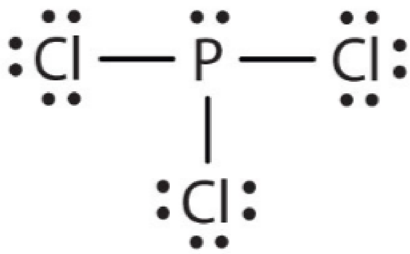
Sample Problems

1. Just as Lewis structures map out the valence electrons in atoms, Lewis acids accept electrons and Lewis bases donate electrons.

   a. H is the Lewis acid and H_2O is the Lewis base.
   b. $AlCl_3$ is the Lewis acid and Cl^- is the Lewis base.

2. PCl_3 will behave as a Lewis base, donating electrons to a coordinate covalent bond because it has a non-bonding pair of electrons.



You do NOT have to do #7

3. An acid gives hydrogen ions. A base gives hydroxide ions.

4. Acids are hydrogen-ion donors and bases are hydrogen-ion acceptors.

5. A Lewis acid is an electron-pair acceptor; a Lewis base is an electron-pair donor.

6. Both are electrolytes and change the color of an acid-base indicator. Acids have a sour taste; bases taste bitter.

7. **a.** diprotic
b. triprotic
c. monoprotic
d. diprotic

8. $HNO_3 + H_2O \rightarrow H_3O^+ + NO_3^-$; HNO_3 is the hydrogen-ion donor; its conjugate base is NO_3^-. H_2O is the hydrogen-ion acceptor; its conjugate acid is H_3O^+.
$CO_3^{2-} + H_2O \rightleftharpoons HCO_3^- + OH^-$; CO_3^{2-} is the hydrogen ion acceptor; its conjugate acid is HCO_3^-. H_2O is the hydrogen-ion donor; its conjugate base is OH^-.

9. **BIGIDEA** The reaction of NaOH with aluminum generates heat, which softens greases and oils and hydrogen, which agitates the mixture.

Sample Problems

10. Arrhenius acids have an $[H^+] > 10^{-7}$ M and bases have an $[H^+] < 10^{-7}$ M. Arrhenius bases have an $[OH^-] > 10^{-7}$ M and acids have an $[OH^-] < 10^{-7}$ M. A neutral solution has $[H^+] = 10^{-7}$ M and $[OH^-] = 10^{-7}$ M. Therefore, $[H^+] = [OH^-] = 10^{-7}$ M.

    a. $[H^+] < 10^{-7}$ M is a base
    b. $[OH^-] > 10^{-7}$ M is a base
    c. $[H^+] > 10^{-7}$ M is slightly acidic
    d. $[OH^-] = 10^{-7}$ M is neutral

11. Acidity or basicity depends on $k_w = [H^+] \times [OH^-] = 10^{-14}$ M.
    If $[OH^-] = 1.0 \times 10^{-3}$ M, then $[H^+] = 10^{-14}$ M / $[OH^-] = 1.0 \times 10^{-11}$ M (basic)

12. $pH = -\log_{10} [H^+]$

    a. $pH = -\log_{10} [4.5 \times 10^{-2} M] = -\log 4.5 + (-\log 10^{-2}) = -0.65 + 2 = 1.35$

    b. $pH = -\log_{10} [8.7 \times 10^{-6} M] = -\log 8.7 + (-\log 10^{-6}) = -0.94 + 6 = 5.06$

    c. $pH = -\log_{10} [1.5 \times 10^{-3} M] = -\log 1.5 + (-\log 10^{-3}) = -0.18 + 3 = 2.82$

    d. $pH = -\log_{10} [1.2 \times 10^{-3} M] = -\log 1.2 + (-\log 10^{-3}) = -0.08 + 3 = 2.92$

13. $pH = -\log_{10} [H^+]$ … when the number is 1.0, use the negative superscript as the pH

    a. $pH = -\log_{10} [1.0 \times 10^{-12} M] = -\log 1.0 + (-\log 10^{-2}) = 0 + 12 = 12$

    b. $pH = -\log_{10} [1.0 \times 10^{-4} M] = -\log 1.0 + (-\log 10^{-4}) = 0 + 4 = 4$

14. [***This is an optional question.***]
    Since $pH = -\log_{10} [H^+]$, then $[H^+]$ = antilog -pH
    Antilog on your calculator is usually “shift” “log”. Put in the pH, negate it and then take the antilog.

    a. $[H^+]$ = antilog –(5.00) = 1.0×10^{-5} M
    $pH = -\log_{10} [H^+]$ … if the number has no decimal greater than 1,
    $[H^+] = 1.0 \times 10^{-pH}$

b. $[H^+]$ = antilog –(12.83) = 1.5×10^{-13} M

15. [***This is an optional question.***]
Since pH = $-\log_{10} [H^+]$, then $[H^+]$ = antilog -pH
Antilog on your calculator is usually “shift” “log”. Put in the pH, negate it and then take the antilog.

a. $[H^+]$ = antilog –(4.00) = 1.0×10^4 M
pH = $-\log_{10} [H^+]$ … if the number has no decimal greater than 1,
$[H^+] = 1.0 \times 10^{-pH}$

b. $[H^+]$ = antilog –(11.55) = 2.8×10^{-12} M

16. Acidity or basicity depends on $k_w = [H^+] \times [OH^-] = 10^{-14}$ M.
Therefore, $[H^+] = 10^{-14}$ M / $[OH^-]$. Then, pH = $-\log_{10} [H^+]$

a. $[H^+] = 10^{-14}$ M / [4.3×10^{-5} M] = 2.3×10^{-10} M
pH = $-\log_{10}$ [2.3×10^{-10} M] = $-\log 2.3 + (-\log 10^{-10}) = -0.36 + 10 = 9.64$

b. $[H^+] = 10^{-14}$ M / [4.5×10^{-11} M] = 2.2×10^{-4} M
pH = $-\log_{10}$ [2.2×10^{-4} M] = $-\log 2.2 + (-\log 10^{-4}) = -0.34 + 4 = 3.66$

17. Acidity or basicity depends on $k_w = [H^+] \times [OH^-] = 10^{-14}$ M.
Therefore, $[H^+] = 10^{-14}$ M / $[OH^-]$. Then, pH = $-\log_{10} [H^+]$

a. $[H^+] = 10^{-14}$ M / [5.0×10^{-9} M] = 2.0×10^{-6} M
pH = $-\log_{10}$ [2.0×10^{-6} M] = $-\log 2.0 + (-\log 10^{-6}) = -0.3 + 6 = 5.7$

b. $[H^+] = 10^{-14}$ M / [8.3×10^{-4} M] = 1.2×10^{-11} M
pH = $-\log_{10}$ [1.2×10^{-11} M] = $-\log 1.2 + (-\log 10^{-11}) = -0.08 + 11 = 10.92$

19.2 Lesson Check

18. $[H^+] \times [OH^-] = 1.0 \times 10^{-14}$; when $[H^+]$ in a solution increases, the $[OH^-]$ decreases.

19. basic: greater than 7; acidic: less than 7; neutral: 7

20. acid-base indicator or a pH meter

21. $[H^+]$ decreases as pH increases.

22. **a.** 6.0 **b.** 4.00 **c.** 12.0 **d.** 3.0

23. In basic solutions the $[OH^-]$ is greater than $[H^+]$; in acidic solutions $[H^+]$ is greater than $[OH^-]$.

24. **a.** $1.0 \times 10^{-8} M$ **b.** $1.0 \times 10^{-5} M$ **c.** $1.0 \times 10^{-2} M$

22. $pH = -\log_{10} [H^+]$

a. $pH = -\log_{10} [1.0 \times 10^{-6}\ M] = -\log 1.0 + (-\log 10^{-6}) = 0 + 6 = 6.0$

b. $[H^+] = 0.00010\ M = 1.0 \times 10^{-4}\ M$

$pH = -\log_{10} [1.0 \times 10^{-4}\ M] = -\log 1.0 + (-\log 10^{-4}) = 0 + 4 = 4.0$

c. Acidity or basicity depends on $k_w = [H^+] \times [OH^-] = 10^{-14}\ M$.

$[OH^-] = 1.0 \times 10^{-2}\ M$,

then $[H^+] = 10^{-14}\ M / [OH^-] = 1.0 \times 10^{-12}\ M$ (acidic)

$pH = -\log_{10} [1.0 \times 10^{-12}\ M] = -\log 1.0 + (-\log 10^{-12}) = 0 + 12 = 12.0$

d. Acidity or basicity depends on $k_w = [H^+] \times [OH^-] = 10^{-14}\ M$.

$[OH^-] = 1.0 \times 10^{-11}\ M$,

then $[H^+] = 10^{-14}\ M / [OH^-] = 1.0 \times 10^{-3}\ M$ (acidic)

$pH = -\log_{10} [1.0 \times 10^{-3}\ M] = -\log 1.0 + (-\log 10^{-3}) = 0 + 3 = 3.0$

24. $pH = -\log_{10} [H^+]$ … if the number has no decimal greater than 1,
$[H^+] = 1.0 \times 10^{-pH}$
Acidity or basicity depends on $k_w = [H^+] \times [OH^-] = 10^{-14}\ M$.
Therefore, $[OH^-] = 10^{-14}\ M / [H^+]$.

a. $pH = 6.00 = 1.0 \times 10^{-6}\ M = [H^+]$; $[OH^-] = 1.0 \times 10^{-8}\ M$

b. $pH = 9.00 = 1.0 \times 10^{-9}\ M = [H^+]$; $[OH^-] = 1.0 \times 10^{-5}\ M$

c. $pH = 12.00 = 1.0 \times 10^{-12}\ M = [H^+]$; $[OH^-] = 1.0 \times 10^{-2}\ M$

You do NOT have to do #25-26

25. $0.1000M - 4.2 \times 10^{-3}M = 0.0958M$

$$K_a = \frac{(4.2 \times 10^{-3}) \times (4.2 \times 10^{-3})}{(0.0958)} = 1.8 \times 10^{-4}$$

26. $0.2000M - 9.86 \times 10^{-4}M = 0.199M$

$$K_a = \frac{(9.86 \times 10^{-4}) \times (9.86 \times 10^{-4})}{(0.199)} = 4.89 \times 10^{-6}$$

19.3 Lesson Check

You do NOT have to do #32

27. the degree to which they ionize in water

28. A strong acid is completely ionized in aqueous solution and has a large K_a. A weak acid ionizes only slightly in aqueous solution and has a small K_a.

29. hypochlorous acid

30 Substitute the measured concentrations of all the substances present at equilibrium into the expressions for K_a or K_b.

31. [HA] will be much greater than [H⁺] at equilibrium.

~~**32.** $K_a = 6.66 \times 10^{-11}$.~~

33. **a.** $HNO_3 + H_2O \rightarrow H_3O + NO_3^-$

b. $CH_3COOH + H_2O \rightleftharpoons CH_3COO^- + H_3O^+$

c. $NH_3 + H_2O \rightleftharpoons NH_4^+ + H_2O$

d. $Mg(OH)_2 \rightarrow Mg^{2} + 2OH^-$

34. It is concentrated based on its molarity, but weak based on its K_a.

Sample Problems

35. $H_3PO_4(aq) + 3KOH(aq) \longrightarrow K_3PO_4(aq) + 3H_2O(l)$

$$1.56\ \cancel{\text{mol } H_3PO_4} \times \frac{3\text{ mol KOH}}{1\ \cancel{\text{mol } H_3PO_4}} = 4.68\text{ mol KOH}$$

36. $HNO_3(aq) + NaOH(aq) \longrightarrow NaNO_3(aq) + H_2O(l)$

$$0.20\ \cancel{\text{mol } HNO_3} \times \frac{1\text{ mol NaOH}}{1\ \cancel{\text{mol } HNO_3}} = 0.20\text{ mol NaOH}$$

37. $$25.0\ \cancel{\text{mL KOH}} \times \frac{1.00\ \cancel{\text{mol KOH}}}{1000\ \cancel{\text{mL KOH}}} \times \frac{1\ \cancel{\text{mol HCl}}}{1\ \cancel{\text{mol KOH}}} \times \frac{1000\text{ mL HCl}}{0.45\ \cancel{\text{mol HCl}}} = 56\text{ mL HCl}$$

38. $$38.5\ \cancel{\text{mL NaOH}} \times \frac{0.150\ \cancel{\text{mol NaOH}}}{1000\ \cancel{\text{mL NaOH}}} \times \frac{1\text{ mol } H_3PO_4}{3\ \cancel{\text{mol NaOH}}} = 0.00193\text{ mol } H_3PO_4$$

$$\frac{0.00193\text{ mol } H_3PO_4}{0.0150\text{ L } H_3PO_4} = 0.129M\ H_3PO_4$$

Lesson Check Answers

39. a salt and water

40. At that point, the number of moles of hydrogen ions is equal to the number of moles of hydroxide ions.

41. **a.** 0.03 mol

b. 2 mol

c. 0.2 mol

42. **a.** $H_2SO_4 + 2KOH \rightarrow 2H_2O + K_2SO_4$

b. $2H_3PO_4 + 3Ca(OH)_2 \rightarrow 6H_2O + Ca_3(PO_4)_2$

c. $2HNO_3 + Mg(OH)_2 \rightarrow 2H_2O + Mg(NO_3)_2$

43. Acid-base neutralizations are double displacement reactions. The positive ions are exchanged between the acid and base reactants; the products are a salt and water.

41. Write a balanced chemical equation for each:

a. $HCl + KOH \rightarrow KCl + HOH$ 1:1 mol ratio … 0.03 mol

b. $HCl + NH_3 \rightarrow NH_4^+Cl + HOH$ 1:1 mol ratio … 2 mol

c. $2HCl + Ca(OH)_2 \rightarrow CaCl_2 + 2HOH$ 2:1 mol ratio … 0.2 mol

Sample Problems

44. **a.** $HPO_4^{2-}(aq) + H^+(aq) \longrightarrow H_2PO_4^-(aq)$

b. $H_2PO_4^-(aq) + OH^-(aq) \longrightarrow HPO_4^{2-}(aq) + H_2O(l)$

45. $HCOO^-(aq) + H^+(aq) \longrightarrow HCOOH(aq)$

19.5 Lesson Check

Lesson Check Answers

46. acidic solution: salt of a strong acid and weak base; basic solution: salt of a weak acid and a strong base

47. a weak acid and one of its salts or a weak base and one of its salts

48. d

49. Pair (b) because it consists of a weak acid and one of its salts.

50. $NH_3 + H \rightarrow NH_4^+$

$NH_4^+ + OH^- \rightarrow NH_3 + H_2O$

51 According to Le Châtelier, when a system in equilibrium is disturbed, the response is an attempt to restore the equilibrium. When the stress is the addition of an acid or a base, a buffer system helps to restore the equilibrium.

48. $(NH_4)_2SO_4$ is the product of a strong acid (H_2SO_4) and a weak base (NH_3), yielding an acidic solution.