Solved Problems

NTRODUCTION

12.1. What is the difference between gas and gasoline?

12.2. What is a fluid?

PRESSURE OF GASES

12.3. Change 806 mmHg to (*a*) torr, (*b*) atmospheres, and (*c*) kilopascals.

12.4. Change (*a*) 703 torr to atmospheres, (*b*) 1.25 atm to torr, (*c*) 743 mmHg to torr, and (*d*) 1.01 atm to millimeters of mercury (mmHg).

12.5. How many pounds force does 1 atm pressure exert on the side of a metal can that measures 6.0 in. by 9.0 in.? (1 atm = 14.7 $lb/in.^2$)

BOYLE'S LAW

12.6. What law may be stated qualitatively as "When you squeeze a gas, it gets smaller"?

12.7. Calculate the product of the pressure and volume for each point in Table 12-1. What can you conclude?

12.8. If 4.00 L of gas at 1.22 atm is changed to 876 torr at constant temperature, what is its final volume?

12.9. If 1.25 L of gas at 780 torr is changed to 965 mL at constant temperature, what is its final pressure?

12.10. A 1.00-L sample of gas at 25°C and 1.00-atm pressure is changed to 3.00-atm pressure at 25°C. What law may be used to determine its final volume?

12.11. A sample of gas occupies 2.48 L. What will be its new volume if its pressure is doubled at constant temperature?

GRAPHICAL REPRESENTATION OF DATA

12.12. On a graph, what criteria represent direct proportionality?

12.13. What is the pressure of the gas described in Table 12-1 at 10.0 L? Answer first by calculating with Boyle's law, second by reading from Fig. 12-2, and third by reading from Fig. 12-3. Which determination is easiest (assuming that the graphs have already been drawn)?

12.14. Plot the following data:

V (L)	*P* (atm)
1.50	4.00
3.00	2.00
6.00	1.00
12.00	0.500

Replot the data, using the volume and the reciprocal of the pressure. Do these values fall on a straight line? Are volume and the reciprocal of pressure directly proportional? Are volume and pressure directly proportional?

CHARLES' LAW

12.15. Plot the following data:

V (L)	*t* (°C)
4.92	100
4.26	50
3.60	0
2.94	−50

Do the values fall on a straight line? Are volume and Celsius temperature directly proportional? Replot the volume versus the Kelvin temperature. Are volume and Kelvin temperature directly proportional?

12.16. If 42.3 mL of gas at 22°C is changed to 44°C at constant pressure, what is its final volume?

12.17. If 0.979 L of gas at 0°C is changed to 737 mL at constant pressure, what is its final temperature?

THE COMBINED GAS LAW

12.18. Calculate the missing value for each set of data in the following table:

	P_1	V_1	T_1	P_2	V_2	T_2
(*a*)	—	29.1 L	45°C	780 torr	2.22 L	77°C
(*b*)	12.0 atm	—	28°C	12.0 atm	750 mL	53°C
(*c*)	721 torr	200 mL	—	1.21 atm	0.850 L	100°C
(*d*)	1.00 atm	4.00 L	273 K	1.00 atm	2.00 L	—
(*e*)	7.00 atm	—	333 K	3.10 atm	6.00 L	444 K
(*f*)	1.00 atm	3.65 L	130°C	—	5.43 L	130°C

12.19. A 9.00-L sample of gas has its pressure tripled while its absolute temperature is increased by 50%. What is its new volume?

THE IDEAL GAS LAW

12.20. Calculate R, the gas law constant, (*a*) in units of L·torr/(mol·K) and (b) in units of mL·atm/(mol·K).

12.21. How can you recognize an ideal gas law problem?

12.22. Calculate the absolute temperature of 0.118 mol of a gas that occupies 10.0 L at 0.933 atm.

12.23. Calculate the pressure of 0.0303 mol of a gas that occupies 1.24 L at 22°C.

12.24. Calculate the value of R if 1.00 mol of gas occupies 22.4 L at STP.

12.25. Calculate the volume of 0.193 mol of a gas at 27°C and 825 torr.

12.26. Calculate the volume of 1.00 mol of H_2O at 1.00-atm pressure and a temperature of 25°C.

12.27. Calculate the number of moles of a gas that occupies 4.00 L at 303 K and 1.12 atm.

12.28. Calculate the number of moles of gas present in Problem 12.18*d*.

12.29. (*a*) Compare qualitatively the volumes at STP of 6.8 mol N_2 and 6.8 mol H_2. (*b*) Compare the volumes at STP of 6.8 g N_2 and 6.8 g H_2.

12.30. Calculate the volume of 7.07 g of helium at 27°C and 1.00 atm.

12.31. Repeat the prior problem, using hydrogen gas instead of helium. Explain why the occurrence of hydrogen gas in diatomic molecules is so important.

DALTON'S LAW OF PARTIAL PRESSURES

12.32. What is the total pressure of a gas mixture containing He at 0.173 atm, Ne at 0.201 atm, and Ar at 0.550 atm?

12.33. What is the pressure of H_2 if 0.250 mol of H_2 and 0.120 mol of He are placed in a 10.0-L vessel at 27°C?

12.34. What is the total pressure of a gas mixture containing H_2 at 0.173 atm, N_2 at 0.201 atm, and NO at 0.550 atm?

12.35. What is the difference between Problems 12.32 and 12.34?

12.36. The total pressure of a 125-mL sample of oxygen collected over water at 25°C is 1.030 atm. (*a*) How many moles of gas are present? (*b*) How many moles of water vapor are present? (*c*) How many moles of oxygen are present?

12.37. In Dalton's law problems, what is the difference in behavior of water vapor mixed with air compared to helium mixed with air?

Supplementary Problems

12.38. Explain why gas law problems are not given with data to four or five significant figures.

12.39. A mixture of gases contains He, Ne, and Ar. The pressure of He is 0.300 atm. The volume of Ne is 4.00 L. The temperature of Ar is 27°C. What value can be calculated from these data? Explain.

12.40. Under what conditions of temperature and pressure do the gas laws work best?

12.41. If 0.223 g of a gas occupies 2.87 L at 17°C and 700 torr, what is the identity of the gas?

12.42. Which temperature scale must be used in (*a*) Boyle's law problems, (*b*) ideal gas law problems, (*c*) combined gas law problems, and (*d*) Charles' law problems?

12.43. The total pressure of a mixture of gases is 1.50 atm. The mixture contains 0.10 mol of N_2 and 0.20 mol of O_2. What is the partial pressure of O_2?

12.44. Calculate the mass of $KClO_3$ required to decompose to provide 0.728 L of O_2 at 20°C and 1.02 atm.

12.45. Two samples of gas at equal pressures and temperatures are held in containers of equal volume. What can be stated about the comparative number of molecules in each gas sample?

12.46. P is inversely proportional to V. Write three mathematical expressions that relate this fact.

12.47. Write an equation for the combined gas law, using temperature in degrees Celsius. Explain why the Kelvin scale is more convenient.

12.48. Calculate the molar mass of a gas if 12.6 g occupies 5.11 L at 1.12 atm and 22°C.

12.49. What volume of a CO_2 and H_2O mixture at 1.00 atm and 450 K can be prepared by the thermal decomposition of 0.950 g $NaHCO_3$?

$$2\ NaHCO_3 \xrightarrow{\text{heat}} Na_2CO_3 + CO_2 + H_2O(g)$$

12.50. What volume of O_2 at STP can be prepared by the thermal decomposition of 0.500 g of Hg_2O?

12.51. What volume will 14.3 g of CH_4 occupy at 38°C and 0.888 atm?

12.52. Show that the volumes of individual gases involved in a chemical reaction (before they are mixed or after they are separated), all measured at the same temperature and pressure, are in proportion to their numbers of moles. Use the following reaction as an example:

$$2\ CO + O_2 \longrightarrow 2\ CO_2$$

12.53. What is the ratio of volume of CO_2 produced to O_2 used up, both at the same temperature and pressure, for the reaction in the prior problem?

12.54. (*a*) Calculate the initial volume of 3.00 L of gas whose pressure has been increased from 1.00 atm to 2.00 atm.
(*b*) Calculate the final volume of 3.00 L of gas whose pressure has been increased from 1.00 atm to 2.00 atm.

12.55. Calculate the final volume of 1.20 L of gas whose pressure is halved at constant temperature.

12.56. In a certain experiment, when 2.500 g of $KClO_3$ was heated, some O_2 was driven off. After the experiment, 2.402 g of solid was left. (Not all the $KClO_3$ decomposed.) (*a*) Write a balanced chemical equation for the reaction. (*b*) What compounds make up the solid? (*c*) What causes the loss of mass of the sample? (*d*) Calculate the volume of oxygen produced at STP.

12.57. Calculate the ratios of volume to Celsius temperature for the data in Table 12-3. Are the ratios the same for all the temperatures?

12.58. Replot the data of Table 12-1, using P and $1/V$. Is the result a straight line? Explain.

12.59. What is the final temperature or pressure in each of the following parts: (*a*) The pressure of a sample of gas at STP is raised 2 atm. (*b*) The pressure of a sample of gas at STP is raised to 2 atm. (*c*) The temperature of a sample of gas at STP is raised 20°C. (*d*) The temperature of a sample of gas at STP is raised to 20°C?

12.60. For a mixture of two gases, gas 1 and gas 2, put in equals signs or plus signs in the appropriate places in the equations below:

$$P_{\text{total}} = P_1 \quad P_2 \qquad T_{\text{total}} = T_1 \quad T_2$$
$$V_{\text{total}} = V_1 \quad V_2 \qquad n_{\text{total}} = n_1 \quad n_2$$

12.61. Draw a figure to represent all the conversions learned so far, including volumes of gases. Start with the figure you did for Problem 11.31.

CHAPTER 13

Kinetic Molecular Theory

13.1 INTRODUCTION

In Chap. 12 laws governing the behavior of gases were presented. The fact that gases exert pressure was stated, but no reasons why gases should exhibit such behavior were given. The *kinetic molecular theory* (KMT) explains all the gas laws that we have studied and some additional ones. It describes gases in terms of the behavior of the molecules that make them up. (The noble gases exist as individual atoms, but for purposes of explaining the KMT, they will be included and treated as molecules. They may be thought of as monatomic molecules.)

13.2 POSTULATES OF THE KINETIC MOLECULAR THEORY

All gases, under ordinary conditions of temperature and pressure, are made of molecules (including one-atom molecules such as are present in samples of the noble gases). That is, ionic substances do not form gases under conditions prevalent on earth. The molecules of a gas act according to the following postulates:

1. Molecules are in constant random motion.
2. The molecules exhibit negligible intermolecular attractions or repulsions except when they collide.
3. Their collisions are elastic, which means that although the molecules transfer energy from one to another, they do not lose kinetic energy when they collide with one another or with the walls of their container.
4. The molecules occupy a negligible fraction of the volume occupied by the gas as a whole.
5. The average *kinetic energy* of the gas molecules is directly proportional to the *absolute* temperature of the

$$\overline{\text{KE}} = \tfrac{3}{2}kT = \tfrac{1}{2}m\overline{v^2}$$

mean "average." The k in the proportionality constant is called the *Boltzmann constant.* It is ideal gas law constant (with unfamiliar units), divided by Avogadro's number. Note that this for all gases.

eans that molecules move in any direction whatsoever until they collide with another molecule or a wall, whereupon they bounce off and move in another direction until their next collision. Postulate 2 means that the molecules move in a straight line at constant speed between collisions. Postulate 3 means that there is no friction in molecular collisions. The molecules have the same total kinetic energy after each collision as before. Postulate 4 concerns the volume of the molecule themselves versus the volume of the container they occupy. The individual particles do not occupy the entire container. If the molecules of gas had zero volumes and zero intermolecular attractions and repulsions, the gas would obey the ideal gas law exactly. Postulate 5 means that if two gases are at the same temperature, their molecules will have the same average kinetic energies.

EXAMPLE 13.1. Calculate the value for k, the Boltzmann constant, using the following value for R:

$$R = 8.31 \text{ J/(mol·K)}$$

Ans.

$$k = \frac{R}{N} = \frac{8.31 \text{ J/(mol·K)}}{6.02 \times 10^{23} \text{ molecules/mol}} = \frac{1.38 \times 10^{-23} \text{ J}}{\text{molecule·K}}$$

EXAMPLE 13.2. Calculate the average kinetic energy of H_2 molecules at 1.00 atm and 300 K.

Ans.

$$\overline{\text{KE}} = \tfrac{3}{2}kT = 1.5[1.38 \times 10^{-23} \text{ J/(molecule·K)}](300 \text{ K}) = 6.21 \times 10^{-21} \text{ J}$$

13.3. EXPLANATION OF GAS PRESSURE, BOYLE'S LAW, AND CHARLES' LAW

Kinetic molecular theory explains why gases exert pressure. The constant bombardment of the walls of the vessel by the gas molecules, like the hitting of a target by machine gun bullets, causes a constant force to be applied to the wall. The force applied, divided by the area of the wall, is the pressure of the gas.

Boyle's law may be explained using the kinetic molecular theory by considering the box illustrated in Fig. 13-1. If a sample of gas is placed in the left half of the box shown in the figure, it will exert a certain pressure. If the volume is doubled by extending the right wall to include the entire box shown in the figure, the pressure should fall to one-half its original value. Why should that happen? In an oversimplified picture, the molecules bouncing back and forth between the right and left walls now have twice as far to travel, and thus they hit each wall only one-half as often in a given time. Therefore, the pressure is only one-half what it was. How about the molecules that are traveling up and down or in and out? There are as many such molecules as there were before, and they hit the walls as often; but they are now striking an area twice as large, and so the pressure is one-half what it was originally. Thus, doubling the volume halves the pressure. This can be shown to be true no matter what the shape of the container.

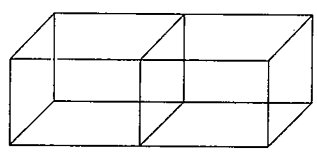

Fig. 13-1. Explanation of Boyle's law

Charles' law governs the volume of a gas at constant pressure when its temperature is changed. When the absolute temperature of a gas is multiplied by 4, for example, the average kinetic energy of its molecules is also multiplied by 4 (postulate 5). The kinetic energy of any particle is given by $\text{KE} = \frac{1}{2}mv^2$, where m is the mass of the particle and v is its velocity. When the kinetic energy is multiplied by 4, what happens to the velocity? It is doubled. (Since the v term is being squared, to effect a fourfold increase, you need only to double the velocity: $2^2 = 4$.)

$$\text{KE}_1 = \tfrac{1}{2}m{v_1}^2$$

$$\text{KE}_2 = \tfrac{1}{2}m{v_2}^2 = 4\text{KE}_1 = \tfrac{1}{2}m(2v_1)^2$$

The velocity v_2 is equal to $2v_1$. On average, in a sample of gas the molecules are going twice as fast at the higher temperature. They therefore hit the walls (1) twice as often per unit time and (2) twice as hard each time they hit, for a combined effect of 4 times the pressure (in a given volume). If we want a constant pressure, we have to expand the volume to 4 times what it was before, and we see that multiplying the absolute temperature by 4 must be accompanied by a fourfold increase in volume if the pressure is to be kept constant.

13.4. GRAHAM'S LAW

An experimental law not yet discussed is *Graham's law*, which states that the rate of effusion or diffusion of a gas is inversely proportional to the square root of its molar mass. *Effusion* is the passage of a gas through small holes in its container, such as the pores in a porous cup. The deflation of a helium-filled party balloon over several days results from the helium atoms effusing through the tiny pores of the balloon wall. *Diffusion* is the passage of a gas through another gas. For example, if a bottle of ammonia water is spilled in one corner of a room, the odor of ammonia is soon apparent throughout the room. The ammonia molecules have diffused through the air molecules. Consider two gases with molar masses MM_1 and MM_2. The ratio of their rates of effusion or diffusion is given by

$$\frac{r_1}{r_2} = \sqrt{\frac{MM_2}{MM_1}}$$

That is, the heavier a molecule of the gas, the more slowly it effuses or diffuses.

The rate of effusion or diffusion of a gas is directly proportional to the "average" velocity of its molecules.

EXAMPLE 13.3. A sample of nitrogen and a sample of neon are both at the same temperature. What is the ratio of the "average" velocities of their molecules?

Ans. Since the temperatures are the same, so are the average kinetic energies of their molecules. From Graham's law,

$$\frac{v_{Ne}}{v_{nitrogen}} = \sqrt{\frac{MM_{nitrogen}}{MM_{Ne}}} = \sqrt{\frac{28.0}{20.2}} = 1.18$$

The neon atoms are moving 1.18 times as fast as the (heavier) nitrogen molecules.

Graham's law may be explained in terms of the kinetic molecular theory as follows: Since two gases are at the same temperature, their average kinetic energies are the same:

$$\overline{KE_1} = \overline{KE_2} = \tfrac{1}{2}m_1\overline{v_1^2} = \tfrac{1}{2}m_2\overline{v_2^2}$$

Multiplying the last of these equations by 2 yields

$$m_1\overline{v_1^2} = m_2\overline{v_2^2} \qquad \text{or} \qquad \frac{m_1}{m_2} = \frac{\overline{v_2^2}}{\overline{v_1^2}}$$

Since the masses of the molecules are proportional to their molar masses, and the average velocity of the molecules is a measure of the rate of effusion or diffusion, all we have to do to this equation to get Graham's law is to take its square root. (The square root of $\overline{v^2}$ is not quite equal to the average velocity, but is a quantity called the *root mean square velocity*. See Problem 13.17.)

Solved Problems

POSTULATES OF THE KINETIC MOLECULAR THEORY

13.1. (*a*) Calculate the volume at 100°C of 18.0 g of liquid water, assuming the density to be 1.00 g/mL. (*b*) Calculate the volume of 18.0 g of water vapor at 100°C and 1.00-atm pressure, using the ideal gas law. (*c*) Assuming that the volume of the liquid is the total volume of the molecules themselves, calculate the percentage of the gas volume occupied by molecules.

13.2. If two different gases are at the same temperature, which of the following must also be equal, (*a*) their pressures, (*b*) their average molecular velocities, or (*c*) the average kinetic energies of their molecules?

13.3. Does the kinetic molecular theory state that all the molecules of a given sample of gas have the same velocity since they are all at one temperature?

13.4. Calculate the temperature at which CO_2 molecules have the same "average" velocity as nitrogen molecules have at 273 K.

13.5. If the molecules of a gas are compressed so that their average distance of separation gets smaller, what should happen to the forces between them? To their ideal behavior?

EXPLANATION OF GAS PRESSURE, BOYLE'S LAW, AND CHARLES' LAW

13.6. Suppose that we double the length of each side of a rectangular box containing a gas. (*a*) What will happen to the volume? (*b*) What will happen to the pressure? (*c*) Explain the effect on the pressure on the basis of the kinetic molecular theory.

GRAHAM'S LAW

13.7. (*a*) If the velocity of a single gas molecule doubles, what happens to its kinetic energy? (*b*) If the average velocity of the molecules of a gas doubles, what happens to the temperature of the gas?

13.8. Would it be possible to separate isotopes by using the principle of Graham's law? Explain what factors would be important.

13.9. Calculate the ratio of rates of effusion of $^{238}UF_6$ and $^{235}UF_6$. Fluorine is 100% ^{19}F.

13.10. List the different molecular masses possible in UCl_3 with ^{238}U and ^{235}U as well as ^{35}Cl and ^{37}Cl.

13.11. What possible complications would there be in trying to separate hydrogen into ^{1}H and ^{2}H by gaseous diffusion?

Supplementary Problems

13.12. (*a*) Is the ratio

$$\frac{\text{Total volume of gas molecules}}{\text{Volume of gas sample}}$$

smaller for a given sample of gas at constant pressure at 300 K or at 400 K? (*b*) Will the gas exhibit more ideal behavior at 300 K or at 400 K?

13.13. (*a*) Is the ratio

$$\frac{\text{Total volume of gas molecules}}{\text{Volume of gas sample}}$$

smaller for a given sample of gas at constant temperature at 1.00 atm or at 2.00 atm? (*b*) Will the gas exhibit more ideal behavior at 1.00 atm or at 2.00 atm?

13.14. (*a*) Calculate the average kinetic energy of O_2 molecules at 1.00 atm and 300 K. (*b*) Does the pressure matter? (*c*) Does the identity of the gas matter?

13.15. (*a*) Calculate the "average" velocity of O_2 molecules at 1.00 atm and 300 K. (*b*) Does the pressure matter? (*c*) Does the identity of the gas matter?

13.16. Explain why neon atoms obey the gas laws the same as nitrogen molecules.

13.17. (*a*) Calculate the square of each of the following numbers: 1, 2, 3, 4, and 5. (*b*) Calculate the average of the numbers. (*c*) Calculate the average of the squares. (*d*) Is the square root of the average of the squares equal to the average of the numbers? (*e*) Explain why quotation marks are used with "average" velocity in the text for the velocity of molecules with average kinetic energies.

13.18. Contrast the motions of the molecules of a sample of gas at rest to those in a hurricane wind.

13.19. Oxygen gas and sulfur dioxide gas are at the same temperature. What is the ratio of the "average" velocities of their molecules?

Solved Problems

INTRODUCTION

14.1. (*a*) What is the formula of a compound of two ions: X^+ and Y^{2-}? (*b*) What is the formula of a covalent compound of two elements: "W" with an oxidation state of +1 and "Z" with an oxidation state of −2?

ASSIGNING OXIDATION NUMBERS

14.2. Show that rules 2 and 3 (Sec. 14.2) are corollaries of rule 1.

14.3. Draw an electron dot diagram for H_2O_2. Assign an oxidation number to oxygen on this basis. Compare this number with that assigned by rule 6 (Sec. 14.2).

14.4. What is the sum of the oxidation numbers of all the *atoms* in the following compounds or ions? (*a*) $PO_4{}^{3-}$, (*b*) $VO_2{}^+$, (*c*) $ClO_2{}^-$, (*d*) $Cr_2O_7{}^{2-}$, (*e*) $SiCl_4$, and (*f*) NaCl.

14.5. Determine the oxidation numbers for the underlined elements: $(\underline{V}O_2)_3\underline{P}O_4$.

14.6. What is the oxidation number of chlorine in each of the following? (*a*) Cl_2O_3, (*b*) $ClO_4{}^-$, and (*c*) ClF_5.

14.7. Determine the oxidation number for the underlined element: (*a*) $\underline{P}OCl_3$, (*b*) $H\underline{N}O_2$, (*c*) $Na_2\underline{S}O_4$, (*d*) $\underline{P}Cl_5$, and (*e*) $\underline{N}_2O_3$.

14.8. Determine the oxidation number for the underlined element: (*a*) $\underline{Cl}O^-$, (*b*) $\underline{P}O_4{}^{3-}$, (*c*) $\underline{S}O_4{}^{2-}$, and (*d*) $\underline{V}O^{2+}$.

14.9. What is the oxidation number of Si in $Si_6O_{18}{}^{12-}$?

14.10. What oxyacid of nitrogen can be prepared by adding water to N_2O_5? *Hint:* Both compounds have nitrogen in the same oxidation state.

PERIODIC RELATIONSHIPS OF OXIDATION NUMBERS

14.11. Predict the formulas of two compounds of each of the following pairs of elements: (*a*) S and O, (*b*) Cl and O, (*c*) P and S, (*d*) P and F, (*e*) I and F, and (*f*) S and F.

14.12. Predict the formulas of four fluorides of iodine.

14.13. Write the formulas for two monatomic ions for each of the following metals: (*a*) Co, (*b*) Tl, (*c*) Sn, and (*d*) Cu.

OXIDATION NUMBERS IN INORGANIC NOMENCLATURE

14.14. Name NO_2 and N_2O_4, using the Stock system. Explain why the older system using prefixes is still useful.

BALANCING OXIDATION-REDUCTION EQUATIONS

14.15. Why is it possible to add H^+ and/or H_2O to an equation for a reaction carried out in aqueous acid solution when none visibly appears or disappears?

14.16. How many electrons are involved in a reaction of one atom with a change of oxidation number from (*a*) +2 to −3 and (*b*) +5 to −2?

14.17. Identify (*a*) the oxidizing agent, (*b*) the reducing agent, (*c*) the element oxidized, and (*d*) the element reduced in the following reaction:

$$8\,H^+ + CrO_4^{\,2-} + 3\,Co^{2+} \longrightarrow Cr^{3+} + 3\,Co^{3+} + 4\,H_2O$$

14.18. Balance the equation for the reduction of HNO_3 to NH_4NO_3 by Mn by the oxidation number change method. Add other compounds as needed.

ELECTROCHEMISTRY

14.19. Even if sodium metal were produced by the electrolysis of aqueous NaCl, what would happen to the sodium produced in the water?

14.20. Explain why Al cannot be produced from its salts in aqueous solution.

14.21. Can a Daniell cell be recharged?

14.22. Explain why a Daniell cell cannot be placed in a single container, like the lead storage cell.

14.23. Write a balanced chemical equation for (*a*) the direct reaction of zinc with copper(II) sulfate and (*b*) the overall reaction in a Daniell cell containing $ZnSO_4$ and $CuSO_4$.

14.24. The electrolysis of brine (concentrated NaCl solution) produces hydrogen at the cathode and chlorine at the anode. Write a net ionic equation for each half-reaction and the total reaction. What other chemical is produced in this commercially important process?

14.25. Must two oxidation-reduction half-reactions be carried out (*a*) in the same location and (*b*) at the same time?

Supplementary Problems

14.26. Explain why we were able to use the charges on the monatomic cations in names in Chap. 6 instead of the required oxidation numbers.

14.27. Determine the oxidation number of oxygen in (*a*) Na_2O_2, (*b*) RbO_2, and (*c*) OF_2.

14.28. Since the reactions in Problem 14.23 are the same, and the Daniell cell produces electric energy, what kind of energy does the direct reaction produce?

14.29. Complete and balance the following redox equations:

(*a*) $MnO_4^- + Hg_2^{2+} \longrightarrow Hg^{2+} + Mn^{2+}$

(*b*) $I^- + H_2O_2 \longrightarrow I_2 + H_2O$

(*c*) $Br_2 + OH^- \longrightarrow Br^- + BrO_3^-$

(*d*) $P_4 + OH^- \longrightarrow PH_3 + HPO_3^{2-}$

(*e*) $H_2C_2O_4 + MnO_4^- \longrightarrow Mn^{2+} + CO_2$

(*f*) $Cr(OH)_2 + H_2O_2 \longrightarrow Cr(OH)_3$

14.30. Which of the equations in Problem 14.29 represent reactions in basic solution? How can you tell?

14.31. Complete and balance the following equations:

(*a*) $Zn + H^+ + NO_3^- \longrightarrow NH_4^+ + Zn^{2+}$

(*b*) $Zn + HNO_3 \longrightarrow NH_4NO_3 + Zn(NO_3)_2$

(*c*) How are these equations related? Which is easier to balance?

14.32. Balance the following equation by the oxidation number change method:

$$NH_3 + O_2 \longrightarrow NO + H_2O$$

14.33. Complete and balance the following equations:

(*a*) $H_2O_2 + Sn^{2+} \longrightarrow Sn^{4+} + H_2O$

(*b*) $Br^- + BrO_3^- \longrightarrow Br_2 + H_2O$

(*c*) $Ce^{4+} + SO_3^{2-} \longrightarrow Ce^{3+} + SO_4^{2-}$

(*d*) $Zn + OH^- \longrightarrow Zn(OH)_4^{2-} + H_2$

(*e*) $Cu_2O + H^+ \longrightarrow Cu + Cu^{2+} + H_2O$

(*f*) $H_2SO_4(conc) + Zn \longrightarrow H_2S + Zn^{2+}$

(*g*) $HNO_3 + Pb(NO_3)_2 \longrightarrow Pb(NO_3)_4 + NO$

(*h*) $H^+ + NO_3^- + Sn^{2+} \longrightarrow NO + Sn^{4+}$

(*i*) $SO_4^{2-} + I^- + H^+ \longrightarrow SO_2 + I_2 + H_2O$

(*j*) $Cr^{3+} + MnO_4^- \longrightarrow Cr_2O_7^{2-} + Mn^{2+}$

(*k*) $Ag^+ + S_2O_3^{2-} \longrightarrow S_4O_6^{2-} + Ag$

14.34. Consider the following part of an equation:

$$NO_2 + MnO_4^- \longrightarrow Mn^{2+} +$$

(*a*) If one half-reaction is a reduction, what must the other half-reaction be? (*b*) To what oxidation state can the nitrogen be changed? (*c*) Complete and balance the equation.

14.35. What is the oxidation number of sulfur in $S_2O_8^{2-}$, the peroxydisulfate ion?

14.36. Which of the following reactions (indicated by unbalanced equations) occur in acid solution and which occur in basic solution?

(*a*) $HNO_3 + Cu \longrightarrow NO + Cu^{2+}$

(*b*) $CrO_4^{2-} + Fe(OH)_2 \longrightarrow Cr(OH)_3 + Fe(OH)_3$

(*c*) $N_2H_4 + I^- \longrightarrow NH_4^+ + I_2$

14.37. Complete and balance the following equations:

(*a*) $Cl^- + MnO_2 + H^+ \longrightarrow Mn^{2+} + H_2O + Cl_2$

(*b*) $ClO^- \longrightarrow Cl^- + ClO_3^-$

(*c*) $Pb + PbO_2 + SO_4^{2-} \longrightarrow PbSO_4 + H_2O$

(*d*) $I_3^- + Co(CN)_6^{4-} \longrightarrow Co(CN)_6^{3-} + I^-$

(*e*) $V^{2+} + H_3AsO_4 \longrightarrow HAsO_2 + VO^{2+}$

(*f*) $Hg_2^{2+} + CN^- \longrightarrow C_2N_2 + Hg$

(*g*) $VO^{2+} + AsO_4^{3-} \longrightarrow AsO_2^- + VO_2^+$

14.38. The oxidizing ability of H_2SO_4 depends on its concentration. Which element is reduced by reaction of Zn on H_2SO_4 in each of the following reactions?

$$Zn + H_2SO_4(\text{dilute}) \longrightarrow ZnSO_4 + H_2$$
$$4\,Zn + 5\,H_2SO_4(\text{conc}) \longrightarrow H_2S + 4\,H_2O + 4\,ZnSO_4$$

14.39. Complete and balance the following equation:

$$Sb_2S_3 + HNO_3 \longrightarrow H_3SbO_4 + SO_2 + NO$$

14.40. What is the maximum oxidation state of fluorine in any compound?

14.41. Calculate the oxidation number of carbon in (*a*) CH_2O and (*b*) CH_2F_2.

14.42. What is the more likely formula for bismuth in the +5 oxidation state—Bi^{5+} or BiO_3^-?

14.43. Explain why direct current (dc) rather than alternating current (ac) is used for electrolysis. Why is direct current used in cars?