Fundamentals of Engineering (FE) General Exam
Review Program
March 3rd ~ April 12th, 2010

Chemistry

David K. Ryan, Ph.D.

Department of Chemistry
University of Massachusetts Lowell

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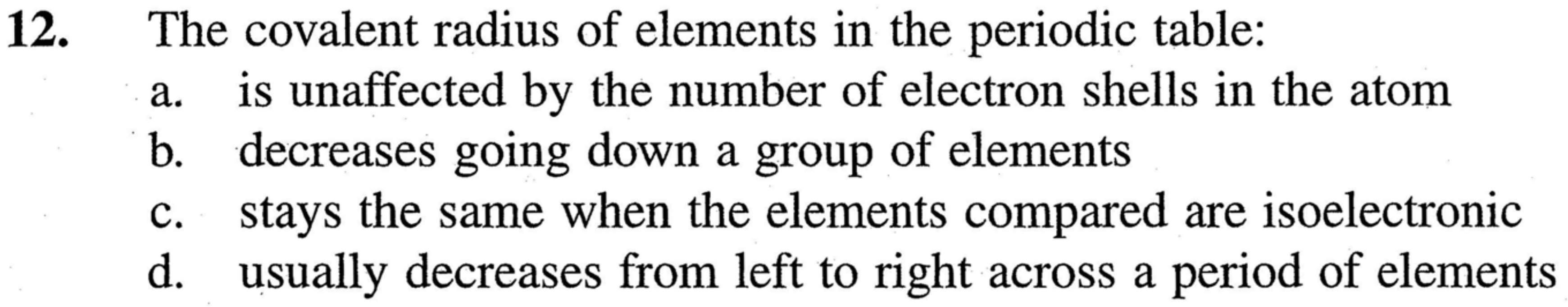
http://faculty.uml.edu/david_ryan/

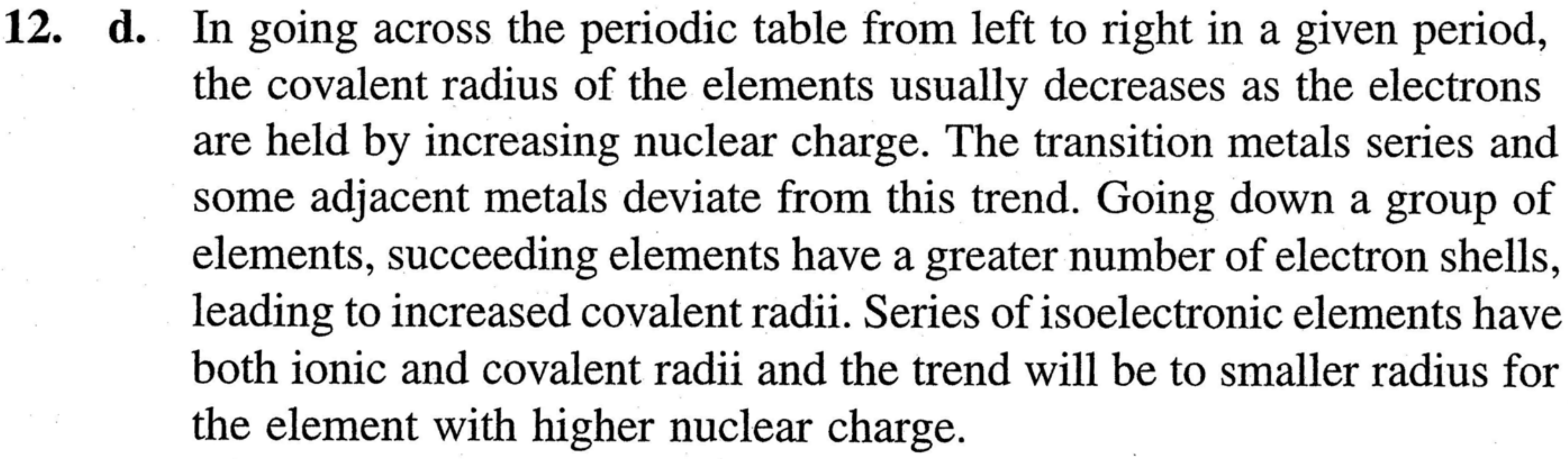
Problems



12. The covalent radius of elements in the periodic table:

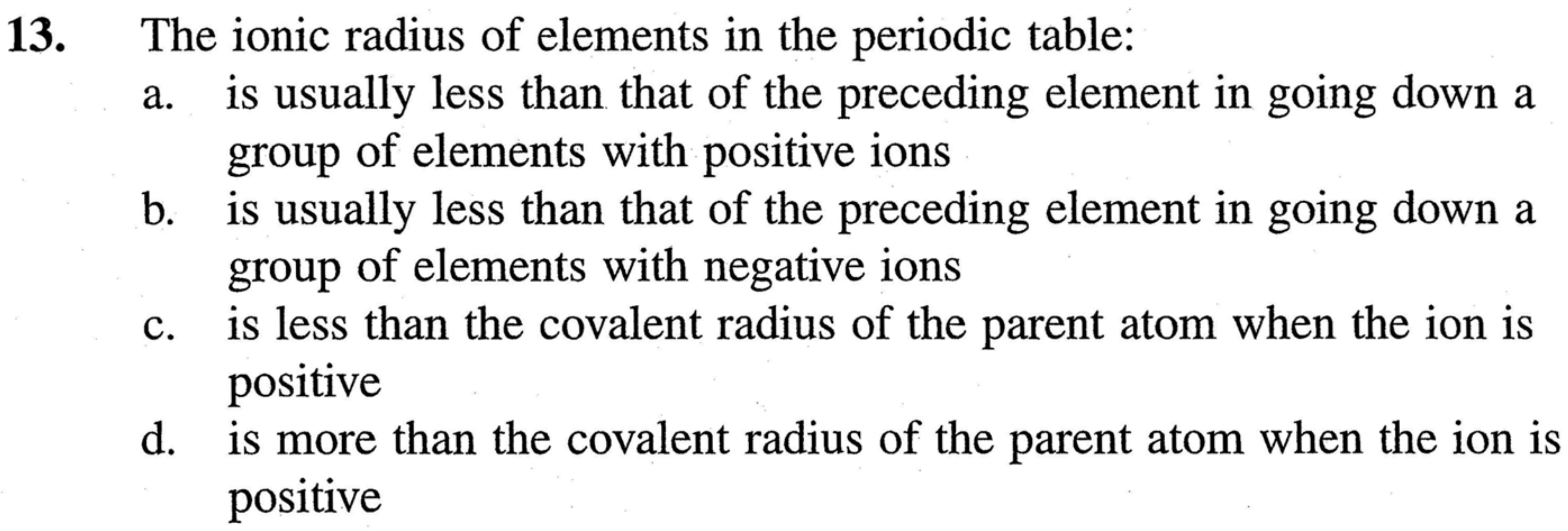
a. is unaffected by the number of electron shells in the atom
b. decreases going down a group of elements
c. stays the same when the elements compared are isoelectronic
d. usually decreases from left to right across a period of elements

Problems



12. **d.** In going across the periodic table from left to right in a given period, the covalent radius of the elements usually decreases as the electrons are held by increasing nuclear charge. The transition metals series and some adjacent metals deviate from this trend. Going down a group of elements, succeeding elements have a greater number of electron shells, leading to increased covalent radii. Series of isoelectronic elements have both ionic and covalent radii and the trend will be to smaller radius for the element with higher nuclear charge.

Problems



13. The ionic radius of elements in the periodic table:

a. is usually less than that of the preceding element in going down a group of elements with positive ions

b. is usually less than that of the preceding element in going down a group of elements with negative ions

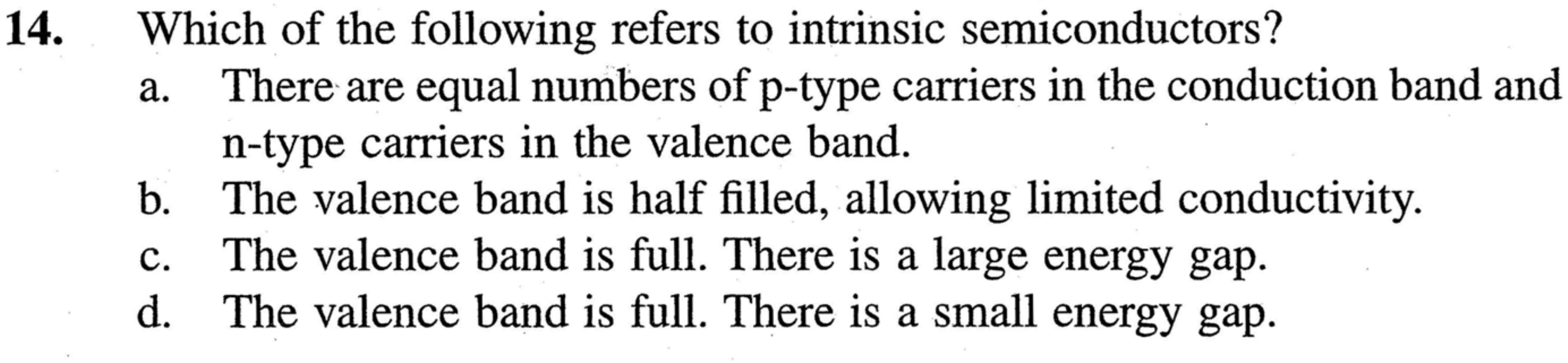
c. is less than the covalent radius of the parent atom when the ion is positive

d. is more than the covalent radius of the parent atom when the ion is positive

Problems

13. **c.** Ionic radii are larger than covalent radii for negative ions and smaller for positive ions. Going down a group both positive and negative ions tend to have larger radii because of the larger electron shells. In an isolectronic series the ions with higher atomic number (greater nuclear charge) tend to have smaller ionic radii.

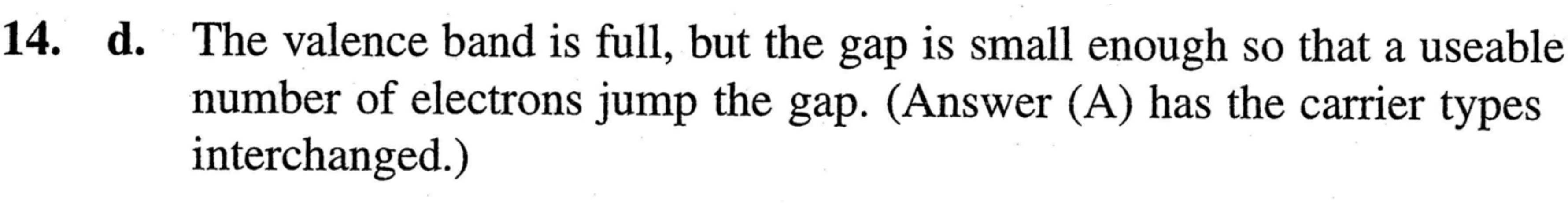
Problems



14. Which of the following refers to intrinsic semiconductors?

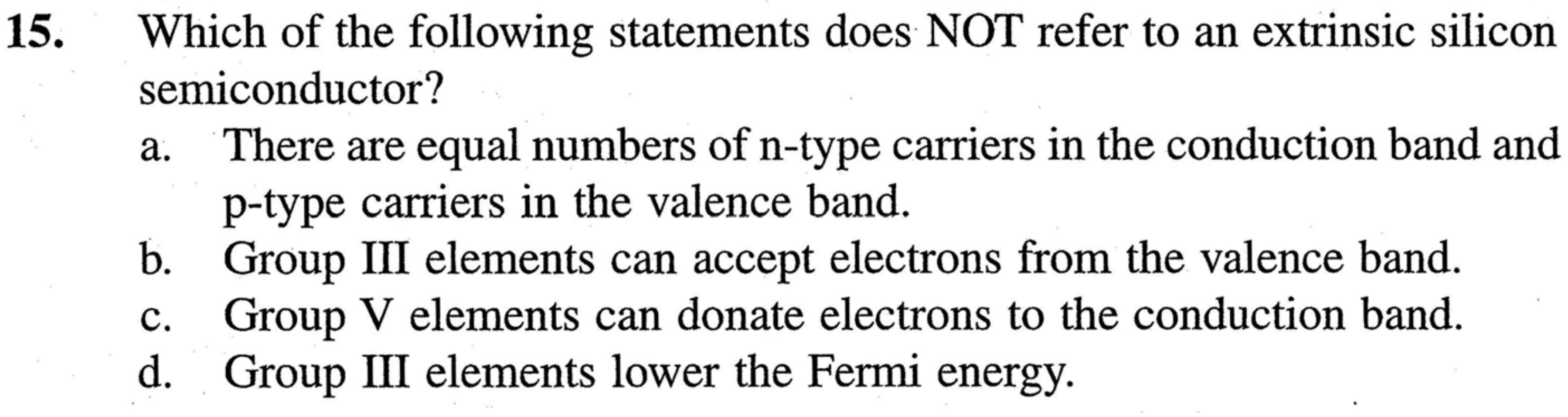
a. There are equal numbers of p-type carriers in the conduction band and n-type carriers in the valence band.
b. The valence band is half filled, allowing limited conductivity.
c. The valence band is full. There is a large energy gap.
d. The valence band is full. There is a small energy gap.

Problems



14. **d.** The valence band is full, but the gap is small enough so that a useable number of electrons jump the gap. (Answer (A) has the carrier types interchanged.)

Problems



15. Which of the following statements does NOT refer to an extrinsic silicon semiconductor?

a. There are equal numbers of n-type carriers in the conduction band and p-type carriers in the valence band.

b. Group III elements can accept electrons from the valence band.

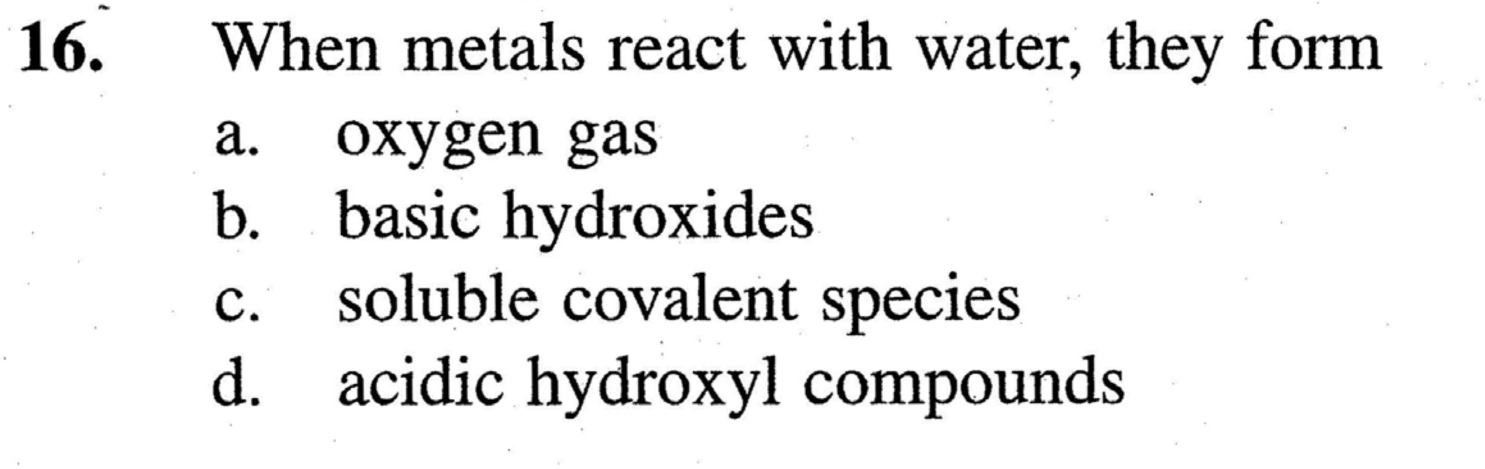
c. Group V elements can donate electrons to the conduction band.

d. Group III elements lower the Fermi energy.

Problems

15. **a.** This question pertains to extrinsic semiconductors and NOT to intrinsic semiconductors.

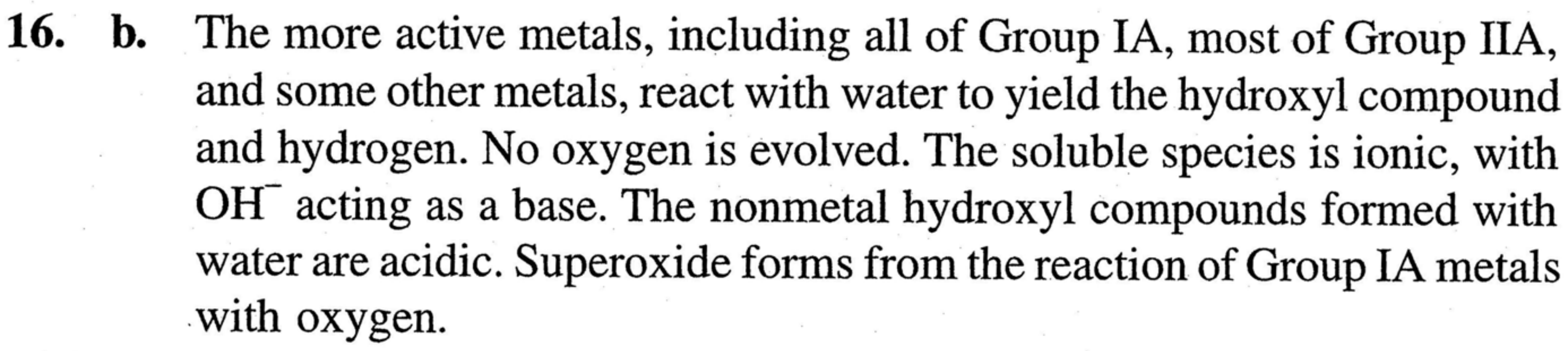
Problems



16. When metals react with water, they form

a. oxygen gas
b. basic hydroxides
c. soluble covalent species
d. acidic hydroxyl compounds

Problems



16. **b.** The more active metals, including all of Group IA, most of Group IIA, and some other metals, react with water to yield the hydroxyl compound and hydrogen. No oxygen is evolved. The soluble species is ionic, with OH^- acting as a base. The nonmetal hydroxyl compounds formed with water are acidic. Superoxide forms from the reaction of Group IA metals with oxygen.

Problems

17. A 0.4157 g sample containing sodium carbonate (Na_2CO_3) and inert impurities requires 37.46 ml of 0.1023 molar HCl for titration. The percentage by weight of sodium carbonate in the sample is most nearly

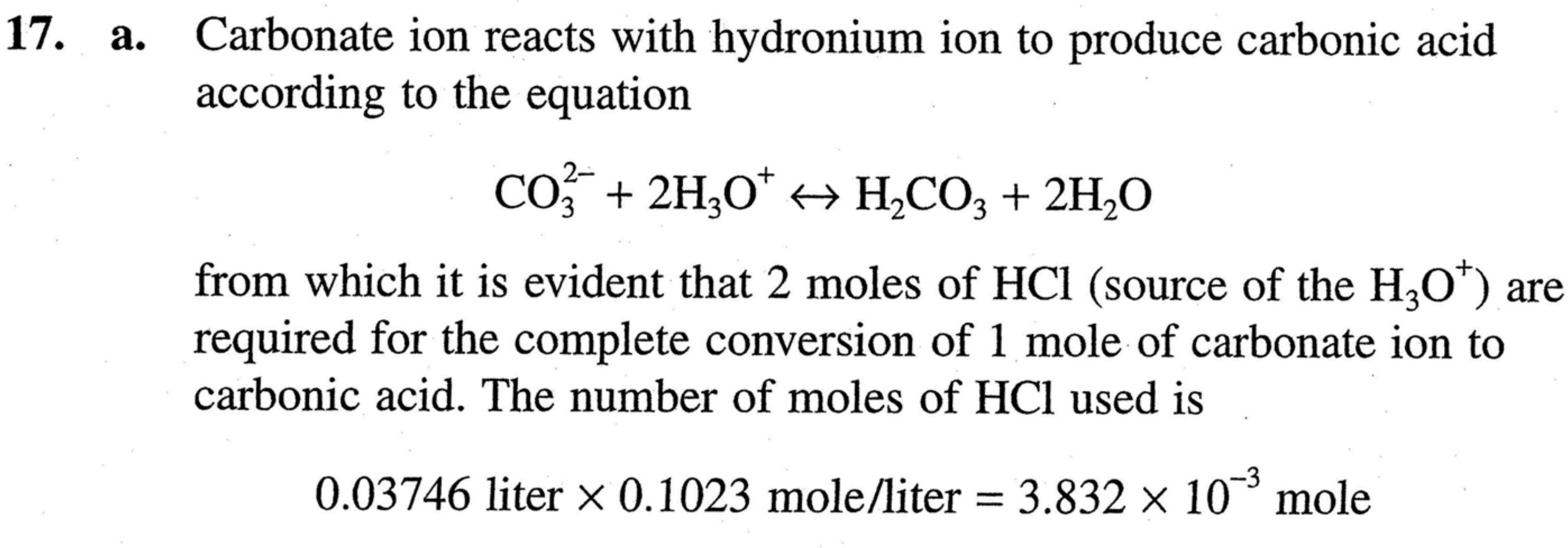
a. 48.86%

b. 100%

c. 97.72%

d. 24.43%

Problems



17. **a.** Carbonate ion reacts with hydronium ion to produce carbonic acid according to the equation

$$CO_3^{2-} + 2H_3O^+ \leftrightarrow H_2CO_3 + 2H_2O$$

from which it is evident that 2 moles of HCl (source of the H_3O^+) are required for the complete conversion of 1 mole of carbonate ion to carbonic acid. The number of moles of HCl used is

$$0.03746 \text{ liter} \times 0.1023 \text{ mole/liter} = 3.832 \times 10^{-3} \text{ mole}$$

Problems

Since 2 moles of HCl are required to react with 1 mole of CO_3^{2-}, the number of moles of sodium carbonate present is half this amount, or

$$3.832 \times 10^{-3} \text{ mole}/2 = 1.916 \times 10^{-3} \text{ mole}$$

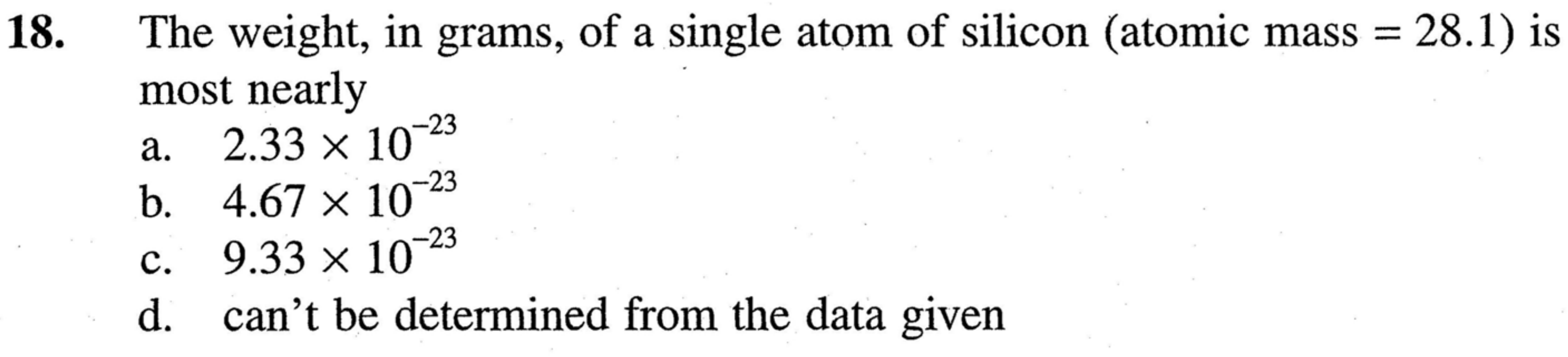
The formula weight of Na_2CO_3 is 106.0, and the weight of Na_2CO_3 in the sample is therefore

$$1.916 \times 10^{-3} \text{ mole} \times 106.0 \text{ g/mole} = 0.2031 \text{ g}$$

The percentage by weight of Na_2CO_3 in the sample is thus

$$0.2031 \text{ g}/0.4157 \text{ g} = 0.4886 = 48.86\%$$

Problems



18. The weight, in grams, of a single atom of silicon (atomic mass = 28.1) is most nearly

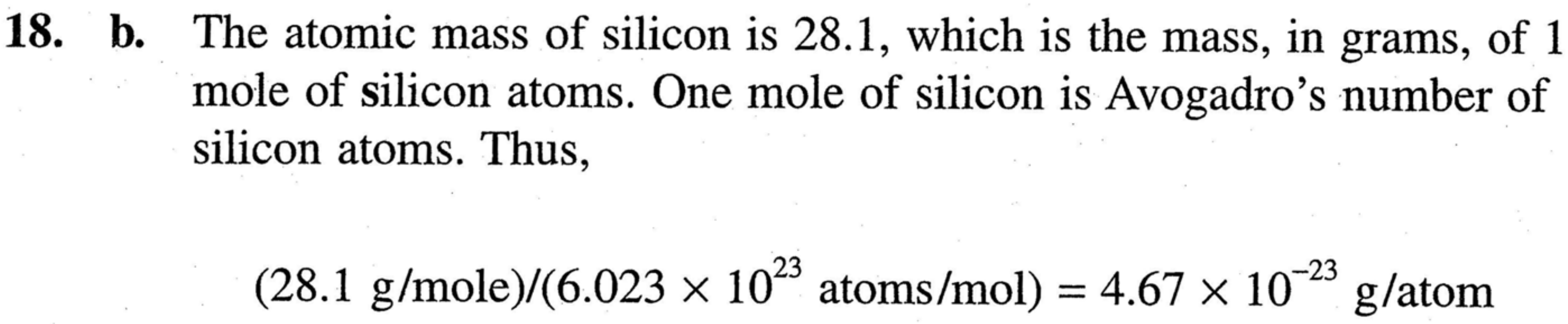
a. 2.33×10^{-23}

b. 4.67×10^{-23}

c. 9.33×10^{-23}

d. can't be determined from the data given

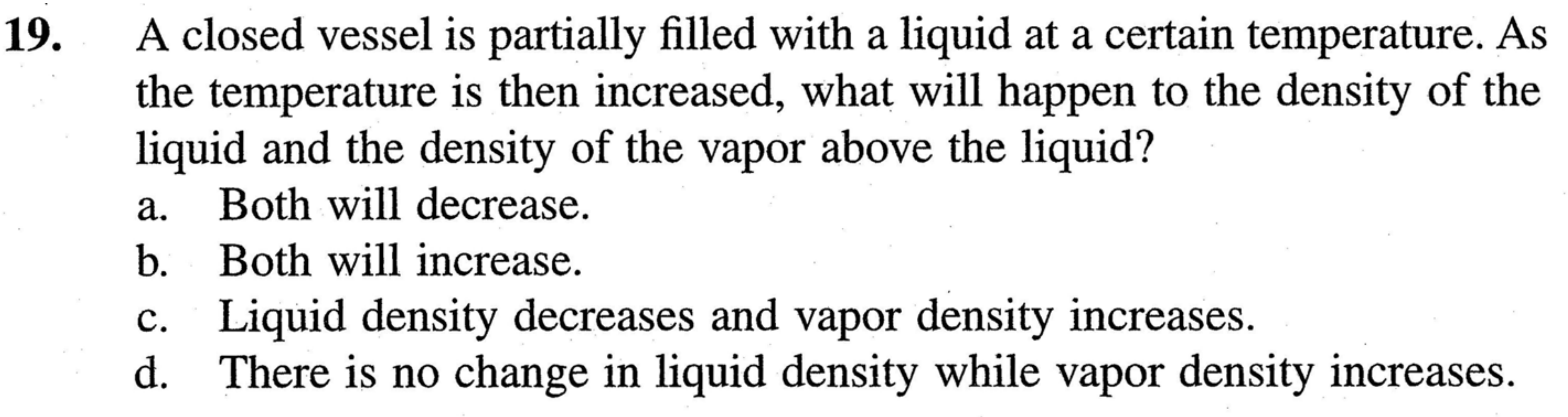
Problems



18. **b.** The atomic mass of silicon is 28.1, which is the mass, in grams, of 1 mole of silicon atoms. One mole of silicon is Avogadro's number of silicon atoms. Thus,

$$(28.1 \text{ g/mole})/(6.023 \times 10^{23} \text{ atoms/mol}) = 4.67 \times 10^{-23} \text{ g/atom}$$

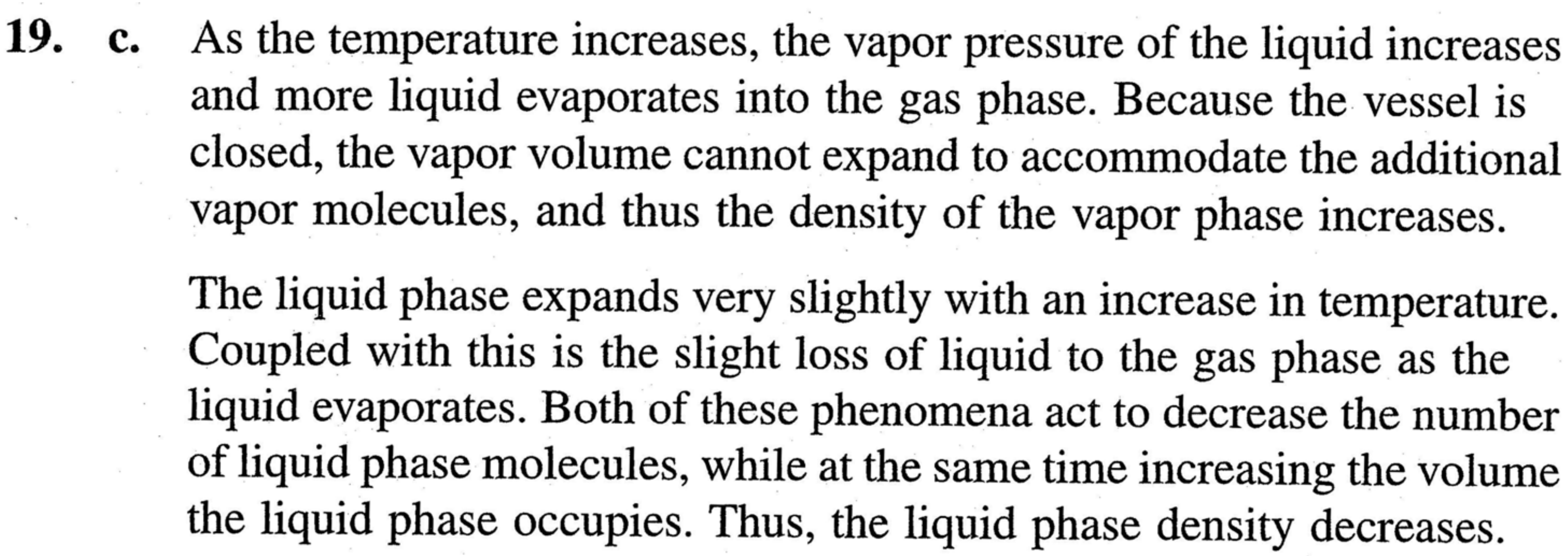
Problems



19. A closed vessel is partially filled with a liquid at a certain temperature. As the temperature is then increased, what will happen to the density of the liquid and the density of the vapor above the liquid?

a. Both will decrease.
b. Both will increase.
c. Liquid density decreases and vapor density increases.
d. There is no change in liquid density while vapor density increases.

Problems



19. **c.** As the temperature increases, the vapor pressure of the liquid increases and more liquid evaporates into the gas phase. Because the vessel is closed, the vapor volume cannot expand to accommodate the additional vapor molecules, and thus the density of the vapor phase increases.

The liquid phase expands very slightly with an increase in temperature. Coupled with this is the slight loss of liquid to the gas phase as the liquid evaporates. Both of these phenomena act to decrease the number of liquid phase molecules, while at the same time increasing the volume the liquid phase occupies. Thus, the liquid phase density decreases.

Problems

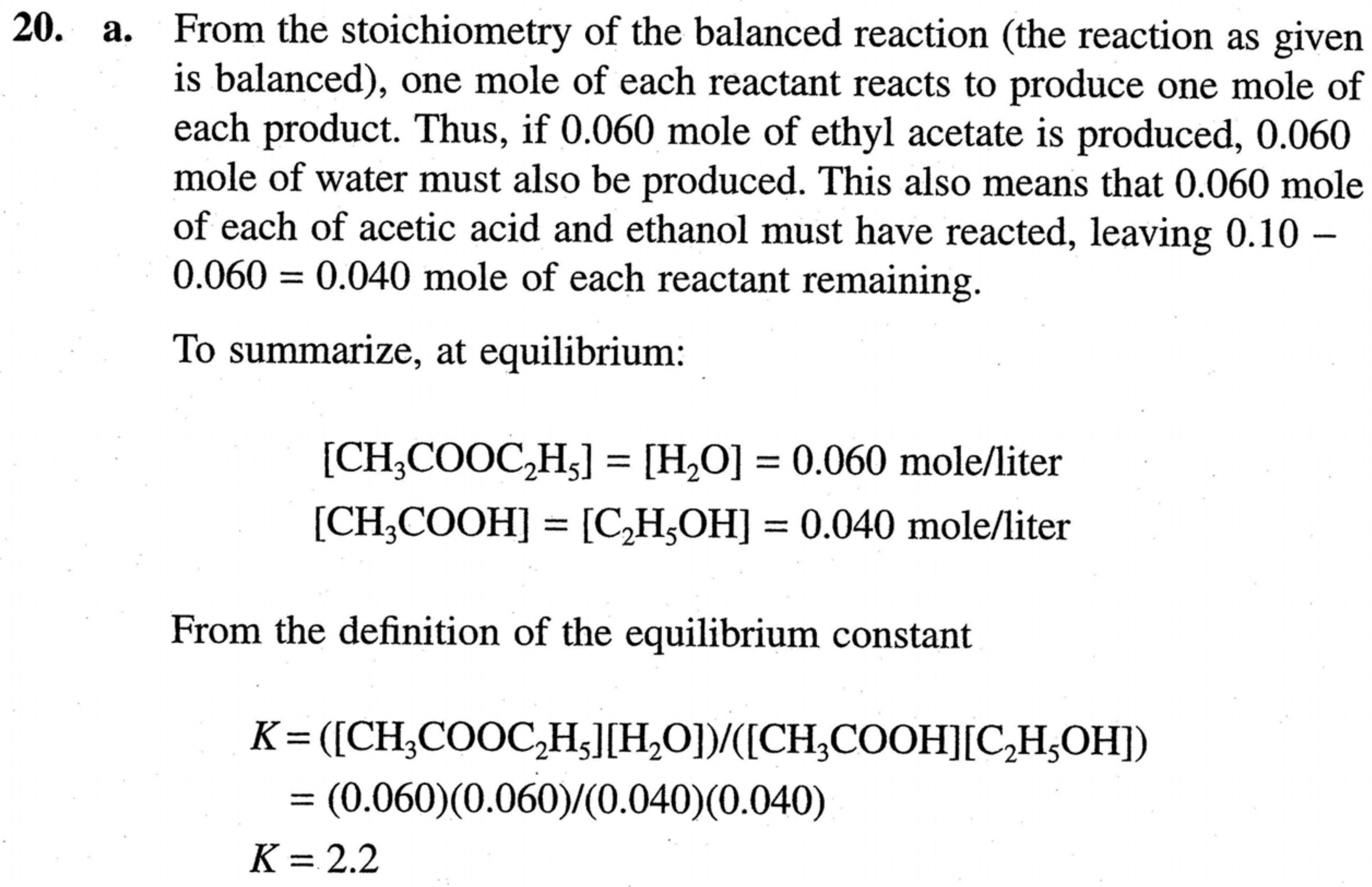
20. Ethyl acetate can be prepared from acetic acid and ethanol according to the equation

$$CH_3COOH + C_2H_5OH - CH_3COOC_2H_5 + H_2O$$

At some temperature, 0.10 mole of acetic acid is mixed with 0.10 mole of ethyl alcohol in 1.0 liter of solution (a nonreactive solvent is used), and when equilibrium has been established 0.060 mole of ethyl acetate is present. The equilibrium constant, *K*, for this reaction is most nearly.

a. 2.2
b. 0.44
c. 1.0
d. can't be determined from the data given

Problems



20. **a.** From the stoichiometry of the balanced reaction (the reaction as given is balanced), one mole of each reactant reacts to produce one mole of each product. Thus, if 0.060 mole of ethyl acetate is produced, 0.060 mole of water must also be produced. This also means that 0.060 mole of each of acetic acid and ethanol must have reacted, leaving 0.10 – 0.060 = 0.040 mole of each reactant remaining.

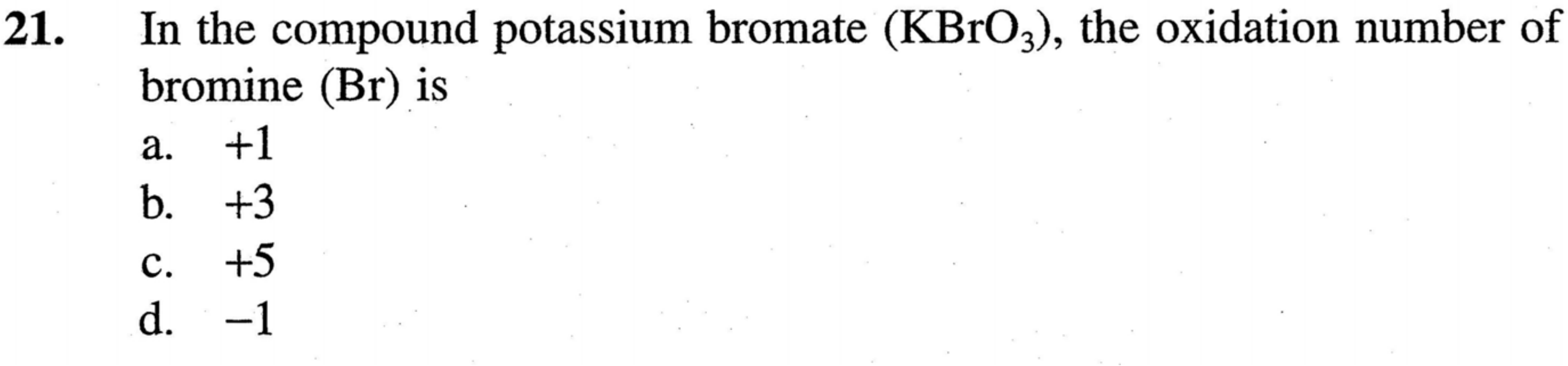
To summarize, at equilibrium:

$$[CH_3COOC_2H_5] = [H_2O] = 0.060 \text{ mole/liter}$$
$$[CH_3COOH] = [C_2H_5OH] = 0.040 \text{ mole/liter}$$

From the definition of the equilibrium constant

$$K = ([CH_3COOC_2H_5][H_2O])/([CH_3COOH][C_2H_5OH])$$
$$= (0.060)(0.060)/(0.040)(0.040)$$
$$K = 2.2$$

Problems



21. In the compound potassium bromate ($KBrO_3$), the oxidation number of bromine (Br) is

a. +1

b. +3

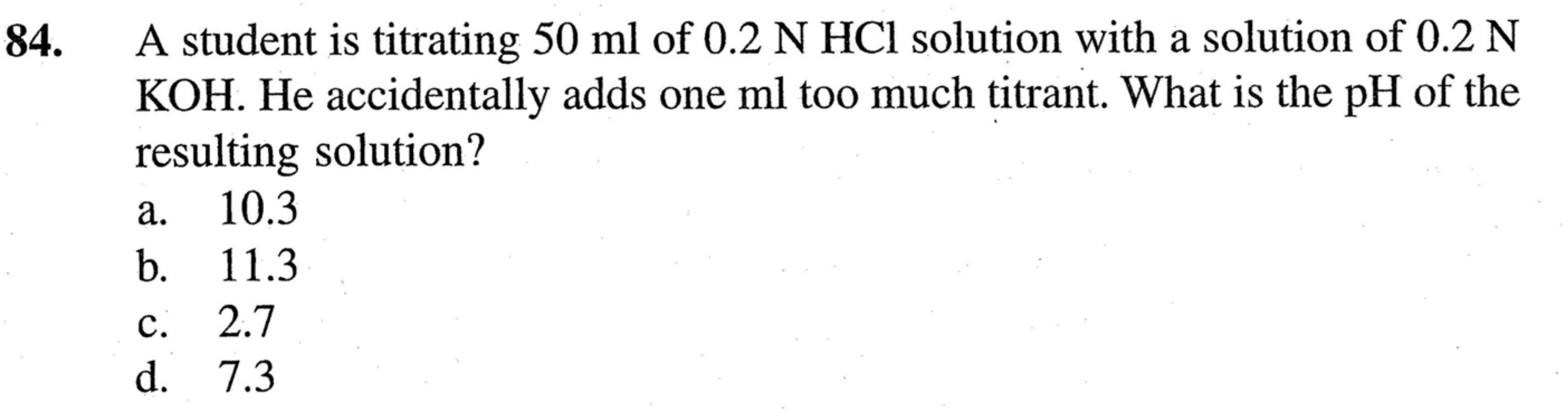
c. +5

d. −1

Problems

21. **c.** We begin by assigning oxidation numbers to the potassium and the oxygen. Potassium, being an alkali metal, will always have an oxidation number of +1, while oxygen, unless it is a part of a peroxide, will almost always have an oxidation number of –2. Thus, with three oxygen atoms present, the total oxidation number of the oxygen is –6, while that of the potassium is +1. To make the molecule neutral as a whole, bromine must be +5.

Problems



84. A student is titrating 50 ml of 0.2 N HCl solution with a solution of 0.2 N KOH. He accidentally adds one ml too much titrant. What is the pH of the resulting solution?

a. 10.3
b. 11.3
c. 2.7
d. 7.3

Problems

84. **b.** Since both solutions are the same normality, the equivalence point will be at 50 ml KOH solution. If one ml excess is added, there will be 51 ml of KOH but all will be neutralized by the HCl except for one ml. The resulting solution will have an excess of one ml of 0.2 N KOH in 101 ml solution. Since we are limited to one significant figure by the data, 100 ml is a valid approximation. Then

$$\frac{0.2\text{ equivalents}}{\text{liter}} \times \frac{1\text{ liter}}{1000\text{ ml}} = 0.2 \times 10^{-3}\text{ equivalents in ml excess}$$

$$\frac{0.2\times10^{-3}\text{ equivalents}}{0.1\text{ L solution liter}} = \frac{0.2\times10^{-4}\text{equivalents}}{1\text{ L solution}}$$

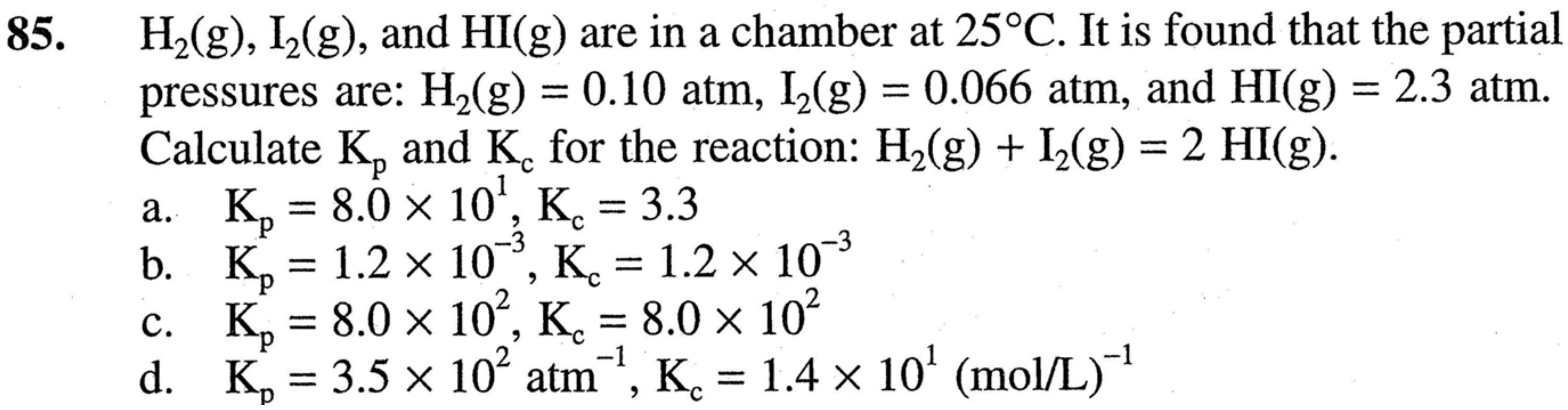
Since equivalent KOH = moles KOH,

$$[\text{OH}-] = 0.2 \times 10^{-4} = 2 \times 10^{-3}\text{ mol/L}$$

$$\text{pOH} = -\log(2 \times 10^{-3}) = -(-3 + 0.30) = 2.7$$

$$\text{pH} = 14 - \text{pOH} = 11.3$$

Problems



85. $H_2(g)$, $I_2(g)$, and HI(g) are in a chamber at 25°C. It is found that the partial pressures are: $H_2(g) = 0.10$ atm, $I_2(g) = 0.066$ atm, and $HI(g) = 2.3$ atm. Calculate K_p and K_c for the reaction: $H_2(g) + I_2(g) = 2\ HI(g)$.

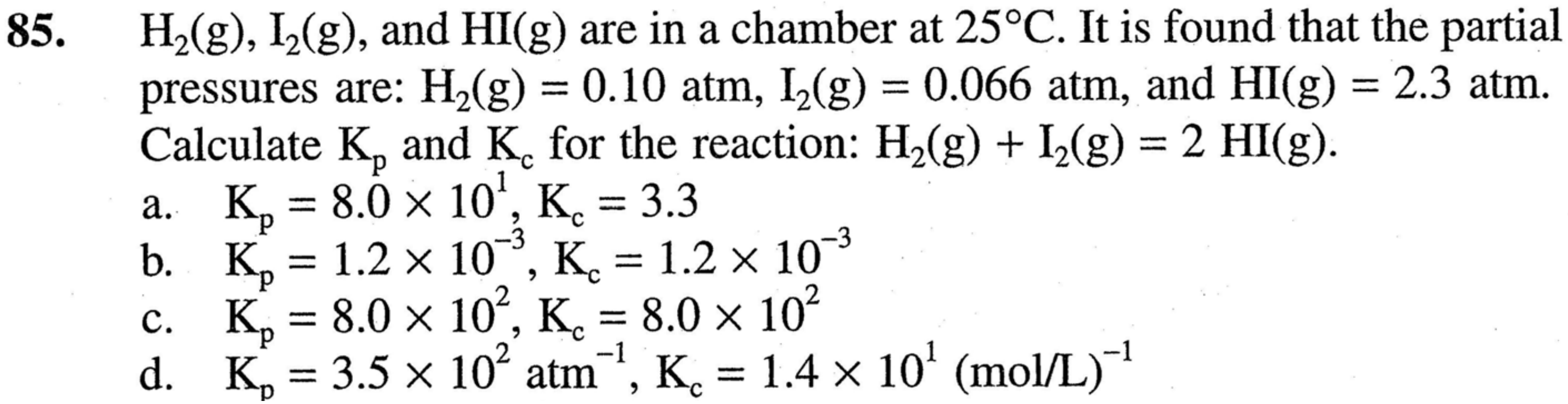
a. $K_p = 8.0 \times 10^1$, $K_c = 3.3$

b. $K_p = 1.2 \times 10^{-3}$, $K_c = 1.2 \times 10^{-3}$

c. $K_p = 8.0 \times 10^2$, $K_c = 8.0 \times 10^2$

d. $K_p = 3.5 \times 10^2\ \text{atm}^{-1}$, $K_c = 1.4 \times 10^1\ (\text{mol/L})^{-1}$

Problems



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a. $K_p = 8.0 \times 10^1$, $K_c = 3.3$
b. $K_p = 1.2 \times 10^{-3}$, $K_c = 1.2 \times 10^{-3}$
c. $K_p = 8.0 \times 10^2$, $K_c = 8.0 \times 10^2$
d. $K_p = 3.5 \times 10^2\ \text{atm}^{-1}$, $K_c = 1.4 \times 10^1\ (\text{mol/L})^{-1}$

Problems

Relating Kc and Kp

The ideal gas equation $PV = nRT$

can be manipulated into a form that reflects a gas concentration term

$P = (n/V)\,RT$ where n/V = molarity

For any gas, A; $P = [A]\,RT$

Problems

Consider the Rxn: $N_2O_4(g) = 2NO_2(g)$

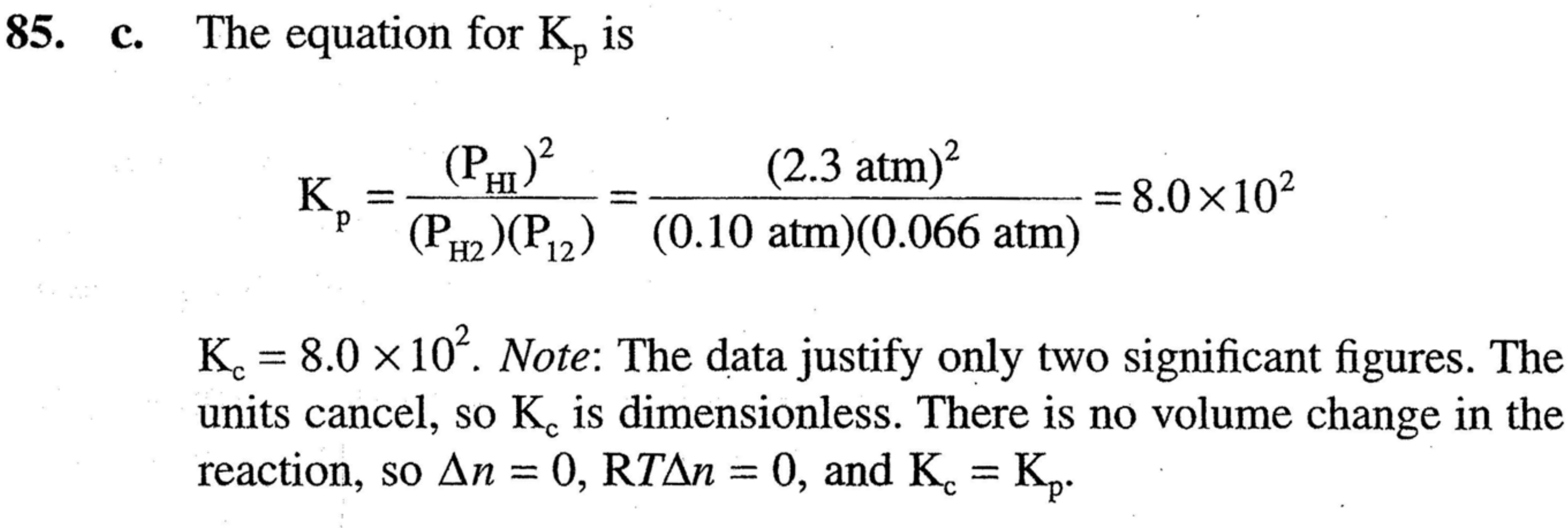
Kp for this rxn follows as:

$$K_P = \frac{P_{NO_2}^2}{P_{N_2O_4}} = \frac{[NO_2]^2(RT)^2}{[N_2O_4](RT)} = K_c(RT)$$

As a general expression: $K_P = K_C(RT)^{\Delta n}$

where Δn = (number moles of gaseous product - number moles of gaseous reactant)

Problems

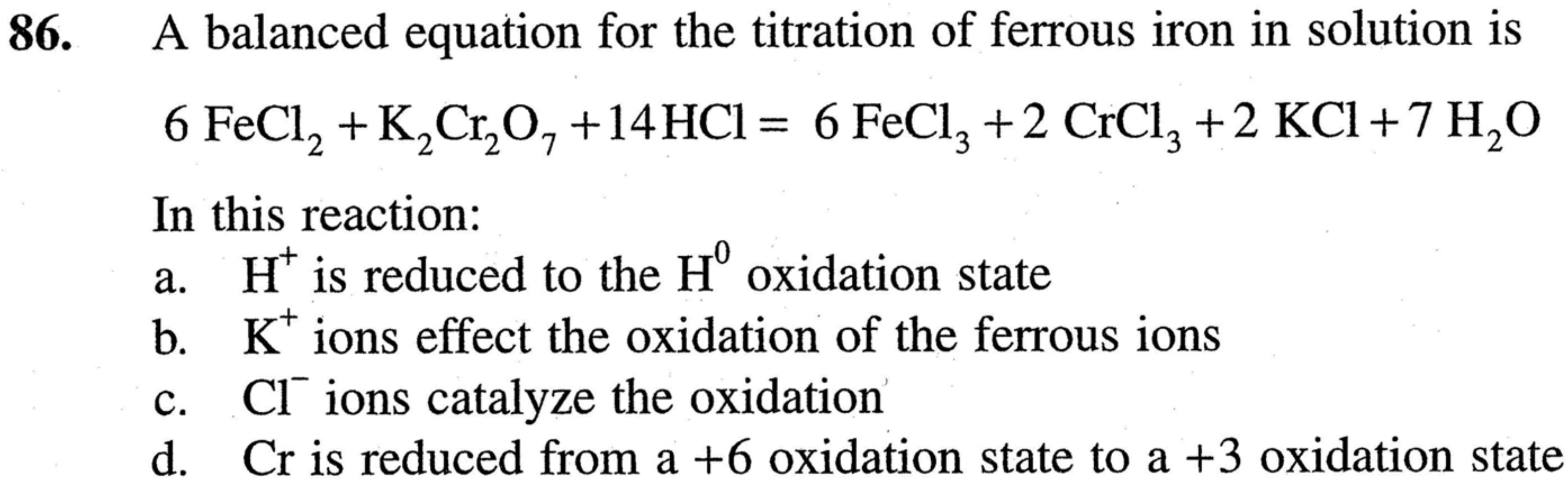


85. **c.** The equation for K_p is

$$K_p = \frac{(P_{HI})^2}{(P_{H2})(P_{12})} = \frac{(2.3\ \text{atm})^2}{(0.10\ \text{atm})(0.066\ \text{atm})} = 8.0 \times 10^2$$

$K_c = 8.0 \times 10^2$. *Note*: The data justify only two significant figures. The units cancel, so K_c is dimensionless. There is no volume change in the reaction, so $\Delta n = 0$, $RT\Delta n = 0$, and $K_c = K_p$.

Problems



86. A balanced equation for the titration of ferrous iron in solution is

$$6\ FeCl_2 + K_2Cr_2O_7 + 14 HCl = 6\ FeCl_3 + 2\ CrCl_3 + 2\ KCl + 7\ H_2O$$

In this reaction:

a. H^+ is reduced to the H^0 oxidation state

b. K^+ ions effect the oxidation of the ferrous ions

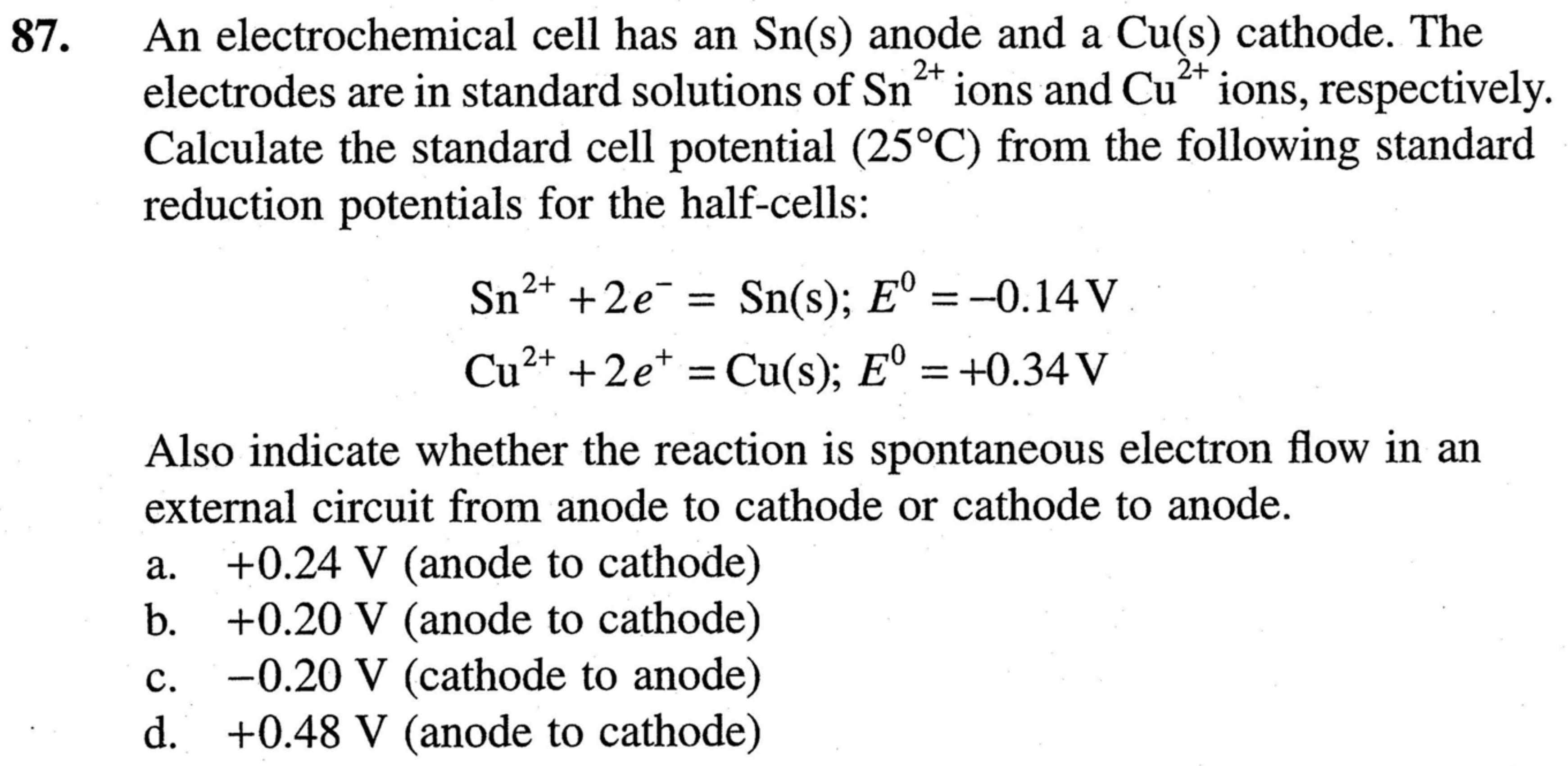
c. Cl^- ions catalyze the oxidation

d. Cr is reduced from a +6 oxidation state to a +3 oxidation state

Problems

86. **d.** The equation represents an oxidation-reduction reaction used for the determination of iron. The H^+ is not reduced as no H_2 is formed. The K^+ and CI^- ions do not enter into the reaction or affect its rate. Cr in $K_2Cr_2O_7$ is in a +6 oxidation state, and Cr in $CrCI_3$ is in a +3 oxidation state.

Problems



87. An electrochemical cell has an Sn(s) anode and a Cu(s) cathode. The electrodes are in standard solutions of Sn^{2+} ions and Cu^{2+} ions, respectively. Calculate the standard cell potential (25°C) from the following standard reduction potentials for the half-cells:

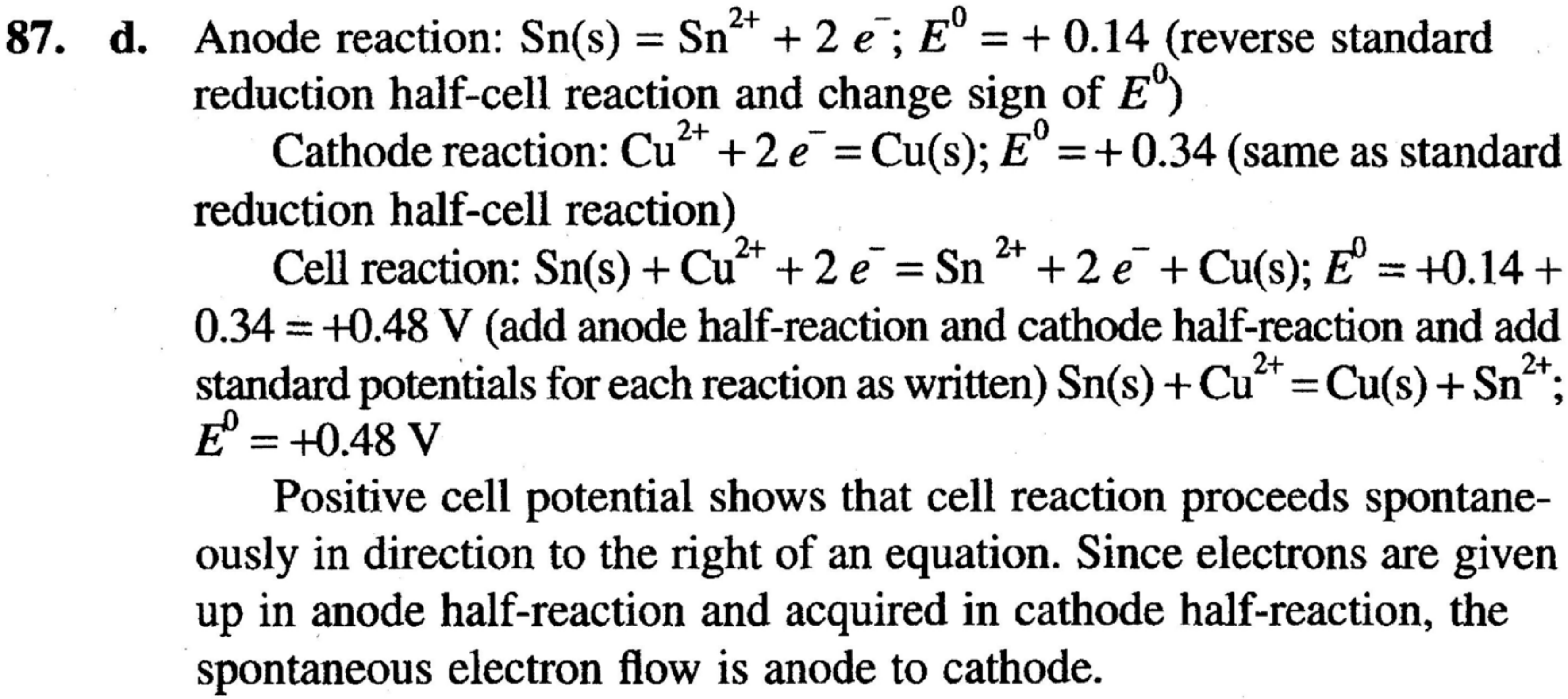
$$Sn^{2+} + 2e^{-} = Sn(s);\ E^{0} = -0.14\,V$$

$$Cu^{2+} + 2e^{+} = Cu(s);\ E^{0} = +0.34\,V$$

Also indicate whether the reaction is spontaneous electron flow in an external circuit from anode to cathode or cathode to anode.

a. +0.24 V (anode to cathode)
b. +0.20 V (anode to cathode)
c. −0.20 V (cathode to anode)
d. +0.48 V (anode to cathode)

Problems



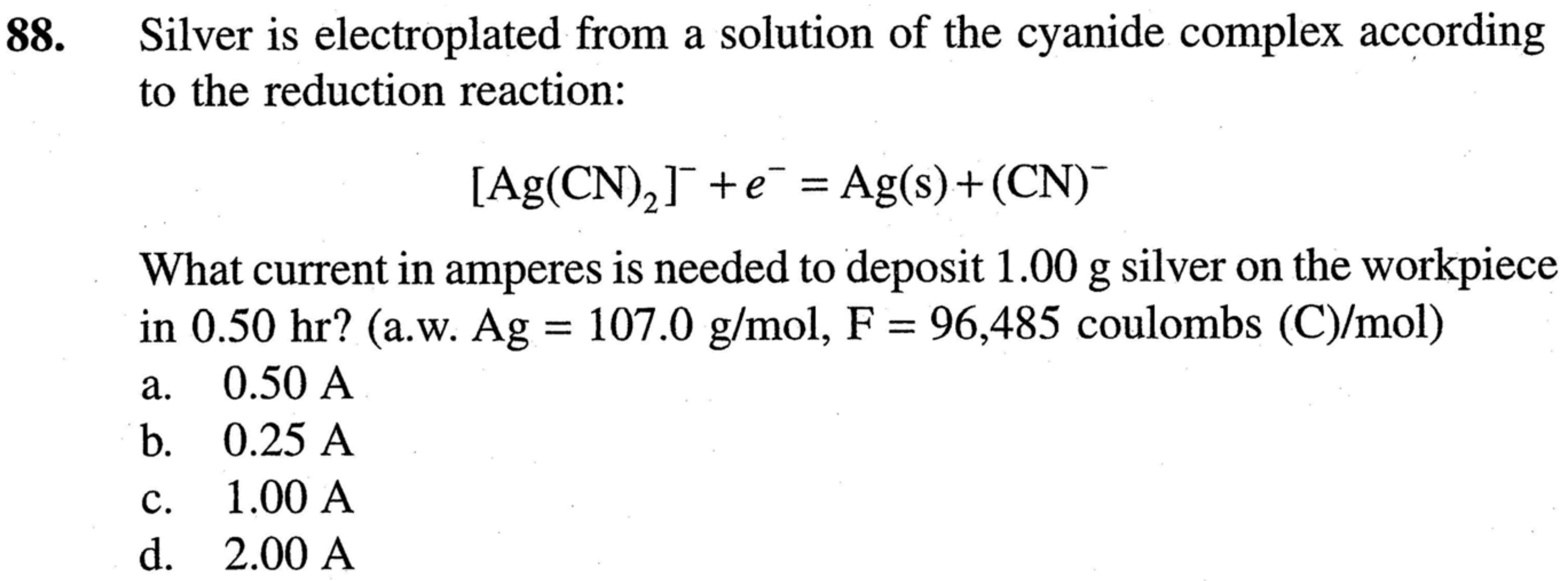
87. d. Anode reaction: $\mathrm{Sn(s)} = \mathrm{Sn}^{2+} + 2\,e^{-}$; $E^{0} = +\,0.14$ (reverse standard reduction half-cell reaction and change sign of E^{0})

Cathode reaction: $\mathrm{Cu}^{2+} + 2\,e^{-} = \mathrm{Cu(s)}$; $E^{0} = +\,0.34$ (same as standard reduction half-cell reaction)

Cell reaction: $\mathrm{Sn(s)} + \mathrm{Cu}^{2+} + 2\,e^{-} = \mathrm{Sn}^{2+} + 2\,e^{-} + \mathrm{Cu(s)}$; $E^{0} = +0.14 + 0.34 = +0.48$ V (add anode half-reaction and cathode half-reaction and add standard potentials for each reaction as written) $\mathrm{Sn(s)} + \mathrm{Cu}^{2+} = \mathrm{Cu(s)} + \mathrm{Sn}^{2+}$; $E^{0} = +0.48$ V

Positive cell potential shows that cell reaction proceeds spontaneously in direction to the right of an equation. Since electrons are given up in anode half-reaction and acquired in cathode half-reaction, the spontaneous electron flow is anode to cathode.

Problems



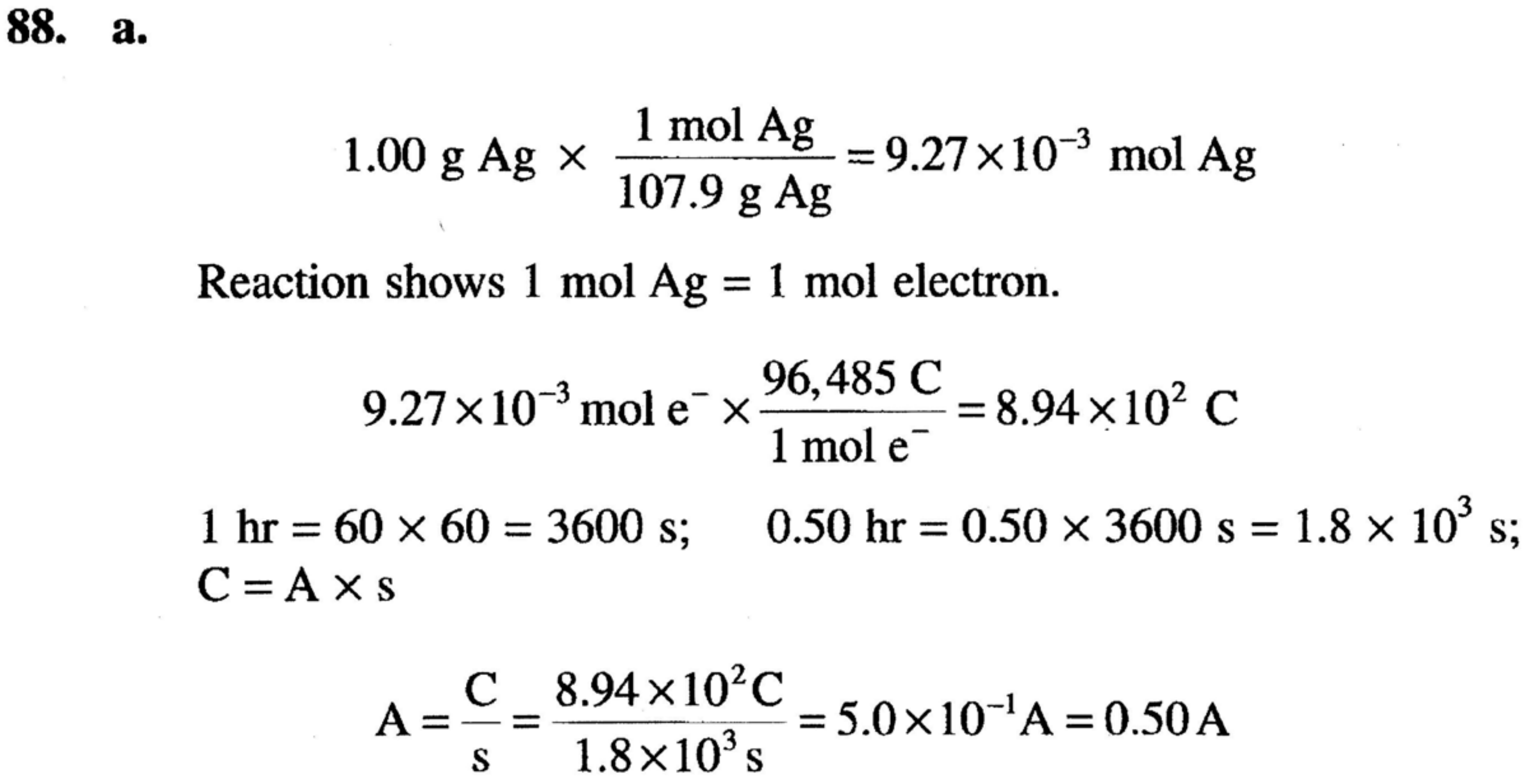
88. Silver is electroplated from a solution of the cyanide complex according to the reduction reaction:

$$[Ag(CN)_2]^- + e^- = Ag(s) + (CN)^-$$

What current in amperes is needed to deposit 1.00 g silver on the workpiece in 0.50 hr? (a.w. Ag = 107.0 g/mol, F = 96,485 coulombs (C)/mol)

a. 0.50 A
b. 0.25 A
c. 1.00 A
d. 2.00 A

Problems



88. a.

$$1.00\text{ g Ag} \times \frac{1\text{ mol Ag}}{107.9\text{ g Ag}} = 9.27\times10^{-3}\text{ mol Ag}$$

Reaction shows 1 mol Ag = 1 mol electron.

$$9.27\times10^{-3}\text{ mol e}^{-} \times \frac{96{,}485\text{ C}}{1\text{ mol e}^{-}} = 8.94\times10^{2}\text{ C}$$

1 hr = 60 × 60 = 3600 s; 0.50 hr = 0.50 × 3600 s = 1.8×10^3 s;
C = A × s

$$\text{A} = \frac{\text{C}}{\text{s}} = \frac{8.94\times10^{2}\text{C}}{1.8\times10^{3}\text{s}} = 5.0\times10^{-1}\text{A} = 0.50\text{A}$$

Problems

89. The gaseous reaction

$$F_2 + 2\ NO_2 \rightarrow 2\ NO_2F$$

is known to be first order in $[F_2]$. In an experiment, the initial concentration $[F_2]$ is held constant and the concentration $[NO_2]$ is doubled. The rate of reaction (R) is found to double. What is the rate law, and what is the overall order of the reaction?

a. $R = k\ [F_2]\ [NO_2]^2$, third order

b. $R = k\ [NO_2]$, first order

c. $R = k\ [F_2]\ [NO_2]$, second order

d. $R = k\ [F_2]^2\ [NO_2]$, third order

Problems

89. **c.** The rate equation is: $R = k\,[F_2]^m\,[NO_2]^n$, where m and n must be experimentally determined. However, since it is given that the reaction is first order in $[F_2]$, $m = 1$. So: $R = k\,[F_2]\,[NO_2]^n$. In the experiment, $[F_2]$ is held constant, and a doubling in $[NO_2]$ gives a doubling in R. Then

$$\frac{R_1}{R_2} = \left(\frac{[NO_2]_1}{[NO_2]_2}\right)^n$$

Since $R_2 = 2 \times R_1$ and $[NO_2]_2 = 2 \times [NO_2]_1$,

$$\frac{1}{2} = \left(\frac{1}{2}\right)^n$$

and $n = 1$. The rate law is $R = [F_2][NO_2]$. Since $m = 1$ and $n = 1$, the overall order of the reaction is $n + m = 2$, or second order.

Problems

90. A 10-g sample of a hydrocarbon is analyzed by combustion and found to contain 9.37 g carbon and 0.63 g hydrogen. The molecular weight of the hydrocarbon sample is estimated from a freezing-point-lowering experiment to be 128 g/mol (a.w. C = 12.01, H = 1.01). The molecular formula of the sample is

a. C_8H_{18}
b. $C_{10}H_8$
c. C_9H_4O
d. C_5H_4

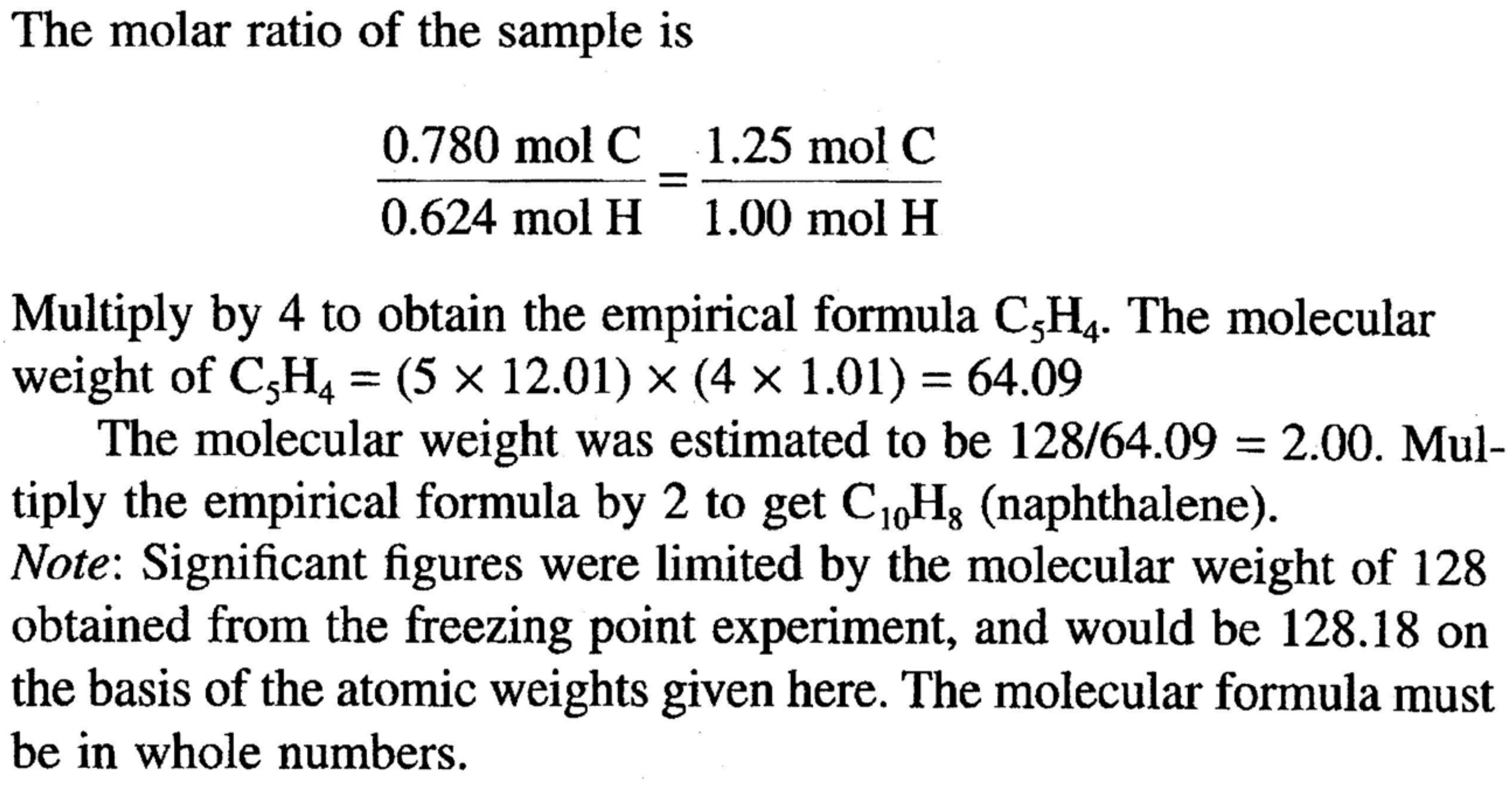
Problems

90. **b.** The number of moles of carbon and hydrogen in the sample are:

$$\text{moles C} = \frac{9.37\text{ g}}{\text{sample}} \times \frac{\text{moles}}{12.01\text{ g}} = 0.780\text{ mol}$$

$$\text{moles H} = \frac{0.63\text{g}}{\text{sample}} \times \frac{\text{moles}}{1.01\text{ g}} = 0.624\text{ mol}$$

Problems



The molar ratio of the sample is

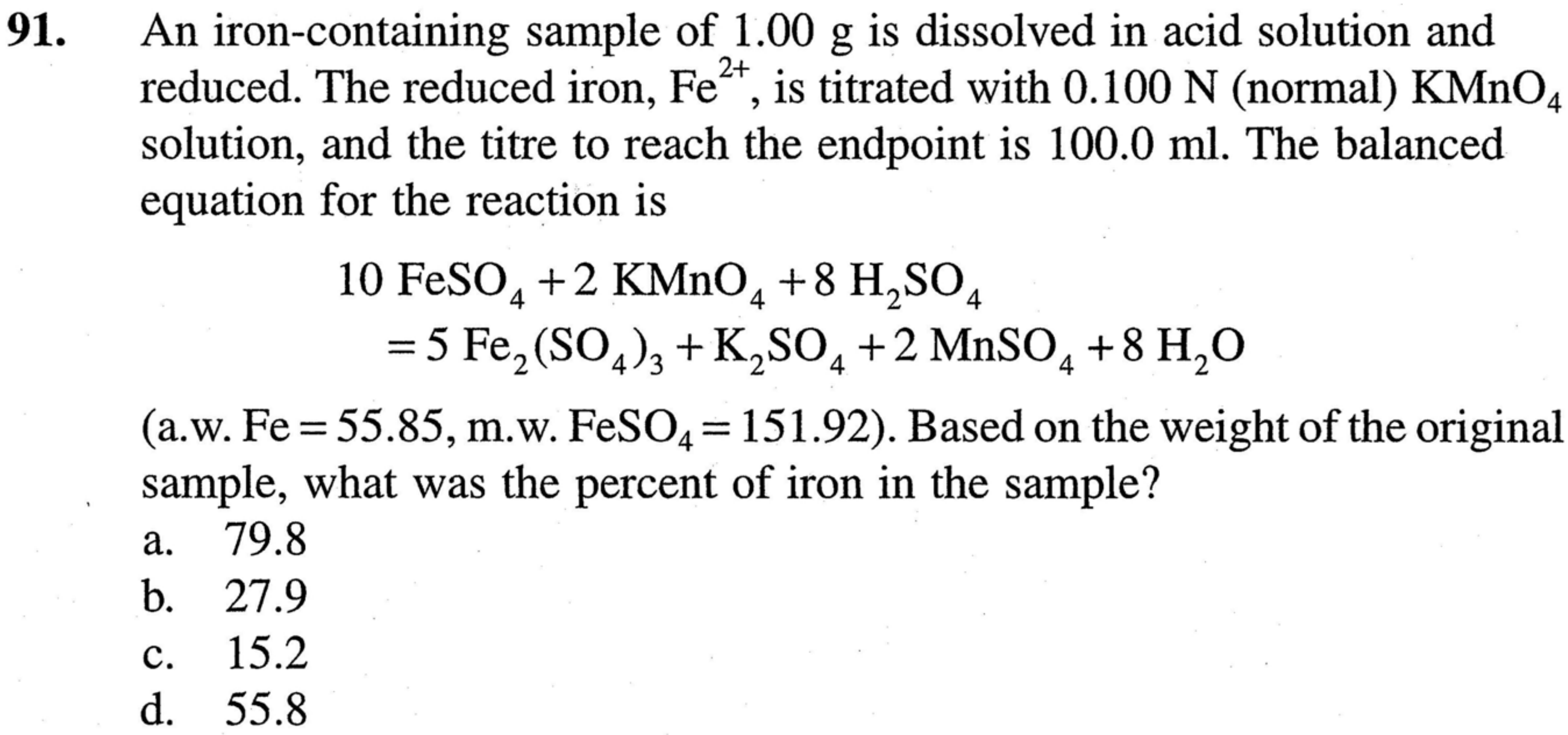
$$\frac{0.780 \text{ mol C}}{0.624 \text{ mol H}} = \frac{1.25 \text{ mol C}}{1.00 \text{ mol H}}$$

Multiply by 4 to obtain the empirical formula C_5H_4. The molecular weight of $C_5H_4 = (5 \times 12.01) \times (4 \times 1.01) = 64.09$

The molecular weight was estimated to be 128/64.09 = 2.00. Multiply the empirical formula by 2 to get $C_{10}H_8$ (naphthalene).

Note: Significant figures were limited by the molecular weight of 128 obtained from the freezing point experiment, and would be 128.18 on the basis of the atomic weights given here. The molecular formula must be in whole numbers.

Problems



91. An iron-containing sample of 1.00 g is dissolved in acid solution and reduced. The reduced iron, Fe^{2+}, is titrated with 0.100 N (normal) $KMnO_4$ solution, and the titre to reach the endpoint is 100.0 ml. The balanced equation for the reaction is

$$10\ FeSO_4 + 2\ KMnO_4 + 8\ H_2SO_4$$
$$= 5\ Fe_2(SO_4)_3 + K_2SO_4 + 2\ MnSO_4 + 8\ H_2O$$

(a.w. Fe = 55.85, m.w. $FeSO_4$ = 151.92). Based on the weight of the original sample, what was the percent of iron in the sample?

a. 79.8
b. 27.9
c. 15.2
d. 55.8

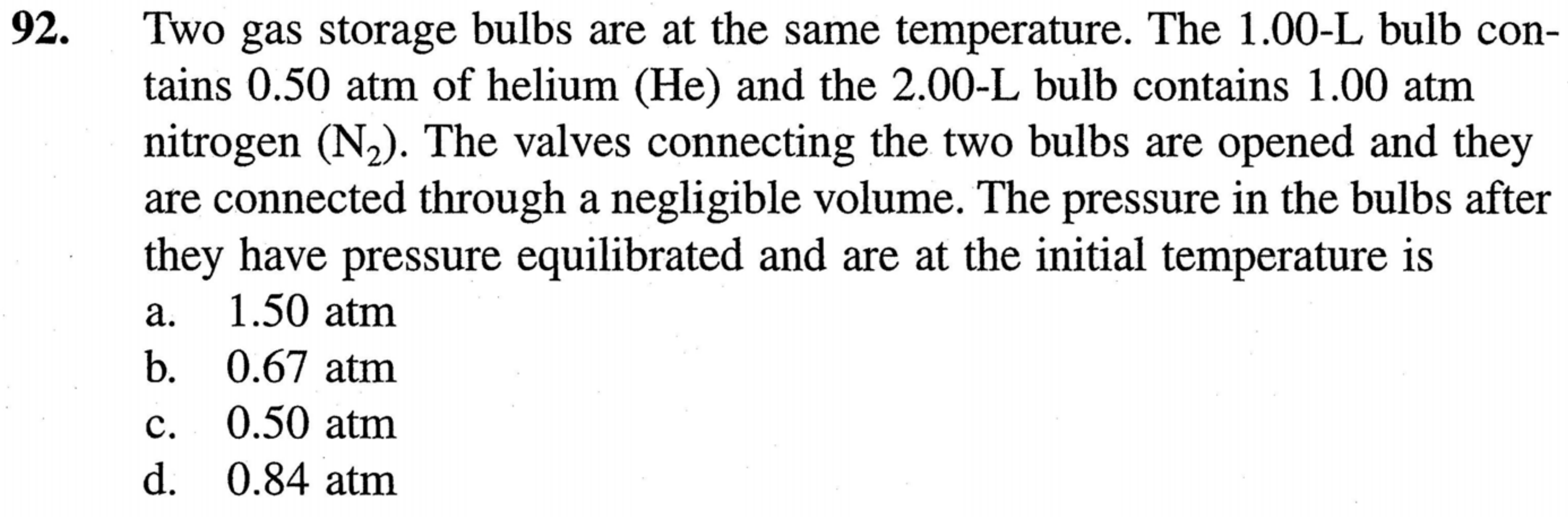
Problems

91. d. Since the normality and the volume of the $KMnO_4$ titre are given, the amount of equivalents used in the titration can be calculated from: N(normality) × L(volume) = equivalents 0.100 N × 0.1000 L = 0.0100 equivalents. The equivalents of Fe must be the same. Since Fe in the reaction is oxidized from Fe^{2+} to Fe^{3+}, there is only one mole of electrons transferred per mole of Fe, and the equivalent weight is the same as the formula weight. Then

$$\text{g Fe} = 0.01 \times 55.8 = .558; \qquad \text{g sample} = 1.00$$

$$\%\ \text{FE} = \frac{0.558}{1.00} \times 100 = 55.8\%$$

Problems



92. Two gas storage bulbs are at the same temperature. The 1.00-L bulb contains 0.50 atm of helium (He) and the 2.00-L bulb contains 1.00 atm nitrogen (N_2). The valves connecting the two bulbs are opened and they are connected through a negligible volume. The pressure in the bulbs after they have pressure equilibrated and are at the initial temperature is

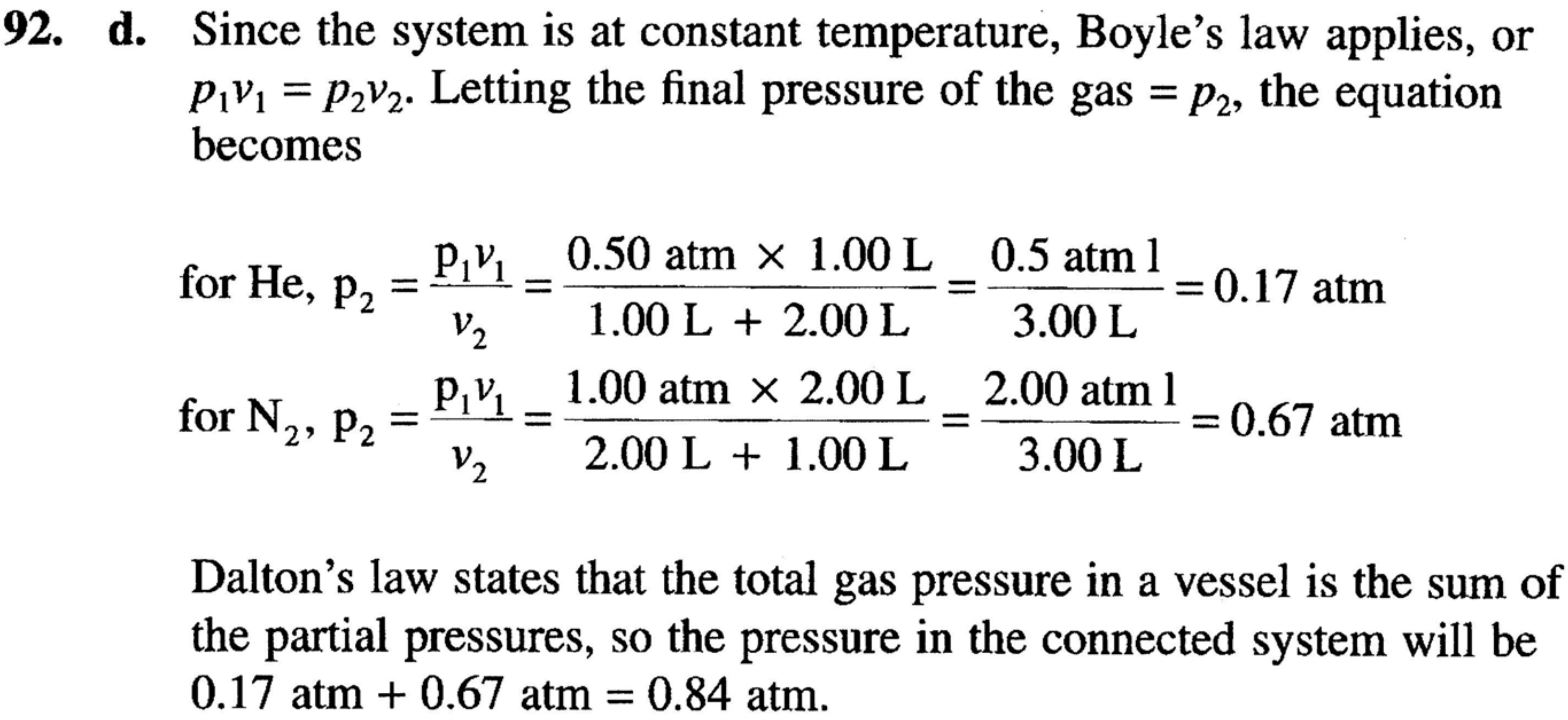
a. 1.50 atm

b. 0.67 atm

c. 0.50 atm

d. 0.84 atm

Problems



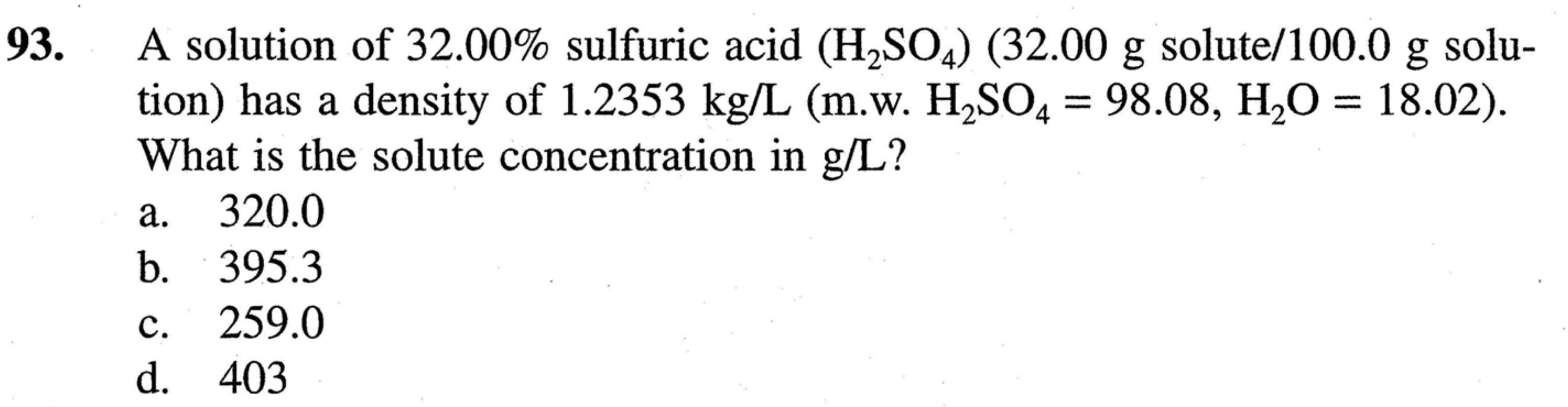
92. d. Since the system is at constant temperature, Boyle's law applies, or $p_1v_1 = p_2v_2$. Letting the final pressure of the gas = p_2, the equation becomes

$$\text{for He, } p_2 = \frac{p_1v_1}{v_2} = \frac{0.50\ \text{atm} \times 1.00\ \text{L}}{1.00\ \text{L} + 2.00\ \text{L}} = \frac{0.5\ \text{atm l}}{3.00\ \text{L}} = 0.17\ \text{atm}$$

$$\text{for } N_2,\ p_2 = \frac{p_1v_1}{v_2} = \frac{1.00\ \text{atm} \times 2.00\ \text{L}}{2.00\ \text{L} + 1.00\ \text{L}} = \frac{2.00\ \text{atm l}}{3.00\ \text{L}} = 0.67\ \text{atm}$$

Dalton's law states that the total gas pressure in a vessel is the sum of the partial pressures, so the pressure in the connected system will be 0.17 atm + 0.67 atm = 0.84 atm.

Problems



93. A solution of 32.00% sulfuric acid (H_2SO_4) (32.00 g solute/100.0 g solution) has a density of 1.2353 kg/L (m.w. H_2SO_4 = 98.08, H_2O = 18.02). What is the solute concentration in g/L?

a. 320.0

b. 395.3

c. 259.0

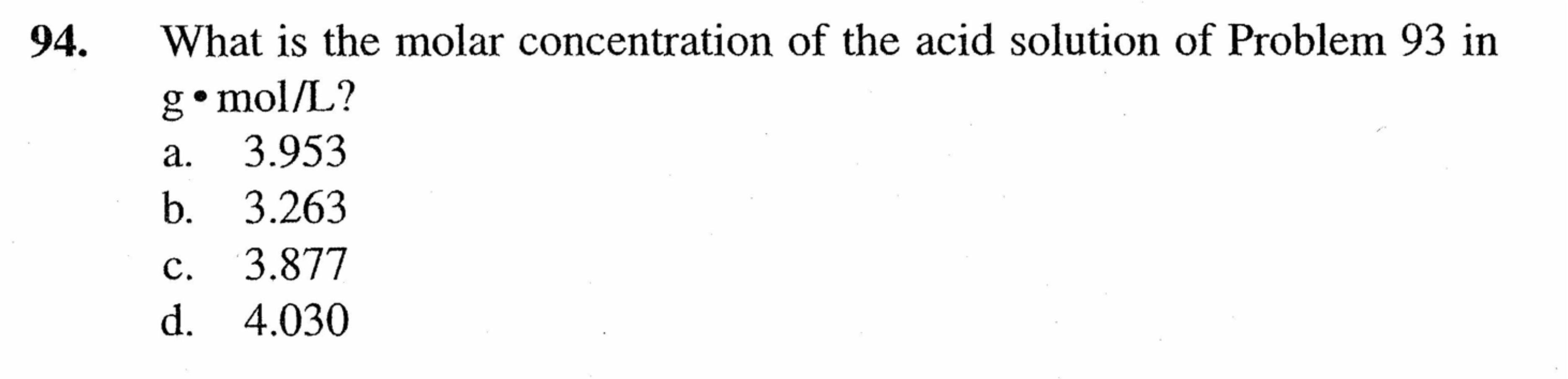
d. 403

Problems

93. **b.** The density in kg/L multiplied by the fractional weight of the solute gives the g/L of solute.

$$\frac{1.2353\text{ kg}}{\text{L}} = \frac{32.00\text{ g solute}}{100\text{ g solution}} = \frac{0.3953\text{ kg solute}}{\text{L solution}} \times \frac{1000\text{ g}}{\text{kg}} = \frac{395.3\text{ g solute}}{\text{L solution}}$$

Problems



94. What is the molar concentration of the acid solution of Problem 93 in g • mol/L?

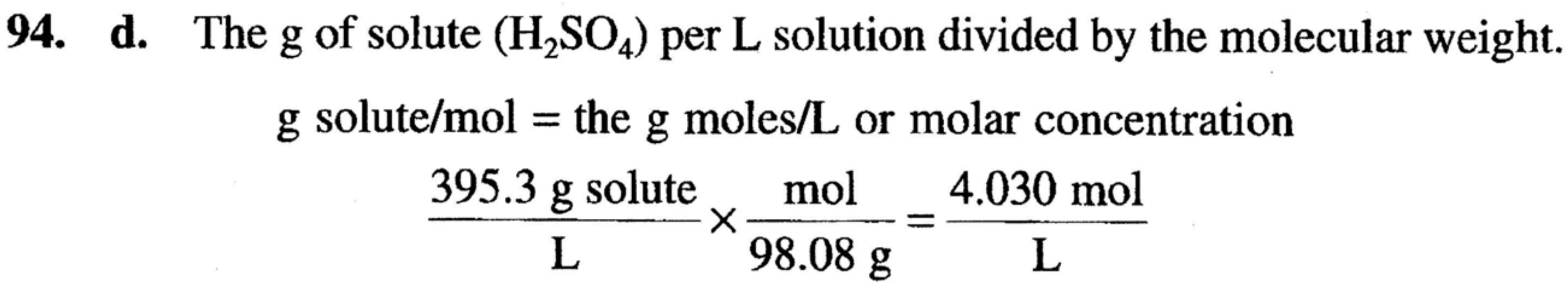
a. 3.953

b. 3.263

c. 3.877

d. 4.030

Problems



94. **d.** The g of solute (H_2SO_4) per L solution divided by the molecular weight.

g solute/mol = the g moles/L or molar concentration

$$\frac{395.3 \text{ g solute}}{\text{L}} \times \frac{\text{mol}}{98.08 \text{ g}} = \frac{4.030 \text{ mol}}{\text{L}}$$

Problems

End of Problems Section

Problems

Problems