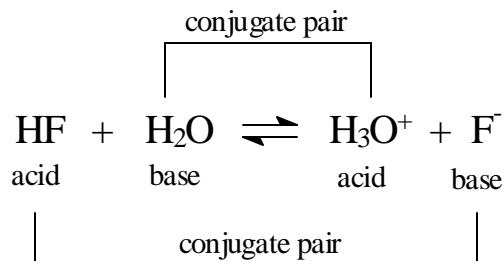
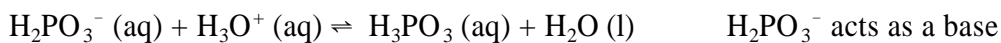
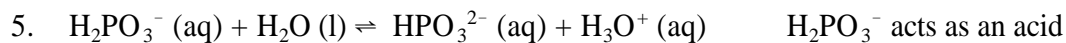


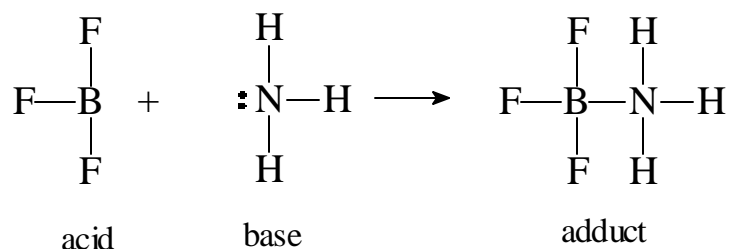
- Of the acids listed, HCl and HNO<sub>3</sub> are strong acids, and the remaining acids are weak.
- According to the Brønsted-Lowry concept, an acid is a species that donates a proton (H<sup>+</sup>) to another species in a proton transfer reaction, while a base is a species that accepts a proton (H<sup>+</sup>) from another species in a proton transfer reaction.



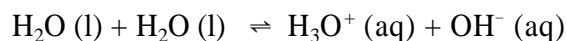
- The conjugate acid of a base is the chemical species which results when a proton (H<sup>+</sup>) is added to the base. In the equation above, H<sub>3</sub>O<sup>+</sup> is the conjugate acid of H<sub>2</sub>O and HF is the conjugate acid of F<sup>-</sup>.



- According to the Lewis concept, an acid is a species that accepts an electron pair while a base is a species that donates an electron pair. The equation below is an example of a Lewis acid-base reaction.



- The self-ionization or autoionization of water is a reaction between two water molecules in which a proton is transferred from one molecule to the other, as shown in the equation below.



The expression for  $K_w$  is:  $K_w = [\text{H}_3\text{O}^+][\text{OH}^-]$  and  $K_w = 1.0 \times 10^{-14}$  at 25°C.

11. The rigorous definition of pH is

$$\text{pH} = -\log a_{\text{H}^+}$$

where  $a_{\text{H}^+}$  = the hydrogen ion activity. We can approximate the hydrogen ion activity with the concentration of the hydronium ion, so we say

$$\text{pH} = -\log [\text{H}_3\text{O}^+]$$

pH may be measured by using a pH meter and electrodes or by using acid-base indicators.

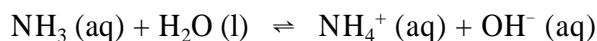
12. The lower the pH the more acidic the solution. A solution with pH = 4 is more acidic than a solution with pH = 5.
13. An aqueous solution at 37°C has  $[\text{H}_3\text{O}^+][\text{OH}^-] = 2.5 \times 10^{-14}$ .  
A neutral solution has  $[\text{H}_3\text{O}^+] = [\text{OH}^-]$ , so, at 37°C in a neutral solution

$$[\text{H}_3\text{O}^+] = \sqrt{2.5 \times 10^{-14}} = 1.6 \times 10^{-7} \text{ M} \text{ and } \text{pH} = -\log(1.6 \times 10^{-7}) = 6.80.$$

14.  $\text{pH} + \text{pOH} = \text{pK}_w$ . At 37°C,  $\text{K}_w = 2.5 \times 10^{-14}$  and  $\text{pK}_w = 13.60$ ,

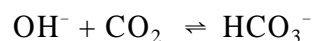
so  $\text{pH} + \text{pOH} = 13.60$ .

15. When  $\text{NH}_3$  reacts with water, the  $\text{NH}_3$  is the Brønsted base and the  $\text{H}_2\text{O}$  is the Brønsted acid as shown.



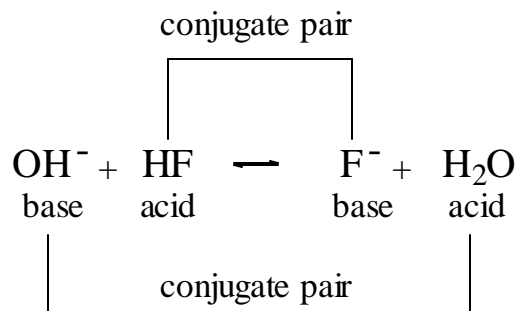
$\text{NH}_3$  does not fit the Arrhenius definition of a base, but it does fit the Brønsted-Lowry definition.

17. In this reaction,  $\text{OH}^-$  acts as a Lewis base (electron-pair donor) and the  $\text{CO}_2$  acts as the Lewis acid (electron-pair acceptor). The adduct (or product of the reaction) is the bicarbonate (hydrogen carbonate) ion.



19.  $\text{K}_w$  decreases with decreasing temperature, so the concentration of  $[\text{H}_3\text{O}^+]$  will decrease with decreasing temperature. As  $[\text{H}_3\text{O}^+]$  decreases, the value of its negative log increases, so the pH of pure water increases with decreasing temperature.

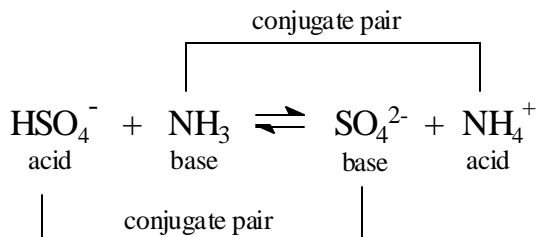
23.



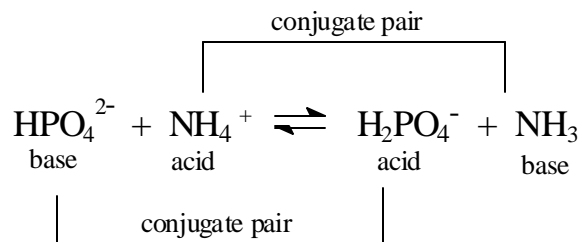
25. a. acid =  $\text{HPO}_4^{2-}$       conjugate base =  $\text{PO}_4^{3-}$   
 b. acid =  $\text{H}_2\text{S}$       conjugate base =  $\text{HS}^-$   
 c. acid =  $\text{HNO}_2$       conjugate base =  $\text{NO}_2^-$   
 d. acid =  $\text{HAsO}_4^{2-}$       conjugate base =  $\text{AsO}_4^{3-}$

27. a. base =  $\text{ClO}^-$       conjugate acid =  $\text{HClO}$   
 b. base =  $\text{AsH}_3$       conjugate acid =  $\text{AsH}_4^+$   
 c. base =  $\text{H}_2\text{PO}_4^-$       conjugate acid =  $\text{H}_3\text{PO}_4$   
 d. base =  $\text{TeO}_3^{2-}$       conjugate acid =  $\text{HTeO}_3^-$

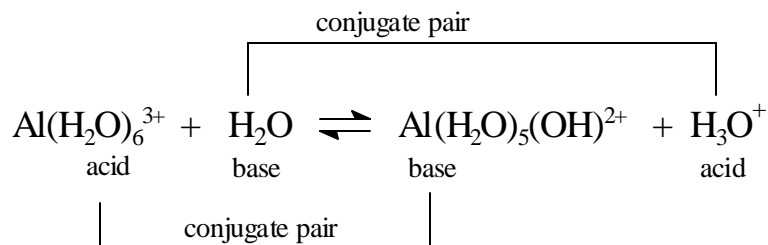
29. a.

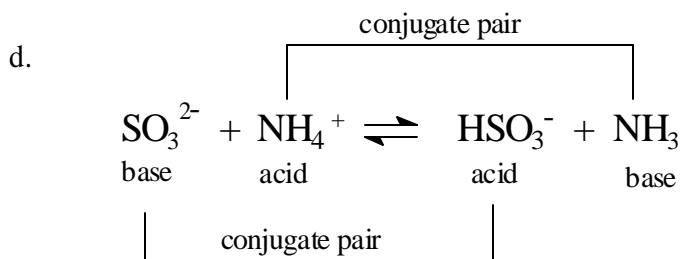


b.

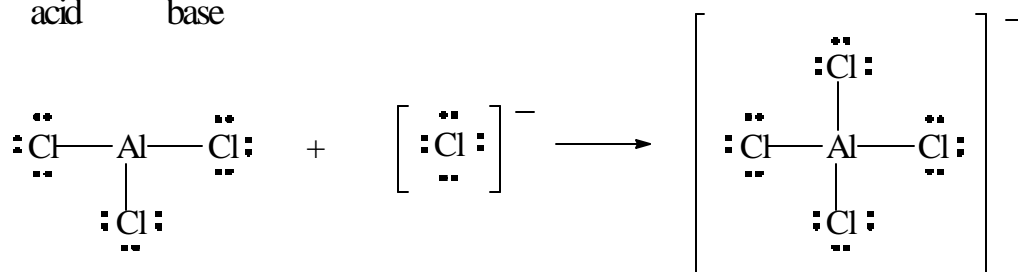
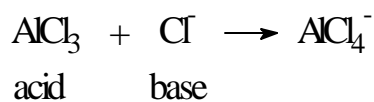


c.

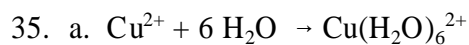
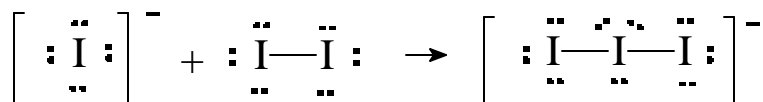
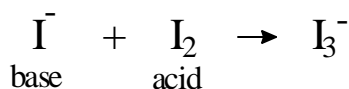




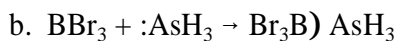
33. a.



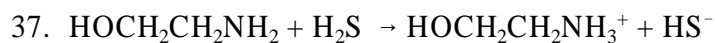
b.



Each water molecule donates an electron pair to  $\text{Cu}^{2+}$ .  $\text{H}_2\text{O}$  acts as the Lewis base while  $\text{Cu}^{2+}$  acts as the Lewis acid.

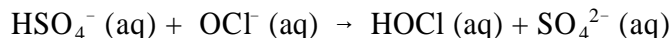


$\text{AsH}_3$  donates an electron pair, making it the Lewis base.  $\text{BBr}_3$  accepts the electron pair, making it the Lewis acid.

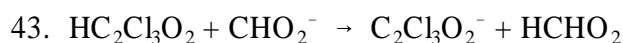


Here, an  $\text{H}^+$  from the  $\text{H}_2\text{S}$  can be thought of as the Lewis acid and the  $\text{HOCH}_2\text{CH}_2\text{NH}_2$  with the lone pair on N is the Lewis base.

39. Looking at the table of acid strengths,  $\text{HSO}_4^-$  is a stronger acid than  $\text{HOCl}$ , so the equation will go from left to right.



41. a.  $\text{H}_3\text{PO}_4$  is a stronger acid than  $\text{NH}_4^+$ , so the species on the left are favored.  
 b.  $\text{H}_2\text{S}$  is a stronger acid than  $\text{HCN}$ , so the species on the left are favored.  
 c.  $\text{HCO}_3^-$  is a stronger acid than  $\text{H}_2\text{O}$ , so the species on the right are favored.  
 d.  $\text{Al}(\text{H}_2\text{O})_6^{3+}$  is a stronger acid than  $\text{H}_2\text{O}$ , so the species on the right are favored.



Since the species on the right are favored, it follows that  $\text{HC}_2\text{Cl}_3\text{O}_2$  is a stronger acid than  $\text{HCHO}_2$ . In general, in an acid base equilibrium, the weaker acid and base are favored.

45. a.  $\text{H}_2\text{S}$  is stronger than  $\text{HS}^-$ . The conjugate base of an acid is always a weaker acid than the acid itself. For polyprotic acids, acid strength decreases as the acid is deprotonated.  
 b.  $\text{H}_2\text{SO}_3$  is a stronger acid than  $\text{H}_2\text{SeO}_3$ . For oxoacids which differ only in the identity of the principal or central atom, the acid strength increases with increasing electronegativity of the central atom. S is more electronegative than Se.  
 c.  $\text{HBr}$  is a stronger acid than  $\text{H}_2\text{Se}$ . For binary acids of elements in the same period, acid strength increases with increasing electronegativity. Br is more electronegative than Se.  
 d.  $\text{HIO}_4$  is a stronger acid than  $\text{HIO}_3$ . For oxoacids of the same element, acid strength increases with increasing number of oxygen atoms bonded to the central atom.  
 e.  $\text{H}_2\text{S}$  is stronger than  $\text{H}_2\text{O}$ . For binary acids of elements in the same family, acid strength increases as the H-X bond strength decreases. H-X bond strength decreases with increasing size of X.

47. a.  $[\text{H}_3\text{O}^+] = 1.2 \text{ M}$   $[\text{OH}^-] = \frac{K_w}{[\text{H}_3\text{O}^+]} = \frac{1.0 \times 10^{-14}}{1.2} = 8.3 \times 10^{-15} \text{ M}$

b.  $[\text{OH}^-] = 0.32 \text{ M}$   $[\text{H}_3\text{O}^+] = \frac{K_w}{[\text{OH}^-]} = \frac{1.0 \times 10^{-14}}{0.32} = 3.1 \times 10^{-14} \text{ M}$

c.  $[\text{OH}^-] = 2(0.085) \text{ M} = 0.17 \text{ M}$   $[\text{H}_3\text{O}^+] = \frac{K_w}{[\text{OH}^-]} = \frac{1.0 \times 10^{-14}}{0.17} = 5.9 \times 10^{-14} \text{ M}$

$$d. [\text{H}_3\text{O}^+] = 0.38 \text{ M} \quad [\text{OH}^-] = \frac{K_w}{[\text{H}_3\text{O}^+]} = \frac{1.0 \times 10^{-14}}{0.38} = 2.6 \times 10^{-14} \text{ M}$$

$$49. [\text{H}_3\text{O}^+] = 0.050 \text{ M} \quad [\text{OH}^-] = \frac{K_w}{[\text{H}_3\text{O}^+]} = \frac{1.0 \times 10^{-14}}{0.050} = 2.0 \times 10^{-13} \text{ M}$$

$$51. [\text{OH}^-] = 2(0.0085) \text{ M} = 0.017 \text{ M} \quad [\text{H}_3\text{O}^+] = \frac{K_w}{[\text{OH}^-]} = \frac{1.0 \times 10^{-14}}{0.017} = 5.9 \times 10^{-12} \text{ M}$$

53. We will assume the temperature is  $25^\circ\text{C}$  where  $K_w = 1.0 \times 10^{-14}$  and a neutral solution has  $[\text{H}_3\text{O}^+] = [\text{OH}^-] = 1.0 \times 10^{-7} \text{ M}$ .

- $[\text{H}_3\text{O}^+] > 1.0 \times 10^{-7} \text{ M}$  so the solution is acidic.
- $[\text{OH}^-] < 1.0 \times 10^{-7} \text{ M}$  so  $[\text{H}_3\text{O}^+] > 1.0 \times 10^{-7} \text{ M}$  and the solution is acidic.
- $[\text{OH}^-] = 1.0 \times 10^{-7} \text{ M}$  so  $[\text{H}_3\text{O}^+] = 1.0 \times 10^{-7} \text{ M}$  and the solution is neutral.
- $[\text{H}_3\text{O}^+] < 1.0 \times 10^{-7} \text{ M}$  so the solution is basic.

55.  $[\text{OH}^-] < 1.0 \times 10^{-7} \text{ M}$  so  $[\text{H}_3\text{O}^+] > 1.0 \times 10^{-7} \text{ M}$  and the solution is acidic.

$$[\text{H}_3\text{O}^+] = \frac{K_w}{[\text{OH}^-]} = \frac{1.0 \times 10^{-14}}{1.5 \times 10^{-9}} = 6.7 \times 10^{-6} \text{ M}$$

57. At  $25^\circ\text{C}$ , a pH of 7 is neutral.  $\text{pH} < 7$  indicates an acidic solution and  $\text{pH} > 7$  indicates a basic solution.

- $4.6 < 7$ , the solution is acidic.
- $\text{pH} = 7$ , the solution is neutral.
- $1.6 < 7$ , the solution is acidic.
- $10.5 > 7$ , the solution is basic.

59. 

- $3.5 < 7$ , the solution is acidic.
- $\text{pH} = 7$ , the solution is neutral.
- $9.0 > 7$ , the solution is basic.
- $5.5 < 7$ , the solution is acidic.

61. 

- $\text{pH} = -\log(1.0 \times 10^{-8}) = 8.00$
- $\text{pH} = -\log(5.0 \times 10^{-12}) = 11.30$
- $\text{pH} = -\log(7.5 \times 10^{-3}) = 2.12$
- $\text{pH} = -\log(6.35 \times 10^{-9}) = 8.197$

63.  $\text{pH} = -\log(7.5 \times 10^{-3}) = 2.12$

65. The exercise can be approached in either of two ways illustrated in parts a and b. We assume the temperature is  $25^\circ\text{C}$  where  $K_w = 1.0 \times 10^{-14}$

a. 
$$[\text{H}_3\text{O}^+] = \frac{K_w}{[\text{OH}^-]} = \frac{1.0 \times 10^{-14}}{5.25 \times 10^{-9}} = 1.9 \times 10^{-6} \text{ M}$$

$$\text{pH} = -\log(1.9 \times 10^{-6}) = 5.72$$

b.  $\text{pOH} = -\log(8.3 \times 10^{-3}) = 2.08$

$$\text{pH} = 14.00 - 2.08 = 11.92$$

c.  $\text{pOH} = -\log(3.6 \times 10^{-12}) = 11.44$

$$\text{pH} = 14.00 - 11.44 = 2.56$$

d.  $\text{pOH} = -\log(2.1 \times 10^{-8}) = 7.68$

$$\text{pH} = 14.00 - 7.68 = 6.32$$

67.  $\text{pOH} = -\log(0.0040) = 2.40$

$$\text{pH} = 14.00 - 2.40 = 11.60$$

69.  $\text{pH} = 5.12$        $[\text{H}_3\text{O}^+] = 10^{-\text{pH}} = 10^{-5.12} = 7.6 \times 10^{-6} \text{ M}$

71.  $\text{pOH} = 14.00 - \text{pH} = 14.00 - 11.63 = 2.37$

$$[\text{OH}^-] = 10^{-\text{pOH}} = 10^{-2.37} = 4.36 \times 10^{-3} \text{ M}$$

73.  $5.80 \text{ g NaOH} \times \frac{1 \text{ mol NaOH}}{40.00 \text{ g}} = 0.145 \text{ mol NaOH} = 0.145 \text{ mol OH}^-$

$$[\text{OH}^-] = \frac{0.145 \text{ mol}}{1.00 \text{ L}} = 0.145 \text{ M}$$

$$[\text{H}_3\text{O}^+] = \frac{K_w}{[\text{OH}^-]} = \frac{1.0 \times 10^{-14}}{0.145} = 6.9 \times 10^{-14} \text{ M}$$

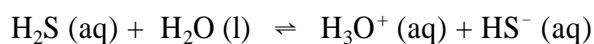
$$\text{pH} = -\log(6.9 \times 10^{-14}) = 13.16$$

75. The yellow color of methyl red indicates that the pH is greater than about 5.6. The yellow color of bromthymol blue indicates that the pH is below about 6.4. The pH is therefore between approximately 5.6 and 6.4.

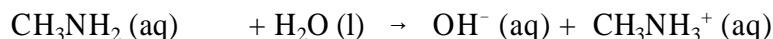
77. a. Strictly speaking, BaO is not an Arrhenius base since it does not contain OH<sup>-</sup>. However, O<sup>2-</sup> will react with water to give OH<sup>-</sup>, making it a Brønsted base.



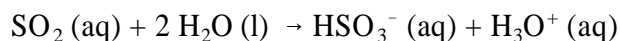
b. H<sub>2</sub>S is an acid



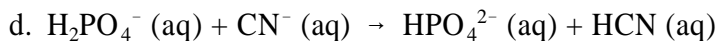
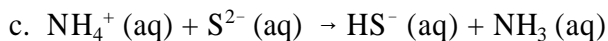
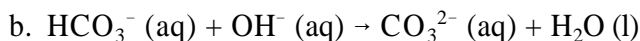
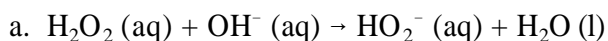
c. Strictly speaking, CH<sub>3</sub>NH<sub>2</sub> is not an Arrhenius base since it does not contain OH<sup>-</sup>. However, CH<sub>3</sub>NH<sub>2</sub> will react with water to give OH<sup>-</sup>, making it a Brønsted base.



d. SO<sub>2</sub> is not an acid according to the Arrhenius definition, since it does not contain H<sup>+</sup>. However it is a Lewis acid which reacts with water to yield H<sub>3</sub>O<sup>+</sup> as shown.

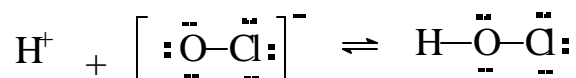


79. There are several possible answers. The acid has to react with the conjugate base of a weaker acid.



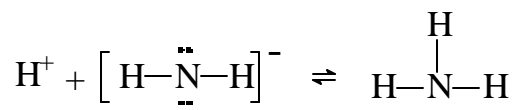
81. a.  $\text{ClO}^{-} (\text{aq}) + \text{H}_2\text{O} (\text{l}) \rightleftharpoons \text{HClO} (\text{aq}) + \text{OH}^{-} (\text{aq})$

This is a Brønsted acid base reaction in which ClO<sup>-</sup> is the base and H<sub>2</sub>O is the acid. The reactants are favored since HClO is a stronger acid than H<sub>2</sub>O. It can also be viewed as a Lewis acid-base reaction in which H<sup>+</sup> from the water acts as the acid and OCl<sup>-</sup> acts as the base

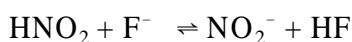




This is a Brønsted acid base reaction in which  $\text{NH}_2^-$  is the base and  $\text{NH}_4^+$  is the acid. The products are favored since  $\text{NH}_3$  is only very slightly ionized. It can also be viewed as a Lewis acid-base reaction in which  $\text{H}^+$  from the ammonium ion acts as the acid and  $\text{NH}_2^-$  acts as the base.



83.  $\text{HNO}_2$  is slightly stronger than HF, so the equilibrium lies slightly farther to the products side.



85.  $\text{H}_2\text{S} < \text{H}_2\text{Se} < \text{HBr}$ .

$\text{H}_2\text{Se}$  is stronger than  $\text{H}_2\text{S}$  because in a group, the strength of a binary acid increases with increasing size of the non-hydrogen atom.  $\text{HBr}$  is stronger than  $\text{H}_2\text{Se}$ , because in a given period for binary acids, acid strength increases as the electronegativity of the non-hydrogen atom increases.

87.  $\text{KOH}$  is a strong base and completely dissociated in solution. We assume the temperature is  $25^\circ\text{C}$  where  $K_w = 1.0 \times 10^{-14}$

$$[\text{OH}^-] = 0.25 \text{ M} \quad [\text{H}_3\text{O}^+] = \frac{K_w}{[\text{OH}^-]} = \frac{1.0 \times 10^{-14}}{0.25} = 4.0 \times 10^{-14} \text{ M}$$

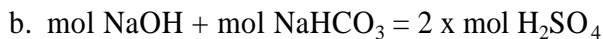
89.  $\text{pH} = -\log(1.5 \times 10^{-3}) = 2.82$

91.  $\text{pH} = 3.15$        $[\text{H}_3\text{O}^+] = 10^{-\text{pH}} = 10^{-3.15} = 7.1 \times 10^{-4} \text{ M}$

$$[\text{OH}^-] = \frac{K_w}{[\text{H}_3\text{O}^+]} = \frac{1.0 \times 10^{-14}}{7.1 \times 10^{-4}} = 1.4 \times 10^{-11} \text{ M}$$

93. a. molecular:  $2 \text{NaHCO}_3 + \text{H}_2\text{SO}_4 \rightarrow \text{Na}_2\text{SO}_4 + 2 \text{CO}_2 (\text{g}) + 2 \text{H}_2\text{O} (\text{l})$

net ionic:  $\text{HCO}_3^- + \text{H}_3\text{O}^+ \rightarrow \text{CO}_2 (\text{g}) + 2 \text{H}_2\text{O} (\text{l})$



$$\text{mol OH}^- = 0.0354 \text{ L} \times \frac{0.108 \text{ mol}}{\text{L}} = 0.00382 \text{ mol OH}^-$$

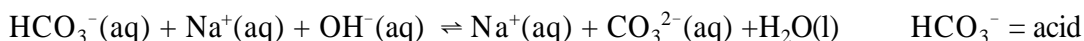
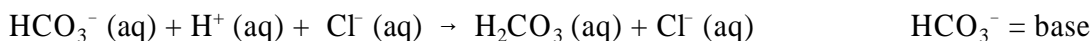
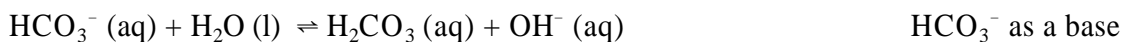
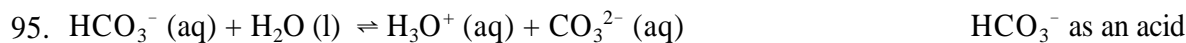
$$\text{mol H}_2\text{SO}_4 = 0.02500 \text{ L} \times \frac{0.437 \text{ mol}}{\text{L}} = 0.0109 \text{ mol H}_2\text{SO}_4$$

$$\text{mol NaHCO}_3 = 2(0.0109) \text{ mol} - 0.00382 \text{ mol} = 0.0180 \text{ mol NaHCO}_3.$$

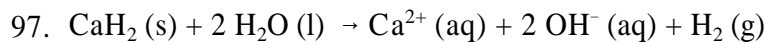
c.  $0.0180 \text{ mol NaHCO}_3 \times \frac{84.01 \text{ g}}{\text{mol}} = 1.51 \text{ g NaHCO}_3$

$$\% \text{ NaHCO}_3 = \frac{1.51 \text{ g}}{2.500 \text{ g}} \times 100 = 60.5 \% \text{ NaHCO}_3$$

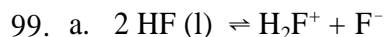
$$100.0\% - 60.5\% = 39.5\% \text{ KCl}$$



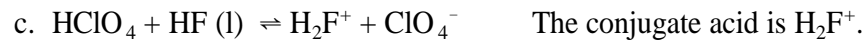
Note:  $\text{H}_2\text{CO}_3 (\text{aq})$  is unstable and decomposes to  $\text{CO} (\text{g})$  and  $\text{H}_2\text{O} (\text{l})$



$\text{H}^-$  is a stronger base than  $\text{OH}^-$ , because the  $\text{H}^-$  ion reacts with  $\text{H}_2\text{O}$  to yield  $\text{OH}^-$ . The statement that the “strongest base that can exist in water is the hydroxide ion” is a consequence of the leveling effect. Any base which is stronger than  $\text{OH}^-$  will react with water, abstracting  $\text{H}^+$  from a water molecule and leaving  $\text{OH}^-$ . A base stronger than  $\text{OH}^-$  cannot survive in water.

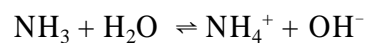


b.  $\text{NaF}$  is a source of  $\text{F}^-$  which is the conjugate base of  $\text{HF}$ .  $\text{NaF}$  is therefore a base in liquid  $\text{HF}$ . One acid-base definition which was not mentioned in your book is the solvent system definition. It says that an acid is any solute which increases the concentration of the cation derived from the solvent, and a base is any solvent which increases the concentration of the anion derived from the solvent. since  $\text{NaF}$  increases the  $\text{F}^-$  concentration, it is a base by the solvent system definition.



$$101. \quad 4.25 \text{ g NH}_3 \times \frac{1 \text{ mol NH}_3}{17.03 \text{ g}} = 0.250 \text{ mol NH}_3$$

$$[\text{NH}_3] = \frac{0.250 \text{ mol NH}_3}{0.250 \text{ L}} = 1.00 \text{ M NH}_3$$



If 0.42% of the  $\text{NH}_3$  reacts with water,  $[\text{OH}^-] = 0.0042 \text{ M}$ .

$$\text{pOH} = -\log[\text{OH}^-] = -\log(0.0042) = 2.38$$

$$\text{pH} = 14.00 - 2.38 = 11.62$$