Acids and bases (Brønsted)

Definition of acids and bases given by Brønsted in 1923:
Acids: substances, donating protons
Bases: substances, accepting protons

Easy examples:
HCl, H_2SO_4 can donate a proton
→ acids
NH_3, OH^- can accept a proton
→ bases

Pairs of acids and bases:
Example:

$HCl + H_2O \rightleftharpoons H_3O^+ \quad Cl^-$

According to Brønsted´s definition:
HCl : acid Cl^- : conjugate base

H_2O : base H_3O^+ : conjugate acid

Therefore the acid/base pairs are HCl/Cl^- and H_3O^+/H_2O
For the protolyse reaction in thermal equilibrium one can formulate the law of mass action:

$$K = \frac{c(H_3O^+)c(Cl^-)}{c(H_2O)c(HCl)}$$

$$K_c(H_2O) = \frac{c(H_3O^+)c(Cl^-)}{c(HCl)}$$

$$K_a = \frac{c(H_3O^+)c(Cl^-)}{c(HCl)}$$

Because the concentration of water in an aqueous solution is nearly constant one can calculate a new constant K_a including the concentration of water. Because of the large numeric range it is useful to define the $pK_a = -\log K_a$. If the strength of an acid is high the strength of their conjugated base is low.

The strength of acids can be derived by their position in the periodic table:

With an increasing electronegativity the splitting of the bond becomes more easy, so that within a period the acidity increases as follows:

$NH_3 < H_2O < HF$

Within a group the increasing radius of the ions has a stronger effect than the electro-negativity. (better allocation of the charge on the increasing surface). Therefore the acidity increases as follows:

HF < HBr < HCl < HF

Literature:
Shriver, D. F., Atkins, P. W., Lanford, C. H., Anorganische Chemie (1992), VCH, Weinheim.