

Chapter 1: Matter and Measurements

22. Refer to Section 1.2.

- (a) $4020.6 \text{ mL} = 4.0206 \times 10^3 \text{ mL}$
- (b) 1.006 g (This is already in proper scientific notation.)
- (c) $100.1^\circ\text{C} = 1.001 \times 10^2^\circ\text{C}$

32. Refer to Section 1.2 and Table 1.2.

Convert one of the numbers to the units of the other. Once the numbers are expressed in common units, they can be compared directly.

(a) $37.12 \text{ g} \times \frac{1 \text{ kg}}{1000 \text{ g}} = 0.03712 \text{ kg}$, thus: $37.12 \text{ g} < 0.03712 \text{ kg}$.

(b) $28 \text{ m}^3 \times \left(\frac{100 \text{ cm}}{1 \text{ m}}\right)^3 = 28 \times 10^6 \text{ cm}^3$, thus: $28 \text{ m}^3 > 28 \times 10^2 \text{ cm}^3$.

(c) $525 \text{ mm} \times \frac{1 \text{ m}}{1000 \text{ mm}} \times \frac{1 \text{ nm}}{1 \times 10^{-9} \text{ m}} = 525 \times 10^6 \text{ nm}$,

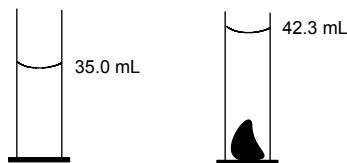
thus $525 \text{ mm} = 525 \times 10^6 \text{ nm}$

48. Refer to Section 1.3.

Calculate the volume of the object from the change in the volume of the graduated cylinder, then calculate the density.

Volume of object = $42.3 \text{ mL} - 35.0 \text{ mL} = 7.3 \text{ mL}$

density = $\frac{\text{mass (g)}}{\text{volume (mL)}} = \frac{11.33 \text{ g}}{7.3 \text{ mL}} = 1.6 \text{ g/mL}$



Chapter 2: Atoms, Molecules, and Ions

16. Refer to Section 2.2.

- (a) Na-21 has 11 protons and $21 - 11 = 10$ neutrons. Adding one neutron makes the number 22, thus the symbol is ${}^{22}_{11}\text{Na}$.
- (b) An isobar of Na-21 with 10 protons would have an atomic number of 10, thus the element symbol would be Ne. The nuclear symbol is ${}^{21}_{10}\text{Ne}$.
- (c) A nucleus with 11 protons is sodium. The mass number is $11 + 12 = 23$. The nuclear symbol is ${}^{23}_{11}\text{Na}$. This is indeed an isotope of Na-21.

38. Refer to Section 2.5.

The sum of the charges must equal zero; if not, add extras of the needed ion to “balance” the charges.

- (a) Ba^{2+} and I^- give BaI_2 ($+2 + 2(-1) = 0$)
 Ba^{2+} and N^{3-} give Ba_3N_2 ($3(+2) + 2(-3) = 0$)
- (b) O^{2-} and Fe^{2+} give FeO ($-2 + (+2) = 0$)
 O^{2-} and Fe^{3+} give Fe_2O_3 ($3(-2) + 2(+3) = 0$)

52. Refer to Sections 2.4 and 2.5.

- (a) “Compounds containing carbon atoms are molecular” is **usually true**. Carbon, being a nonmetal can combine with other non-metals to form molecular compounds such as those mentioned in Section 2.4 (indeed, a whole field of chemistry, called organic chemistry, is dedicated to the study of such compounds). Carbon can, however, also form ionic compounds such as calcium carbide (used in old miners lamps).
- (b) “A molecule is made up of nonmetal atoms” is **always true**, by definition.
- (c) “An ionic compound has at least one metal atom” is **usually true**. Most ionic compounds do contain a metal, but there are ionic compounds in which the cation (positive ion) is not a metal, such as in ammonium chloride, NH_4Cl .

Chapter 3: Mass Relations in Chemistry; Stoichiometry

26. Refer to Section 3.2 and Example 3.3.

Multiply the number of each atom in the molecule by the atomic mass of that atom, and then add up the masses to get the molar mass of the molecule.

(a) Ga: $1(69.723) = 69.723 \text{ g/mol}$.

(b) $\text{CaSO}_4 \cdot \frac{1}{2} \text{H}_2\text{O}$: $1(40.078) + 1(32.066) + 4(15.9994) + \frac{1}{2} [2(1.00794) + 1(15.9994)]$
 $= 136.1416 + \frac{1}{2} (18.01528) = 145.149 \text{ g/mol}$.

(c) $\text{C}_{14}\text{H}_{10}\text{O}_4$: $14(12.011) + 10(1.0079) + 4(15.9994)$
 $= 168.154 + 10.0794 + 63.9976 = 242.231 \text{ g/mol}$.

38. Refer to Section 3.3.

Mass of 1 mol. $\text{C}_7\text{H}_5\text{BiO}_4 = 7(12.01) + 5(1.008) + 209.0 + 4(16.00) = 362.1 \text{ g}$

$$346 \text{ mg Bi} \times \frac{1 \text{ g}}{1000 \text{ mg}} \times \frac{1 \text{ mol. C}_7\text{H}_5\text{BiO}_4}{209.0 \text{ g Bi}} \times \frac{362.1 \text{ g}}{1 \text{ mol. C}_7\text{H}_5\text{BiO}_4} = 0.599 \text{ g C}_7\text{H}_5\text{BiO}_4$$

$$\frac{0.599 \text{ g C}_7\text{H}_5\text{BiO}_4}{1.500 \text{ g Pepto - Bismal}} \times 100\% = 39.9\%$$

49. Refer to Section 3.3.

$$\text{mass C} = 8.692 \text{ g CO}_2 \times \frac{12.01 \text{ g C}}{44.01 \text{ g CO}_2} = 2.372 \text{ g}$$

$$\text{mass H} = 1.142 \text{ g H}_2\text{O} \times \frac{2(1.008) \text{ g H}}{18.02 \text{ g H}_2\text{O}} = 0.1278 \text{ g}$$

$$\text{mass Cl} = 2.571 \text{ g HCl} \times \frac{35.45 \text{ g Cl}}{36.46 \text{ g HCl}} = 2.500 \text{ g}$$

$$\text{mol C} = \frac{2.372 \text{ g}}{12.01 \text{ g/mol}} = 0.1975$$

$$\text{mol H} = \frac{0.1278 \text{ g}}{1.008 \text{ g/mol}} = 0.1268$$

$$\text{mol Cl} = \frac{2.500 \text{ g}}{35.45 \text{ g/mol}} = 0.07052$$

$$1 \text{ Cl} : 1.8 \text{ H} : 2.8 \text{ C}$$

$$\text{Hence, C}_{14}\text{H}_9\text{Cl}_5$$

68. Refer to Section 3.4 and Example 3.9.

(a) First, the mass of the sample of tin, Sn, must be determined. Note: the thickness is given in units of millimeters and must first be converted into cm.

$$0.600 \text{ mm} \times \frac{\text{cm}}{10 \text{ mm}} = 0.0600 \text{ cm}$$

$$(8.25 \text{ cm} \times 21.5 \text{ cm} \times 0.0600 \text{ cm}) \times \frac{7.28 \text{ g Sn}}{\text{cm}^3} = 77.5 \text{ g Sn}$$

$$77.5 \text{ g Sn} \times \frac{1 \text{ mol. Sn}}{118.71 \text{ g Sn}} \times \frac{1 \text{ mol. SnO}_2}{1 \text{ mol. Sn}} \times \frac{150.71 \text{ g SnO}_2}{1 \text{ mol. SnO}_2} = 98.4 \text{ g SnO}_2$$

$$(b) 98.4 \text{ g SnO}_2 \times \frac{31.999 \text{ g O}_2}{150.71 \text{ g SnO}_2} \times \frac{1 \text{ L O}_2}{1.309 \text{ g O}_2} \times \frac{100 \text{ L air}}{21 \text{ L O}_2} = 76 \text{ L air}$$

Chapter 4: Reactions in Aqueous Solution

4. Refer to Section 4.1.

$$(a) M = \frac{\text{mol}}{\text{L}} = \frac{15.00 \text{ g ScI}_3 \left(\frac{1 \text{ mol}}{425.7 \text{ g}} \right)}{0.655 \text{ L}} = 0.0538 \text{ M}$$

$$\begin{aligned} 0.0538 \text{ M Sc}^{3+} (\text{aq}) \\ 0.0538 \text{ M} * 3 = 0.161 \text{ M I}^{-} (\text{aq}) \end{aligned}$$

$$(b) M = \frac{\text{mol}}{\text{L}} = \frac{15.00 \text{ g Na}_2\text{CO}_3 \left(\frac{1 \text{ mol}}{106.0 \text{ g}} \right)}{0.655 \text{ L}} = 0.216 \text{ M}$$

$$\begin{aligned} 0.216 \text{ M Na}^{+} (\text{aq}) \\ 0.216 \text{ M CO}_3^{2-} (\text{aq}) \end{aligned}$$

$$(c) M = \frac{\text{mol}}{\text{L}} = \frac{15.00 \text{ g Mg}_3(\text{PO}_4)_2 \left(\frac{1 \text{ mol}}{262.84 \text{ g}} \right)}{0.655 \text{ L}} = 0.0871 \text{ M}$$

$$\begin{aligned} 0.0871 \text{ M} * 3 = 0.261 \text{ M Mg}^{2+} (\text{aq}) \\ 0.0871 \text{ M} * 2 = 0.174 \text{ M PO}_4^{3-} (\text{aq}) \end{aligned}$$

$$(d) M = \frac{\text{mol}}{\text{L}} = \frac{15.00 \text{ g K}_2\text{O} \left(\frac{1 \text{ mol}}{94.20 \text{ g}} \right)}{0.655 \text{ L}} = 0.243 \text{ M}$$

$$\begin{aligned} 0.243 \text{ M} * 2 = 0.486 \text{ M K}^{+} (\text{aq}) \\ 0.243 \text{ M O}^{2-} (\text{aq}) \end{aligned}$$

18. Refer to Sections 2.6, 4.1, and 4.2, Figures 4.3 and 4.4, and Example 4.4.

Recall that soluble salts ionize when dissolved. Write the reactions for the ionizations. Look at the resulting ions. If there are pairs that would result in insoluble salts, these salts would form and precipitate from solution.

- (a) Ions present: Cu^{2+} , SO_4^{2-} , Na^+ , Cl^-
Possible precipitates: CuCl_2 , Na_2SO_4
Both compounds are soluble, thus no precipitate forms.
- (b) Ions present: Mn^{2+} , NO_3^- , Na^+ , OH^-
Possible precipitates: $\text{Mn}(\text{OH})_2$, NaNO_3
 NaNO_3 is soluble, but $\text{Mn}(\text{OH})_2$ is not.
 $\text{Mn}^{2+}(\text{aq}) + 2\text{OH}^-(\text{aq}) \rightarrow \text{Mn}(\text{OH})_2(\text{s})$
- (c) Ions present: Ag^+ , NO_3^- , H^+ , Cl^-
Possible precipitates: AgCl , HNO_3
 HNO_3 is soluble, but AgCl is not.
 $\text{Ag}^+(\text{aq}) + \text{Cl}^-(\text{aq}) \rightarrow \text{AgCl}(\text{s})$
- (d) Ions present: Co^{2+} , SO_4^{2-} , Ba^{2+} , OH^-
Possible precipitates: $\text{Co}(\text{OH})_2$, BaSO_4
Both compounds are insoluble.
 $\text{Co}^{2+}(\text{aq}) + 2\text{OH}^-(\text{aq}) \rightarrow \text{Co}(\text{OH})_2(\text{s})$
 $\text{Ba}^{2+}(\text{aq}) + \text{SO}_4^{2-}(\text{aq}) \rightarrow \text{BaSO}_4(\text{s})$
- (e) Ions present: NH_4^+ , K^+ , OH^- , CO_3^{2-}
Possible precipitates: NH_4OH , K_2CO_3
Both compounds are soluble, thus no precipitate forms.

36. Refer to Sections 2.6 and 4.3 and Table 4.2.

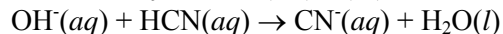
For the net ionic equation $\text{OH}^-(\text{aq}) + \text{HB}(\text{aq}) \rightarrow \text{B}^-(\text{aq}) + \text{H}_2\text{O}(\text{l})$ to be correct, the reactants must be a weak acid and a strong base.

- (a) The equation is **not** correct.
Hydrochloric acid (HCl) is a strong acid, the reacting species is H^+ .
Pyridine ($\text{C}_5\text{H}_5\text{N}$) is a weak base, the reacting species is $\text{C}_5\text{H}_5\text{N}$.
 $\text{H}^+(\text{aq}) + \text{C}_5\text{H}_5\text{N}(\text{aq}) \rightarrow \text{C}_5\text{H}_5\text{NH}^+(\text{aq})$
- (b) The equation is **not** correct.
Sulfuric acid (H_2SO_4) is a strong acid, the reacting species is H^+ .
Rubidium hydroxide (RbOH) is a strong base, the reacting species is OH^- .
 $\text{H}^+(\text{aq}) + \text{OH}^-(\text{aq}) \rightarrow \text{H}_2\text{O}(\text{l})$
- (c) The equation **is** correct.
Hydrofluoric acid (HF) is a weak acid, the reacting species is HF .
Potassium hydroxide (KOH) is a strong base, the reacting species is OH^- .
 $\text{OH}^-(\text{aq}) + \text{HF}(\text{aq}) \rightarrow \text{F}^-(\text{aq}) + \text{H}_2\text{O}(\text{l})$
- (d) The equation is **not** correct.
Hydroiodic acid (HI) is a strong acid, the reacting species is H^+ .
Ammonia (NH_3) is a weak base, the reacting species is NH_3 .
 $\text{H}^+(\text{aq}) + \text{NH}_3(\text{aq}) \rightarrow \text{NH}_4^+(\text{aq})$

(e) The equation **is** correct.

Hydrocyanic acid (HCN) is a weak acid, the reacting species is HCN.

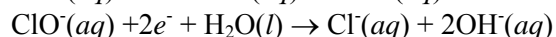
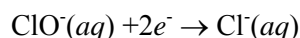
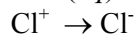
Strontium hydroxide ($\text{Sr}(\text{OH})_2$) is a strong base, the reacting species is OH^- .



56. Refer to Section 4.4 and Example 4.9.

Begin by determining the oxidation numbers of the element being reduced or oxidized. Balance the element being oxidized or reduced and then balance the oxidation number by adding electrons. Balance the charge by adding H^+ (if acidic) or OH^- (if basic), and H_2O to balance H and O.

(a) $\text{ClO}^-(aq) \rightarrow \text{Cl}^-(aq)$ [basic]



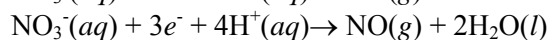
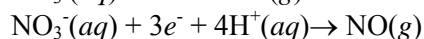
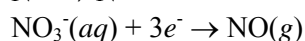
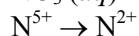
Reduction

Cl and electrons balanced

charges balanced

O and H balanced

(b) $\text{NO}_3^-(aq) \rightarrow \text{NO}(g)$ [acidic]



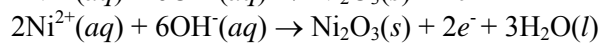
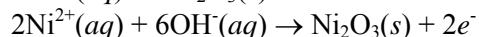
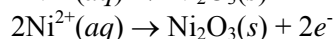
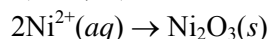
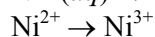
Reduction

N and electrons balanced

charges balanced

O and H balanced

(c) $\text{Ni}^{2+}(aq) \rightarrow \text{Ni}_2\text{O}_3(s)$ [basic]



Oxidation

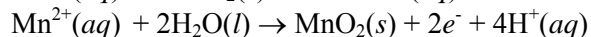
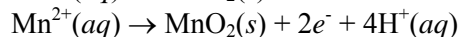
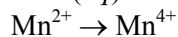
Ni balanced

electrons balanced

charges balanced

O and H balanced

(d) $\text{Mn}^{2+}(aq) \rightarrow \text{MnO}_2(s)$ [acidic]



Oxidation

Mn and electrons balanced

charges balanced

O and H balanced

Chapter 5: Gases

12. Refer to Section 5.1 and Example 5.3.

$$V_1 = 25.0 \text{ mL} = 0.0250 \text{ L}$$

$$T_1 = 23^\circ\text{C} = 23 + 273 = 296 \text{ K}$$

$$P_1 = 745 \text{ mm Hg} \times \frac{1 \text{ atm}}{760.0 \text{ mm Hg}} = 0.980 \text{ atm}$$

Since volume increased by 8%, multiply V_1 by 1.08 to get V_2 .

$$V_2 = (0.0250 \text{ L})(1.08) = 0.0270 \text{ L}$$

$$T_2 = 82^\circ\text{C} = 82 + 273 = 355 \text{ K}$$

$$P_2 = ?$$

$$\frac{P_1 V_1}{T_1} = \frac{P_2 V_2}{T_2}$$

$$\frac{(0.980 \text{ atm})(0.0250 \text{ L})}{(296 \text{ K})} = \frac{(P_2)(0.0270 \text{ L})}{355 \text{ K}}$$

$$P_2 = 1.09 \text{ atm} = 828 \text{ mm Hg}$$

Yes, the pressure increased.

20. Refer to Section 5.3 and Example 5.3.

Start by converting all units to those used in the ideal gas law. Calculated moles using the ideal gas equation if P , V and T are given, otherwise use the molar mass and the given mass. Then apply the ideal gas law to fill in any other missing data.

Pressure	Volume	Temperature	Moles	Grams
22.7 atm	1.75 L	19°C	1.66	96.5
0.895 atm	6.17 L	6°C	0.241	14.0
433 mm Hg	92.4 mL	1.62 K	0.395	23.0
1.74 bar	8.66 L	98°F	0.585	34.0

$$T = 19^\circ\text{C} + 273 = 292 \text{ K}$$

$$P = \frac{nRT}{V} = \frac{(1.66 \text{ mol.})(0.0821 \text{ L} \cdot \text{atm/mol} \cdot \text{K})(292 \text{ K})}{(1.75 \text{ L})} = 22.7 \text{ atm}$$

$$1.66 \text{ mol. C}_4\text{H}_{10} \times \frac{58.12 \text{ g}}{1 \text{ mol.}} = 96.5 \text{ g}$$

$$T = 6^\circ\text{C} + 273 = 279 \text{ K}$$

$$14.0 \text{ g} \times \frac{1 \text{ mol.}}{58.12 \text{ g}} = 0.241 \text{ mol.}$$

$$V = \frac{nRT}{P} = \frac{(0.241 \text{ mol.})(0.0821 \text{ L} \cdot \text{atm/mol} \cdot \text{K})(279 \text{ K})}{(0.895 \text{ atm})} = 6.17 \text{ L}$$

$$433 \text{ mm Hg} \times \frac{1 \text{ atm}}{760 \text{ mm Hg}} = 0.570 \text{ atm}$$

$$T = \frac{PV}{nR} = \frac{(0.570 \text{ atm})(0.0924 \text{ L})}{(0.0821 \text{ L} \cdot \text{atm/mol} \cdot \text{K})(0.395 \text{ mol.})} = 1.62 \text{ K}$$

$$1.74 \text{ bar} \times \frac{1 \text{ atm}}{1.013 \text{ bar}} = 1.72 \text{ atm}$$

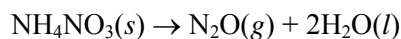
$$T = \frac{98^\circ\text{F} - 32}{1.8} + 273 = 310 \text{ K}$$

$$n = \frac{PV}{RT} = \frac{(1.72 \text{ atm})(8.66 \text{ L})}{(0.0821 \text{ L} \cdot \text{atm/mol} \cdot \text{K})(310 \text{ K})} = 0.585 \text{ mol.}$$

$$0.585 \text{ mol.} \times \frac{58.12 \text{ g}}{1 \text{ mol.}} = 34.0 \text{ g}$$

36. Refer to Section 5.4 and Example 5.5.

(a) This equation can be balanced by inspection.



(b) $T = 250^\circ\text{C} + 273 = 523 \text{ K}$

$$V = \frac{nRT}{P} = \frac{(0.0625 \text{ mol.})(0.0821 \text{ L} \cdot \text{atm/mol} \cdot \text{K})(298 \text{ K})}{1.0 \text{ atm}} = 2.7 \text{ L}$$

52. Refer to Section 5.6.

(a) $P(\text{H}_2) = P(\text{Total}) - P(\text{H}_2\text{O}) = 748 \text{ mm Hg} - 23.8 \text{ mm Hg} = 724.2 \text{ mm Hg}$

$$(b) \ n = \frac{PV}{RT} = \frac{\left[\frac{23.8 \text{ mm Hg}}{760 \text{ mm Hg}} \right] (0.250 \text{ L})}{(0.08216 \frac{\text{L} \cdot \text{atm}}{\text{mol} \cdot \text{K}})(298 \text{ K})} = 0.000320 \text{ moles H}_2\text{O}$$

$$(c) \ n = \frac{PV}{RT} = \frac{\left[\frac{724.2 \text{ mm Hg}}{760 \text{ mm Hg}} \right] (0.250 \text{ L})}{(0.08216 \frac{\text{L} \cdot \text{atm}}{\text{mol} \cdot \text{K}})(298 \text{ K})} = 0.00973 \text{ moles H}_2$$

$$(d) \ 0.0186 \text{ g He} \left(\frac{1 \text{ mol}}{4.0026 \text{ g}} \right) = 0.00465 \text{ mol He are added to the flask}$$

New Total Pressure:

$$P = \frac{nRT}{V}$$

$$P = \frac{(0.000320 + 0.00973 + 0.00465 \text{ moles})(0.08216 \frac{\text{L} \cdot \text{atm}}{\text{mol} \cdot \text{K}})(298 \text{ K})}{0.250 \text{ L}}$$

$$P = 1.44 \text{ atm or } 1094 \text{ mm Hg}$$

$$P(\text{He}) = X(\text{He}) \cdot P(\text{Total}) =$$

$$\left(\frac{(0.00465 \text{ moles})}{(0.000320 + 0.00973 + 0.00465 \text{ moles})} \right) (1094 \text{ mm Hg})$$

$$= 346 \text{ mm Hg}$$

$$(e) \ P(\text{Total}) = 1.44 \text{ atm or } 1094 \text{ mm Hg (from above)}$$

Chapter 6: Electronic Structure and the Periodic Table

8. Refer to Section 6.1 and Example 6.2.

Calculate the amount of energy in one photon. Use that and the total amount of energy to calculate the number of photons emitted.

$$\lambda = 633 \text{ nm} \times \frac{1 \text{ m}}{1 \times 10^9 \text{ nm}} = 6.33 \times 10^{-7} \text{ m}$$

$$E = \frac{hc}{\lambda} = \frac{(6.626 \times 10^{-34} \text{ J} \cdot \text{s})(2.998 \times 10^8 \text{ m/s})}{6.33 \times 10^{-7} \text{ m}} = 3.14 \times 10^{-19} \text{ J}$$

$$12 \text{ kJ} \times \frac{1000 \text{ J}}{1 \text{ kJ}} \times \frac{1 \text{ photon}}{3.14 \times 10^{-19} \text{ J}} = 3.8 \times 10^{22} \text{ photons}$$

22. Refer to Section 6.3.

The type of orbital is determined by the ℓ value.

- (a) $\ell = 1$, therefore: p-orbital.
- (b) $\ell = 0$, therefore: s-orbital.
- (c) $\ell = 2$, therefore: d-orbital

48. Refer to Section 6.6.

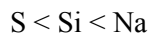
The only main group metals in the 4th period are K, Ca and Ga, the remainder of the elements are transition metals, nonmetals or metalloids.

	4s	4p
K	(↑)	() () ()
Ca	(↑↓)	() () ()
Ga	(↑↓)	(↑) () ()

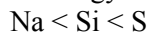
- (a) Ca
- (b) K and Ga
- (c) none
- (d) none

54. Refer to Section 6.8, Figures 6.13 and 6.15, and Example 6.11.

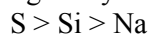
(a) Atomic radius increases from right to left across a period, therefore:



(b) Ionization energy increases from left to right across a period, therefore:



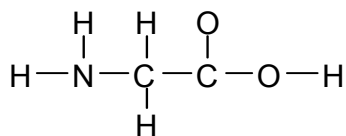
(c) Electronegativity decreases from right to left across a period, therefore:



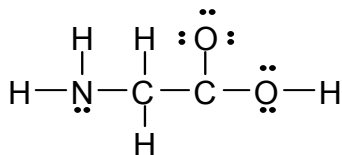
Chapter 7: Covalent Bonding

12. Refer to Section 7.1 and Example 7.1.

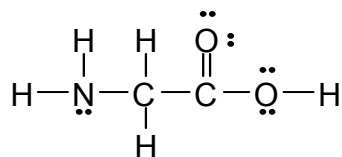
2C: 2 x 4 valence electrons
 5H: 5 x 1 valence electrons
 2O: 2 x 6 valence electrons
 N: 5 valence electrons
 total: 30 electrons



From the information given, the structure must be that on the left.



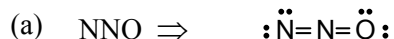
Note that this structure does not have a complete octet around one of the carbon atoms. To complete the octet, move a pair of electrons from the terminal oxygen to form a C=O double bond.



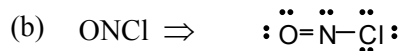
This structure provides octets for all the atoms and provides that all formal charges are zero.

34. Refer to Section 7.2, Examples 7.5 and 7.6, Figures 7.4 and 7.5, and Table 7.3.

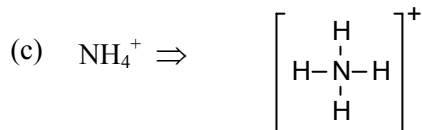
Draw the Lewis structure of the compound and determine the number of bonded groups and the number of electron pairs. Then use Table 7.3 to assign the geometry.



2 bonded groups, no electron pairs, thus AX_2 and **linear**.



2 bonded groups, one electron pair, thus AX_2E and **bent**.



4 bonded groups, no electron pairs, thus AX₄ and **tetrahedron**.



2 bonded groups, one electron pair, thus AX₂E and **bent**.

52. Refer to Section 7.4, Table 7.4, Example 7.9, and Problem 34 (above).

Recall that the total number of groups (bonded atoms and electron pairs) around the central atom is equal to the number of orbitals that hybridized. Furthermore, the sum of the superscripts in the hybrid orbital notation gives the total number of hybrid orbitals.

- | | | | | |
|-----|------------------------------|-------------------|----------|-----------------|
| (a) | NNO | AX ₂ | 2 groups | sp |
| (b) | ONCl | AX ₂ E | 3 groups | sp ² |
| (c) | NH ₄ ⁺ | AX ₄ | 4 groups | sp ³ |
| (d) | O ₃ | AX ₂ E | 3 groups | sp ² |

Chapter 8: Thermochemistry

8. Refer to Problem 1.16, Section 8.1, and Example 8.1.

- (a) In the solution process, heat is absorbed. Thus, it is **not** exothermic, it is endothermic.
(b) Heat is absorbed from the water. Thus, heat exits the water and the process for the water is exothermic. Thus, $q_{\text{H}_2\text{O}} = -27.6 \text{ kJ}$.

$$\begin{aligned} \text{(c)} \quad 250.0 \text{ mL} \times \frac{1.00 \text{ g}}{1 \text{ mL}} &= 250.0 \text{ g} \\ -27.6 \text{ kJ} \times \frac{1000 \text{ J}}{1 \text{ kJ}} &= -27600 \text{ J} \end{aligned}$$

$$q = mc\Delta t$$

$$-27600 \text{ J} = (250.0 \text{ g})(4.18 \text{ J/g}^\circ\text{C}) \Delta t$$

$$\Delta t = -26.4^\circ\text{C}$$

$$\Delta t = t_f - t_i$$

$$-26.4^\circ\text{C} = t_f - 30.0^\circ\text{C}$$

$$t_f = 3.6^\circ\text{C}$$

$$\text{(d)} \quad t_{\text{F}} = 1.8t_{\text{C}} + 32$$

$$t_i = 30.0^\circ\text{C}$$

$$t_{\text{F}} = 1.8(30.0^\circ\text{C}) + 32$$

$$t_i = 86.0^\circ\text{F}$$

$$t_f = 3.6^\circ\text{C}$$

$$t_{\text{F}} = 1.8(3.6^\circ\text{C}) + 32$$

$$t_f = 38.5^\circ\text{F}$$

16. Refer to Section 8.3.

Calculate q_{reaction} for 5.00 g caffeine using 1 mol. caffeine per $4.96 \times 10^3 \text{ kJ}$ as a conversion factor. Then use q_{reaction} to calculate C_{cal} .

Since heat is evolved, q is negative. $q_{\text{reaction}} = -4.96 \times 10^3 \text{ kJ}$

$$5.00 \text{ g caffeine} \times \frac{1 \text{ mol. caffeine}}{194.20 \text{ g caffeine}} \times \frac{-4.96 \times 10^3 \text{ kJ}}{1 \text{ mol. caffeine}} = -128 \text{ kJ}$$

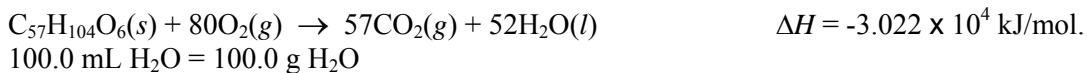
$$q_{\text{reaction}} = -C_{\text{cal}} \times \Delta t$$

$$-128 \text{ kJ} = -C_{\text{cal}} \times 11.37^\circ\text{C}$$

$$C_{\text{cal}} = 11.2 \text{ kJ/}^\circ\text{C} = 1.12 \times 10^4 \text{ J/}^\circ\text{C}$$

28. Refer to Sections 8.1 - 8.4.

Calculate the amount of heat needed to raise the temperature of the water 3°C. Recall that the heat absorbed by the water will be equal to the heat released by the fat. Finally, use the enthalpy of the reaction to calculate the amount of fat needed to produce the required heat.



$$q_{\text{H}_2\text{O}} = (100.0 \text{ g})(4.18 \text{ J/g}^\circ\text{C})(25.00^\circ\text{C} - 22.00^\circ\text{C}) = 1.25 \times 10^3 \text{ J}$$

$$q_{\text{fat}} = -q_{\text{H}_2\text{O}} = -1.25 \times 10^3 \text{ J}$$

$$-1.25 \times 10^3 \text{ J} \times \frac{1 \text{ kJ}}{1000 \text{ J}} \times \frac{1 \text{ mol. fat}}{-3.022 \times 10^4 \text{ kJ}} \times \frac{885.4 \text{ g fat}}{1 \text{ mol. fat}} = 3.67 \times 10^{-2} \text{ g fat}$$

Chapter 9: Liquids and Solids

8. Refer to Section 9.1 and Example 9.2.

$$(a) \ln\left(\frac{P_2}{P_1}\right) = \frac{+\Delta H_{\text{vap}}}{R} \left(\frac{1}{T_1} - \frac{1}{T_2}\right)$$

$$\ln\left(\frac{760 \text{ atm}}{203 \text{ atm}}\right) = \frac{+\Delta H_{\text{vap}}}{8.31 \text{ J/mol}\cdot\text{K}} \left(\frac{1}{308 \text{ K}} - \frac{1}{337.8 \text{ K}}\right)$$

$$1.32 = \frac{+\Delta H_{\text{vap}}}{8.31 \text{ J/mol}\cdot\text{K}} (2.86 \times 10^{-4} \text{ K}^{-1})$$

$$\Delta H_{\text{vap}} = 3.84 \times 10^4 \text{ J/mol.}$$

$$\Delta H_{\text{vap}} = 38.4 \text{ kJ/mol.}$$

$$(b) \ln\left(\frac{P_2}{P_1}\right) = \frac{+\Delta H_{\text{vap}}}{R} \left(\frac{1}{T_1} - \frac{1}{T_2}\right)$$

$$\ln\left(\frac{760 \text{ atm}}{P_1}\right) = \frac{3.84 \times 10^4 \text{ J/mol.}}{8.31 \text{ J/mol}\cdot\text{K}} \left(\frac{1}{313.2 \text{ K}} - \frac{1}{337.8 \text{ K}}\right)$$

$$\ln\left(\frac{760 \text{ atm}}{P_1}\right) = 4.62 \times 10^3 \text{ K} (2.32 \times 10^{-4} \text{ K}^{-1})$$

$$\ln(760 \text{ atm}) - \ln P_1 = 1.07$$

$$5.56 = \ln P_1$$

$$P_1 = 260 \text{ mm Hg}$$

28. Refer to Section 9.3, Examples 9.4 and 9.6, and Chapter 7.

All molecules have dispersion forces. One can consider each molecule to be roughly spherical. Those that have an even distribution of charge over the surface (b and c) will not have a dipole, while those that have an uneven distribution (a and d) will have a dipole.

(a) PH_3 : **Dispersion and dipole.** The electrons are not symmetrically distributed, so this molecule has a dipole. Each H bears a slight negative charge and the P bears a slight positive charge. Recall from Chapter 7 that PH_3 is trigonal pyramid in shape.

(b) N_2 : **Dispersion.** The electrons are symmetrically distributed, so this molecule does not have a dipole. Each N has identical, zero, charge.

- (c) CH₄: **Dispersion.** The electrons are symmetrically distributed, so this molecule does not have a dipole. Each H bears an identical, slight positive charge, thus, the surface of the "sphere" has an even distribution of charge.
- (d) H₂O: **Dispersion and dipole.** The electrons are not symmetrically distributed, so this molecule has a dipole. The O bears a slight negative charge, while the H's each bear a slight positive charge. Recall from Chapter 7 that water is bent.

48. Refer to Section 9.4.

- (a) Graphite is a network covalent compound, so the C atoms are held together in an extended network of covalent bonds.
- (b) Silicon carbide is a network covalent compound, so the Si and C atoms are held together in an extended network of covalent bonds.
- (c) FeCl₂ is an ionic solid. The structural units are composed of individual Fe²⁺ and Cl⁻ ions.
- (d) Acetylene (C₂H₂) is a molecular solid, composed of individual C₂H₂ molecules held together by dispersion forces.

Chapter 10: Solutions

10. Refer to Section 10.1 and Examples 10.4 and 10.5.

	Molality	Mass Percent of Solvent	Ppm Solute	Mole Fraction of Solvent
(a)	2.577	86.58%	1.340 x 10⁵	0.9556
(b)	20.4	45.0	5.50 x 10⁵	0.731
(c)	0.07977	99.5232%	4768	0.9986
(d)	12.6	57.0%	4.30 x 10⁵	0.815

Since these are aqueous solutions, the solvent is water. The solute is urea, CO(NH₂)₂.
Molar mass of urea: (12.01) + 16.00 + 2(14.01 + 2(1.008)) = 60.06 g/mol.

- (a) 2.577 *m* indicates that there are 2.577 mol. of solute per 1000 g of solvent and the calculations are based on that ratio.

$$2.577 \text{ mol.} \times \frac{60.06 \text{ g urea}}{1 \text{ mol. urea}} = 154.8 \text{ g urea}$$

$$1 \text{ kg} \times \frac{1000 \text{ g}}{1 \text{ kg}} \times \frac{1 \text{ mol. H}_2\text{O}}{18.02 \text{ g H}_2\text{O}} = 55.49 \text{ mol. H}_2\text{O}$$

$$\text{mass \%} = \frac{1000 \text{ g}}{(1000 \text{ g} + 154.8 \text{ g})} \times 100\% = \frac{1000 \text{ g}}{1155 \text{ g}} = \mathbf{86.58\%}$$

$$\text{ppm solute} = \frac{154.8 \text{ g}}{(1000 \text{ g} + 154.8 \text{ g})} \times 10^6 = \mathbf{1.340 \times 10^5}$$

$$X_{\text{H}_2\text{O}} = \frac{55.49 \text{ mol.}}{(55.49 \text{ mol.} + 2.577 \text{ mol.})} = \mathbf{0.9556}$$

- (b) 45.0 mass % of solvent means that 45.0 g of solvent are present for each 100.0 g of solution. Calculations are then based on this ratio.

$$45.0 \text{ g} \times \frac{1 \text{ mol. H}_2\text{O}}{18.02 \text{ g}} = 2.50 \text{ mol. H}_2\text{O}$$

$$\text{mass of solute} = 100.0 \text{ g (total)} - 45.0 \text{ g (water)} = 55.0 \text{ g CO(NH}_2)_2$$

$$55.0 \text{ g urea} \times \frac{1 \text{ mol. urea}}{60.06 \text{ g urea}} = 0.916 \text{ mol. urea}$$

$$\text{molality} = \frac{0.916 \text{ mol. urea}}{0.0450 \text{ kg H}_2\text{O}} = \mathbf{20.4 \text{ m}}$$
 (remember to convert mass of solvent to kg)

$$\text{ppm solute} = \frac{55.0 \text{ g urea}}{100 \text{ g}} \times 10^6 = \mathbf{5.50 \times 10^5}$$

$$X_{\text{H}_2\text{O}} = \frac{2.50 \text{ mol.}}{(2.50 \text{ mol.} + 0.916 \text{ mol.})} = \mathbf{0.731}$$

$$(c) \text{ ppm} = \frac{\text{mass of solute}}{\text{mass of solution}} \times 10^6 = 4768 \text{ g}$$

If we assume a total mass of 10^6 g, then the mass of solute = 4768 g by definition. Any assumption for mass is valid here, this one was chosen for simplicity.

$$4768 \text{ g urea} \times \frac{1 \text{ mol. urea}}{60.06 \text{ g urea}} = 79.39 \text{ mol. urea}$$

$$10^6 \text{ g} - 4768 \text{ g} = 995232 \text{ g}; 995232 \text{ g H}_2\text{O} \times \frac{1 \text{ mol.}}{18.0152 \text{ g}} = 55244.0 \text{ mol. H}_2\text{O}$$

$$X_{\text{H}_2\text{O}} = \frac{55244.0 \text{ mol.}}{(55244.0 \text{ mol.} + 79.39 \text{ mol.})} = \mathbf{0.9986}$$

$$\text{mass \%} = \frac{9.95232 \times 10^5 \text{ g}}{10^6 \text{ g}} \times 100\% = \mathbf{99.5232\%}$$

$$\text{molality} = \frac{79.39 \text{ mol.}}{995.232 \text{ kg}} = \mathbf{0.07977 \text{ m}}$$

(d) $X_{\text{solvent}} = 0.815$ indicates that there are 0.815 moles H_2O per 1 mole of solvent and solute combined. Consequently, there must be 0.185 mol. (1-0.815) urea.

$$0.185 \text{ mol. urea} \times \frac{60.06 \text{ g urea}}{1 \text{ mol. urea}} = 11.1 \text{ g urea}$$

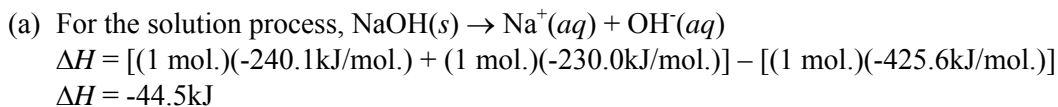
$$0.815 \text{ mol. H}_2\text{O} \times \frac{18.02 \text{ g}}{1 \text{ mol.}} = 14.7 \text{ g H}_2\text{O}$$

$$\frac{0.185 \text{ mol. urea}}{0.0147 \text{ kg H}_2\text{O}} = \mathbf{12.6 \text{ m}}$$
 (remember to convert mass of solvent to kg)

$$\text{Mass \%} = \frac{\text{mass of solvent}}{\text{total mass}} \times 100\% = \frac{14.7 \text{ g}}{(11.1 \text{ g} + 14.7 \text{ g})} \times 100\% = \mathbf{57.0\%}$$

$$\text{ppm} = \frac{\text{mass of solute}}{\text{mass of solution}} \times 10^6 = \frac{11.1 \text{ g urea}}{(11.1 \text{ g} + 14.7 \text{ g})} \times 10^6 = \mathbf{4.30 \times 10^5}$$

26. Refer to Section 10.2 and Chapter 8.



(b) Since the solution process is exothermic ($\Delta H < 0$), an increase in temperature will decrease the solubility of NaOH.

52. Refer to Section 10.3 and Example 10.8.

$$4.60 \text{ mm Hg} \times \frac{1 \text{ atm}}{760 \text{ mm Hg}} = 6.05 \times 10^{-3} \text{ atm}$$

$$\pi = MRT$$

$$6.05 \times 10^{-3} \text{ atm} = (M)(0.0821 \text{ L}\cdot\text{atm/mol}\cdot\text{K})(293 \text{ K})$$

$$M = 2.52 \times 10^{-4} \text{ mol./L}$$

$$0.200 \text{ L} \times \frac{2.52 \times 10^{-4} \text{ mol.}}{1 \text{ L}} = 5.03 \times 10^{-5} \text{ mol.}$$

$$\frac{3.27 \text{ g hemoglobin}}{5.03 \times 10^{-5} \text{ mol.}} = 6.50 \times 10^4 \text{ g/mol.}$$

Chapter 11: Rate of Reaction

6. Refer to Section 11.1.

$$\text{rate} = \frac{-\Delta[\text{Br}^-]}{5\Delta t} = \frac{-\Delta[\text{BrO}_3^-]}{\Delta t} = \frac{-\Delta[\text{H}^+]}{6\Delta t} = \frac{\Delta[\text{Br}_2]}{3\Delta t} = \frac{\Delta[\text{H}_2\text{O}]}{3\Delta t}$$

$$(a) \frac{\Delta[\text{Br}_2]}{\Delta t} = 0.039 \text{ mol./L}\cdot\text{s} \Rightarrow \frac{\Delta[\text{Br}_2]}{3\Delta t} = \frac{0.039 \text{ mol./L}\cdot\text{s}}{3}$$

$$\frac{\Delta[\text{Br}_2]}{3\Delta t} = \frac{\Delta[\text{H}_2\text{O}]}{3\Delta t} \Rightarrow \frac{\Delta[\text{H}_2\text{O}]}{3\Delta t} = \frac{0.039 \text{ mol./L}\cdot\text{s}}{3}$$

$$\frac{\Delta[\text{H}_2\text{O}]}{\Delta t} = \frac{3(0.039 \text{ mol./L}\cdot\text{s})}{3} = 0.039 \text{ mol./L}\cdot\text{s}$$

$$(b) \frac{-\Delta[\text{Br}^-]}{5\Delta t} = \frac{\Delta[\text{Br}_2]}{3\Delta t}$$

$$\frac{-\Delta[\text{Br}^-]}{5\Delta t} = \frac{0.039 \text{ mol./L}\cdot\text{s}}{3}$$

$$\frac{-\Delta[\text{Br}^-]}{\Delta t} = \frac{5(0.039 \text{ mol./L}\cdot\text{s})}{3} = 0.065 \text{ mol./L}\cdot\text{s}$$

$$(c) \frac{-\Delta[\text{H}^+]}{6\Delta t} = \frac{\Delta[\text{Br}_2]}{3\Delta t} \Rightarrow \frac{-\Delta[\text{H}^+]}{6\Delta t} = \frac{0.039 \text{ mol./L}\cdot\text{s}}{3}$$

$$\frac{-\Delta[\text{H}^+]}{\Delta t} = \frac{6(0.039 \text{ mol./L}\cdot\text{s})}{3} = 0.078 \text{ mol./L}\cdot\text{s}$$

16. Refer to Section 11.2.

(a) Zero order

(b) rate = $k[\text{Y}]^0$

(c) rate = $k[\text{Y}]^0$

$$0.045 \frac{\text{M}}{\text{hr}} = k(1)$$

$$k = 0.045 \frac{\text{M}}{\text{hr}}$$

26. Refer to Section 11.1 and Example 11.2.

- (a) To determine the reaction order in $[\text{BF}_3]$, select two experiments in which $[\text{NH}_3]$ is constant (such as experiments 1 and 4, as used below).

$$\frac{\text{rate}_1}{\text{rate}_4} = \frac{k[\text{BF}_3]_1^m [\text{NH}_3]_1^n}{k[\text{BF}_3]_4^m [\text{NH}_3]_4^n} \Rightarrow \frac{\text{rate}_1}{\text{rate}_4} = \left(\frac{[\text{BF}_3]_1}{[\text{BF}_3]_4} \right)^m \left(\frac{[\text{NH}_3]_1}{[\text{NH}_3]_4} \right)^n$$

$$\frac{0.0341 \text{ mol./L} \cdot \text{s}}{0.102 \text{ mol./L} \cdot \text{s}} = \left(\frac{0.100 \text{ mol./L}}{0.300 \text{ mol./L}} \right)^m \left(\frac{0.100 \text{ mol./L}}{0.100 \text{ mol./L}} \right)^n$$

$$0.334 = (0.333)^m (1)^n$$

$$0.334 = (0.333)^m$$

$$m = 1 \quad (\text{The reaction is 1st order in } \text{BF}_3.)$$

To determine the reaction order in $[\text{NH}_3]$, select two experiments in which $[\text{BF}_3]$ is constant (such as experiments 2 and 3) and repeat the process above.

$$\frac{0.159 \text{ mol./L} \cdot \text{s}}{0.0512 \text{ mol./L} \cdot \text{s}} = \left(\frac{0.200 \text{ mol./L}}{0.200 \text{ mol./L}} \right)^m \left(\frac{0.233 \text{ mol./L}}{0.750 \text{ mol./L}} \right)^n$$

$$0.311 = (1)^m (0.311)^n$$

$$0.311 = (0.311)^n$$

$$n = 1 \quad (\text{The reaction is 1st order in } \text{NH}_3.)$$

$$\text{Overall reaction order} = m + n = 1 + 1 = 2$$

(b) $\text{Rate} = k[\text{BF}_3][\text{NH}_3]$

- (c) Substitute values from any of the experiments into the rate equation and solve for k .

$$0.0341 \text{ M/s} = k(0.100 \text{ M})(0.100 \text{ M})$$

$$0.0341 \text{ M/s} = k(0.01 \text{ M}^2)$$

$$k = 3.41 \text{ s}^{-1} \cdot \text{M}^{-1} = 3.41 \text{ L/mol} \cdot \text{s}$$

(d) $\text{Rate} = k[\text{BF}_3][\text{NH}_3]$

$$\text{Rate} = (3.41 \text{ L/mol} \cdot \text{s})(0.553 \text{ M})(0.300 \text{ M})$$

$$\text{Rate} = 0.566 \text{ M/s}$$

Chapter 12: Gaseous Chemical Equilibrium

4. Refer to Section 12.1.

Time (min)	0	1	2	3	4	5	6
P_A (atm)	1.000	0.778	0.580	0.415	0.355	0.325	0.325
P_B (atm)	0.400	0.326	0.260	0.205	0.185	0.175	0.175
P_C (atm)	0.000	0.148	0.280	0.390	0.430	0.450	0.450

1 min: $1.000 \text{ atm A} - 0.778 \text{ atm A} = 0.222 \text{ atm A}$

$$0.222 \text{ atm A} \times \frac{1B}{3A} = 0.0740 \text{ atm B}$$

$0.400 \text{ atm B} - 0.0740 \text{ atm B} = 0.326 \text{ atm B}$

$$0.222 \text{ atm A} \times \frac{2C}{3A} = 0.148 \text{ atm C}$$

2 min: $0.400 \text{ atm B} - 0.260 \text{ atm B} = 0.140 \text{ atm B}$

$$0.140 \text{ atm B} \times \frac{3A}{1B} = 0.420 \text{ atm A}$$

$1.000 \text{ atm A} - 0.420 \text{ atm A} = 0.580 \text{ atm A}$

$$0.140 \text{ atm B} \times \frac{2C}{1B} = 0.280 \text{ atm C}$$

3 min: $0.390 \text{ atm C} \times \frac{1B}{2C} = 0.195 \text{ atm B}$

$0.400 \text{ atm B} - 0.195 \text{ atm B} = 0.205 \text{ atm B}$

$$0.390 \text{ atm C} \times \frac{3A}{2C} = 0.585 \text{ atm A}$$

$1.000 \text{ atm A} - 0.585 \text{ atm A} = 0.415 \text{ atm A}$

4 min: $0.400 \text{ atm B} - 0.185 \text{ atm B} = 0.215 \text{ atm B}$

$$0.215 \text{ atm B} \times \frac{3A}{1B} = 0.645 \text{ atm A}$$

$1.000 \text{ atm A} - 0.645 \text{ atm A} = 0.355 \text{ atm A}$

$$0.215 \text{ atm B} \times \frac{2C}{1B} = 0.430 \text{ atm C}$$

5 min: $1.000 \text{ atm A} - 0.325 \text{ atm A} = 0.675 \text{ atm A}$

$$0.675 \text{ atm A} \times \frac{1\text{B}}{3\text{A}} = 0.225 \text{ atm B}$$

$0.400 \text{ atm B} - 0.225 \text{ atm B} = 0.175 \text{ atm B}$

$$0.675 \text{ atm A} \times \frac{2\text{C}}{3\text{A}} = 0.450 \text{ atm C}$$

6 min: Since P_B has not changed from 5 min, P_A and P_C will also equal the values from 5 min.

6. Refer to Section 12.2 and Examples 12.1 and 12.2.

Recall that liquids and solids do not show up in equilibrium expressions.

$$(a) K = \frac{(P_{\text{H}_2})^3(P_{\text{CO}})}{(P_{\text{CH}_4})}$$

$$(b) K = \frac{(P_{\text{NO}})^4(P_{\text{H}_2\text{O}})^6}{(P_{\text{NH}_3})^4(P_{\text{O}_2})^5}$$

$$(c) K = P_{\text{CO}_2}$$

$$(d) K = \frac{1}{(P_{\text{NH}_3})(P_{\text{HCl}})}$$

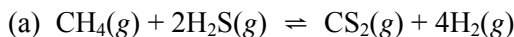
14. Refer to Section 12.2.

$$(a) K_1 = \frac{[\text{O}_2][\text{SO}_2]^2}{[\text{SO}_3]^2}$$

$$(b) K_2 = \frac{[\text{O}_2]^{\frac{1}{2}}[\text{SO}_2]}{[\text{SO}_3]}$$

$$(c) K_2 = K_1^{1/2}$$

24. Refer to Section 12.2 and Chapter 5.



(b) Use the ideal gas law equation to calculate the pressure of each gas, then substitute those values into the equation for the equilibrium constant.

$$P_{\text{CH}_4} = \frac{n}{V}RT = \frac{0.00142 \text{ mol.}}{1 \text{ L}}(0.0821 \text{ L} \cdot \text{atm/mol} \cdot \text{K})(1123 \text{ K}) = 0.131 \text{ atm}$$

$$P_{\text{H}_2\text{S}} = \frac{n}{V} RT = \frac{6.14 \times 10^{-4} \text{ mol.}}{1 \text{ L}} (0.0821 \text{ L} \cdot \text{atm/mol} \cdot \text{K})(1123 \text{ K}) = 0.0566 \text{ atm}$$

$$P_{\text{CS}_2} = \frac{n}{V} RT = \frac{0.00266 \text{ mol.}}{1 \text{ L}} (0.0821 \text{ L} \cdot \text{atm/mol} \cdot \text{K})(1123 \text{ K}) = 0.245 \text{ atm}$$

$$P_{\text{H}_2} = \frac{n}{V} RT = \frac{0.00943 \text{ mol.}}{1 \text{ L}} (0.0821 \text{ L} \cdot \text{atm/mol} \cdot \text{K})(1123 \text{ K}) = 0.869 \text{ atm}$$

$$K = \frac{(P_{\text{H}_2})^4 (P_{\text{CS}_2})}{(P_{\text{CH}_4})(P_{\text{H}_2\text{S}})^2} = \frac{(0.869 \text{ atm})^4 (0.245 \text{ atm})}{(0.131 \text{ atm})(0.0566 \text{ atm})^2} = 333$$

38. Refer to Section 12.4.

$$P_{\text{total}} = P_{\text{NO}} + P_{\text{NO}_2} + P_{\text{O}_2} \Rightarrow 1.25 = P_{\text{NO}} + P_{\text{NO}_2} + 0.515$$

$$0.735 \text{ atm} = P_{\text{NO}} + P_{\text{NO}_2}$$

$$\text{let } x = P_{\text{NO}}; \text{ then } 0.735 - x = P_{\text{NO}_2}$$

$$K = \frac{(P_{\text{NO}})^2 (P_{\text{O}_2})}{(P_{\text{NO}_2})^2} \Rightarrow 0.87 = \frac{(x)^2 (0.515)}{(0.735 - x)^2}$$

$$\sqrt{1.69} = \sqrt{\frac{(x)^2}{(0.735 - x)^2}}$$

$$1.3 = \frac{x}{0.735 - x} \Rightarrow x = 0.42$$

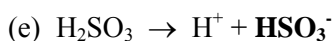
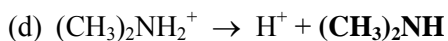
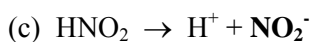
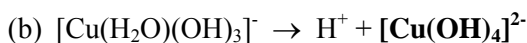
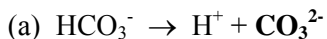
$$P_{\text{NO}_2} = 0.32 \text{ atm}$$

$$P_{\text{NO}} = 0.42 \text{ atm}$$

Chapter 13: Acids and Bases

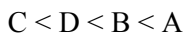
6. Refer to Section 13.1.

Write the equation for the dissociation of the species (HB) into H^+ and B^- . The resulting B^- will be the conjugate base of the acid HB.



38. Refer to Section 13.4.

(a) The stronger the acid, the greater the degree of ionization and thus the larger the K_a .



(b) The largest K_a corresponds to the smallest pK_a .

A has the smallest pK_a .

44. Refer to Section 13.4.

pH of 2.642 M HB is 5.32

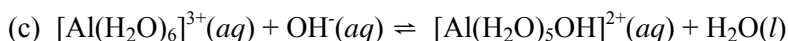
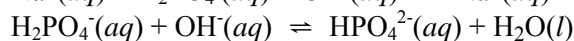
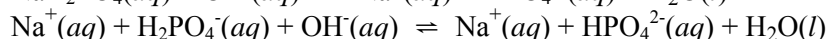
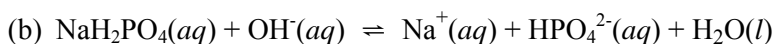
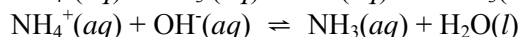
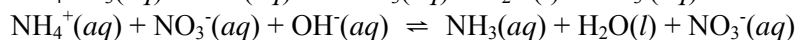
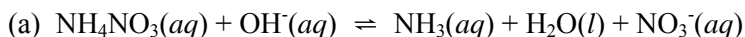
$$[H^+] = x = 10^{-pH} = 4.79 \times 10^{-6} \text{ M}$$

$$K_a = \frac{[H^+][B^-]}{[HB]} = \frac{x^2}{0.129 - x} = \frac{(4.79 \times 10^{-6})^2}{0.129 - 4.79 \times 10^{-6}} = 1.78 \times 10^{-10}$$

Chapter 14: Equilibria in Acid-Base Solutions

4. Refer to Chapters 4 and 13 and Problem 2 above.

Write the reactions, eliminating the spectator ions. Bear in mind that weak acids and bases (Table 13.2) do not ionize significantly, thus they exist in solution as the undissociated acid or base.



26. Refer to Section 14.1, Example 14.2 and Table 13.2.

Calculate the moles of the acid and base, then calculate the $[\text{H}^+]$ and pH.

(a) Calculate $[\text{H}^+]$ from pH, then calculate the acid/base ratio of the buffer.

$$[\text{H}^+] = 10^{-\text{pH}} = 10^{-3.0} = 1 \times 10^{-3} \text{ M}$$

$$[\text{H}^+] = K_a \times \frac{[\text{HB}]}{[\text{B}^-]} \Rightarrow 1 \times 10^{-3} = 1.9 \times 10^{-4} \times \frac{[\text{HCHO}_2]}{[\text{CHO}_2^-]}$$

$$\frac{[\text{HCHO}_2]}{[\text{CHO}_2^-]} = 5$$

(b) NaCHO_2 will dissociate to Na^+ and CHO_2^- . Use this fact and the equation from part (a) to calculate $[\text{HCHO}_2]$.

$$\frac{[\text{HCHO}_2]}{[\text{CHO}_2^-]} = 5 \Rightarrow \frac{[\text{HCHO}_2]}{0.139} = 5$$

$[\text{HCHO}_2] = 0.7 \text{ M}$, thus 0.7 mol. would need to be added to the liter of solution.

(c) $\frac{[\text{HCHO}_2]}{[\text{CHO}_2^-]} = 5 \Rightarrow \frac{0.159}{[\text{CHO}_2^-]} = 5$

$$[\text{CHO}_2^-] = 0.03 \text{ M}$$

$$0.3500 \text{ L} \times \frac{0.03 \text{ mol. CHO}_2^-}{1 \text{ L}} \times \frac{1 \text{ mol. NaCHO}_2}{1 \text{ mol. CHO}_2^-} \times \frac{68.01 \text{ g NaCHO}_2}{1 \text{ mol. NaCHO}_2} = 0.7 \text{ g NaCHO}_2$$

$$(d) \frac{[\text{HCHO}_2]}{[\text{CHO}_2^-]} = 5 \Rightarrow \frac{[\text{HCHO}_2]}{0.500} = 5$$

$[\text{HCHO}_2] = 2 \text{ M}$, since the volume of the formate solution is one liter, 2 mol of HCHO_2 is needed.

$$2 \text{ mol. HCHO}_2 \times \frac{1 \text{ L}}{0.236 \text{ mol. HCHO}_2} = 1 \times 10^1 \text{ L}$$

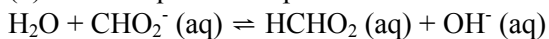
44. Refer to Section 14.3.

$$(a) M_B V_B = M_A V_A$$

$$(M_A)(0.02500 \text{ L}) = (0.117 \text{ M})(0.03974 \text{ L})$$

$$M_A = 0.186 \text{ M}$$

(b) At the equivalence point



$$K_b = K_w / K_a = 1 \times 10^{-14} / 1.9 \times 10^{-4} = 5.26 \times 10^{-11}$$

$$\frac{[\text{OH}^-][\text{HCHO}_2]}{[\text{CHO}_2^-]} = 5.26 \times 10^{-11} = \frac{x^2}{0.0783 - x}$$

$$X = 1.94 \times 10^{-6}$$

$$[\text{HCHO}_2] = 1.94 \times 10^{-6}$$

$$[\text{CHO}_2^-] = 7.83 \times 10^{-2}$$

$$[\text{OH}^-] = 1.94 \times 10^{-6}$$

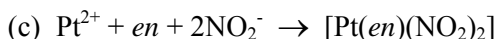
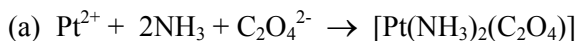
$$[\text{K}^+] = 7.83 \times 10^{-2}$$

$$(c) \text{pH} = 8.29$$

Chapter 15: Complex Ions

4. Refer to Section 15.1.

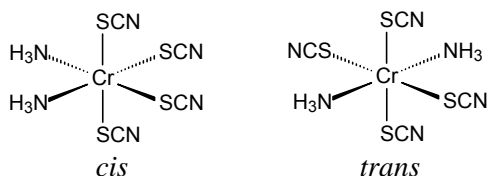
Put the metal and the ligands together and add up the charges to get the overall charge.



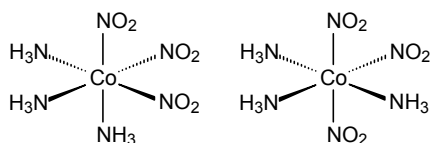
* *en* = ethylenediamine

22. Refer to Section 15.2, Example 15.3 and Figures 15.5 and 15.6.

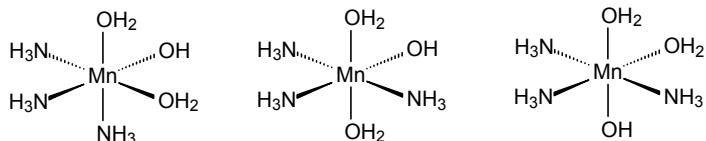
- (a) As in Figure 15.5, there are four of one ligand (SCN^-) and two of another (NH_3). There are two possible geometric isomers (NH_3 *cis* and NH_3 *trans*).



- (b) As in Figure 15.6, there are three of one ligand (NO_2^-), and three of another (NH_3). There are two possible geometric isomers.

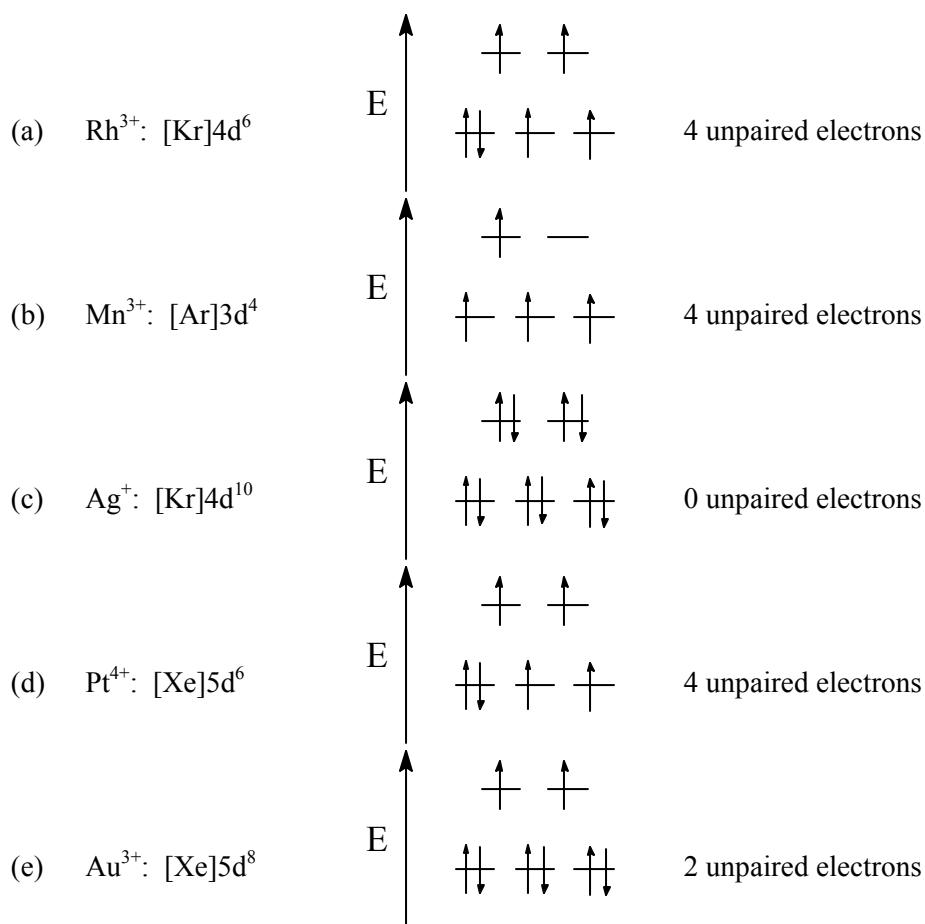


- (c) This complex has three NH_3 , two H_2O and one OH^- .



36. Refer to Section 15.3 and Problem 34.

For weak field ligands, Δ_0 is small, resulting in high spin complexes.



52. Refer to Chapter 15.

- (a) **Fals(e)** en occupies two coordination sites. Thus, the coordination number is 6.
- (b) **Fals(e)** $[\text{Ni}(\text{CN})_6]^{4-}$ is a low spin complex with a large Δ_o , thus it absorbs at a short wavelength. $[\text{Ni}(\text{NH}_3)_6]^{2+}$ is a high spin complex with a small Δ_o , thus it absorbs at a long wavelength.
- (c) **Tru(e)** Cr^{3+} has a partially filled d orbital. Thus there will be crystal field splitting energy and thus its complexes will be colored. Zn^{2+} has a d^{10} configuration. Thus there is no crystal field splitting energy and its complexes are not colored.
- (d) **Fals(e)** Ions with eight or more d electrons cannot form both high- and low-spin octahedral complexes.

Chapter 16: Precipitation Equilibria

10. Refer to Section 16.1.

	Compound	[cation]	[anion]	K_{sp}
(a)	$BaCO_3$	1.3×10^{-3}	1.3×10^{-3}	1.6×10^{-6}
(b)	$Cr(OH)_3$	2.7×10^{-8}	2.9×10^{-8}	6.3×10^{-31}
(c)	$Pb_3(PO_4)_2$	1×10^{-8}	8×10^{-6}	1×10^{-34}

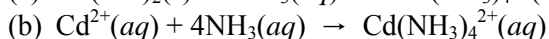
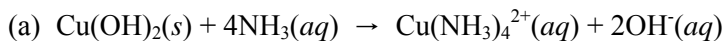
30. Refer to Section 16.2.

$$\frac{0.00865 \text{ g}}{100.0 \text{ mL}} \left(\frac{1 \text{ mol}}{86.0 \text{ g}} \right) \left(\frac{1000 \text{ mL}}{1 \text{ L}} \right) = 1.000 \times 10^{-3} \text{ M} = [Cr^{2+}]$$

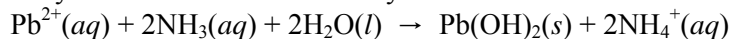
$$pH = 8.5 \text{ so } pOH = 5.5 \text{ or } [OH^-] = 1 \times 10^{-5.5}$$

$$K_{sp} = [Cr^{2+}][OH^-]^2 = 1.00 \times 10^{-14}$$

40. Refer to Section 16.2, Table 16.2, and Example 16.9.



(c) The lead must be combining with an anion to give an electrically neutral precipitate. A likely candidate is OH^- formed by the reaction of ammonia and water.



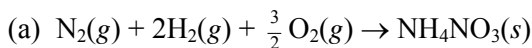
Chapter 17: Spontaneity of Reaction

8. Refer to Section 17.2.

- (a) O₂ (g) at 758 mm Hg
- (b) glucose (s)
- (c) Hg (l)

30. Refer to Section 17.4 and Example 17.5.

Determine ΔH_f° from the table in Appendix 1. Write the equation for the formation of the compounds from their elements. Calculate ΔS° and then ΔG_f° .



$$\Delta H_{\text{reaction}} = \Delta H_f^\circ = -365.6 \text{ kJ}$$

$$\Delta S^\circ = \sum S^\circ_{\text{products}} - \sum S^\circ_{\text{reactants}}$$

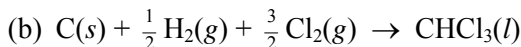
$$\Delta S^\circ = [(1 \text{ mol.})(0.1511 \text{ kJ/mol}\cdot\text{K})] - [(1 \text{ mol.})(0.1915 \text{ kJ/mol}\cdot\text{K}) \\ + (2 \text{ mol.})(0.1306 \text{ kJ/mol}\cdot\text{K}) + (\frac{3}{2} \text{ mol.})(0.2050 \text{ kJ/mol}\cdot\text{K})]$$

$$\Delta S^\circ = -0.6091 \text{ kJ/K}$$

$$\Delta G_{\text{reaction}} = \Delta G_f^\circ = \Delta H^\circ - T\Delta S^\circ$$

$$\Delta G_f^\circ = -365.6 \text{ kJ} - (298 \text{ K})(-0.6091 \text{ kJ/K}) = -184.1 \text{ kJ}$$

Note that ΔG_f° is defined for 1 mole, thus $\Delta G_f^\circ = -184.1 \text{ kJ/mol}$.



$$\Delta H_{\text{reaction}} = \Delta H_f^\circ = -134.5 \text{ kJ}$$

$$\Delta S^\circ = \sum S^\circ_{\text{products}} - \sum S^\circ_{\text{reactants}}$$

$$\Delta S^\circ = [(1 \text{ mol.})(0.2017 \text{ kJ/mol}\cdot\text{K})] - [(1 \text{ mol.})(0.0057 \text{ kJ/mol}\cdot\text{K}) \\ + (\frac{1}{2} \text{ mol.})(0.1306 \text{ kJ/mol}\cdot\text{K}) + (\frac{3}{2} \text{ mol.})(0.2230 \text{ kJ/mol}\cdot\text{K})]$$

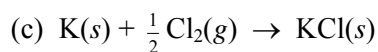
$$\Delta S^\circ = -0.2038 \text{ kJ/K}$$

$$\Delta G_{\text{reaction}} = \Delta G_f^\circ = \Delta H^\circ - T\Delta S^\circ$$

$$\Delta G_f^\circ = -134.5 \text{ kJ} - (298 \text{ K})(-0.2038 \text{ kJ/K})$$

$$\Delta G_f^\circ = -73.8 \text{ kJ}$$

Note that ΔG_f° is defined for 1 mole, thus $\Delta G_f^\circ = -73.8 \text{ kJ/mol}$.



$$\Delta H_{\text{reaction}} = \Delta H_f^\circ = -436.7 \text{ kJ}$$

$$\Delta S^\circ = \sum S^\circ_{(\text{prod})} - \sum S^\circ_{(\text{react.})}$$

$$\Delta S^\circ = [(1 \text{ mol.})(0.0826 \text{ kJ/mol}\cdot\text{K})] \\ - [(1 \text{ mol.})(0.0642 \text{ kJ/mol}\cdot\text{K}) + (\frac{1}{2} \text{ mol.})(0.2230 \text{ kJ/mol}\cdot\text{K})]$$

$$\Delta S^\circ = -0.0931 \text{ kJ/K}$$

$$\Delta G_{\text{reaction}} = \Delta G_f^\circ = \Delta H^\circ - T\Delta S^\circ$$

$$\Delta G_f^\circ = -436.7 \text{ kJ} - (298 \text{ K})(-0.0931 \text{ kJ/K})$$

$$\Delta G_f^\circ = -409.0 \text{ kJ}$$

Note that ΔG_f° is defined for 1 mole, thus $\Delta G_f^\circ = -409.0 \text{ kJ/mol}$.

66. Refer to Section 17.5.

$$(a) \Delta G^\circ = -RT \ln K$$

$$\Delta G^\circ = -(0.00831 \text{ kJ/mol}\cdot\text{K})(721 \text{ K})\ln(50.0)$$

$$\Delta G^\circ = -23.4 \text{ kJ/mol.}$$

$$(b) \Delta G^\circ = \sum \Delta G_f^\circ_{\text{products}} - \sum \Delta G_f^\circ_{\text{reactants}}$$

$$\Delta G^\circ = (2 \text{ mol.})(1.7 \text{ kJ/mol.}) - [(1 \text{ mol.})(0.0 \text{ kJ/mol.}) + (1 \text{ mol.})(19.4 \text{ kJ/mol.})]$$

$$\Delta G^\circ = -16 \text{ kJ/mol.}$$

$$\Delta G^\circ = -RT \ln K$$

$$-16 \text{ kJ} = -(0.00831 \text{ kJ/mol}\cdot\text{K})(298 \text{ K})\ln K$$

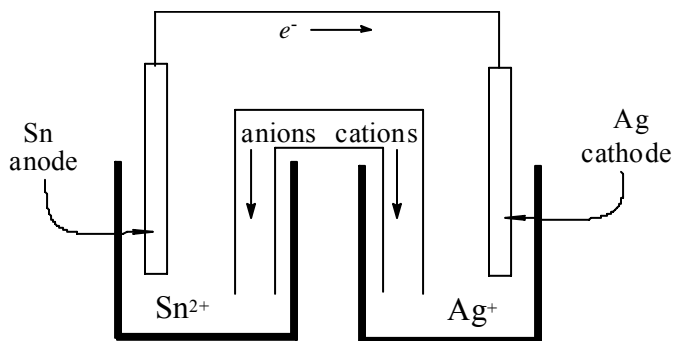
$$\ln K = 6.46$$

$$K = 6.4 \times 10^2$$

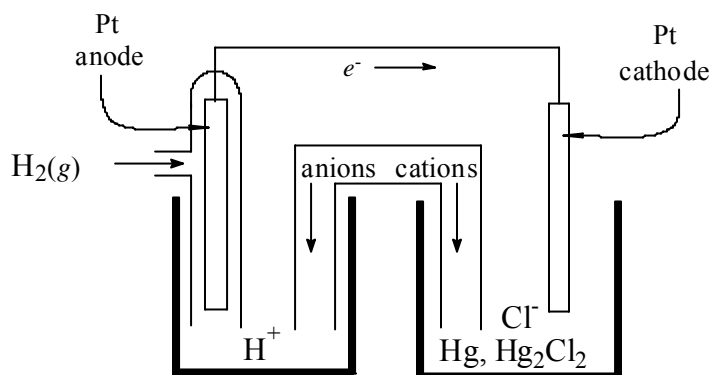
Chapter 18: Electrochemistry

4. Refer to Section 18.1 and Example 18.1.

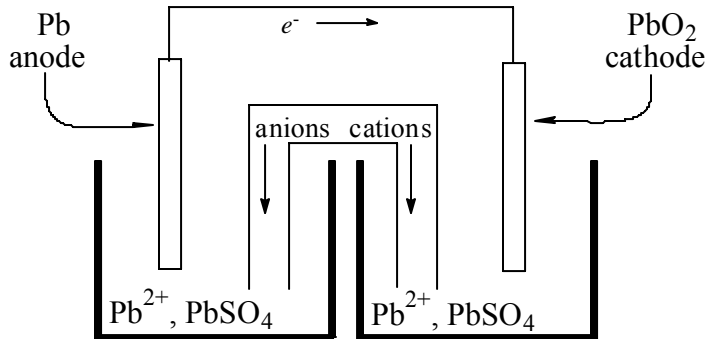
- (a) $\text{Sn}(s)$ is oxidized at the anode to $\text{Sn}^{2+}(aq)$.
 $\text{Ag}^+(aq)$ is reduced at the cathode to $\text{Ag}(s)$.



- (b) $\text{H}_2(g)$ is oxidized at the anode to $\text{H}^+(aq)$.
 $\text{Hg}_2^{2+}(\text{Hg}_2\text{Cl}_2)$ is reduced at the cathode to $\text{Hg}(l)$.



- (c) $\text{Pb}(s)$ is oxidized at the anode to $\text{Pb}^{2+}(\text{PbSO}_4(s))$.
 $\text{Pb}^{4+}(\text{PbO}_2(s))$ is reduced at the cathode to $\text{Pb}^{2+}(\text{PbSO}_4(s))$.



16. Refer to Section 18.1.

- (a) $\text{Co}^{3+} (\text{aq})$
- (b) $\text{Mn} (\text{s})$
- (c) $\text{Co}^{3+} (\text{aq})$
- (d) $\text{Mn}^{2+} (\text{aq})$
- (e) Yes
- (f) No
- (g) $\text{Fe}^{3+} (\text{aq})$ and $\text{Co}^{3+} (\text{aq})$
- (h) $\text{Pb} (\text{s})$, $\text{Cd} (\text{s})$, and $\text{Mn} (\text{s})$

74. Refer to Sections 18.5, 18.6, and Example 18.8.

Use the equation on page 502 for the reaction of a lead storage battery to determine the moles of e^- per mole of Pb.

$$(a) \quad 12.0 \text{ lb Ca} \times \frac{453.6 \text{ g}}{1 \text{ lb}} \times \frac{1 \text{ mol.}}{40.078 \text{ g Ca}} \times \frac{2 \text{ mol. } e^-}{1 \text{ mol. Ca}} = 272 \text{ mol. } e^-$$

$$272 \text{ mol. } e^- \times \frac{9.648 \times 10^4 \text{ C}}{1 \text{ mol. } e^-} = 2.62 \times 10^7 \text{ C}$$

$$3.2 \text{ V} = \frac{\text{J}}{2.62 \times 10^7 \text{ C}}$$

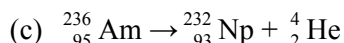
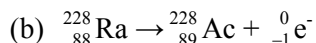
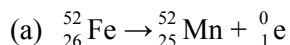
$$\text{J} = (3.2 \text{ V})(2.62 \times 10^7 \text{ C}) = 8.4 \times 10^7 \text{ J}$$

(b) $\$ 0.09 = 1 \text{ kWh} = 3.600 \times 10^6 \text{ J}$

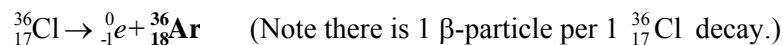
$$8.4 \times 10^7 \text{ J} \times \frac{1 \text{ kWh}}{3.600 \times 10^6 \text{ J}} \times \frac{\$ 0.09}{1 \text{ kWh}} = \$ 2.1$$

Chapter 19: Nuclear Chemistry

4. Refer to Section 19.1.



26. Refer to Sections 19.1 and 19.2.



$$\frac{2.3 \times 10^{-6}}{\text{yr}} \times \frac{1 \text{ yr}}{365 \text{ d}} \times \frac{1 \text{ d}}{24 \text{ hr}} \times \frac{1 \text{ hr}}{60 \text{ min}} = 4.4 \times 10^{-12} / \text{min}$$

$$1.00 \text{ mg Cl-36} \times \frac{1 \text{ g}}{1000 \text{ mg}} \times \frac{1 \text{ mol. Cl-36}}{36.0 \text{ g}} = 2.78 \times 10^{-5} \text{ mol. Cl-36}$$

$$2.78 \times 10^{-5} \text{ mol. Cl-36} \times \frac{1 \text{ mol. } \beta \text{- particles}}{1 \text{ mol. } {}_{17}^{36}\text{Cl}} \times \frac{6.022 \times 10^{23}}{1 \text{ mol.}} = 1.67 \times 10^{19} \beta \text{- particles}$$

$$A = kN = (4.4 \times 10^{-12} / \text{min})(1.67 \times 10^{19} \beta \text{- particles}) = 7.3 \times 10^7 \text{ particles/min}$$

$$\frac{7.3 \times 10^7 \text{ particles}}{1 \text{ min}} \times \frac{1 \text{ min}}{60 \text{ s}} \times \frac{1 \text{ Ci}}{3.700 \times 10^{10} \text{ particles/s}} = 3.3 \times 10^{-5} \text{ Ci}$$

38. Refer to Section 19.3 and Example 19.4.



(b) $\Delta m = 2(0.00055 \text{ g/mol.}) + 89.8824 \text{ g/mol.} - 89.8869 \text{ g/mol.} = -0.0034 \text{ g/mol.}$

(c) Calculate ΔE for one mole, then convert to kJ/g.

$$\Delta E = 9.00 \times 10^{10} \text{ kJ/g} \times \Delta m = (9.00 \times 10^{10} \text{ kJ/g})(-0.0034 \text{ g/mol.}) = -3.1 \times 10^8 \text{ kJ/mol.}$$

$$6.50 \text{ mg} \times \frac{1 \text{ g}}{1000 \text{ mg}} \times \frac{1 \text{ mol. Sr}}{89.8869 \text{ g}} \times \frac{-3.1 \times 10^8 \text{ kJ}}{1 \text{ mol.}} = -2.2 \times 10^4 \text{ kJ}$$

Chapter 20: Chemistry of the Metals

10. Refer to Section 20.1 and Chapter 18.

one metric ton = 1×10^3 kg

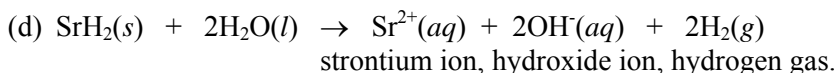
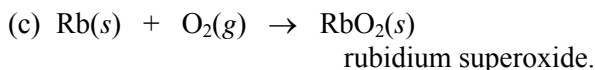
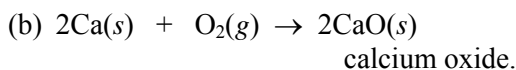
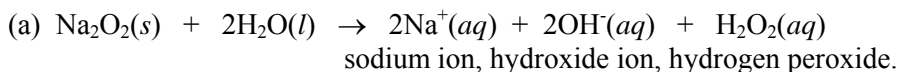
$$1.000 \times 10^3 \text{ kg} \times \frac{1000 \text{ g}}{1 \text{ kg}} \times \frac{1 \text{ mol. Zn}}{65.39 \text{ g}} = 1.529 \times 10^4 \text{ mol. Zn}$$

$$1.529 \times 10^4 \text{ mol. Zn} \times \frac{2 \text{ mol. } e^-}{1 \text{ mol. Zn}} \times \frac{9.648 \times 10^4 \text{ C}}{1 \text{ mol. } e^-} = 2.950 \times 10^9 \text{ C}$$

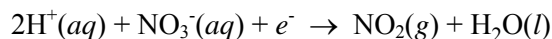
$$J = VC = (3.0 \text{ V})(2.950 \times 10^9 \text{ C}) = 8.9 \times 10^9 \text{ J}$$

$$8.9 \times 10^9 \text{ J} \times \frac{1 \text{ kWh}}{3.600 \times 10^6 \text{ J}} = 2.5 \times 10^3 \text{ kWh}$$

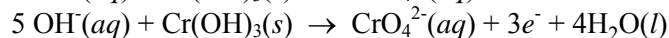
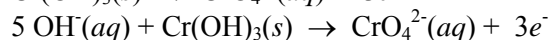
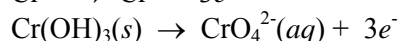
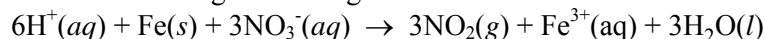
16. Refer to Section 20.2 and Table 20.1.

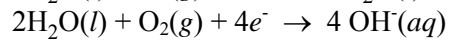
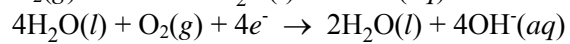
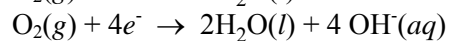
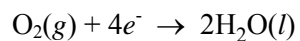


24. Refer to Section 20.3 and Chapter 4.

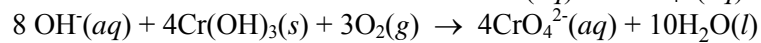
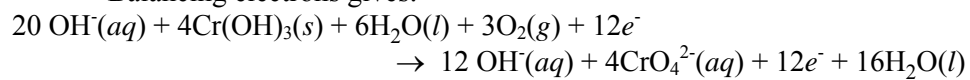


Balancing electrons gives:





Balancing electrons gives:



Chapter 21: Chemistry of the Nonmetals

20. Refer to Section 21.1 and Table 18.1.

- | | |
|---|--|
| <p>(a) $2\text{Br}^-(aq) \rightarrow \text{Br}_2(l) + 2e^-$
 $\text{Cl}_2(g) + 2e^- \rightarrow 2\text{Cl}^-(aq)$
 $\text{Cl}_2(g) + 2\text{Br}^-(aq) \rightarrow 2\text{Cl}^-(aq) + \text{Br}_2(l)$</p> | <p>$E_{\text{ox}}^\circ = -1.077 \text{ V}$
 $E_{\text{red}}^\circ = +1.360 \text{ V}$
 E° is (+) thus reaction occurs.</p> |
| <p>(b) $2\text{Cl}^-(aq) \rightarrow \text{Cl}_2(g) + 2e^-$
 $\text{I}_2(s) + 2e^- \rightarrow 2\text{I}^-(aq)$
 $\text{I}_2(s) + \text{Cl}^-(aq) \rightarrow \text{N.R.}$</p> | <p>$E_{\text{ox}}^\circ = -1.360 \text{ V}$
 $E_{\text{red}}^\circ = +0.534 \text{ V}$
 E° is (-) thus no reaction occurs.</p> |
| <p>(c) $2\text{Br}^-(aq) \rightarrow \text{Br}_2(l) + 2e^-$
 $\text{I}_2(s) + 2e^- \rightarrow 2\text{I}^-(aq)$
 $\text{I}_2(s) + \text{Br}^-(aq) \rightarrow \text{N.R.}$</p> | <p>$E_{\text{ox}}^\circ = -1.077 \text{ V}$
 $E_{\text{red}}^\circ = +0.534 \text{ V}$
 E° is (-) thus no reaction occurs.</p> |
| <p>(d) $2\text{Cl}^-(aq) \rightarrow \text{Cl}_2(g) + 2e^-$
 $\text{Br}_2(l) + 2e^- \rightarrow 2\text{Br}^-(aq)$
 $\text{Br}_2(l) + \text{Cl}^-(aq) \rightarrow \text{N.R.}$</p> | <p>$E_{\text{ox}}^\circ = -1.360 \text{ V}$
 $E_{\text{red}}^\circ = +1.077 \text{ V}$
 E° is (-) thus no reaction occurs.</p> |

32. Refer to Section 21.4 and Chapters 7 and 13.

- | | | |
|-------------------------------|--|---|
| (a) NO_3^- | $\left[\begin{array}{c} \text{:O:} \\ \text{:}\ddot{\text{O}}\text{---N---}\ddot{\text{O}}\text{:} \\ \text{:}\ddot{\text{O}}\text{:} \end{array} \right]^-$ | |
| (b) HSO_4^- | $\left[\begin{array}{c} \text{:O:} \\ \text{H---}\ddot{\text{O}}\text{---S---}\ddot{\text{O}}\text{:} \\ \text{:}\ddot{\text{O}}\text{:} \end{array} \right]^-$ | This Lewis structure has a -1 formal charge on the most electronegative atom (O). |
| (c) H_2PO_4^- | $\left[\begin{array}{c} \text{:}\ddot{\text{O}}\text{:} \\ \text{H---}\ddot{\text{O}}\text{---P---}\ddot{\text{O}}\text{---H} \\ \text{:}\ddot{\text{O}}\text{:} \end{array} \right]^-$ | This Lewis structure has a -1 formal charge on the most electronegative atom (O). |

70. Refer to Sections 21.3 and 21.4.

- (a) Increasing oxidation number corresponds to an increase in oxygen atoms around the central atom. These oxygen atoms stabilize the negative charge of the resulting conjugate base by i) delocalizing the negative charge (consider resonance structures) and ii) by the electronegative oxygens pulling electron density away from the atom bearing the charge.
- (b) NO_2 has an odd number of electrons and will thus have an unpaired electron.
- (c) In general, the reduction reactions of the oxoanions (oxoanions acting as oxidizing agents) involve H^+ as a reactant. Therefore higher $[\text{H}^+]$ (lower pH) causes the reaction to be more spontaneous.
- (d) The sugar is oxidized to carbon, which is black.

Chapter 22: Organic Chemistry

12. Refer to Section 22.2.

The structures are printed in Appendix 6, p670. Thus, this space will be used to explain the answer.

- (a) "pent" = five carbon chain
"-yne" = triple bond
"2-" = triple bond starts at carbon 2
- (b) "pent" = five carbon chain
"-yne" = triple bond
"2-" = triple bond starts at carbon 2
"4-methyl" = methyl group attached to carbon 4
- (c) "hex" = six carbon chain
"-yne" = triple bond
"3-" = triple bond starts at carbon 3
"2-methyl" = methyl group attached to carbon 2
- (d) "but" = four carbon chain
"-yne" = triple bond
"1-" = triple bond starts at carbon 1
"3,3-dimethyl" = two methyl groups attached to carbon 3

Note the nomenclature priority system. The multiple bond has a lower number than the functional groups.

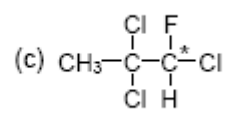
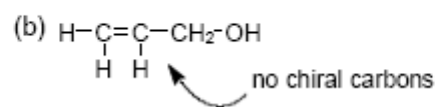
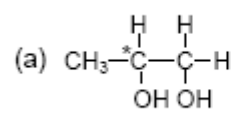
32. Refer to Section 22.5 and Example 22.9.

This type of problem is best approached by systematically moving the Cl atoms, first keeping two to a carbon, then moving them one at a time. Be aware that the isomer with both Cl's on the 1st carbon is identical to that with both on the last carbon.



46. Refer to Section 22.5 and Example 22.11.

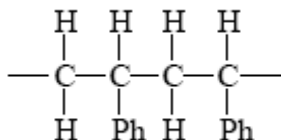
The chiral carbons (denoted with an *) are those that have four different groups attached.



Chapter 23: Organic Polymers, Natural and Synthetic

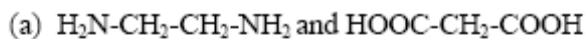
4. Refer to Section 23.1 and Example 23.2.

This problem is solved the same way as Example 23.2(a) by replacing CH_3 with a phenyl group (abbreviated as Ph)

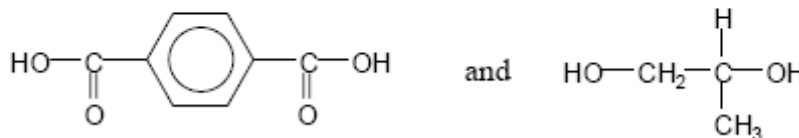


14. Refer to Section 23.2 and Example 23.4.

Condensation polymers are made by removing OH from one monomer and H from the other. To identify the monomer, locate the bond forming the polymer linkage and add OH to one end of the bond and H to the other.



(b)



26. Refer to Table 23.3.

Calculate the possible combinations of the three amino acids. This can be done mathematically or by inspection.

Lys	Phe	Val
Lys	Val	Phe
Phe	Lys	Val
Phe	Val	Lys
Val	Lys	Phe
Val	Phe	Lys

Thus, there are 6 possible tripeptides.