ANSWER KEY

Section Review 21.1

Part A Completion

1. acid
2. hydroxide
3. water
4. neutralization
5. titration
6. end point
7. equivalence
8. mass
9. one mole
10. 2 equivalents

Part B True-False

11. AT
12. AT
13. ST
14. AT

Part C Matching

15. c
16. e
17. a
18. b
19. d

Part D Questions and Problems

20. **a.** $H_3PO_4 + Al(OH)_3 \rightarrow AlPO_4 + 3H_2O$
    b. $2HI + Ca(OH)_2 \rightarrow CaI_2 + 2H_2O$

Section Review 21.2

Part A Completion

1. salt
2. acidic
3. basic
4. neutral
5. hydrolyze
6. strong
7. weak
8. buffer
9. capacity
10. solubility product constant

Part B True-False

11. NT
12. NT
13. AT
14. AT

Part C Matching

15. a
16. d
17. b
18. c

Part D Questions and Problems

19. $[CO_3^{2-}] = 0.00070M$

$[Ba^{2+}] = 0.0015M$

$[CO_3^{2-}] \times [Ba^{2+}]$

$= (7.0 \times 10^{-4}M) \times (1.5 \times 10^{-3}M)$

$= 1.1 \times 10^{-6}$

Precipitation occurs because the ion product (1.1×10^{-6}) is greater than the K_{sp} of $BaCO_3$ (5.0×10^{-9}).

Practice Problems

Section 21.1

1. $2NaOH(aq) + H_2SO_4(aq) \rightarrow Na_2SO_4(aq) + 2H_2O(l)$

$$0.014\ \cancel{\text{L } H_2SO_4} \times \frac{0.75\ \cancel{\text{mol } H_2SO_4}}{1\ \cancel{\text{L } H_2SO_4}} \times \frac{2\ \text{mol NaOH}}{1\ \cancel{\text{mol } H_2SO_4}} = 0.021\ \text{mol NaOH}$$

$$\text{Molarity} = \frac{\text{moles}}{\text{liters}} = \frac{0.021\ \text{mol NaOH}}{0.038\ \text{L NaOH}}$$

$$= 0.55M\ \text{NaOH}$$

2. $Ca(OH)_2(aq) + 2HC_2H_3O_2(aq) \rightarrow Ca(C_2H_3O_2)_2(aq) + 2H_2O(l)$

$$0.0142\ \cancel{\text{L } HC_2H_3O_2} \times \frac{0.0140\ \cancel{\text{mol } HC_2H_3O_2}}{1\ \cancel{\text{L } HC_2H_3O_2}} \times \frac{1\ \text{mol } Ca(OH)_2}{2\ \cancel{\text{mol } HC_2H_3O_2}}$$

$$= 9.94 \times 10^{-5}\ \text{mol } Ca(OH)_2$$

$$= 9.94 \times 10^{-5}\ \text{mol } Ca(OH)_2$$

$$\text{Molarity} = \frac{9.94 \times 10^{-5}\ \text{mol } Ca(OH)_2}{0.0246\ \text{L } Ca(OH)_2}$$

$$= 0.00404M$$

3. $Ca(OH)_2(aq) + H_2SO_4(aq) \rightarrow CaSO_4(aq) + 2H_2O(l)$

$$0.0198\ \cancel{\text{L } Ca(OH)_2} \times \frac{0.0100\ \cancel{\text{mol } Ca(OH)_2}}{1\ \cancel{\text{L } Ca(OH)_2}} \times \frac{1\ \text{mol } H_2SO_4}{1\ \cancel{\text{mol } Ca(OH)_2}} = 0.000198\ \text{mol } H_2SO_4$$

$$= 0.000198\ \text{mol } H_2SO_4$$

$$\text{Molarity} = \frac{\text{moles}}{\text{liters}} = \frac{0.000198\ \text{mol } H_2SO_4}{0.0124\ \text{L } H_2SO_4}$$

$$= 0.0160M\ H_2SO_4$$

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4. $Ba(OH)_2(aq) + 2HCl(aq) \rightarrow BaCl_2(aq) + 2H_2O(l)$

$$0.0122\ \text{L}\ \cancel{\text{HCl}} \times \frac{0.25\ \text{mol}\ \cancel{\text{HCl}}}{1\ \text{L}\ \cancel{\text{HCl}}} \times \frac{1\ \text{mol Ba(OH)}_2}{2\ \text{mol HC}}$$

$$= 0.0015\ \text{mol Ba(OH)}_2$$

$$\text{liters} = \frac{\text{moles}}{\text{molarity}} = \frac{0.0015\ \text{mol Ba(OH)}_2}{0.12M\ \text{Ba(OH)}_2}$$

$$= 0.0125\ \text{L Ba(OH)}_2 = 13\ \text{mL Ba(OH)}_2$$

5. $Al(OH)_3(aq) + 3HCl(aq) \rightarrow AlCl_3(aq) + 3H_2O(l)$

$$0.0550\ \text{g}\ \cancel{\text{Al(OH)}_3} \times \frac{1\ \text{mol}\ \cancel{\text{Al(OH)}_3}}{78.0\ \text{g}\ \cancel{\text{Al(OH)}_3}} \times \frac{3\ \text{mol HCl}}{1\ \text{mol}\ \cancel{\text{Al(OH)}_3}} = 0.00212\ \text{mol HCl}$$

$$\text{liters} = \frac{\text{moles}}{\text{molarity}} = \frac{0.00212\ \text{mol HCl}}{0.200M\ \text{HCl}}$$

$$= 0.0106\ \text{L HCl} = 10.6\ \text{mL HCl}$$

6. $4.20\ \cancel{\text{L}} \times \frac{1.20\ \text{equiv}}{1\ \cancel{\text{L}}} = 5.04\ \text{equiv KOH}$

7. H_3PO_4 contains 3 equiv/mol;
normality = $1.2N\ H_3PO_4$

8. $\frac{24.6\ \cancel{\text{g}}\ \text{HCl}}{36.5\ \cancel{\text{g}}/\text{equiv}} = 0.674\ \text{equiv HCl}$

$$\text{Normality} = \frac{0.674\ \text{equiv}}{1.50\ \text{L}} = 0.449N\ \text{HCl}$$

9. $V_A = \frac{N_B \times V_B}{N_A} = \frac{0.110\ \cancel{N} \times 68.0\ \text{mL}}{0.250\cancel{N}}$

$= 29.9\ \text{mL}\ H_2SO_4$

Section 21.2

1. $CHO_2^- + H^+ \rightleftharpoons HCHO_2$
$HCHO_2 + OH^- \rightleftharpoons CHO_2^- + H_2O$

2. **a.** $Ca(OH)_2(s) \rightleftharpoons Ca^{2+}(aq) + 2OH^-(aq)$
$K_{sp} = [Ca^{2+}] \times [OH^-]^2$

b. $Ag_2CO_3(s) \rightleftharpoons 2Ag^+(aq) + CO_3^{2-}(aq)$
$K_{sp} = [Ag^+]^2 \times [CO_3^{2-}]$

3. $Ag_2CO_3(s) \rightleftharpoons 2Ag^+(aq) + CO_3^{2-}(aq)$
$K_{sp} = [Ag^+]^2 \times [CO_3^{2-}]$
$K_{sp} = (2x)^2(x) = 8.1 \times 10^{-12}$
$4x^3 = 8.1 \times 10^{-12}$
$x^3 = 2.0 \times 10^{-12}$
$x = 1.3 \times 10^{-4}M = [CO_3^{2-}]$
$[Ag^+] = 2x = 2.6 \times 10^{-4}M$

4. $Fe(OH)_2(s) \rightleftharpoons Fe^{2+}(aq) + 2OH^-(aq)$
$K_{sp} = [Fe^{2+}] \times [OH^-]^2$
$[Fe^{2+}] = 0.5[OH^-] = 6.0 \times 10^{-6}$
$K_{sp} = (6.0 \times 10^{-6})(1.2 \times 10^{-5})^2$
$K_{sp} = 8.6 \times 10^{-16}$

5. $SrCO_3 \rightleftharpoons Sr^{2+} + CO_3^{2-}$
$K_{sp} = [Sr^{2+}] \times [CO_3^{2-}]$
Let $x = [Sr^{2+}] = [CO_3^{2-}]$
$K_{sp} = x^2 = 9.3 \times 10^{-10}$
$x = [Sr^{2+}] = 3.0 \times 10^{-5}M$

6. $K_{sp} = [Ag^+]^2 \times [CO_3^{2-}]$
Let $[Ag^+] = 2x$ and $[CO_3^{2-}] = x$
assume $x << 0.20$ mol
$K_{sp} = (2x)^2(0.20) = 8.1 \times 10^{-12}$
$x^2 = 1.0 \times 10^{-11}$
$x = 3.2 \times 10^{-6}M = [CO_3^{2-}]$
$[Ag^+] = 6.4 \times 10^{-6}M$

7. $K_{sp}(PbSO_4) = 6.3 \times 10^{-7} = [Pb^{2+}] \times [SO_4^{2-}]$
The total volume of solution is 1000 mL, so
$[Pb^{2+}] = 0.0020\ \text{mol/L} = 0.0020M$
$[SO_4^{2-}] = 0.0012\ \text{mol/L} = 0.0012M$
$[Pb^{2+}] \times [SO_4^{2-}] = (0.0020)(0.0012)$
$= 2.4 \times 10^{-6}$
Because this product exceeds the K_{sp} value, precipitation will occur.

8. $K_{sp}(CaCO_3) = 4.5 \times 10^{-9} = [Ca^{2+}] \times [CO_3^{2-}]$
The total volume is 1000 mL, so
$[Ca^{2+}] = 0.0021\ \text{mol/L} = 0.0021M$
$[CO_3^{2-}] = 0.0013\ \text{mol/L} = 0.0013M$
$[Ca^{2+}] \times [CO_3^{2-}] = (0.0021)(0.0013)$
$= 2.7 \times 10^{-6}$
Because this product exceeds the K_{sp} value, precipitation will occur.

9. NaCl has no ion in common with $Mg(OH)_2$.

Interpreting Graphics 21

1. $C_6H_5COOH + NaOH \rightarrow C_6H_5COONa + H_2O$
One mole of sodium hydroxide will neutralize one mole of benzoic acid.

2. To determine the equivalence point, find the area of the titration curve where the pH changes abruptly when a small volume of NaOH is added. Locate the point on this steep portion of the curve equidistant between the two plateaus. The pH at the equivalence point is approximately 8.5; the solution is slightly basic.

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3. Benzoic acid is a weak acid. The neutralization of a weak acid with a strong base, such as NaOH, produces a basic solution at the equivalence point.

4. $0.025 \cancel{L} \times 0.10 \text{ mol}/\cancel{L} \text{ NaOH}$
$= 0.0025 \text{ mol NaOH}$

5. The equivalence point occurs when the number of moles of NaOH added equals the number of moles of C_6H_5COOH originally present. Because NaOH is a strong base, each mole of NaOH added reacts with each mole of C_6H_5COOH present. Thus, at the equivalence point,
$[C_6H_5COOH] = [NaOH] = 0M$ and
$$[C_6H_5COONa] = \frac{0.0025 \text{ mol}}{0.050 \text{ L}} = 0.050M.$$

6. Based on the answers to questions 4 and 5, 0.0025 mol C_6H_5COOH were originally present in a volume of 25 mL.
$$\text{Thus, } [C_6H_5COOH] = \frac{0.0025 \text{ mol}}{0.025 \text{ L}} = 0.10M$$

7. Because the equivalence point occurs between pH 6 and pH 11, phenolphthalein would be a good choice. A faint pink color should be detected at the equivalence point. Thymol blue might also be a good candidate. Students should draw a horizontal band on the graph encompassing the pH range 8-10 to show the region of the curve where phenolphthalein would be an effective indicator of neutralization.

8. $C_6H_5COO^- + H_2O \rightleftharpoons C_6H_5COOH + OH^-$
At the equivalence point, the benzoate ion establishes the equilibrium shown. The resulting solution is slightly basic because $[OH^-] > [H^+]$.

9. $$K_b = \frac{[C_6H_5COOH][OH^-]}{[C_6H_5COO^-]}$$

10. $pH + pOH = 14$
$8.5 + pOH = 14$
$pOH = 5.5$
$[OH^-] = 3.2 \times 10^{-6}M$
At the equivalence point $[OH^-] =$
$[C_6H_5COOH] = 3.2 \times 10^{-6}M$
$[C_6H_5COO^-] = 0.050M$
$$K_b = \frac{(3.2 \times 10^{-6})^2}{(0.050)} = 2.0 \times 10^{-10}$$

Vocabulary Review 21

1. hydrolyzing salts
2. solubility product constant (K_{sp})
3. neutral
4. equivalence point
5. normality (N)
6. neutral salts
7. buffer
8. one equivalent

Quiz for Chapter 21

1. d	4. a	7. b
2. a	5. c	8. c
3. c	6. b	

Chapter 21 Test A

A. Matching

1. d	4. f	7. h
2. c	5. a	8. g
3. b	6. e	

B. Multiple Choice

9. d	12. a	15. c
10. c	13. a	16. b
11. b	14. d	

C. True-False

17. ST	19. ST	21. ST
18. NT	20. AT	

D. Problems

22. $\text{gfm } H_2SO_3 = 82.1 \text{ g}$
$$20.5 \cancel{\text{g } H_2SO_3} \times \frac{1.00 \cancel{\text{mol } H_2SO_3}}{82.1 \cancel{\text{g } H_2SO_3}} \times \frac{2 \text{ equiv } H_2SO_3}{1 \cancel{\text{mol } H_2SO_3}} = 0.499 \text{ equiv } H_2SO_3$$

23. a. $2H_3PO_4 + 3Ca(OH)_2 \rightarrow Ca_3(PO_4)_2 + 6H_2O$
b. $2HBr + Mg(OH)_2 \rightarrow MgBr_2 + 2H_2O$
c. $3H_2SO_4 + 2Al(OH)_3 \rightarrow Al_2(SO_4)_3 + 6H_2O$
d. $HCl + NH_4OH \rightarrow NH_4Cl + H_2O$

24. $V_1 \times N_1 = V_2 \times N_2$
$50.0 \text{ mL} \times N_1 = 20.0 \text{ mL} \times 1.50N$
$$N_1 = \frac{20.0 \cancel{\text{mL}} \times 1.50N}{50.0 \cancel{\text{mL}}}$$
$N_1 = 0.600N$

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25. $V_1 \times N_1 = V_2 \times N_2$

$V_1 \times 0.600N = 90.0\text{ mL} \times 0.40N$

$$V_1 = \frac{90\text{ mL} \times 0.40\cancel{N}}{0.600\cancel{N}}$$

$V_1 = 60.0\text{ mL}$

26. $V_1 \times N_1 = V_2 \times N_2$

$V_1 \times (5.0N) = 400\text{ mL} \times 0.20N$

$$V_1 = \frac{400\text{ mL} \times 0.20\cancel{N}}{5.0\cancel{N}}$$

$V_1 = 16\text{ mL}$

27. **a.** basic **c.** acidic

b. neutral **d.** basic

E. Essay

28. A measured amount of the base is placed in a beaker. An acid–base indicator is added. A burett is used to add acid of a known concentration (the standard solution). Acid is added until the indicator shows that neutralization has occurred. This is the end point of the titration. The amount of acid added is recorded. Based on the concentration and amount of acid added, the number of moles of acid added are calculated. The equation for the neutralization reaction is used to determine the number of moles of base needed to neutralize this amount of acid. The concentration of the base can be calculated from the number of moles of base required and the volume. The calculation can also be carried out in terms of normality.

F. Additional Questions

29. **a.** greater than 7 **d.** 7

b. less than 7 **e.** greater than 7

c. 7

30. **a.** $NH_3 + H^+ \rightarrow NH_4^+$

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

31. $0.96\text{ g } \cancel{(NH_4)_2CO_3} \times \frac{1\text{ mol } (NH_4)_2CO_3}{96\text{ g } \cancel{(NH_4)_2CO_3}}$

$= 0.010\text{ mol } (NH_4)_2CO_3$

$$\frac{0.010\text{ mol}}{20.0\text{ L}} = 5.0 \times 10^{-4} M\ (NH_4)_2CO_3$$

$= 5.0 \times 10^{-4} M\ CO_3^{2-}$

$0.20\text{ g } \cancel{CaBr_2} \times \frac{1\text{ mol } CaBr_2}{200\text{ g } \cancel{CaBr_2}}$

$= 1.0 \times 10^{-3}\text{ mol } CaBr_2$

$$\frac{1.0 \times 10^{-3}\text{ mol}}{20.0\text{ L}} = 5.0 \times 10^{-5} M\ CaBr_2$$

$= 5.0 \times 10^{-5} M\ Ca^{2+}$

$[Ca^{2+}] \times [CO_3^{2-}] = 5.0 \times 10^{-5} \times 5.0 \times 10^{-4}$

$= 2.5 \times 10^{-8}$

Yes, a precipitate will form because the ion product is larger than the K_{sp}.

32. $CaSO_4(s) \rightleftharpoons Ca^{2+}(aq) + SO_4^{2-}(aq)$

gfm $CaSO_4 = 136.1$ g/mol

$$\frac{0.67\text{ g}}{1\text{ L}} \times \frac{1\text{ mol}}{136.1\text{ g}} = 4.9 \times 10^{-3}\text{ mol/L}$$

$K_{sp} = [Ca^{2+}] \times [SO_4^{2-}]$

$= (4.9 \times 10^{-3}) \times (4.9 \times 10^{-3})$

$= 2.4 \times 10^{-5}$

Chapter 21 Test B

A. Matching

1. h	**5.** i	**9.** e
2. a	**6.** j	**10.** d
3. g	**7.** c	
4. f	**8.** b	

B. Multiple Choice

11. d	**16.** c	**21.** c
12. d	**17.** b	**22.** c
13. b	**18.** a	**23.** d
14. a	**19.** b	**24.** d
15. a	**20.** d	

C. True-False

25. ST	**27.** ST	**29.** NT
26. AT	**28.** AT	

D. Problems

30. **a.** $HF(aq) + KOH(aq) \rightarrow KF(aq) + H_2O(l)$

b. $2HNO_3(aq) + Ba(OH)_2(aq) \rightarrow Ba(NO_3)_2(aq) + 2H_2O(l)$

c. $H_2SO_4(aq) + 2LiOH(aq) \rightarrow Li_2SO_4(aq) + 2H_2O(l)$

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

31. $H_2SO_4(aq) + 2KOH(aq) \rightarrow K_2SO_4(aq) + 2H_2O(l)$

1 mol 2 mol 1 mol 2 mol

$$\frac{1\text{ mol } H_2SO_4}{2\text{ } \cancel{\text{mol KOH}}} \times 0.35\text{ } \cancel{\text{mol KOH}}$$

$= 0.18\text{ mol } H_2SO_4$

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32. $2HCl(aq) + Ba(OH)_2(aq) \rightarrow BaCl_2(aq) + 2H_2O(l)$

2 mol 1 mol 1 mol 2 mol

$$\frac{2 \text{ mol HCl}}{1 \text{ mol Ba(OH)}_2} \times \frac{(0.0154 \text{ mol Ba(OH)}_2}{1 \text{ L}} \times \frac{(1 \text{ L})}{1000 \text{ mL}} \times (27.4 \text{ mL})$$

$$= 8.44 \times 10^{-4} \text{ mol HCl}$$

$$\text{Molarity} = \frac{\text{mol HCl}}{\text{L}} = \frac{8.44 \times 10^{-4} \text{ mol}}{0.020 \text{ L}}$$

$$= 0.042M$$

33. $N_1V_1 = N_2V_2$

$$V_1 = \frac{N_2V_2}{N_1}$$

$$V_1 = \frac{(0.350)(50.0 \text{ mL})}{(0.750)}$$

$$V_1 = 23.3 \text{ mL}$$

34. $CuCl(s) \rightarrow Cu^+(aq) + Cl^-(aq)$

$$K_{sp} = [Cu^+][Cl^-]$$

$$3.2 \times 10^{-7} = [Cu^+][Cu^+]$$

$$3.2 \times 10^{-7} = [Cu^+]^2$$

$$5.7 \times 10^{-4}M = [Cu^+]$$

E. Essay

35. A $0.25M$ H_3PO_4 solution contains 0.25 moles of solute per liter of total solution. Since H_3PO_4 is triprotic, each mole of the acid contains 3 moles of hydrogen ions, or 3 equivalents. The normality of this solution is thus $0.75N$. This means that the $0.75N$ solution contains 0.75 equivalents per liter of solution. For H_3PO_4, $N = 3 \times M$.

F. Additional Questions and Problems

36. $$15.0 \text{ g } H_3PO_4 \times \frac{1 \text{ mol } H_3PO_4}{98.0 \text{ g } H_3PO_4} \times 3 \text{ equiv/1 mol} = 0.459 \text{ equiv } H_3PO_4$$

37. $N_1V_1 = N_2V_2$

$$N_1 = \frac{N_2V_2}{V_1}$$

$$N_1 = \frac{(4.0)(150.0 \text{ mL})}{(250.0 \text{ mL})}$$

$$N_1 = 2.4N$$

38. $CaSO_4(s) \rightleftharpoons Ca^{2+}(aq) + SO_4^{2-}(aq)$

$$K_{sp} = [Ca^{2+}][SO_4^{2-}]$$

$$6.10 \times 10^{-5} = [0.450][SO_4^{2-}]$$

$$1.36 \times 10^{-4}M = [SO_4^{2-}]$$

39. $[Ag^+] = 5.0 \times 10^{-8}M$ and $[Cl^-] = 1.0 \times 10^{-9}M$.

$$K_{sp} = [Ag^+][Cl^-]$$

$$K_{sp} = [5.0 \times 10^{-8}][1.0 \times 10^{-9}]$$

$$K_{sp} = [5.0 \times 10^{-17}]$$

Since 5.0×10^{-17} is less than the K_{sp} of 1.8×10^{-10}, no precipitate will be produced.

40. $[Pb^{2+}] = 0.200M$ and $[SO_4^{2-}] = 0.100M$.

$$K_{sp} = [Pb^{2+}][SO_4^{2-}]$$

$$K_{sp} = [0.200][0.100]$$

$$K_{sp} = 2.00 \times 10^{-2}$$

Since 2.0×10^{-2} is greater than the K_{sp} of 1.8×10^{-8}, a precipitate will be produced.

Section Review 22.1

Part A Completion

1. redox
2. away
3. toward
4. reduction
5. oxidizing
6. reduced
7. reducing
8. oxidized

Part B True-False

9. AT
10. AT
11. NT
12. NT

Part C Matching

13. e
14. b
15. f
16. c
17. a
18. d

Part D Questions and Problems

19. Oxidation is the complete or partial loss of electrons. Reduction is the complete or partial gain of electrons.

20. The zinc metal, Zn, was oxidized and is the reducing agent. The copper ion, Cu^{2+}, was reduced and is the oxidizing agent.

21. When oxygen and water attack iron, the iron atoms lose electrons as the iron begins to be oxidized. Since aluminum and zinc are better reducing agents than iron and are more easily oxidized, they immediately transfer electrons to the iron ions, reducing them back to neutral iron atoms.

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Section Review 22.2

Part A Completion

1. zero
2. sign
3. charge
4. zero
5. charge on the ion
6. electron
7. oxidation
8. decrease

Part B True-False

9. AT
10. AT
11. NT
12. NT
13. NT
14. NT
15. AT
16. AT

Part C Matching

17. e
18. h
19. d
20. a
21. g
22. c
23. f
24. i
25. b

Part D Questions and Problems

26. An increase in the oxidation number of an atom indicates oxidation. A decrease in the oxidation number indicates reduction.

27. N is reduced (+5 to +2);
Br is oxidized (−1 to 0)

Mn is reduced (+7 to +2);
Cl is oxidized (−1 to 0)

N is reduced (+5 to +2);
Sb is oxidized (0 to +5)

S is reduced (+6 to +4);
C is oxidized (0 to +4)

Section Review 22.3

Part A Completion

1. oxidation number
2. half-reaction
3. balanced
4. ionic
5. two
6. added
7. ionic

Part B True-False

8. AT
9. AT
10. AT
11. NT
12. NT
13. NT

Part C Matching

14. c
15. a
16. f
17. b
18. e
19. d
20. g

Part D Questions and Problems

21. **a.** $HNO_3 \rightarrow NO$; N changes +5 to +2, a gain of $3e^-$; multiply by 2

$2HI \rightarrow I_2$; I changes from −1 to 0, a loss of $2e^-$ for I_2; multiply by 3

$2HNO_3 + 6HI \rightarrow 2NO + 3I_2 + 4H_2O$

b. $HNO_3 \rightarrow NO_2$; N changes +5 to +4, a gain of $1e^-$; multiply by 10

$I_2 \rightarrow 2HIO_3$; I changes from 0 to +5, a loss of $10e^-$ for I_2

$10HNO_3 + I_2 \rightarrow 2HIO_3 + 10NO_2 + 4H_2O$

22. **a.** $S^{2-} \rightarrow S + 2e^-$ and
$3e^- + 4H^+ + NO_3^- \rightarrow NO + 2H_2O$;
Multiply the oxidation reaction by 3 and the reduction reaction by 2.

$3H_2S + 2HNO_3 \rightarrow 3S + 2NO + 4H_2O$

b. $Fe^{2+} \rightarrow Fe^{3+} + e^-$ and $6e^- + 14H^+ + Cr_2O_7^{2-} \rightarrow 2Cr^{3+} + 7H_2O$; multiply the oxidation reaction by 6

$14H^+ + 6Fe^{2+} + Cr_2O_7^{2-} \rightarrow 6Fe^{2+} + 2Cr^{3+} + 7H_2O$

Practice Problems

Section 22.1

1. Sr: oxidized (reducing agent)
   O_2: reduced (oxidizing agent)
2. Li: oxidized (reducing agent)
   S: reduced (oxidizing agent)
3. Cs: oxidized (reducing agent)
   Br_2: reduced (oxidizing agent)
4. Mg: oxidized (reducing agent)
   N_2: reduced (oxidizing agent)
5. Fe: oxidized (reducing agent)
   O_2: reduced (oxidizing agent)
6. Br^-: oxidized (reducing agent)
   Cl_2: reduced (oxidizing agent)
7. Si: oxidized (reducing agent)
   F_2 reduced (oxidizing agent)
8. Ca: oxidized (reducing agent)
   O_2 reduced (oxidizing agent)
9. Mg: oxidized (reducing agent)
   H^+: reduced (oxidizing agent)
10. Na: oxidized (reducing agent)
    H_2O: reduced (oxidizing agent)

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Section 22.2

1. **a.** Sn is tin in an uncombined state. The oxidation number is 0.
   b. The ionic charge on potassium is 1+, thus the oxidation number is +1.
   c. The ionic charge on sulfur is 2−, thus the oxidation number is −2.
   d. The ionic charge on iron is 3+, thus the oxidation number is +3.
   e. Se is selenium in an uncombined state. The oxidation number is 0
   f. The ionic charge on magnesium is 2+, thus the oxidation number is +2.
   g. The ionic charge on tin is 4+, thus the oxidation number is +4.
   h. The ionic charge on bromine is 1−, thus the oxidation number is −1.
2. **a.** +3 **c.** +2
   b. +6 **d.** +6
3. **a.** $\overset{0}{C} + \overset{+1\ +6-2}{H_2SO_4} \rightarrow \overset{+4\ -2}{CO_2} + \overset{+4\ -2}{SO_2} + \overset{+1\ -2}{H_2O}$
   Carbon is oxidized ($0 \rightarrow +4$).
   Sulfur is reduced ($+6 \rightarrow +4$).
   b. $\overset{+1+5\ -2}{HNO_3} + \overset{+1-1}{HI} \rightarrow \overset{+2\ -2}{NO} + \overset{0}{I_2} + \overset{+1\ -2}{H_2O}$
   Nitrogen is reduced ($+5 \rightarrow +2$).
   Iodide ion is oxidized ($-1 \rightarrow 0$).
   c. $\overset{+1\ +7\ -2}{KMnO_4} + \overset{+1-1}{HCl} \rightarrow \overset{+2\ -1}{MnCl_2} + \overset{0}{Cl_2} + \overset{+1\ -2}{H_2O} + \overset{+1-1}{KCl}$
   Manganese is reduced ($+7 \rightarrow +2$).
   Chloride ion is oxidized ($-1 \rightarrow 0$).
   d. $\overset{0}{Sb} + \overset{+1+5\ -2}{HNO_3} \rightarrow \overset{+5\ -2}{Sb_2O_5} + \overset{+2\ -2}{NO} + \overset{+1\ -2}{H_2O}$
   Antimony is oxidized ($0 \rightarrow +5$).
   Nitrogen is reduced ($+5 \rightarrow +2$).
4. **a.** Oxidizing agent is sulfur;
   Reducing agent is carbon.
   b. Oxidizing agent is nitrogen;
   Reducing agent is iodine.
   c. Oxidizing agent is manganese;
   Reducing agent is chlorine.
   d. Oxidizing agent is nitrogen;
   Reducing agent is antimony.

Section 22.3

1. **a.** Increase in oxidation number of carbon = +4; decrease in oxidation number of sulfur = −2.
   $C + 2H_2SO_4 \rightarrow CO_2 + 2SO_2 + 2H_2O$
   b. Increase in oxidation number of sulfur = +2; decrease in oxidation number of nitrogen = −3.
   $3H_2S + 2HNO_3 \rightarrow 3S + 2NO + 4H_2O$
   c. Increase in oxidation number of iodine ion = +1; decrease in oxidation number of nitrogen = −3.
   $2HNO_3 + 6HI \rightarrow 2NO + 3I_2 + 4H_2O$
   d. Increase in oxidation number of antimony = +5; decrease in oxidation number of nitrogen = −3.
   $6Sb + 10HNO_3 \rightarrow 3Sb_2O_5 + 10NO + 5H_2O$
   e. Increase in oxidation number of chlorine ion = +1; decrease in oxidation number of manganese = −5.
   $2KMnO_4 + 16HCl \rightarrow 2MnCl_2 + 5Cl_2 + 8H_2O + 2KCl$
   f. Increase in oxidation number of iodine ion = +1; decrease in oxidation number of iodine = −7.
   $KIO_4 + 7KI + 8HCl \rightarrow 8KCl + 4I_2 + 4H_2O$
   g. Increase in oxidation number of zinc = +2; decrease in oxidation number of chromium = −3.
   $3Zn + 2Cr_2O_7^{2-} + 28H^+ \rightarrow 3Zn^{2+} + 4Cr^{3+} + 14H_2O$
2. **a.** $Fe^{2+} \rightarrow Fe^{3+} + e^-$
   $5e^- + 8H^+ + MnO_4^- \rightarrow Mn^{2+} + 4H_2O$
   b. $Sn^{2+} \rightarrow Sn^{4+} + 2e^-$
   $6H^+ + 6e^- + IO_3^- \rightarrow I^- + 3H_2O$
   c. $S^{2-} \rightarrow S + 2e^-$
   $3e^- + 4H^+ + NO_3^- \rightarrow NO + 2H_2O$
   d. $4OH^- + Mn^{2+} \rightarrow MnO_2 + 2H_2O + 2e^-$
   $2e^- + H_2O + H_2O_2 \rightarrow H_2O + 2OH^-$
3. **a.** $2OH^- + Zn + HgO \rightarrow ZnO_2^{2-} + Hg + H_2O$
   b. $8H^+ + 5Fe^{2+} + MnO_4^- \rightarrow 5Fe^{3+} + Mn^{2+} + 4H_2O$
   c. $6H^+ + 3Sn^{2+} + IO_3^- \rightarrow 3Sn^{4-} + I^- + 3H_2O$
   d. $8H^+ + 3S^{2-} + 2NO_3^- \rightarrow 3S + 2NO + 4H_2O$
   e. $2OH^- + Mn^{2+} + H_2O_2 \rightarrow MnO_2 + 2H_2O$
   f. $2OH^- + CrO_2 + ClO^- \rightarrow CrO_4^{2-} + Cl^- + H_2O$

Interpreting Graphics 22

1. **a.** 3 **d.** 2
   b. 2 **e.** 4
   c. 3 **f.** 1

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2. $MnO_4^- + 8H^+ + 5Fe^{2+} \rightarrow Mn^{2+} + 4H_2O + 5Fe^{3+}$

3. The end point occurs when the number of equivalents of MnO_4^- added equals the number of equivalents of Fe^{2+} originally present in the reaction flask. One equivalent is the amount of reducing agent (or oxidizing agent) that can give (or accept) one mole of electrons. When all the Fe^{2+} in the flask is oxidized, the next drop of MnO_4^- remains unreacted, and the solution in the flask turns light purple, signaling the end point of the titration.

4. Volume $KMnO_4$ = Initial Volume − Final Volume = 48.65 mL − 23.35 mL = 25.30 mL

$$\text{Moles } MnO_4^- = 0.02530 \cancel{\text{L}} \times \frac{0.0200 \text{ mol}}{\cancel{\text{L}}}$$
$$= 5.06 \times 10^{-4} \text{ mol } MnO_4^-$$

$$\text{Moles iron(II)} = 5.06 \times 10^{-4} \cancel{\text{mol } MnO_4^-} \times \frac{5 \text{ mol } Fe^{2+}}{1 \cancel{\text{mol } MnO_4^-}} = 2.53 \times 10^{-3} \text{ mol } Fe^{2+}$$

$$\text{Mass Fe} = 2.53 \times 10^{-3} \cancel{\text{mol } Fe^{2+}} \times \frac{55.85 \text{ g } Fe^{2+}}{1 \cancel{\text{mol } Fe^{2+}}} = 0.141 \text{ g}$$

$$\% \text{ Fe in ore} = \frac{0.141 \text{ g}}{2.938 \text{ g}} \times 100\% = 4.80\%$$

Vocabulary Review 22

1. oxidizing agent
2. oxidation-number-change method
3. reduction
4. half-reaction method
5. oxidation number
6. reducing agent
7. half-reaction
8. oxidation-reduction reaction
9. oxidation
10. redox reaction

Quiz for Chapter 22

1. b	4. b	7. c
2. a	5. a	8. d
3. b	6. a	

Chapter 22 Test A

A. Matching

1. d	5. i	9. c
2. j	6. g	10. h
3. f	7. e	
4. a	8. b	

B. Multiple Choice

11. b	17. c	23. d
12. d	18. c	24. c
13. c	19. a	25. c
14. b	20. a	26. c
15. a	21. c	
16. b	22. b	

C. Questions

27. **a.** Na oxidized, reducing agent; Br_2 reduced, oxidizing agent
    b. S reduced, oxidizing agent; K oxidized, reducing agent
28. $2Cr + 3Br_2 \rightarrow 2Cr^{3+} + 6Br^-$
29. **a.** Li +1, Al +3, F −1
    b. Na +1, O −2
    c. S 0 (element)
30. Oxidation-number change method:

$$\overset{+3\ -2}{Fe_2O_3} + \overset{+2\ -2}{CO} \rightarrow \overset{0}{Fe} + \overset{+4\ -2}{CO_2}$$

(2 × (−3) = −6; 3 × (+2) = +6)

$Fe_2O_3 + 3CO \rightarrow 2Fe + 3CO_2$

Half-reaction method:

$6H^+ + Fe_2O_3 + 6e^- \rightarrow 2Fe + 3H_2O$

$3(H_2O + CO \rightarrow CO_2 + 2H^+ + 2e^-)$

$6H^+ + Fe_2O_3 + 6e^- + 3H_2O + 3CO \rightarrow 2Fe + 3H_2O + 3CO_2 + 6H^+ + 6e^-$

$Fe_2O_3 + 3CO \rightarrow 2Fe + 3CO_2$

D. Essay

31. An oxidation number is assigned to an element in a compound according to a set of arbitrary rules. The oxidation number of an element in an uncombined state is zero. The oxidation number of a monatomic ion is the same in magnitude and sign as the ionic charge. The sum of the oxidation numbers of the elements in a neutral compound is zero. In a polyatomic ion, however, the sum is equal to the charge on the ion. Oxidation

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numbers help keep track of electrons in redox reactions. An oxidation-number increase is oxidation. A decrease is reduction.

Chapter 22 Test B

A. Matching

1. e	5. i	9. b
2. f	6. a	10. j
3. d	7. c	
4. g	8. h	

B. Multiple Choice

11. d	17. a	23. d
12. a	18. b	24. c
13. a	19. c	25. d
14. d	20. c	26. b
15. d	21. b	27. b
16. b	22. b	28. a

C. Questions

29. **a.** K; I; I_2; K
b. Na; H; H_2O; Na
c. H; Cu; CuO; H_2
d. Mg; Cu; $Cu(NO_3)_2$; Mg

30. **a.** K_2SO_4 = +1, +6, −2
b. $Cu(NO_3)_2$ = +2, +5, −2
c. $HAsO_3$ = +1, +5, −2
d. MnO_4^- = +7, −2

31. **a.** $4HNO_3 + 3Ag \rightarrow 3AgNO_3 + NO + 2H_2O$
($3 \times (+1) = +3$; $1 \times (-3) = -3$)
b. $Br_2 + SO_2 + 2H_2O \rightarrow H_2SO_4 + 2HBr$
($1 \times (+2) = +2$; $2 \times (-1) = -2$)

32. **a.** $HNO_2 + HI \rightarrow I_2 + NO + H_2O$
$H^+(aq) + NO_2^-(aq) + H^+(aq) + I^-(aq) \rightarrow I_2(aq) + NO(g) + H_2O(l)$
Oxidation: $2I^-(aq) \rightarrow I_2 + 2e^-$
Reduction: $2[2H^+(aq) + NO_2^-(aq) + 1e^- \rightarrow NO + H_2O]$
$4H^+ + 2NO_2^- + 2e^- \rightarrow 2NO + 2H_2O$
$4H^+ + 2I^- + 2NO_2^- \rightarrow I_2 + 2NO + H_2O$
Final: $2HNO_2 + 2HI \rightarrow I_2 + 2NO + 2H_2O$

b. $K_2Cr_2O_7 + FeCl_2 + HCl \rightarrow CrCl_3 + KCl + FeCl_3 + H_2O$
$2K^+(aq) + Cr_2O_7^{2-}(aq) + Fe^{2+}(aq) + 2Cl^- + H^+(aq) + Cl^-(aq) \rightarrow Cr^{3+}(aq) + 3Cl^-(aq) + K^+(aq) + Cl^-(aq) + Fe^{3+}(aq) + 3Cl^-(aq) + H_2O$
Oxidation: $6[Fe^{2+} \rightarrow Fe^{2+} + 1e^-]$
Reduction: $2Cr^{6+} + 6e^- \rightarrow 2Cr^{3+}$
$6Fe^{2+} + 2Cr^{6+} \rightarrow 6Fe^{3+} + 2Cr^{3+}$
Final: $K_2Cr_2O_7 + 6FeCl_2 + 14HCl \rightarrow 2CrCl_3 + 2KCl + 6FeCl_3 + 7H_2O$

D. Essay

33. Since oxidation is the loss of electrons, it can only occur in the presence of another substance that will accept the lost electrons. The accepting substance gains electrons, and thus, undergoes reduction. In other words, a loss of electrons can only occur if a gain takes place concurrently.

Section Review 23.1

Part A Completion

1. electrochemical process
2. electrons
3. voltaic cells
4. salt bridge
5. ions
6. anode
7. cathode

Part B True-False

8. NT	10. NT
9. AT	11. ST

Part C Matching

12. g	15. b	17. e
13. f	16. c	18. a
14. d		

Part D Questions and Problems

19. The shorthand notation $Mg(s) \mid MgSO_4(aq) \parallel PbSO_4(aq) \mid Pb(s)$ represents a magnesium-lead voltaic cell. The single vertical lines indicate boundaries of phases that are in contact, and the double vertical lines represent the salt bridge that separates the anode compartment from the cathode compartment. In this electrochemical cell,

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Mg is oxidized to Mg^{2+} at the anode (the negative electrode) and Pb^{2+} is reduced to Pb at the cathode (the positive electrode). Electrons flow from the anode, through an external circuit (connected to a light bulb or voltmeter), to the cathode. To complete the circuit, sulfate (SO_4^{2-}) anions move from the cathode compartment to the anode compartment, and magnesium and sodium cations move from the anode compartment to the cathode compartment. Check students' diagrams.

Section Review 23.2

Part A Completion

1. electric potential
2. electrons
3. electrons
4. reduction
5. cell potential
6. standard hydrogen electrode
7. 0.00 V
8. less
9. spontaneous

Part B True-False

10. NT
11. ST
12. NT
13. NT

Part C Matching

14. b
15. d
16. f
17. c
18. a
19. e

Part D Questions and Problems

20. Oxidation: $Mg \rightarrow Mg^{2+} + 2e^-$

Reduction: $2e^- + Cl_2 \rightarrow 2Cl^-$

Redox: $Mg + Cl_2 \rightarrow Mg^{2+} + 2Cl^-$

$E^0_{cell} = E^0_{red} - E^0_{oxid}$

$= E^0_{Cl_2} - E^0_{Mg}$

$= +1.36\,V - (-2.37\,V)$

$= +3.73\,V$

Section Review 23.3

Part A Completion

1. electrolysis
2. electrolytic cell
3. electrons
4. battery
5. electrolyte
6. hydrogen/oxygen
7. oxygen/hydrogen
8. sodium metal

Part B True-False

9. AT
10. ST
11. ST
12. AT

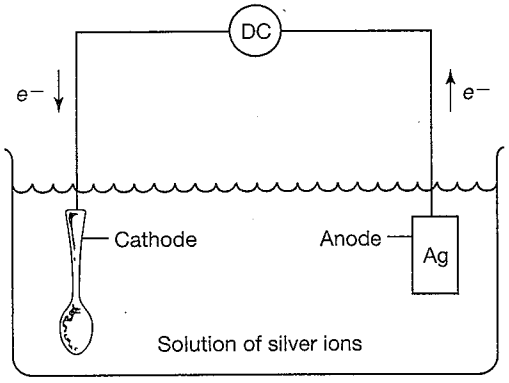
Part C Matching

13. b
14. d
15. e
16. c
17. a

Part D Questions and Problems

18. In electrolytic cells, electrical energy is used to bring about a normally nonspontaneous chemical reaction. In a voltaic cell, chemical energy is converted to electrical energy by a spontaneous redox reaction. Electrolytic cells are used in electroplating, in refining metals, and in the production of substances such as sodium hydroxide, aluminum, sodium, and chlorine. Voltaic cells are used in pacemakers, hearing aids, and cameras.

19.



Anode (oxidation): $Ag(s) \rightarrow Ag^+(aq) + e^-$

Cathode (reduction): $Ag^+(aq) + e^- \rightarrow Ag(s)$

Practice Problems

Section 23.2

1. a. $Cl_2(g) + Mg(s) \rightarrow 2Cl^-(aq) + Mg^{2+}(aq)$

$E^0_{cell} = 1.36\,V - (-2.37\,V) = 3.73\,V$

cathode: $Cl_2(g) + 2e^- \rightarrow 2Cl^-(aq)$

b. $2Ag^+(aq) + Ni(s) \rightarrow Ni^{2+}(aq) + 2Ag(s)$

$E^0_{cell} = 0.80\,V - (-0.25\,V) = 1.05\,V$

cathode: $Ag^+(aq) + e^- \rightarrow Ag(s)$

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c. $2MnO_4^-(aq) + 16H^+(aq) + 5Cd(s) \rightarrow 5Cd^{2+}(aq) + 2Mn^{2+}(aq) + 8H_2O(l)$

$E^0_{cell} = 1.51\,V - (-0.40\,V) = 1.91\,V$

cathode: $MnO_4^-(aq) + 8H^+(aq) + 5e^- \rightarrow Mn^{2+}(aq) + 4H_2O(l)$

d. $Br_2 + 2Na(s) \rightarrow 2Na^+(aq) + 2Br^-(aq)$

$E^0_{cell} = 1.07\,V - (-2.71\,V) = 3.78\,V$

cathode: $Br_2(l) + 2e^- \rightarrow 2Br^-(aq)$

e. $MnO_2(s) + 4H^+(aq) + H_2(g) \rightarrow 2H^+(aq) + Mn^{2+}(aq) + 2H_2O(l)$

$E^0_{cell} = 1.28\,V - 0.00\,V = 1.28\,V$

cathode: $MnO_2(s) + 4H^+(aq) + 2e^- \rightarrow Mn^{2+}(aq) + 2H_2O(l)$

2. **a.** $E^0_{cell} = -0.14\,V - (-2.90\,V) = +2.76\,V$; spontaneous

b. $E^0_{cell} = +0.80\,V - 1.36\,V = -0.56\,V$; nonspontaneous

c. $E^0_{cell} = +2.87\,V - (-0.76\,V) = +3.63\,V$; spontaneous

d. $E^0_{cell} = -0.28\,V - (-3.05\,V) = +2.77\,V$; spontaneous

e. $E^0_{cell} = -2.93\,V - 0.54\,V = -3.47\,V$; nonspontaneous

Interpreting Graphics 23

1. anode(+)
2. cathode(−)
3. electrorefining
4. **a.** The anode(+) of the electrolytic cell should be connected to the positive(+) terminal of the battery. The cathode(−) of the electrolytic cell should be connected to the negative(−) terminal of the battery.

   b. The anode of the electrolytic cell is connected to the cathode of the battery. The cathode of the electrolytic cell is connected to the anode of the battery.
5. Oxidation occurs at the anode, labeled number 1 in the diagram. Reduction occurs at the cathode, labeled number 2 in the diagram.
6. Students should indicate the flow of electrons out of the anode(+) and into the cathode(−).
7. The voltage should be great enough to oxidize copper metal at the anode and reduce copper(II) ions at the cathode, but not high enough to oxidize other metals at the anode and reduce them at the cathode. The voltage should be greater than 0.34 V but less than 0.44 V.
8. **a.** gold, silver, and platinum

   b. zinc 2+ and iron 2+

   c. copper

Vocabulary Review 23

1. voltaic cell
2. fuel cell
3. electrochemical cell
4. electrochemical process
5. cathode
6. reduction potential

Solution: aluminum

Quiz for Chapter 23

1. NT	7. NT	13. NT
2. AT	8. AT	14. AT
3. NT	9. AT	15. NT
4. AT	10. ST	16. AT
5. ST	11. NT	17. AT
6. NT	12. AT	18. NT

Chapter 23 Test A

A. Matching

1. b	5. j	9. f
2. h	6. g	10. a
3. d	7. i	
4. c	8. e	

B. Multiple Choice

11. d	16. a	21. c
12. b	17. d	22. c
13. c	18. c	23. a
14. c	19. a	24. a
15. c	20. a	25. c

C. True-False

26. NT	28. AT	30. NT
27. NT	29. NT	

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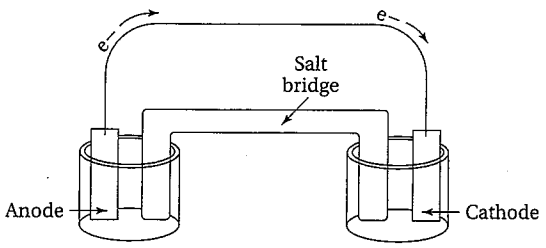
D. Question

31.



E. Essay

32. In both voltaic and electrolytic cells, oxidation occurs at the anode and reduction occurs at the cathode. In the voltaic cell, the cathode is positive and the anode is negative. In the electrolytic cell, the anode is positive and the cathode is negative.

F. Additional Questions

33. The negative value means that the tendency for zinc ions to be reduced is less than that of hydrogen ions to be reduced, so zinc metal is oxidized when paired with the standard hydrogen half-cell.

34. The reduction potential of a half-cell is a measure of the tendency of a given half-reaction to occur as a reduction.

35. **a.** $Ni^{2+}(aq) + 2e^- \rightarrow Ni(s)$ $\quad E^0 = -0.25\,V$

$Zn^{2+}(aq) + 2e^- \rightarrow Zn(s)$ $\quad E^0 = -0.76\,V$

$E^0_{cell} = E^0_{red} - E^0_{oxid}$

$= -0.76\,V - (-0.25\,V)$

$= -0.51\,V$

This reaction is not spontaneous.

b. $2[Al^{3+}(aq) + 3e^- \rightarrow Al(s)]$ $\quad E^0 = -1.66\,V$

$3[Co^{2+}(aq) + 2e^- \rightarrow Co(s)]$ $\quad E^0 = -0.28\,V$

$E^0_{cell} = E^0_{red} - E^0_{oxid}$

$= -1.66\,V - (-0.28\,V)$

$= -1.38\,V$

This reaction is not spontaneous.

Chapter 23 Test B

A. Matching

1. h	**5.** c	**9.** b
2. d	**6.** g	**10.** a
3. f	**7.** j	
4. i	**8.** e	

B. Multiple Choice

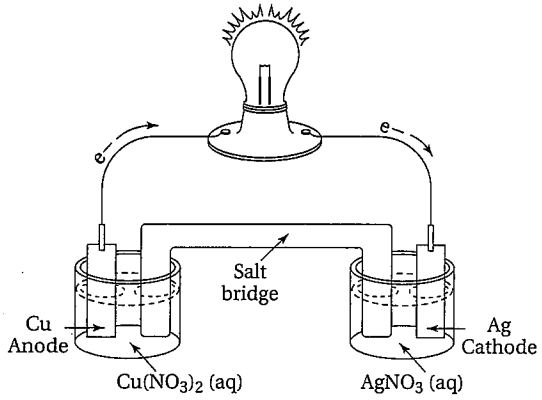
11. c	**17.** d	**23.** b
12. d	**18.** a	**24.** a
13. b	**19.** d	**25.** b
14. c	**20.** c	**26.** c
15. d	**21.** b	**27.** d
16. a	**22.** c	

C. True-False

28. NT	**30.** AT	**32.** AT
29. NT	**31.** AT	**33.** ST

D. Question

34.



35. $Cu(s) \rightarrow Cu^{2+}(aq) + 2e^-$

$2Ag^+(aq) + 2e^- \rightarrow 2Ag(s)$

Net: $Cu(s) + 2Ag^+(aq) \rightarrow Cu^{2+}(aq) + 2Ag(s)$

$E^0_{cell} = E^0_{red} - E^0_{oxid}$

$= +0.80\,V - (+0.34\,V)$

$= +0.46\,V$

E. Essay

36. In both voltaic and electrolytic cells, electrons flow from the anode to the cathode through the external circuit, reduction occurs at the cathode, and oxidation occurs at the anode. However, while the flow of electrons in a voltaic cell is caused by a spontaneous chemical reaction, in an electrolytic cell the flow of electrons is being pushed by an outside source such as a battery. Additionally, while the anode is the negative electrode and the cathode is the positive electrode in a voltaic cell, the reverse is true in an electrolytic cell—the anode in an electrolytic cell is the positive electrode and the cathode is negative.

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F. Additional Questions

37. Since Al is above Pb in the reduction potential table, Al is oxidized and Pb is reduced. Thus, the two half-cell reactions are as follows:

$2[Al(s) \rightarrow Al^{3+}(aq) + 3e^-]$

$3[Pb^{2+}(aq) + 2e^- \rightarrow Pb(s)]$

Net: $2Al(s) + 3Pb^{2+}(aq) \rightarrow 2Al^{3+}(aq) + 3Pb(s)$

$E^0_{cell} = E^0_{red} - E^0_{oxid}$

$= -0.13\,V - (-1.66\,V)$

$= +1.53\,V$

38. a. The half-reactions are:

Oxidation: $Na(s) \rightarrow Na^+(aq) + e^-$

$E^0 = -2.71\,V$

Reduction: $Cu^{2+}(aq) + 2e^- \rightarrow Cu(s)$

$E^0 = +0.34\,V$

$E^0_{cell} = E^0_{red} - E^0_{oxid}$

$E^0_{cell} = +0.34\,V - (-2.71\,V)$

$E^0_{cell} = +3.05\,V$

Since the standard cell potential is positive, the redox reaction will be spontaneous.

b. The half-reactions are:

Oxidation: $Ag(s) \rightarrow Ag^+(aq) + e^-$

$E^0 = +0.80\,V$

Reduction: $Mg^{2+}(aq) + 2e^- \rightarrow Mg(s)$

$E^0 = -2.37\,V$

$E^0_{cell} = E^0_{red} - E^0_{oxid}$

$= -2.37\,V - (+0.80\,V)$

$= -3.17\,V$

Since the standard cell potential is negative, the redox reaction will be nonspontaneous.

c. The half-reactions are:

Oxidation: $Al(s) \rightarrow Al^{3+}(aq) + 3e^-$

$E^0 = -1.66\,V$

Reduction: $Zn^{2+}(aq) + 2e^- \rightarrow Zn(s)$

$E^0 = -0.76\,V$

$E^0_{cell} = E^0_{red} - E^0_{oxid}$

$= -0.76\,V - (-1.66\,V)$

$= +0.90\,V$

Since the standard cell potential is positive, the redox reaction will be spontaneous.

39. Reaction 1: since Z is oxidized and X^{2+} is reduced, Z appears above X in the activity series.

Reaction 2: since W is not oxidized in the presence of X^{2+}, W should appear below X.

Reaction 3: since Y is oxidized and Z^{2+} is reduced, Y appears above Z.

The elements should be listed as follows: Y, Z, X, and W.

Section Review 24.1

Part A Completion

1. alkali
2. alkali salts
3. one
4. sodium hydroxide
5. hydrogen gas
6. electrical
7. sodium chloride
8. Chlorine gas
9. alkaline-earth
10. Magnesium
11. calcium
12. slaked lime

Part B True-False

13. AT **15.** AT **17.** AT
14. NT **16.** AT

Part C Matching

18. e **20.** d **22.** b
19. c **21.** a

Part D Questions and Problems

23. a. $2Na(s) + 2H_2O(l) \rightarrow 2NaOH(aq) + H_2(g)$

b. $Ca(s) + 2H_2O(l) \rightarrow Ca(OH)_2(aq) + H_2(g)$

24. The chemical reactivity increases as you go down each group. Of the alkali metals, lithium is the least reactive and cesium is the most reactive. Of the alkaline-earth metals, beryllium is the least active and barium is the most active.

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Section Review 24.2

Part A Completion

1. boron
2. borax
3. aluminum
4. bauxite
5. Al_2O_3 or aluminum oxide
6. diamond/graphite
7. diamond/graphite
8. nitrogen
9. N_2/a diatomic molecule
10. ammonia
11. Oxygen
12. ozone (O_3)

Part B True-False

13. AT
14. ST
15. ST
16. AT
17. NT

Part C Matching

18. c
19. f
20. e
21. a
22. b
23. d

Part D Questions and Problems

24. The halogens exist as compounds in nature. Fluorine and chlorine are greenish-yellow reactive gases, bromine is a dark-red liquid, and iodine is a purple-black solid. Most of the compounds of the halogens are soluble in water. Fluorine is the most reactive of all nonmetals and the strongest elemental oxidizing agent known. Chlorine is used to purify drinking water and swimming pools, and in sewage treatment. Chlorine is also used to make vinyl chloride, which is used to make the plastic polyvinyl chloride (PVC). Chloride ions are an important component of the blood and other body fluids. Iodide ions are necessary for thyroid function.

Section Review 24.3

Part A Completion

1. ductile/malleable
2. ductile/malleable
3. oxidation state
4. tungsten
5. mercury
6. copper
7. Zinc
8. Iron
9. steel
10. Gold
11. inner transition elements
12. lanthanide
13. actinides

Part B True-False

14. AT
15. ST
16. ST
17. AT

Part C Matching

18. e
19. b
20. d
21. a
22. c

Part D Questions and Problems

23. There are two types of steels: carbon steels and alloy steels. Both types actually contain carbon, but carbon steels contain no other metal other than iron. Carbon steel, with less than 0.2% carbon, is called mild steel and is used when load-bearing ability is not important. Medium steels, which contain from 0.2% to 0.6% carbon, are used for structural materials. High carbon steels contain between 0.8% and 1.5% carbon and are used to make items for which hardness is important. Alloy steels are called stainless steels and contain high percentages of chromium and nickel. The most common stainless steel contains 18% chromium and 8% nickel.

24. $\frac{18}{24} \times 100\% = 75\%$ gold

Section Review 24.4

Part A Completion

1. hydrogen
2. alkali metal
3. halogen
4. hydrogen bromide (HBr)
5. sodium hydride (NaH)
6. Water
7. ammonia
8. rocket fuel
9. noble gases

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10. separate
11. xenon tetrafluoride (XeF_4)

Part B True-False

12. AT
13. NT
14. AT
15. AT

Part C Matching

16. d
17. a
18. c
19. e
20. f
21. b

Part D Questions and Problems

22. **a.** $H_2(g) + Cl_2(g) \rightarrow 2HCl(g)$
    b. $H_2(g) + 2Na(s) \rightarrow 2NaH(s)$
23. Hydrogen, like the halogens, has one fewer valence electron than the noble gas it precedes. Like the halogens, it reacts with alkali metals to form compounds in which it has an oxidation state of −1. It is a gas and the gases appear in the upper right-hand section of the periodic table. These properties make it reasonable to place hydrogen in Group 7A. Hydrogen is more often placed above Group 1A because, like an alkali metal, it reacts with halogens to form compounds in which it has an oxidation state of +1.

Interpreting Graphics 24

1. Check students' tables.
2. Check students' tables.
3. Check students' tables. Most of the 92 naturally occurring elements are solids at room temperature.
4. Cesium is the most reactive alkali metal. Cesium is expected to be much less electronegative than fluorine. As a group, the alkali metals are the least electronegative elements in the periodic table. Because electronegativity decreases as you move down a group, cesium, and francium, a rare and radioactive element, are expected to be the least electronegative elements of all. Fluorine, on the other hand, is the most electronegative element in the periodic table.
5. Beryllium is chemically the most stable element in Group 2A.
6. All of the noble gases are essentially inert (except in rare instances).
7. Group 7A elements are called halogens; they behave as oxidizing agents in most chemical reactions. Fluorine is the strongest elemental oxidizing agent known.
8. The metalloids are solids at room temperature; they conduct electricity at high temperatures but act as electrical insulators at low temperatures. The *p*-block consists of Groups 3A, 4A, 5A, 6A, 7A, and 0 with the exception of helium. The *p*-block elements have electrons in their outermost *p* sublevel. In general, the valence electron configuration for these elements could be written as ns^2np^x, where $1 \leq x \geq 6$.
9. Mercury is a liquid at room temperature. Platinum is often used to catalyze chemical reactions.
10. Check students' tables and keys.

Vocabulary Review 24

1. b
2. a
3. d
4. e
5. f
6. c
7. i
8. j
9. g
10. h

Quiz for Chapter 24

1. c
2. b
3. a
4. b
5. ST
6. NT
7. NT
8. ST
9. AT
10. silicon
11. phosphorus
12. sulfur
13. fluorine

Chapter 24 Test A

A. Matching

1. d
2. i
3. f
4. h
5. c
6. a
7. g
8. j
9. e
10. b

B. Multiple Choice

11. a
12. a
13. b
14. c
15. b
16. c
17. c
18. c
19. a
20. d
21. b
22. b
23. c
24. d
25. b

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C. True-False

26. AT
27. AT
28. AT
29. AT
30. NT
31. NT
32. NT
33. AT
34. AT
35. NT
36. AT
37. AT

D. Questions

38. In general, the more reactive an element, the more likely it is to occur in nature in the combined, rather than in the free, state. Thus, while the alkali metals always occur in the combined state, the noble gases are almost always found in the free state.

39. The Group 0 elements are named the noble gases because, like nobility, they do not mix with others—in this case, other atoms. Noble gases are unreactive because they have full outer shell electron configurations and, thus, have no need to lose, gain, or share electrons.

E. Essay

40. The alkali metals are more reactive than the alkaline-earth metals; neither group occurs free in nature, the alkali metals must be stored under oil or kerosene to prevent their reacting with water; no such precaution is needed for the alkaline-earth metals. The alkali metals are more soluble in water than the alkaline-earth metals.

Chapter 24 Test B

A. Matching

1. f
2. j
3. g
4. d
5. i
6. a
7. h
8. c
9. e
10. b

B. Multiple Choice

11. d
12. b
13. c
14. c
15. b
16. d
17. c
18. b
19. c
20. d
21. a
22. d
23. c
24. c
25. b

C. True-False

26. NT
27. AT
28. AT
29. AT
30. AT
31. AT
32. AT
33. ST
34. NT
35. NT
36. AT
37. ST

D. Questions

38. All three elements have low chemical reactivity and, thus, generally occur free in nature. Since all three are shiny in the uncombined state, humans were probably drawn first to their appearance, then to their malleability.

39. Because alkali metals all contain a valence electron, they tend to lose that electron to achieve a stable noble gas configuration. The low first ionization energies of alkali metals result in frequent chemical reactions with other elements.

E. Essay

40. The halogens are very reactive, while the noble gases seldom react; the halogens are found combined in nature, but the noble gases are usually free. The iodide ion is essential for a healthy thyroid, the fluoride ion is needed for healthy teeth, and chlorine is used as a disinfectant in swimming pools. The noble gases are used in weather balloons (He), to fill deep-sea diving tanks (He and Ne, with O_2), and as the inert atmosphere in flash bulbs (Ar, Kr, and Xe).

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ANSWER KEY

Section Review 21.1

Part A Completion

1. acid
2. hydroxide
3. water
4. neutralization
5. titration
6. end point
7. equivalence
8. mass
9. one mole
10. 2 equivalents

Part B True-False

11. AT
12. AT
13. ST
14. AT

Part C Matching

15. c
16. e
17. a
18. b
19. d

Part D Questions and Problems

20. **a.** $H_3PO_4 + Al(OH)_3 \rightarrow AlPO_4 + 3H_2O$
    b. $2HI + Ca(OH)_2 \rightarrow CaI_2 + 2H_2O$

Section Review 21.2

Part A Completion

1. salt
2. acidic
3. basic
4. neutral
5. hydrolyze
6. strong
7. weak
8. buffer
9. capacity
10. solubility product constant

Part B True-False

11. NT
12. NT
13. AT
14. AT

Part C Matching

15. a
16. d
17. b
18. c

Part D Questions and Problems

19. $[CO_3^{2-}] = 0.00070M$

$[Ba^{2+}] = 0.0015M$

$[CO_3^{2-}] \times [Ba^{2+}]$

$= (7.0 \times 10^{-4}M) \times (1.5 \times 10^{-3}M)$

$= 1.1 \times 10^{-6}$

Precipitation occurs because the ion product (1.1×10^{-6}) is greater than the K_{sp} of $BaCO_3$ (5.0×10^{-9}).

Practice Problems

Section 21.1

1. $2NaOH(aq) + H_2SO_4(aq) \rightarrow Na_2SO_4(aq) + 2H_2O(l)$

$$0.014 \cancel{\text{ L } H_2SO_4} \times \frac{0.75 \cancel{\text{ mol } H_2SO_4}}{1 \cancel{\text{ L } H_2SO_4}} \times \frac{2 \text{ mol NaOH}}{1 \cancel{\text{ mol } H_2SO_4}} = 0.021 \text{ mol NaOH}$$

$$\text{Molarity} = \frac{\text{moles}}{\text{liters}} = \frac{0.021 \text{ mol NaOH}}{0.038 \text{ L NaOH}} = 0.55M \text{ NaOH}$$

2. $Ca(OH)_2(aq) + 2HC_2H_3O_2(aq) \rightarrow Ca(C_2H_3O_2)_2(aq) + 2H_2O(l)$

$$0.0142 \cancel{\text{ L } HC_2H_3O_2} \times \frac{0.0140 \cancel{\text{ mol } HC_2H_3O_2}}{1 \cancel{\text{ L } HC_2H_3O_2}} \times \frac{1 \text{ mol } Ca(OH)_2}{2 \cancel{\text{ mol } HC_2H_3O_2}} = 9.94 \times 10^{-5} \text{ mol } Ca(OH)_2$$

$= 9.94 \times 10^{-5}$ mol $Ca(OH)_2$

$$\text{Molarity} = \frac{9.94 \times 10^{-5} \text{ mol } Ca(OH)_2}{0.0246 \text{ L } Ca(OH)_2} = 0.00404M$$

3. $Ca(OH)_2(aq) + H_2SO_4(aq) \rightarrow CaSO_4(aq) + 2H_2O(l)$

$$0.0198 \cancel{\text{ L } Ca(OH)_2} \times \frac{0.0100 \cancel{\text{ mol } Ca(OH)_2}}{1 \cancel{\text{ L } Ca(OH)_2}} \times \frac{1 \text{ mol } H_2SO_4}{1 \cancel{\text{ mol } Ca(OH)_2}} = 0.000198 \text{ mol } H_2SO_4$$

$= 0.000198$ mol H_2SO_4

$$\text{Molarity} = \frac{\text{moles}}{\text{liters}} = \frac{0.000198 \text{ mol } H_2SO_4}{0.0124 \text{ L } H_2SO_4} = 0.0160M \text{ } H_2SO_4$$

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4. $Ba(OH)_2(aq) + 2HCl(aq) \rightarrow BaCl_2(aq) + 2H_2O(l)$

$$0.0122\ \text{L}\cancel{\text{HCl}} \times \frac{0.25\ \text{mol}\cancel{\text{HCl}}}{1\ \text{L}\cancel{\text{HCl}}} \times \frac{1\ \text{mol Ba(OH)}_2}{2\ \text{mol HC}}$$

$$= 0.0015\ \text{mol Ba(OH)}_2$$

$$\text{liters} = \frac{\text{moles}}{\text{molarity}} = \frac{0.0015\ \text{mol Ba(OH)}_2}{0.12M\ \text{Ba(OH)}_2}$$

$$= 0.0125\ \text{L Ba(OH)}_2 = 13\ \text{mL Ba(OH)}_2$$

5. $Al(OH)_3(aq) + 3HCl(aq) \rightarrow AlCl_3(aq) + 3H_2O(l)$

$$0.0550\ \text{g}\cancel{\text{Al(OH)}_3} \times \frac{1\ \text{mol}\cancel{\text{Al(OH)}_3}}{78.0\ \text{g}\cancel{\text{Al(OH)}_3}} \times \frac{3\ \text{mol HCl}}{1\ \text{mol}\cancel{\text{Al(OH)}_3}} = 0.00212\ \text{mol HCl}$$

$$\text{liters} = \frac{\text{moles}}{\text{molarity}} = \frac{0.00212\ \text{mol HCl}}{0.200M\ \text{HCl}}$$

$$= 0.0106\ \text{L HCl} = 10.6\ \text{mL HCl}$$

6. $4.20\ \cancel{\text{L}} \times \frac{1.20\ \text{equiv}}{1\ \cancel{\text{L}}} = 5.04\ \text{equiv KOH}$

7. H_3PO_4 contains 3 equiv/mol;
normality = $1.2N\ H_3PO_4$

8. $\frac{24.6\ \cancel{\text{g}}\ \text{HCl}}{36.5\ \cancel{\text{g}}/\text{equiv}} = 0.674\ \text{equiv HCl}$

$$\text{Normality} = \frac{0.674\ \text{equiv}}{1.50\ \text{L}} = 0.449N\ \text{HCl}$$

9. $V_A = \frac{N_B \times V_B}{N_A} = \frac{0.110\ \cancel{N} \times 68.0\ \text{mL}}{0.250\cancel{N}}$

$= 29.9\ \text{mL}\ H_2SO_4$

Section 21.2

1. $CHO_2^- + H^+ \rightleftharpoons HCHO_2$
$HCHO_2 + OH^- \rightleftharpoons CHO_2^- + H_2O$

2. **a.** $Ca(OH)_2(s) \rightleftharpoons Ca^{2+}(aq) + 2OH^-(aq)$
$K_{sp} = [Ca^{2+}] \times [OH^-]^2$

b. $Ag_2CO_3(s) \rightleftharpoons 2Ag^+(aq) + CO_3^{2-}(aq)$
$K_{sp} = [Ag^+]^2 \times [CO_3^{2-}]$

3. $Ag_2CO_3(s) \rightleftharpoons 2Ag^+(aq) + CO_3^{2-}(aq)$
$K_{sp} = [Ag^+]^2 \times [CO_3^{2-}]$
$K_{sp} = (2x)^2(x) = 8.1 \times 10^{-12}$
$4x^3 = 8.1 \times 10^{-12}$
$x^3 = 2.0 \times 10^{-12}$
$x = 1.3 \times 10^{-4}M = [CO_3^{2-}]$
$[Ag^+] = 2x = 2.6 \times 10^{-4}M$

4. $Fe(OH)_2(s) \rightleftharpoons Fe^{2+}(aq) + 2OH^-(aq)$
$K_{sp} = [Fe^{2+}] \times [OH^-]^2$
$[Fe^{2+}] = 0.5[OH^-] = 6.0 \times 10^{-6}$
$K_{sp} = (6.0 \times 10^{-6})(1.2 \times 10^{-5})^2$
$K_{sp} = 8.6 \times 10^{-16}$

5. $SrCO_3 \rightleftharpoons Sr^{2+} + CO_3^{2-}$
$K_{sp} = [Sr^{2+}] \times [CO_3^{2-}]$
Let $x = [Sr^{2+}] = [CO_3^{2-}]$
$K_{sp} = x^2 = 9.3 \times 10^{-10}$
$x = [Sr^{2+}] = 3.0 \times 10^{-5}M$

6. $K_{sp} = [Ag^+]^2 \times [CO_3^{2-}]$
Let $[Ag^+] = 2x$ and $[CO_3^{2-}] = x$
assume $x << 0.20$ mol
$K_{sp} = (2x)^2(0.20) = 8.1 \times 10^{-12}$
$x^2 = 1.0 \times 10^{-11}$
$x = 3.2 \times 10^{-6}M = [CO_3^{2-}]$
$[Ag^+] = 6.4 \times 10^{-6}M$

7. $K_{sp}(PbSO_4) = 6.3 \times 10^{-7} = [Pb^{2+}] \times [SO_4^{2-}]$
The total volume of solution is 1000 mL, so
$[Pb^{2+}] = 0.0020\ \text{mol/L} = 0.0020M$
$[SO_4^{2-}] = 0.0012\ \text{mol/L} = 0.0012M$
$[Pb^{2+}] \times [SO_4^{2-}] = (0.0020)(0.0012)$
$= 2.4 \times 10^{-6}$
Because this product exceeds the K_{sp} value, precipitation will occur.

8. $K_{sp}(CaCO_3) = 4.5 \times 10^{-9} = [Ca^{2+}] \times [CO_3^{2-}]$
The total volume is 1000 mL, so
$[Ca^{2+}] = 0.0021\ \text{mol/L} = 0.0021M$
$[CO_3^{2-}] = 0.0013\ \text{mol/L} = 0.0013M$
$[Ca^{2+}] \times [CO_3^{2-}] = (0.0021)(0.0013)$
$= 2.7 \times 10^{-6}$
Because this product exceeds the K_{sp} value, precipitation will occur.

9. NaCl has no ion in common with $Mg(OH)_2$.

Interpreting Graphics 21

1. $C_6H_5COOH + NaOH \rightarrow C_6H_5COONa + H_2O$
One mole of sodium hydroxide will neutralize one mole of benzoic acid.

2. To determine the equivalence point, find the area of the titration curve where the pH changes abruptly when a small volume of NaOH is added. Locate the point on this steep portion of the curve equidistant between the two plateaus. The pH at the equivalence point is approximately 8.5; the solution is slightly basic.

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3. Benzoic acid is a weak acid. The neutralization of a weak acid with a strong base, such as NaOH, produces a basic solution at the equivalence point.

4. $0.025 \cancel{\text{L}} \times 0.10 \text{ mol}/\cancel{\text{L}}\ \text{NaOH}$
$= 0.0025 \text{ mol NaOH}$

5. The equivalence point occurs when the number of moles of NaOH added equals the number of moles of C_6H_5COOH originally present. Because NaOH is a strong base, each mole of NaOH added reacts with each mole of C_6H_5COOH present. Thus, at the equivalence point,
$[C_6H_5COOH] = [NaOH] = 0M$ and
$$[C_6H_5COONa] = \frac{0.0025 \text{ mol}}{0.050 \text{ L}} = 0.050M.$$

6. Based on the answers to questions 4 and 5, 0.0025 mol C_6H_5COOH were originally present in a volume of 25 mL.
$$\text{Thus, } [C_6H_5COOH] = \frac{0.0025 \text{ mol}}{0.025 \text{ L}} = 0.10M$$

7. Because the equivalence point occurs between pH 6 and pH 11, phenolphthalein would be a good choice. A faint pink color should be detected at the equivalence point. Thymol blue might also be a good candidate. Students should draw a horizontal band on the graph encompassing the pH range 8-10 to show the region of the curve where phenolphthalein would be an effective indicator of neutralization.

8. $C_6H_5COO^- + H_2O \rightleftharpoons C_6H_5COOH + OH^-$
At the equivalence point, the benzoate ion establishes the equilibrium shown. The resulting solution is slightly basic because $[OH^-] > [H^+]$.

9. $$K_b = \frac{[C_6H_5COOH][OH^-]}{[C_6H_5COO^-]}$$

10. $pH + pOH = 14$
$8.5 + pOH = 14$
$pOH = 5.5$
$[OH^-] = 3.2 \times 10^{-6}M$
At the equivalence point $[OH^-] = [C_6H_5COOH] = 3.2 \times 10^{-6}M$
$[C_6H_5COO^-] = 0.050M$
$$K_b = \frac{(3.2 \times 10^{-6})^2}{(0.050)} = 2.0 \times 10^{-10}$$

Vocabulary Review 21

1. hydrolyzing salts
2. solubility product constant (K_{sp})
3. neutral
4. equivalence point
5. normality (N)
6. neutral salts
7. buffer
8. one equivalent

Quiz for Chapter 21

1. d	4. a	7. b
2. a	5. c	8. c
3. c	6. b	

Chapter 21 Test A

A. Matching

1. d	4. f	7. h
2. c	5. a	8. g
3. b	6. e	

B. Multiple Choice

9. d	12. a	15. c
10. c	13. a	16. b
11. b	14. d	

C. True-False

17. ST	19. ST	21. ST
18. NT	20. AT	

D. Problems

22. $\text{gfm } H_2SO_3 = 82.1 \text{ g}$
$$20.5 \cancel{\text{g } H_2SO_3} \times \frac{1.00 \cancel{\text{mol } H_2SO_3}}{82.1 \cancel{\text{g } H_2SO_3}} \times \frac{2 \text{ equiv } H_2SO_3}{1 \cancel{\text{mol } H_2SO_3}} = 0.499 \text{ equiv } H_2SO_3$$

23. **a.** $2H_3PO_4 + 3Ca(OH)_2 \rightarrow Ca_3(PO_4)_2 + 6H_2O$
b. $2HBr + Mg(OH)_2 \rightarrow MgBr_2 + 2H_2O$
c. $3H_2SO_4 + 2Al(OH)_3 \rightarrow Al_2(SO_4)_3 + 6H_2O$
d. $HCl + NH_4OH \rightarrow NH_4Cl + H_2O$

24. $V_1 \times N_1 = V_2 \times N_2$
$50.0 \text{ mL} \times N_1 = 20.0 \text{ mL} \times 1.50N$
$$N_1 = \frac{20.0 \cancel{\text{mL}} \times 1.50N}{50.0 \cancel{\text{mL}}}$$
$N_1 = 0.600N$

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25. $V_1 \times N_1 = V_2 \times N_2$

$V_1 \times 0.600N = 90.0\ \text{mL} \times 0.40N$

$$V_1 = \frac{90\ \text{mL} \times 0.40\cancel{N}}{0.600\cancel{N}}$$

$V_1 = 60.0\ \text{mL}$

26. $V_1 \times N_1 = V_2 \times N_2$

$V_1 \times (5.0N) = 400\ \text{mL} \times 0.20N$

$$V_1 = \frac{400\ \text{mL} \times 0.20\cancel{N}}{5.0\cancel{N}}$$

$V_1 = 16\ \text{mL}$

27. **a.** basic **c.** acidic

b. neutral **d.** basic

E. Essay

28. A measured amount of the base is placed in a beaker. An acid–base indicator is added. A burett is used to add acid of a known concentration (the standard solution). Acid is added until the indicator shows that neutralization has occurred. This is the end point of the titration. The amount of acid added is recorded. Based on the concentration and amount of acid added, the number of moles of acid added are calculated. The equation for the neutralization reaction is used to determine the number of moles of base needed to neutralize this amount of acid. The concentration of the base can be calculated from the number of moles of base required and the volume. The calculation can also be carried out in terms of normality.

F. Additional Questions

29. **a.** greater than 7 **d.** 7

b. less than 7 **e.** greater than 7

c. 7

30. **a.** $NH_3 + H^+ \rightarrow NH_4^+$

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

31. $0.96\ \cancel{\text{g } (NH_4)_2CO_3} \times \dfrac{1\ \text{mol } (NH_4)_2CO_3}{96\ \cancel{\text{g } (NH_4)_2CO_3}}$

$= 0.010\ \text{mol } (NH_4)_2CO_3$

$$\frac{0.010\ \text{mol}}{20.0\ \text{L}} = 5.0 \times 10^{-4} M\ (NH_4)_2CO_3$$

$= 5.0 \times 10^{-4} M\ CO_3^{2-}$

$$0.20\ \cancel{\text{g } CaBr_2} \times \frac{1\ \text{mol } CaBr_2}{200\ \cancel{\text{g } CaBr_2}}$$

$= 1.0 \times 10^{-3}\ \text{mol } CaBr_2$

$$\frac{1.0 \times 10^{-3}\ \text{mol}}{20.0\ \text{L}} = 5.0 \times 10^{-5} M\ CaBr_2$$

$= 5.0 \times 10^{-5} M\ Ca^{2+}$

$[Ca^{2+}] \times [CO_3^{2-}] = 5.0 \times 10^{-5} \times 5.0 \times 10^{-4}$

$= 2.5 \times 10^{-8}$

Yes, a precipitate will form because the ion product is larger than the K_{sp}.

32. $CaSO_4(s) \rightleftharpoons Ca^{2+}(aq) + SO_4^{2-}(aq)$

gfm $CaSO_4 = 136.1$ g/mol

$$\frac{0.67\ \cancel{\text{g}}}{1\ \text{L}} \times \frac{1\ \text{mol}}{136.1\ \cancel{\text{g}}} = 4.9 \times 10^{-3}\ \text{mol/L}$$

$K_{sp} = [Ca^{2+}] \times [SO_4^{2-}]$

$= (4.9 \times 10^{-3}) \times (4.9 \times 10^{-3})$

$= 2.4 \times 10^{-5}$

Chapter 21 Test B

A. Matching

1. h	**5.** i	**9.** e
2. a	**6.** j	**10.** d
3. g	**7.** c	
4. f	**8.** b	

B. Multiple Choice

11. d	**16.** c	**21.** c
12. d	**17.** b	**22.** c
13. b	**18.** a	**23.** d
14. a	**19.** b	**24.** d
15. a	**20.** d	

C. True-False

25. ST	**27.** ST	**29.** NT
26. AT	**28.** AT	

D. Problems

30. **a.** $HF(aq) + KOH(aq) \rightarrow KF(aq) + H_2O(l)$

b. $2HNO_3(aq) + Ba(OH)_2(aq) \rightarrow Ba(NO_3)_2(aq) + 2H_2O(l)$

c. $H_2SO_4(aq) + 2LiOH(aq) \rightarrow Li_2SO_4(aq) + 2H_2O(l)$

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

31. $H_2SO_4(aq) + 2KOH(aq) \rightarrow K_2SO_4(aq) + 2H_2O(l)$

1 mol 2 mol 1 mol 2 mol

$$\frac{1\ \text{mol } H_2SO_4}{2\ \cancel{\text{mol KOH}}} \times 0.35\ \cancel{\text{mol KOH}}$$

$= 0.18\ \text{mol } H_2SO_4$

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32. $2HCl(aq) + Ba(OH)_2(aq) \rightarrow BaCl_2(aq) + 2H_2O(l)$

2 mol 1 mol 1 mol 2 mol

$$\frac{2 \text{ mol HCl}}{1 \cancel{\text{mol Ba(OH)}_2}} \times \frac{(0.0154 \cancel{\text{mol Ba(OH)}_2}}{1 \cancel{\text{L}}} \times \frac{(1 \cancel{\text{L}})}{1000 \cancel{\text{mL}}} \times (27.4 \cancel{\text{mL}})$$

$$= 8.44 \times 10^{-4} \text{ mol HCl}$$

$$\text{Molarity} = \frac{\text{mol HCl}}{\text{L}} = \frac{8.44 \times 10^{-4} \text{ mol}}{0.020 \text{ L}}$$

$$= 0.042M$$

33. $N_1V_1 = N_2V_2$

$$V_1 = \frac{N_2V_2}{N_1}$$

$$V_1 = \frac{(0.350)(50.0 \text{ mL})}{(0.750)}$$

$$V_1 = 23.3 \text{ mL}$$

34. $CuCl(s) \rightarrow Cu^+(aq) + Cl^-(aq)$

$K_{sp} = [Cu^+][Cl^-]$

$3.2 \times 10^{-7} = [Cu^+][Cu^+]$

$3.2 \times 10^{-7} = [Cu^+]^2$

$5.7 \times 10^{-4}M = [Cu^+]$

E. Essay

35. A $0.25M$ H_3PO_4 solution contains 0.25 moles of solute per liter of total solution. Since H_3PO_4 is triprotic, each mole of the acid contains 3 moles of hydrogen ions, or 3 equivalents. The normality of this solution is thus $0.75N$. This means that the $0.75N$ solution contains 0.75 equivalents per liter of solution. For H_3PO_4, $N = 3 \times M$.

F. Additional Questions and Problems

36. $$15.0 \cancel{\text{g } H_3PO_4} \times \frac{1 \cancel{\text{mol } H_3PO_4}}{98.0 \cancel{\text{g } H_3PO_4}} \times 3 \text{ equiv}/1 \cancel{\text{mol}} = 0.459 \text{ equiv } H_3PO_4$$

37. $N_1V_1 = N_2V_2$

$$N_1 = \frac{N_2V_2}{V_1}$$

$$N_1 = \frac{(4.0)(150.0 \cancel{\text{mL}})}{(250.0 \cancel{\text{mL}})}$$

$$N_1 = 2.4N$$

38. $CaSO_4(s) \rightleftharpoons Ca^{2+}(aq) + SO_4^{2-}(aq)$

$K_{sp} = [Ca^{2+}][SO_4^{2-}]$

$6.10 \times 10^{-5} = [0.450][SO_4^{2-}]$

$1.36 \times 10^{-4}M = [SO_4^{2-}]$

39. $[Ag^+] = 5.0 \times 10^{-8}M$ and $[Cl^-] = 1.0 \times 10^{-9}M$.

$K_{sp} = [Ag^+][Cl^-]$

$K_{sp} = [5.0 \times 10^{-8}][1.0 \times 10^{-9}]$

$K_{sp} = [5.0 \times 10^{-17}]$

Since 5.0×10^{-17} is less than the K_{sp} of 1.8×10^{-10}, no precipitate will be produced.

40. $[Pb^{2+}] = 0.200M$ and $[SO_4^{2-}] = 0.100M$.

$K_{sp} = [Pb^{2+}][SO_4^{2-}]$

$K_{sp} = [0.200][0.100]$

$K_{sp} = 2.00 \times 10^{-2}$

Since 2.0×10^{-2} is greater than the K_{sp} of 1.8×10^{-8}, a precipitate will be produced.

Section Review 22.1

Part A Completion

1. redox
2. away
3. toward
4. reduction
5. oxidizing
6. reduced
7. reducing
8. oxidized

Part B True-False

9. AT
10. AT
11. NT
12. NT

Part C Matching

13. e
14. b
15. f
16. c
17. a
18. d

Part D Questions and Problems

19. Oxidation is the complete or partial loss of electrons. Reduction is the complete or partial gain of electrons.

20. The zinc metal, Zn, was oxidized and is the reducing agent. The copper ion, Cu^{2+}, was reduced and is the oxidizing agent.

21. When oxygen and water attack iron, the iron atoms lose electrons as the iron begins to be oxidized. Since aluminum and zinc are better reducing agents than iron and are more easily oxidized, they immediately transfer electrons to the iron ions, reducing them back to neutral iron atoms.

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Section Review 22.2

Part A Completion

1. zero
2. sign
3. charge
4. zero
5. charge on the ion
6. electron
7. oxidation
8. decrease

Part B True-False

9. AT
10. AT
11. NT
12. NT
13. NT
14. NT
15. AT
16. AT

Part C Matching

17. e
18. h
19. d
20. a
21. g
22. c
23. f
24. i
25. b

Part D Questions and Problems

26. An increase in the oxidation number of an atom indicates oxidation. A decrease in the oxidation number indicates reduction.

27. N is reduced (+5 to +2);
Br is oxidized (−1 to 0)

Mn is reduced (+7 to +2);
Cl is oxidized (−1 to 0)

N is reduced (+5 to +2);
Sb is oxidized (0 to +5)

S is reduced (+6 to +4);
C is oxidized (0 to +4)

Section Review 22.3

Part A Completion

1. oxidation number
2. half-reaction
3. balanced
4. ionic
5. two
6. added
7. ionic

Part B True-False

8. AT
9. AT
10. AT
11. NT
12. NT
13. NT

Part C Matching

14. c
15. a
16. f
17. b
18. e
19. d
20. g

Part D Questions and Problems

21. **a.** $HNO_3 \rightarrow NO$; N changes +5 to +2, a gain of $3e^-$; multiply by 2

$2HI \rightarrow I_2$; I changes from −1 to 0, a loss of $2e^-$ for I_2; multiply by 3

$2HNO_3 + 6HI \rightarrow 2NO + 3I_2 + 4H_2O$

b. $HNO_3 \rightarrow NO_2$; N changes +5 to +4, a gain of $1e^-$; multiply by 10

$I_2 \rightarrow 2HIO_3$; I changes from 0 to +5, a loss of $10e^-$ for I_2

$10HNO_3 + I_2 \rightarrow 2HIO_3 + 10NO_2 + 4H_2O$

22. **a.** $S^{2-} \rightarrow S + 2e^-$ and
$3e^- + 4H^+ + NO_3^- \rightarrow NO + 2H_2O$;
Multiply the oxidation reaction by 3 and the reduction reaction by 2.

$3H_2S + 2HNO_3 \rightarrow 3S + 2NO + 4H_2O$

b. $Fe^{2+} \rightarrow Fe^{3+} + e^-$ and $6e^- + 14H^+ + Cr_2O_7^{2-} \rightarrow 2Cr^{3+} + 7H_2O$; multiply the oxidation reaction by 6

$14H^+ + 6Fe^{2+} + Cr_2O_7^{2-} \rightarrow 6Fe^{2+} + 2Cr^{3+} + 7H_2O$

Practice Problems

Section 22.1

1. Sr: oxidized (reducing agent)
   O_2: reduced (oxidizing agent)
2. Li: oxidized (reducing agent)
   S: reduced (oxidizing agent)
3. Cs: oxidized (reducing agent)
   Br_2: reduced (oxidizing agent)
4. Mg: oxidized (reducing agent)
   N_2: reduced (oxidizing agent)
5. Fe: oxidized (reducing agent)
   O_2: reduced (oxidizing agent)
6. Br^-: oxidized (reducing agent)
   Cl_2: reduced (oxidizing agent)
7. Si: oxidized (reducing agent)
   F_2 reduced (oxidizing agent)
8. Ca: oxidized (reducing agent)
   O_2 reduced (oxidizing agent)
9. Mg: oxidized (reducing agent)
   H^+: reduced (oxidizing agent)
10. Na: oxidized (reducing agent)
    H_2O: reduced (oxidizing agent)

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Section 22.2

1. **a.** Sn is tin in an uncombined state. The oxidation number is 0.
   b. The ionic charge on potassium is 1+, thus the oxidation number is +1.
   c. The ionic charge on sulfur is 2−, thus the oxidation number is −2.
   d. The ionic charge on iron is 3+, thus the oxidation number is +3.
   e. Se is selenium in an uncombined state. The oxidation number is 0
   f. The ionic charge on magnesium is 2+, thus the oxidation number is +2.
   g. The ionic charge on tin is 4+, thus the oxidation number is +4.
   h. The ionic charge on bromine is 1−, thus the oxidation number is −1.
2. **a.** +3 **c.** +2
   b. +6 **d.** +6
3. **a.** $\overset{0}{C} + \overset{+1\ +6-2}{H_2SO_4} \rightarrow \overset{+4-2}{CO_2} + \overset{+4-2}{SO_2} + \overset{+1\ -2}{H_2O}$
   Carbon is oxidized ($0 \rightarrow +4$).
   Sulfur is reduced ($+6 \rightarrow +4$).
   b. $\overset{+1+5-2}{HNO_3} + \overset{+1-1}{HI} \rightarrow \overset{+2-2}{NO} + \overset{0}{I_2} + \overset{+1\ -2}{H_2O}$
   Nitrogen is reduced ($+5 \rightarrow +2$).
   Iodide ion is oxidized ($-1 \rightarrow 0$).
   c. $\overset{+1\ +7\ -2}{KMnO_4} + \overset{+1-1}{HCl} \rightarrow \overset{+2\ -1}{MnCl_2} + \overset{0}{Cl_2} + \overset{+1\ -2}{H_2O} + \overset{+1-1}{KCl}$
   Manganese is reduced ($+7 \rightarrow +2$).
   Chloride ion is oxidized ($-1 \rightarrow 0$).
   d. $\overset{0}{Sb} + \overset{+1+5-2}{HNO_3} \rightarrow \overset{+5\ -2}{Sb_2O_5} + \overset{+2-2}{NO} + \overset{+1\ -2}{H_2O}$
   Antimony is oxidized ($0 \rightarrow +5$).
   Nitrogen is reduced ($+5 \rightarrow +2$).
4. **a.** Oxidizing agent is sulfur; Reducing agent is carbon.
   b. Oxidizing agent is nitrogen; Reducing agent is iodine.
   c. Oxidizing agent is manganese; Reducing agent is chlorine.
   d. Oxidizing agent is nitrogen; Reducing agent is antimony.

Section 22.3

1. **a.** Increase in oxidation number of carbon = +4; decrease in oxidation number of sulfur = −2.
   $C + 2H_2SO_4 \rightarrow CO_2 + 2SO_2 + 2H_2O$
   b. Increase in oxidation number of sulfur = +2; decrease in oxidation number of nitrogen = −3.
   $3H_2S + 2HNO_3 \rightarrow 3S + 2NO + 4H_2O$
   c. Increase in oxidation number of iodine ion = +1; decrease in oxidation number of nitrogen = −3.
   $2HNO_3 + 6HI \rightarrow 2NO + 3I_2 + 4H_2O$
   d. Increase in oxidation number of antimony = +5; decrease in oxidation number of nitrogen = −3.
   $6Sb + 10HNO_3 \rightarrow 3Sb_2O_5 + 10NO + 5H_2O$
   e. Increase in oxidation number of chlorine ion = +1; decrease in oxidation number of manganese = −5.
   $2KMnO_4 + 16HCl \rightarrow 2MnCl_2 + 5Cl_2 + 8H_2O + 2KCl$
   f. Increase in oxidation number of iodine ion = +1; decrease in oxidation number of iodine = −7.
   $KIO_4 + 7KI + 8HCl \rightarrow 8KCl + 4I_2 + 4H_2O$
   g. Increase in oxidation number of zinc = +2; decrease in oxidation number of chromium = −3.
   $3Zn + 2Cr_2O_7^{2-} + 28H^+ \rightarrow 3Zn^{2+} + 4Cr^{3+} + 14H_2O$
2. **a.** $Fe^{2+} \rightarrow Fe^{3+} + e^-$
   $5e^- + 8H^+ + MnO_4^- \rightarrow Mn^{2+} + 4H_2O$
   b. $Sn^{2+} \rightarrow Sn^{4+} + 2e^-$
   $6H^+ + 6e^- + IO_3^- \rightarrow I^- + 3H_2O$
   c. $S^{2-} \rightarrow S + 2e^-$
   $3e^- + 4H^+ + NO_3^- \rightarrow NO + 2H_2O$
   d. $4OH^- + Mn^{2+} \rightarrow MnO_2 + 2H_2O + 2e^-$
   $2e^- + H_2O + H_2O_2 \rightarrow H_2O + 2OH^-$
3. **a.** $2OH^- + Zn + HgO \rightarrow ZnO_2^{2-} + Hg + H_2O$
   b. $8H^+ + 5Fe^{2+} + MnO_4^- \rightarrow 5Fe^{3+} + Mn^{2+} + 4H_2O$
   c. $6H^+ + 3Sn^{2+} + IO_3^- \rightarrow 3Sn^{4-} + I^- + 3H_2O$
   d. $8H^+ + 3S^{2-} + 2NO_3^- \rightarrow 3S + 2NO + 4H_2O$
   e. $2OH^- + Mn^{2+} + H_2O_2 \rightarrow MnO_2 + 2H_2O$
   f. $2OH^- + CrO_2 + ClO^- \rightarrow CrO_4^{2-} + Cl^- + H_2O$

Interpreting Graphics 22

1. **a.** 3 **d.** 2
   b. 2 **e.** 4
   c. 3 **f.** 1

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2. $MnO_4^- + 8H^+ + 5Fe^{2+} \rightarrow Mn^{2+} + 4H_2O + 5Fe^{3+}$

3. The end point occurs when the number of equivalents of MnO_4^- added equals the number of equivalents of Fe^{2+} originally present in the reaction flask. One equivalent is the amount of reducing agent (or oxidizing agent) that can give (or accept) one mole of electrons. When all the Fe^{2+} in the flask is oxidized, the next drop of MnO_4^- remains unreacted, and the solution in the flask turns light purple, signaling the end point of the titration.

4. Volume $KMnO_4$ = Initial Volume − Final Volume = 48.65 mL − 23.35 mL = 25.30 mL

$$\text{Moles } MnO_4^- = 0.02530 \cancel{\text{L}} \times \frac{0.0200 \text{ mol}}{\cancel{\text{L}}}$$

$$= 5.06 \times 10^{-4} \text{ mol } MnO_4^-$$

$$\text{Moles iron(II)} = 5.06 \times 10^{-4} \cancel{\text{mol } MnO_4^-} \times \frac{5 \text{ mol } Fe^{2+}}{1 \cancel{\text{mol } MnO_4^-}} = 2.53 \times 10^{-3} \text{ mol } Fe^{2+}$$

$$\text{Mass Fe} = 2.53 \times 10^{-3} \cancel{\text{mol } Fe^{2+}} \times \frac{55.85 \text{ g } Fe^{2+}}{1 \cancel{\text{mol } Fe^{2+}}} = 0.141 \text{ g}$$

$$\% \text{ Fe in ore} = \frac{0.141 \text{ g}}{2.938 \text{ g}} \times 100\% = 4.80\%$$

Vocabulary Review 22

1. oxidizing agent
2. oxidation-number-change method
3. reduction
4. half-reaction method
5. oxidation number
6. reducing agent
7. half-reaction
8. oxidation-reduction reaction
9. oxidation
10. redox reaction

Quiz for Chapter 22

1. b	4. b	7. c
2. a	5. a	8. d
3. b	6. a	

Chapter 22 Test A

A. Matching

1. d	5. i	9. c
2. j	6. g	10. h
3. f	7. e	
4. a	8. b	

B. Multiple Choice

11. b	17. c	23. d
12. d	18. c	24. c
13. c	19. a	25. c
14. b	20. a	26. c
15. a	21. c	
16. b	22. b	

C. Questions

27. **a.** Na oxidized, reducing agent; Br_2 reduced, oxidizing agent
    b. S reduced, oxidizing agent; K oxidized, reducing agent
28. $2Cr + 3Br_2 \rightarrow 2Cr^{3+} + 6Br^-$
29. **a.** Li +1, Al +3, F −1
    b. Na +1, O −2
    c. S 0 (element)
30. Oxidation-number change method:

$$\overset{+3\ -2}{Fe_2O_3} + \overset{+2\ -2}{CO} \rightarrow \overset{0}{Fe} + \overset{+4\ -2}{CO_2}$$

$2 \times (-3) = -6$; $3 \times (+2) = +6$

$Fe_2O_3 + 3CO \rightarrow 2Fe + 3CO_2$

Half-reaction method:

$6H^+ + Fe_2O_3 + 6e^- \rightarrow 2Fe + 3H_2O$

$3(H_2O + CO \rightarrow CO_2 + 2H^+ + 2e^-)$

$6H^+ + Fe_2O_3 + 6e^- + 3H_2O + 3CO \rightarrow 2Fe + 3H_2O + 3CO_2 + 6H^+ + 6e^-$

$Fe_2O_3 + 3CO \rightarrow 2Fe + 3CO_2$

D. Essay

31. An oxidation number is assigned to an element in a compound according to a set of arbitrary rules. The oxidation number of an element in an uncombined state is zero. The oxidation number of a monatomic ion is the same in magnitude and sign as the ionic charge. The sum of the oxidation numbers of the elements in a neutral compound is zero. In a polyatomic ion, however, the sum is equal to the charge on the ion. Oxidation

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numbers help keep track of electrons in redox reactions. An oxidation-number increase is oxidation. A decrease is reduction.

Chapter 22 Test B

A. Matching

1. e	5. i	9. b
2. f	6. a	10. j
3. d	7. c	
4. g	8. h	

B. Multiple Choice

11. d	17. a	23. d
12. a	18. b	24. c
13. a	19. c	25. d
14. d	20. c	26. b
15. d	21. b	27. b
16. b	22. b	28. a

C. Questions

29. **a.** K; I; I_2; K
    b. Na; H; H_2O; Na
    c. H; Cu; CuO; H_2
    d. Mg; Cu; $Cu(NO_3)_2$; Mg

30. **a.** K_2SO_4 = +1, +6, −2
    b. $Cu(NO_3)_2$ = +2, +5, −2
    c. $HAsO_3$ = +1, +5, −2
    d. MnO_4^- = +7, −2

31. **a.** $4HNO_3 + 3Ag \rightarrow 3AgNO_3 + NO + 2H_2O$
    ($3 \times (+1) = +3$; $1 \times (-3) = -3$)
    b. $Br_2 + SO_2 + 2H_2O \rightarrow H_2SO_4 + 2HBr$
    ($1 \times (+2) = +2$; $2 \times (-1) = -2$)

32. **a.** $HNO_2 + HI \rightarrow I_2 + NO + H_2O$
    $H^+(aq) + NO_2^-(aq) + H^+(aq) + I^-(aq) \rightarrow I_2(aq) + NO(g) + H_2O(l)$
    Oxidation: $2I^-(aq) \rightarrow I_2 + 2e^-$
    Reduction: $2[2H^+(aq) + NO_2^-(aq) + 1e^- \rightarrow NO + H_2O]$
    $4H^+ + 2NO_2^- + 2e^- \rightarrow 2NO + 2H_2O$
    $4H^+ + 2I^- + 2NO_2^- \rightarrow I_2 + 2NO + H_2O$
    Final: $2HNO_2 + 2HI \rightarrow I_2 + 2NO + 2H_2O$

    b. $K_2Cr_2O_7 + FeCl_2 + HCl \rightarrow CrCl_3 + KCl + FeCl_3 + H_2O$
    $2K^+(aq) + Cr_2O_7^{2-}(aq) + Fe^{2+}(aq) + 2Cl^- + H^+(aq) + Cl^-(aq) \rightarrow Cr^{3+}(aq) + 3Cl^-(aq) + K^+(aq) + Cl^-(aq) + Fe^{3+}(aq) + 3Cl^-(aq) + H_2O$
    Oxidation: $6[Fe^{2+} \rightarrow Fe^{2+} + 1e^-]$
    Reduction: $2Cr^{6+} + 6e^- \rightarrow 2Cr^{3+}$
    $6Fe^{2+} + 2Cr^{6+} \rightarrow 6Fe^{3+} + 2Cr^{3+}$
    Final: $K_2Cr_2O_7 + 6FeCl_2 + 14HCl \rightarrow 2CrCl_3 + 2KCl + 6FeCl_3 + 7H_2O$

D. Essay

33. Since oxidation is the loss of electrons, it can only occur in the presence of another substance that will accept the lost electrons. The accepting substance gains electrons, and thus, undergoes reduction. In other words, a loss of electrons can only occur if a gain takes place concurrently.

Section Review 23.1

Part A Completion

1. electrochemical process
2. electrons
3. voltaic cells
4. salt bridge
5. ions
6. anode
7. cathode

Part B True-False

8. NT	10. NT
9. AT	11. ST

Part C Matching

12. g	15. b	17. e
13. f	16. c	18. a
14. d		

Part D Questions and Problems

19. The shorthand notation Mg(*s*) | $MgSO_4(aq)$ || $PbSO_4(aq)$ | Pb(*s*) represents a magnesium-lead voltaic cell. The single vertical lines indicate boundaries of phases that are in contact, and the double vertical lines represent the salt bridge that separates the anode compartment from the cathode compartment. In this electrochemical cell,

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Mg is oxidized to Mg^{2+} at the anode (the negative electrode) and Pb^{2+} is reduced to Pb at the cathode (the positive electrode). Electrons flow from the anode, through an external circuit (connected to a light bulb or voltmeter), to the cathode. To complete the circuit, sulfate (SO_4^{2-}) anions move from the cathode compartment to the anode compartment, and magnesium and sodium cations move from the anode compartment to the cathode compartment. Check students' diagrams.

Section Review 23.2

Part A Completion

1. electric potential
2. electrons
3. electrons
4. reduction
5. cell potential
6. standard hydrogen electrode
7. 0.00 V
8. less
9. spontaneous

Part B True-False

10. NT
11. ST
12. NT
13. NT

Part C Matching

14. b
15. d
16. f
17. c
18. a
19. e

Part D Questions and Problems

20. Oxidation: $Mg \rightarrow Mg^{2+} + 2e^-$

Reduction: $2e^- + Cl_2 \rightarrow 2Cl^-$

Redox: $Mg + Cl_2 \rightarrow Mg^{2+} + 2Cl^-$

$E^0_{cell} = E^0_{red} - E^0_{oxid}$

$= E^0_{Cl_2} - E^0_{Mg}$

$= +1.36\,V - (-2.37\,V)$

$= +3.73\,V$

Section Review 23.3

Part A Completion

1. electrolysis
2. electrolytic cell
3. electrons
4. battery
5. electrolyte
6. hydrogen/oxygen
7. oxygen/hydrogen
8. sodium metal

Part B True-False

9. AT
10. ST
11. ST
12. AT

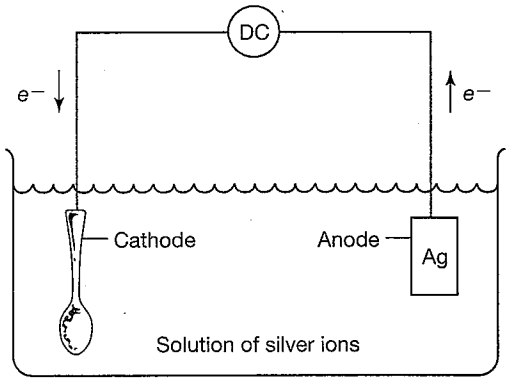
Part C Matching

13. b
14. d
15. e
16. c
17. a

Part D Questions and Problems

18. In electrolytic cells, electrical energy is used to bring about a normally nonspontaneous chemical reaction. In a voltaic cell, chemical energy is converted to electrical energy by a spontaneous redox reaction. Electrolytic cells are used in electroplating, in refining metals, and in the production of substances such as sodium hydroxide, aluminum, sodium, and chlorine. Voltaic cells are used in pacemakers, hearing aids, and cameras.

19.



Anode (oxidation): $Ag(s) \rightarrow Ag^+(aq) + e^-$

Cathode (reduction): $Ag^+(aq) + e^- \rightarrow Ag(s)$

Practice Problems

Section 23.2

1. **a.** $Cl_2(g) + Mg(s) \rightarrow 2Cl^-(aq) + Mg^{2+}(aq)$

$E^0_{cell} = 1.36\,V - (-2.37\,V) = 3.73\,V$

cathode: $Cl_2(g) + 2e^- \rightarrow 2Cl^-(aq)$

b. $2Ag^+(aq) + Ni(s) \rightarrow Ni^{2+}(aq) + 2Ag(s)$

$E^0_{cell} = 0.80\,V - (-0.25\,V) = 1.05\,V$

cathode: $Ag^+(aq) + e^- \rightarrow Ag(s)$

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c. $2MnO_4^-(aq) + 16H^+(aq) + 5Cd(s) \rightarrow 5Cd^{2+}(aq) + 2Mn^{2+}(aq) + 8H_2O(l)$

$E^0_{cell} = 1.51\,V - (-0.40\,V) = 1.91\,V$

cathode: $MnO_4^-(aq) + 8H^+(aq) + 5e^- \rightarrow Mn^{2+}(aq) + 4H_2O(l)$

d. $Br_2 + 2Na(s) \rightarrow 2Na^+(aq) + 2Br^-(aq)$

$E^0_{cell} = 1.07\,V - (-2.71\,V) = 3.78\,V$

cathode: $Br_2(l) + 2e^- \rightarrow 2Br^-(aq)$

e. $MnO_2(s) + 4H^+(aq) + H_2(g) \rightarrow 2H^+(aq) + Mn^{2+}(aq) + 2H_2O(l)$

$E^0_{cell} = 1.28\,V - 0.00\,V = 1.28\,V$

cathode: $MnO_2(s) + 4H^+(aq) + 2e^- \rightarrow Mn^{2+}(aq) + 2H_2O(l)$

2. **a.** $E^0_{cell} = -0.14\,V - (-2.90\,V) = +2.76\,V$; spontaneous

b. $E^0_{cell} = +0.80\,V - 1.36\,V = -0.56\,V$; nonspontaneous

c. $E^0_{cell} = +2.87\,V - (-0.76\,V) = +3.63\,V$; spontaneous

d. $E^0_{cell} = -0.28\,V - (-3.05\,V) = +2.77\,V$; spontaneous

e. $E^0_{cell} = -2.93\,V - 0.54\,V = -3.47\,V$; nonspontaneous

Interpreting Graphics 23

1. anode(+)
2. cathode(−)
3. electrorefining
4. **a.** The anode(+) of the electrolytic cell should be connected to the positive(+) terminal of the battery. The cathode(−) of the electrolytic cell should be connected to the negative(−) terminal of the battery.

   b. The anode of the electrolytic cell is connected to the cathode of the battery. The cathode of the electrolytic cell is connected to the anode of the battery.
5. Oxidation occurs at the anode, labeled number 1 in the diagram. Reduction occurs at the cathode, labeled number 2 in the diagram.
6. Students should indicate the flow of electrons out of the anode(+) and into the cathode(−).
7. The voltage should be great enough to oxidize copper metal at the anode and reduce copper(II) ions at the cathode, but not high enough to oxidize other metals at the anode and reduce them at the cathode. The voltage should be greater than 0.34 V but less than 0.44 V.
8. **a.** gold, silver, and platinum

   b. zinc 2+ and iron 2+

   c. copper

Vocabulary Review 23

1. voltaic cell
2. fuel cell
3. electrochemical cell
4. electrochemical process
5. cathode
6. reduction potential

Solution: aluminum

Quiz for Chapter 23

1. NT	7. NT	13. NT
2. AT	8. AT	14. AT
3. NT	9. AT	15. NT
4. AT	10. ST	16. AT
5. ST	11. NT	17. AT
6. NT	12. AT	18. NT

Chapter 23 Test A

A. Matching

1. b	5. j	9. f
2. h	6. g	10. a
3. d	7. i	
4. c	8. e	

B. Multiple Choice

11. d	16. a	21. c
12. b	17. d	22. c
13. c	18. c	23. a
14. c	19. a	24. a
15. c	20. a	25. c

C. True-False

26. NT	28. AT	30. NT
27. NT	29. NT	

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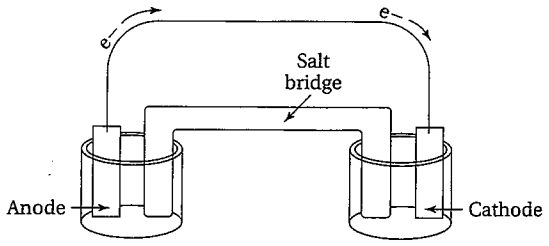
D. Question

31.



E. Essay

32. In both voltaic and electrolytic cells, oxidation occurs at the anode and reduction occurs at the cathode. In the voltaic cell, the cathode is positive and the anode is negative. In the electrolytic cell, the anode is positive and the cathode is negative.

F. Additional Questions

33. The negative value means that the tendency for zinc ions to be reduced is less than that of hydrogen ions to be reduced, so zinc metal is oxidized when paired with the standard hydrogen half-cell.

34. The reduction potential of a half-cell is a measure of the tendency of a given half-reaction to occur as a reduction.

35. **a.** $Ni^{2+}(aq) + 2e^- \rightarrow Ni(s)$ $\quad E^0 = -0.25\,V$

$Zn^{2+}(aq) + 2e^- \rightarrow Zn(s)$ $\quad E^0 = -0.76\,V$

$E^0_{cell} = E^0_{red} - E^0_{oxid}$

$= -0.76\,V - (-0.25\,V)$

$= -0.51\,V$

This reaction is not spontaneous.

b. $2[Al^{3+}(aq) + 3e^- \rightarrow Al(s)]$ $\quad E^0 = -1.66\,V$

$3[Co^{2+}(aq) + 2e^- \rightarrow Co(s)]$ $\quad E^0 = -0.28\,V$

$E^0_{cell} = E^0_{red} - E^0_{oxid}$

$= -1.66\,V - (-0.28\,V)$

$= -1.38\,V$

This reaction is not spontaneous.

Chapter 23 Test B

A. Matching

1. h	**5.** c	**9.** b
2. d	**6.** g	**10.** a
3. f	**7.** j	
4. i	**8.** e	

B. Multiple Choice

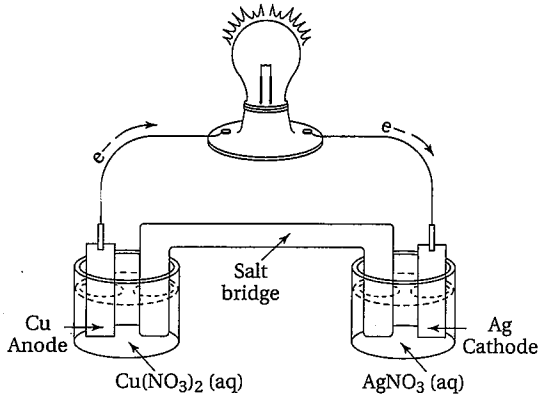
11. c	**17.** d	**23.** b
12. d	**18.** a	**24.** a
13. b	**19.** d	**25.** b
14. c	**20.** c	**26.** c
15. d	**21.** b	**27.** d
16. a	**22.** c	

C. True-False

28. NT	**30.** AT	**32.** AT
29. NT	**31.** AT	**33.** ST

D. Question

34.



35. $Cu(s) \rightarrow Cu^{2+}(aq) + 2e^-$

$2Ag^+(aq) + 2e^- \rightarrow 2Ag(s)$

Net: $Cu(s) + 2Ag^+(aq) \rightarrow Cu^{2+}(aq) + 2Ag(s)$

$E^0_{cell} = E^0_{red} - E^0_{oxid}$

$= +0.80\,V - (+0.34\,V)$

$= +0.46\,V$

E. Essay

36. In both voltaic and electrolytic cells, electrons flow from the anode to the cathode through the external circuit, reduction occurs at the cathode, and oxidation occurs at the anode. However, while the flow of electrons in a voltaic cell is caused by a spontaneous chemical reaction, in an electrolytic cell the flow of electrons is being pushed by an outside source such as a battery. Additionally, while the anode is the negative electrode and the cathode is the positive electrode in a voltaic cell, the reverse is true in an electrolytic cell—the anode in an electrolytic cell is the positive electrode and the cathode is negative.

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F. Additional Questions

37. Since Al is above Pb in the reduction potential table, Al is oxidized and Pb is reduced. Thus, the two half-cell reactions are as follows:

$2[Al(s) \rightarrow Al^{3+}(aq) + 3e^-]$

$3[Pb^{2+}(aq) + 2e^- \rightarrow Pb(s)]$

Net: $2Al(s) + 3Pb^{2+}(aq) \rightarrow 2Al^{3+}(aq) + 3Pb(s)$

$E^0_{cell} = E^0_{red} - E^0_{oxid}$

$= -0.13\,V - (-1.66\,V)$

$= +1.53\,V$

38. a. The half-reactions are:

Oxidation: $Na(s) \rightarrow Na^+(aq) + e^-$

$E^0 = -2.71\,V$

Reduction: $Cu^{2+}(aq) + 2e^- \rightarrow Cu(s)$

$E^0 = +0.34\,V$

$E^0_{cell} = E^0_{red} - E^0_{oxid}$

$E^0_{cell} = +0.34\,V - (-2.71\,V)$

$E^0_{cell} = +3.05\,V$

Since the standard cell potential is positive, the redox reaction will be spontaneous.

b. The half-reactions are:

Oxidation: $Ag(s) \rightarrow Ag^+(aq) + e^-$

$E^0 = +0.80\,V$

Reduction: $Mg^{2+}(aq) + 2e^- \rightarrow Mg(s)$

$E^0 = -2.37\,V$

$E^0_{cell} = E^0_{red} - E^0_{oxid}$

$= -2.37\,V - (+0.80\,V)$

$= -3.17\,V$

Since the standard cell potential is negative, the redox reaction will be nonspontaneous.

c. The half-reactions are:

Oxidation: $Al(s) \rightarrow Al^{3+}(aq) + 3e^-$

$E^0 = -1.66\,V$

Reduction: $Zn^{2+}(aq) + 2e^- \rightarrow Zn(s)$

$E^0 = -0.76\,V$

$E^0_{cell} = E^0_{red} - E^0_{oxid}$

$= -0.76\,V - (-1.66\,V)$

$= +0.90\,V$

Since the standard cell potential is positive, the redox reaction will be spontaneous.

39. Reaction 1: since Z is oxidized and X^{2+} is reduced, Z appears above X in the activity series.

Reaction 2: since W is not oxidized in the presence of X^{2+}, W should appear below X.

Reaction 3: since Y is oxidized and Z^{2+} is reduced, Y appears above Z.

The elements should be listed as follows: Y, Z, X, and W.

Section Review 24.1

Part A Completion

1. alkali
2. alkali salts
3. one
4. sodium hydroxide
5. hydrogen gas
6. electrical
7. sodium chloride
8. Chlorine gas
9. alkaline-earth
10. Magnesium
11. calcium
12. slaked lime

Part B True-False

13. AT
14. NT
15. AT
16. AT
17. AT

Part C Matching

18. e
19. c
20. d
21. a
22. b

Part D Questions and Problems

23. a. $2Na(s) + 2H_2O(l) \rightarrow 2NaOH(aq) + H_2(g)$

b. $Ca(s) + 2H_2O(l) \rightarrow Ca(OH)_2(aq) + H_2(g)$

24. The chemical reactivity increases as you go down each group. Of the alkali metals, lithium is the least reactive and cesium is the most reactive. Of the alkaline-earth metals, beryllium is the least active and barium is the most active.

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Section Review 24.2

Part A Completion

1. boron
2. borax
3. aluminum
4. bauxite
5. Al_2O_3 or aluminum oxide
6. diamond/graphite
7. diamond/graphite
8. nitrogen
9. N_2/a diatomic molecule
10. ammonia
11. Oxygen
12. ozone (O_3)

Part B True-False

13. AT
14. ST
15. ST
16. AT
17. NT

Part C Matching

18. c
19. f
20. e
21. a
22. b
23. d

Part D Questions and Problems

24. The halogens exist as compounds in nature. Fluorine and chlorine are greenish-yellow reactive gases, bromine is a dark-red liquid, and iodine is a purple-black solid. Most of the compounds of the halogens are soluble in water. Fluorine is the most reactive of all nonmetals and the strongest elemental oxidizing agent known. Chlorine is used to purify drinking water and swimming pools, and in sewage treatment. Chlorine is also used to make vinyl chloride, which is used to make the plastic polyvinyl chloride (PVC). Chloride ions are an important component of the blood and other body fluids. Iodide ions are necessary for thyroid function.

Section Review 24.3

Part A Completion

1. ductile/malleable
2. ductile/malleable
3. oxidation state
4. tungsten
5. mercury
6. copper
7. Zinc
8. Iron
9. steel
10. Gold
11. inner transition elements
12. lanthanide
13. actinides

Part B True-False

14. AT
15. ST
16. ST
17. AT

Part C Matching

18. e
19. b
20. d
21. a
22. c

Part D Questions and Problems

23. There are two types of steels: carbon steels and alloy steels. Both types actually contain carbon, but carbon steels contain no other metal other than iron. Carbon steel, with less than 0.2% carbon, is called mild steel and is used when load-bearing ability is not important. Medium steels, which contain from 0.2% to 0.6% carbon, are used for structural materials. High carbon steels contain between 0.8% and 1.5% carbon and are used to make items for which hardness is important. Alloy steels are called stainless steels and contain high percentages of chromium and nickel. The most common stainless steel contains 18% chromium and 8% nickel.

24. $\frac{18}{24} \times 100\% = 75\%$ gold

Section Review 24.4

Part A Completion

1. hydrogen
2. alkali metal
3. halogen
4. hydrogen bromide (HBr)
5. sodium hydride (NaH)
6. Water
7. ammonia
8. rocket fuel
9. noble gases

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10. separate
11. xenon tetrafluoride (XeF_4)

Part B True-False

12. AT
13. NT
14. AT
15. AT

Part C Matching

16. d
17. a
18. c
19. e
20. f
21. b

Part D Questions and Problems

22. **a.** $H_2(g) + Cl_2(g) \rightarrow 2HCl(g)$
    b. $H_2(g) + 2Na(s) \rightarrow 2NaH(s)$
23. Hydrogen, like the halogens, has one fewer valence electron than the noble gas it precedes. Like the halogens, it reacts with alkali metals to form compounds in which it has an oxidation state of −1. It is a gas and the gases appear in the upper right-hand section of the periodic table. These properties make it reasonable to place hydrogen in Group 7A. Hydrogen is more often placed above Group 1A because, like an alkali metal, it reacts with halogens to form compounds in which it has an oxidation state of +1.

Interpreting Graphics 24

1. Check students' tables.
2. Check students' tables.
3. Check students' tables. Most of the 92 naturally occurring elements are solids at room temperature.
4. Cesium is the most reactive alkali metal. Cesium is expected to be much less electronegative than fluorine. As a group, the alkali metals are the least electronegative elements in the periodic table. Because electronegativity decreases as you move down a group, cesium, and francium, a rare and radioactive element, are expected to be the least electronegative elements of all. Fluorine, on the other hand, is the most electronegative element in the periodic table.
5. Beryllium is chemically the most stable element in Group 2A.
6. All of the noble gases are essentially inert (except in rare instances).
7. Group 7A elements are called halogens; they behave as oxidizing agents in most chemical reactions. Fluorine is the strongest elemental oxidizing agent known.
8. The metalloids are solids at room temperature; they conduct electricity at high temperatures but act as electrical insulators at low temperatures. The *p*-block consists of Groups 3A, 4A, 5A, 6A, 7A, and 0 with the exception of helium. The *p*-block elements have electrons in their outermost *p* sublevel. In general, the valence electron configuration for these elements could be written as ns^2np^x, where $1 \leq x \geq 6$.
9. Mercury is a liquid at room temperature. Platinum is often used to catalyze chemical reactions.
10. Check students' tables and keys.

Vocabulary Review 24

1. b
2. a
3. d
4. e
5. f
6. c
7. i
8. j
9. g
10. h

Quiz for Chapter 24

1. c
2. b
3. a
4. b
5. ST
6. NT
7. NT
8. ST
9. AT
10. silicon
11. phosphorus
12. sulfur
13. fluorine

Chapter 24 Test A

A. Matching

1. d
2. i
3. f
4. h
5. c
6. a
7. g
8. j
9. e
10. b

B. Multiple Choice

11. a
12. a
13. b
14. c
15. b
16. c
17. c
18. c
19. a
20. d
21. b
22. b
23. c
24. d
25. b

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C. True-False

26. AT
27. AT
28. AT
29. AT
30. NT
31. NT
32. NT
33. AT
34. AT
35. NT
36. AT
37. AT

D. Questions

38. In general, the more reactive an element, the more likely it is to occur in nature in the combined, rather than in the free, state. Thus, while the alkali metals always occur in the combined state, the noble gases are almost always found in the free state.

39. The Group 0 elements are named the noble gases because, like nobility, they do not mix with others—in this case, other atoms. Noble gases are unreactive because they have full outer shell electron configurations and, thus, have no need to lose, gain, or share electrons.

E. Essay

40. The alkali metals are more reactive than the alkaline-earth metals; neither group occurs free in nature, the alkali metals must be stored under oil or kerosene to prevent their reacting with water; no such precaution is needed for the alkaline-earth metals. The alkali metals are more soluble in water than the alkaline-earth metals.

Chapter 24 Test B

A. Matching

1. f
2. j
3. g
4. d
5. i
6. a
7. h
8. c
9. e
10. b

B. Multiple Choice

11. d
12. b
13. c
14. c
15. b
16. d
17. c
18. b
19. c
20. d
21. a
22. d
23. c
24. c
25. b

C. True-False

26. NT
27. AT
28. AT
29. AT
30. AT
31. AT
32. AT
33. ST
34. NT
35. NT
36. AT
37. ST

D. Questions

38. All three elements have low chemical reactivity and, thus, generally occur free in nature. Since all three are shiny in the uncombined state, humans were probably drawn first to their appearance, then to their malleability.

39. Because alkali metals all contain a valence electron, they tend to lose that electron to achieve a stable noble gas configuration. The low first ionization energies of alkali metals result in frequent chemical reactions with other elements.

E. Essay

40. The halogens are very reactive, while the noble gases seldom react; the halogens are found combined in nature, but the noble gases are usually free. The iodide ion is essential for a healthy thyroid, the fluoride ion is needed for healthy teeth, and chlorine is used as a disinfectant in swimming pools. The noble gases are used in weather balloons (He), to fill deep-sea diving tanks (He and Ne, with O_2), and as the inert atmosphere in flash bulbs (Ar, Kr, and Xe).

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