of substances that cannot accept protons in water. Conversely, basic solvents magnify the acidity of substances that cannot donate a proton to \( \text{(OH}^- \text{)} \).

The acidity and basicity of non-aqueous solvents is difficult to quantify precisely, but one good relative measure is the Hammett acidity function, \( \text{H}_0 \), which is defined analogously to pH according to the Henderson-Hasselbach equation:

\[
\text{H}_0 = pK_a + \log\left( \text{[base]} / \text{[conjugate acid]} \right)
\]

For non-aqueous solvents, or for acidic or basic compounds in dissolved in solvents that do not themselves dissociate, \( \text{H}_0 \) is a rough measure of the pH of the solvent or compound in question. Anhydrous HF and \( \text{H}_2\text{SO}_4 \) have \( \text{H}_0 \) values of approximately -10 and -12 respectively.