Chap 5 Exercises

1. Using information from this chapter, indicate whether each of the following substances in aqueous solution is a nonelectrolyte, weak electrolyte, or strong electrolyte. **(a)** HC_6H_5O; **(b)** Li_2SO_4; **(c)** MgI_2; **(d)** $(CH_3CH_2)_2O$; **(e)** $Sr(OH)_2$.

Ans:

    (a) Because its formula begins with hydrogen, HC_6H_5O is an acid. It is not listed in Table 5-1, so it is a weak acid. A weak acid is a *weak electrolyte*.

    (b) Li_2SO_4 is an ionic compound, that is, a salt. A salt is a *strong electrolyte*.

    (c) MgI_2 also is a salt, a *strong electrolyte*.

    (d) $(CH_3CH_2)_2O$ is a covalent compound whose formula does not begin with H. Thus, it is neither an acid nor a salt. It also is not built around nitrogen, and thus it does not behave as a weak base. This is a *nonelectrolyte*.

    (e) $Sr(OH)_2$ is a *strong electrolyte*, one of the strong bases listed in Table 5-2.

9. A solution is prepared by dissolving 0.132 g $Ba(OH)_2 \cdot 8 H_2O$ in 275 mL of water solution. What is $[OH^-]$ in this solution?

Ans:

Conversion pathway approach:

$$[OH^-] = \frac{0.132 \text{ g Ba}(OH)_2 \cdot 8H_2O}{275 \text{ mL soln}} \times \frac{1000 \text{ mL}}{1 \text{ L}} \times \frac{1 \text{ mol Ba}(OH)_2 \cdot 8H_2O}{315.5\text{g Ba}(OH)_2 \cdot 8H_2O} \times \frac{2 \text{ mol OH}^-}{1 \text{ mol Ba}(OH)_2 \cdot 8 \text{ H}_2O}$$

$$= 3.04 \times 10^{-3} \text{ M OH}^-$$

Stepwise approach:

$$\frac{0.132 \text{ g Ba}(OH)_2 \cdot 8H_2O}{275 \text{ mL soln}} \times \frac{1000 \text{ mL}}{1 \text{ L}} = 0.480 \text{ g/L}$$

$$\frac{0.480 \text{ g}}{L} \times \frac{1 \text{ mol Ba}(OH)_2 \cdot 8H_2O}{315.5\text{g Ba}(OH)_2 \cdot 8H_2O} = \frac{0.00152 \text{ mol Ba}(OH)_2 \times 8H_2O}{L}$$

$$\frac{0.00152 \text{ mol Ba}(OH)_2 \times 8H_2O}{L} \times \frac{2 \text{ mol OH}^-}{1 \text{ mol Ba}(OH)_2 \cdot 8 \text{ H}_2O} = 3.04 \times 10^{-3} \text{ M OH}^-$$

17. Assuming the volumes are additive, what is the $[Cl^-]$ in a solution obtained by mixing 225 mL of 0.625 M KCl and 615 mL of 0.385 M $MgCl_2$?
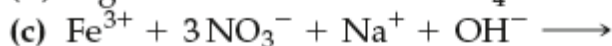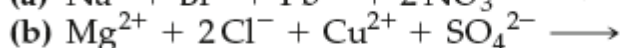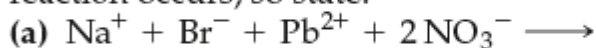
Ans:

(M) moles of chloride ion

$$= \left(0.225 \text{ L} \times \frac{0.625 \text{ mol KCl}}{1 \text{ L soln}} \times \frac{1 \text{ mol Cl}^-}{1 \text{ mol KCl}} \right) + \left(0.615 \text{ L} \times \frac{0.385 \text{ mol MgCl}_2}{1 \text{ L soln}} \times \frac{2 \text{ mol Cl}^-}{1 \text{ mol MgCl}_2} \right)$$

$$= 0.141 \text{ mol Cl}^- + 0.474 \text{ mol Cl}^- = 0.615 \text{ mol Cl}^- \quad [Cl^-] = \frac{0.615 \text{ mol Cl}^-}{0.225 \text{ L} + 0.615 \text{ L}} = 0.732 \text{ M}$$

19. Complete each of the following as a net ionic equation, indicating whether a precipitate forms. If no reaction occurs, so state.

(a) $Na^+ + Br^- + Pb^{2+} + 2 NO_3^- \longrightarrow$
(b) $Mg^{2+} + 2 Cl^- + Cu^{2+} + SO_4^{2-} \longrightarrow$
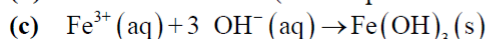(c) $Fe^{3+} + 3 NO_3^- + Na^+ + OH^- \longrightarrow$

Ans:

(E) In each case, each available cation is paired with the available anions, one at a time, to determine if a compound is produced that is insoluble, based on the solubility rules of Chapter 5. Then a net ionic equation is written to summarize this information.
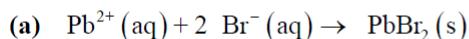
(a) $Pb^{2+}(aq) + 2 Br^-(aq) \rightarrow PbBr_2(s)$
(b) No reaction occurs (all are spectator ions).
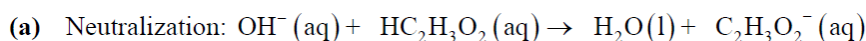(c) $Fe^{3+}(aq) + 3 OH^-(aq) \rightarrow Fe(OH)_3(s)$

27. Complete each of the following as a *net ionic equation*. If no reaction occurs, so state.

(a) $Ba^{2+} + 2 OH^- + CH_3COOH \longrightarrow$
(b) $H^+ + Cl^- + CH_3CH_2COOH \longrightarrow$
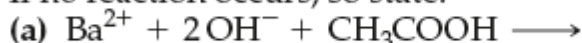
Ans:

(E) The type of reaction is given first, followed by the net ionic equation.

(a) Neutralization: $OH^-(aq) + HC_2H_3O_2(aq) \rightarrow H_2O(l) + C_2H_3O_2^-(aq)$
(b) No reaction occurs. This is the physical mixing of two acids.
(c) Gas evolution: $FeS(s) + 2 H^+(aq) \rightarrow H_2S(g) + Fe^{2+}(aq)$
(d) Gas evolution: $HCO_3^-(aq) + H^+(aq) \rightarrow "H_2CO_3(aq)" \rightarrow H_2O(l) + CO_2(g)$
(e) Redox: $Mg(s) + 2 H^+(aq) \rightarrow Mg^{2+}(aq) + H_2(g)$
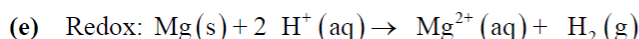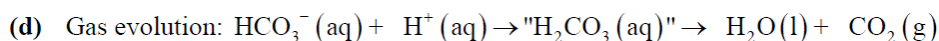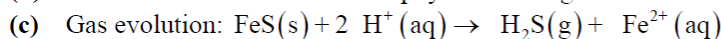
37. Balance these equations for redox reactions occurring in acidic solution.

(a) $MnO_4^- + I^- \longrightarrow Mn^{2+} + I_2(s)$

(b) $BrO_3^- + N_2H_4 \longrightarrow Br^- + N_2$

Ans:

(M)

(a) Oxidation: $\{2\ I^-(aq) \rightarrow\ I_2(s)+2\ e^-$ $\}\times 5$

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

     Net: $10\ I^-(aq)+2\ MnO_4^-(aq)+16\ H^+(aq)\rightarrow 5\ I_2(s)+2\ Mn^{2+}(aq)+8\ H_2O(l)$

(b) Oxidation: $\{N_2H_4(l) \rightarrow\ N_2(g)+4\ H^+(aq)+4\ e^-$ $\}\times 3$

     Reduction: $\{BrO_3^-(aq)+6\ H^+(aq)+6\ e^- \rightarrow\ Br^-(aq)+3\ H_2O(l)$ $\}\times 2$

     Net: $\ 3\ N_2H_4(l)+2\ BrO_3^-(aq)\rightarrow 3\ N_2(g)+2\ Br^-(aq)+6\ H_2O(l)$

(c) Oxidation: $Fe^{2+}(aq) \rightarrow\ Fe^{3+}(aq)+\ e^-$

     Reduction: $VO_4^{3-}(aq)+6\ H^+(aq)+\ e^- \rightarrow\ VO^{2+}(aq)+3\ H_2O(l)$

     Net: $\ Fe^{2+}(aq)+\ VO_4^{3-}(aq)+6\ H^+(aq)\rightarrow\ Fe^{3+}(aq)+\ VO^{2+}(aq)+3\ H_2O(l)$

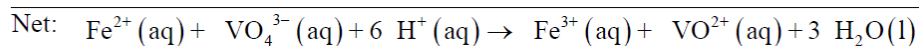(d) Oxidation: $\{UO^{2+}(aq)+\ H_2O(l) \rightarrow\ UO_2^{2+}(aq)+2\ H^+(aq)+2\ e^-$ $\}\times 3$

     Reduction: $\{NO_3^-(aq)+4\ H^+(aq)+3\ e^- \rightarrow\ NO(g)+2\ H_2O(l)$ $\}\times 2$

     Net: $\ 3\ UO^{2+}(aq)+2\ NO_3^-(aq)+2\ H^+(aq)\rightarrow 3\ UO_2^{2+}(aq)+2\ NO(g)+\ H_2O(l)$

47. What are the oxidizing and reducing agents in the fol-
lowing redox reactions?
(a) $5\,SO_3{}^{2-} + 2\,MnO_4{}^- + 6\,H^+ \longrightarrow$
$$5\,SO_4{}^{2-} + 2\,Mn^{2+} + 3\,H_2O$$
(b) $2\,NO_2(g) + 7\,H_2(g) \longrightarrow 2\,NH_3(g) + 4\,H_2O(g)$
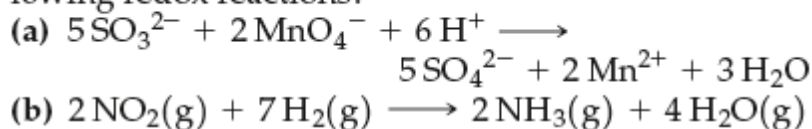
Ans:

(E) The oxidizing agents experience a decrease in the oxidation state of one of their elements,
while the reducing agents experience an increase in the oxidation state of one of their elements.

(a) $SO_3{}^{2-}(aq)$ is the reducing agent; the O.S. of $S = +4$ in $SO_3{}^{2-}$ and $= +6$ in $SO_4{}^{2-}$.
$MnO_4{}^-(aq)$ is the oxidizing agent; the O.S. of $Mn = +7$ in $MnO_4{}^-$ and $+2$ in Mn^{2+}.

(b) $H_2(g)$ is the reducing agent; the O.S. of $H = 0$ in $H_2(g)$ and $= +1$ in $H_2O(g)$.
$NO_2(g)$ is the oxidizing agent; the O.S. of $N = +4$ in $NO_2(g)$ and -3 in $NH_3(g)$.

(c) $\left[Fe(CN)_6\right]^{4-}(aq)$ is the reducing agent; the O.S. of $Fe = +2$ in $\left[Fe(CN)_6\right]^{4-}$
and $= +3$ in $\left[Fe(CN)_6\right]^{3-}$. $H_2O_2(aq)$ is the oxidizing agent; the O.S. of $O = -1$
in H_2O_2 and $= -2$ in H_2O.

49. What volume of 0.0962 M NaOH is required to exactly
neutralize 10.00 mL of 0.128 M HCl?

Ans:

(E) The problem is most easily solved with amounts in millimoles.

$$V_{NaOH} = 10.00\ \text{mL HCl(aq)} \times \frac{0.128\ \text{mmol HCl}}{1\ \text{mL HCl(aq)}} \times \frac{1\ \text{mmol H}^+}{1\ \text{mmol HCl}} \times \frac{1\ \text{mmol OH}^-}{1\ \text{mmol H}^+}$$

$$\times \frac{1\ \text{mmol NaOH}}{1\ \text{mmol OH}^-} \times \frac{1\ \text{mL NaOH(aq)}}{0.0962\ \text{mmol NaOH}} = 13.3\ \text{mL NaOH(aq) soln}$$

65. An iron ore sample weighing 0.9132 g is dissolved in HCl(aq), and the iron is obtained as Fe^{2+}(aq). This solution is then titrated with 28.72 mL of 0.05051 M $K_2Cr_2O_7$. What is the mass percent Fe in the ore sample?

$$6\,Fe^{2+} + 14\,H^+ + Cr_2O_7^{2-} \longrightarrow$$
$$6\,Fe^{3+} + 2\,Cr^{3+} + 7\,H_2O$$
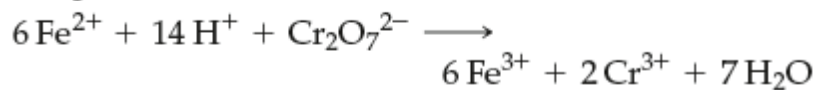
Ans:

(M) First, we will determine the mass of Fe, then the percentage of iron in the ore.

$$\text{mass Fe} = 28.72\ \text{mL} \times \frac{1\ \text{L}}{1000\ \text{mL}} \times \frac{0.05051\ \text{mol Cr}_2\text{O}_7^{2-}}{1\ \text{L soln}} \times \frac{6\ \text{mol Fe}^{2+}}{1\ \text{mol Cr}_2\text{O}_7^{2-}} \times \frac{55.85\ \text{g Fe}}{1\ \text{mol Fe}^{2+}}$$

$$\text{mass Fe} = 0.4861\ \text{g Fe} \quad \% \text{Fe} = \frac{0.4861\,\text{g Fe}}{0.9132\,\text{g ore}} \times 100\% = 53.23\%\ \text{Fe}$$

Chap 6 Exercises

1. Convert each pressure to an equivalent pressure in atmospheres. **(a)** 736 mmHg; **(b)** 0.776 bar; **(c)** 892 Torr; **(d)** 225 kPa.

Ans:

(a) $P = 736 \text{ mmHg} \times \dfrac{1 \text{ atm}}{760 \text{ mmHg}} = 0.968 \text{ atm}$

(b) $P = 0.776 \text{ bar} \times \dfrac{1 \text{ atm}}{1.01325 \text{ bar}} = 0.766 \text{ atm}$

(c) $P = 892 \text{ torr} \times \dfrac{1 \text{ atm}}{760 \text{ torr}} = 1.17 \text{ atm}$

(d) $P = 225 \text{ kPa} \times \dfrac{1000 \text{ Pa}}{1 \text{ kPa}} \times \dfrac{1 \text{ atm}}{101{,}325 \text{ Pa}} = 2.22 \text{ atm}$

11. If 3.0 L of oxygen gas at 177 °C is cooled at constant pressure until the volume becomes 1.50 L, then what is the final temperature?

Ans:

(E) Charles' Law states that $V_1/T_1 = V_2/T_2$. Therefore,

$\dfrac{3.0 \text{ L}}{450 \text{ K}} = \dfrac{1.50 \text{ L}}{T_2}$, and $T_2 = 225 \text{ K}$

23. A sample of gas has a volume of 4.25 L at 25.6 °C and 748 mmHg. What will be the volume of this gas at 26.8 °C and 742 mmHg?

Ans:

(E) Because the number of moles of gas does not change, $\dfrac{P_i \times V_i}{T_i} = nR = \dfrac{P_f \times V_f}{T_f}$ is obtained from the ideal gas equation. This expression can be rearranged as follows.

$V_f = \dfrac{V_i \times P_i \times T_f}{P_f \times T_i} = \dfrac{4.25 \text{ L} \times 748 \text{ mmHg} \times (273.2 + 26.8)\text{K}}{742 \text{ mmHg} \times (273.2 + 25.6)\text{K}} = 4.30 \text{ L}$

29. Kr(g) in a 18.5 L cylinder exerts a pressure of 11.2 atm at 28.2 °C. How many grams of gas are present?

Ans:

(E) $\text{mass} = n \times M = \dfrac{PV}{RT}$

$M = \dfrac{11.2 \text{ atm} \times 18.5 \text{ L} \times 83.80 \text{ g/mol}}{0.08206 \dfrac{\text{L atm}}{\text{mol K}} (28.2+273.2)\,\text{K}} = 702 \text{ g Kr}$

35. A 0.418 g sample of gas has a volume of 115 mL at 66.3 °C and 743 mmHg. What is the molar mass of this gas?

Ans:

(**M**) Use the ideal gas law to determine the amount in moles of the given quantity of gas.

$$M = \frac{mRT}{PV} = \frac{0.418 \text{ g} \times 0.08206\dfrac{\text{L atm}}{\text{mol K}} \times 339.5 \text{ K}}{\left(743 \text{ mmHg} \times \dfrac{1 \text{ atm}}{760 \text{ mmHg}}\right) \times 0.115 \text{ L}} = 104 \text{ g mol}^{-1}$$

Alternatively

$$n = \frac{PV}{RT} = \frac{\left(743 \text{ mmHg} \times \dfrac{1 \text{ atm}}{760 \text{ mmHg}}\right)\left(115 \text{ mL} \times \dfrac{1 \text{ L}}{1000 \text{ mL}}\right)}{0.08206\dfrac{\text{L atm}}{\text{mol K}}(273.2 + 66.3) \text{ K}} = 0.00404 \text{ mol gas} \quad M = \frac{0.418 \text{ g}}{0.00404 \text{ mol}} = 103 \text{ g}/\text{mol}$$

43. In order for a gas-filled balloon to rise in air, the density of the gas in the balloon must be less than that of air.
(a) Consider air to have a molar mass of 28.96 g/mol; determine the density of air at 25 °C and 1 atm, in g/L.

Ans:

(a) $\quad d = \dfrac{MP}{RT} = \dfrac{28.96 \text{ g/mol} \times 1.00 \text{ atm}}{0.08206\dfrac{\text{L atm}}{\text{mol K}} \times (273 + 25)\text{K}} = 1.18 \text{ g/L air}$

(b) $\quad d = \dfrac{MP}{RT} = \dfrac{44.0 \text{ g}/\text{mol CO}_2 \times 1.00 \text{ atm}}{0.08206\dfrac{\text{L atm}}{\text{mol K}} \times (273 + 25)\text{K}} = 1.80 \text{ g}/\text{L CO}_2$

Since this density is greater than that of air, the balloon will not rise in air when filled with CO_2 at $25°C$; instead, it will sink!

53. Calculate the volume of $H_2(g)$, measured at 26 °C and 751 Torr, required to react with 28.5 L $CO(g)$, measured at 0 °C and 760 Torr, in this reaction.

$$3 \text{ CO(g)} + 7 \text{ H}_2\text{(g)} \longrightarrow \text{C}_3\text{H}_8\text{(g)} + 3 \text{ H}_2\text{O(l)}$$

Ans:

(**M**) First we need to find the number of moles of CO(g)
Reaction is $\quad 3 \text{ CO(g)} + 7 \text{ H}_2\text{(g)} \longrightarrow \text{C}_3\text{H}_8\text{(g)} + 3 \text{ H}_2\text{O(l)}$
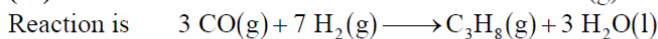
$$n_{CO} = \frac{PV}{RT} = \frac{28.5 \text{ L} \times 760 \text{ torr} \times \dfrac{1 \text{ atm}}{760 \text{ tor}}}{0.08206\dfrac{\text{L atm}}{\text{K mol}} \times 273.15 \text{ K}} = 1.27 \text{ moles CO}$$

$$V_{H_2 \text{ (required)}} = \frac{n_{H_2} RT}{P} = \frac{1.27 \text{ mol CO} \times \dfrac{7 \text{ mol H}_2}{3 \text{ mol CO}} \times 0.08206\dfrac{\text{L atm}}{\text{K mol}} \times 299 \text{ K}}{751 \text{ mmHg} \times \dfrac{1 \text{ atm}}{760 \text{ mmHg}}} = 73.7 \text{ L H}_2$$

55. What is the volume, in liters, occupied by a mixture of 15.2 g Ne(g) and 34.8 g Ar(g) at 7.15 atm pressure and 26.7 °C?

Ans:

(M) Determine the total amount of gas; then use the ideal gas law, assuming that the gases behave ideally.

$$\text{moles gas} = \left(15.2 \text{ g Ne} \times \frac{1 \text{ mol Ne}}{20.18 \text{ g Ne}}\right) + \left(34.8 \text{ g Ar} \times \frac{1 \text{ mol Ar}}{39.95 \text{ g Ar}}\right)$$

$$= 0.753 \text{ mol Ne} + 0.871 \text{ mol Ar} = 1.624 \text{ mol gas}$$

$$V = \frac{nRT}{P} = \frac{1.624 \text{ mol} \times 0.08206 \frac{\text{L atm}}{\text{mol K}} \times (26.7 + 273.2) \text{ K}}{7.15 \text{ atm}} = 5.59 \text{ L gas}$$

67. A 1.65 g sample of Al reacts with excess HCl, and the liberated H_2 is collected over water at 25 °C at a barometric pressure of 744 mmHg. What volume of gaseous mixture, in liters, is collected?

$$2\,Al(s) + 6\,HCl(aq) \longrightarrow 2\,AlCl_3(aq) + 3\,H_2(g)$$

Ans:

(M) The pressure of the liberated $H_2(g)$ is 744 mmHg – 23.8 mmHg = 720. mmHg

$$V = \frac{nRT}{P} = \frac{\left(1.65 \text{ g Al} \times \frac{1 \text{ mol Al}}{26.98 \text{ g}} \times \frac{3 \text{ mol H}_2}{2 \text{ mol Al}}\right) 0.08206 \frac{\text{L atm}}{\text{mol K}} (273 + 25)\text{K}}{720. \text{ mmHg} \times \frac{1 \text{ atm}}{760 \text{ mmHg}}} = 2.37 \text{ L H}_2(g)$$

This is the total volume of both gases, each with a different partial pressure.

74. The u_{rms} of H_2 molecules at 273 K is 1.84×10^3 m/s. At what temperature is u_{rms} for H_2 twice this value?

Ans:

(M)

$$2 = \frac{\sqrt{\dfrac{3RT_1}{M}}}{\sqrt{\dfrac{3RT_2}{M}}} = \sqrt{\frac{T_1}{T_2}} = \sqrt{\frac{T_1}{273 \text{ K}}}$$

Square both sides and solve for T_1. $T_1 = 4 \times 273 \text{ K} = 1092 \text{ K}$

Alternatively, recall that $1 \text{ J} = \text{kg m}^2 \text{ s}^{-2}$

$$u_{rms} = \sqrt{\frac{3RT}{M}} = 1.84 \times 10^3 \text{ m/s}$$

Solve this equation for temperature with u_{rms} doubled.

$$T = \frac{M u_{rms}^2}{3R} = \frac{2.016 \times 10^{-3} \text{ kg/mol } (2 \times 1.84 \times 10^3 \text{ m/s})^2}{3 \times 8.3145 \dfrac{J}{\text{mol K}} \times \dfrac{1 \text{ kg m}^2 \text{ s}^{-2}}{1 \text{ J}}} = 1.09 \times 10^3 \text{ K}$$

83. What are the ratios of the diffusion rates for the pairs of gases **(a)** N_2 and O_2; **(b)** H_2O and D_2O (D = deuterium, i.e., ^2_1H); **(c)** $^{14}CO_2$ and $^{12}CO_2$;

Ans:

    (M)

    (a) $\dfrac{\text{rate } (N_2)}{\text{rate } (O_2)} = \sqrt{\dfrac{M(O_2)}{M(N_2)}} = \sqrt{\dfrac{32.00}{28.01}} = 1.07$ **(b)** $\dfrac{\text{rate } (H_2O)}{\text{rate } (D_2O)} = \sqrt{\dfrac{M(D_2O)}{M(H_2O)}} = \sqrt{\dfrac{20.0}{18.02}} = 1.05$

    (c) $\dfrac{\text{rate } (^{14}CO_2)}{\text{rate } (^{12}CO_2)} = \sqrt{\dfrac{M(^{12}CO_2)}{M(^{14}CO_2)}} = \sqrt{\dfrac{44.0}{46.0}} = 0.978$ **(d)** $\dfrac{\text{rate } (^{235}UF_6)}{\text{rate } (^{238}UF_6)} = \sqrt{\dfrac{M(^{238}UF_6)}{M(^{235}UF_6)}} = \sqrt{\dfrac{352}{349}} = 1.004$

Chap 7 Exercises

9. A 74.8 g sample of copper at 143.2 °C is added to an insulated vessel containing 165 mL of glycerol, $C_3H_8O_3(l)$ $(d = 1.26\,g/mL)$, at 24.8 °C. The final temperature is 31.1°C. The specific heat of copper is $0.385\,J\,g^{-1}\,°C^{-1}$. What is the heat capacity of glycerol in $J\,mol^{-1}\,°C^{-1}$?

Ans:

(M) heat lost by copper = heat gained by glycerol

$$-74.8\ g \times \frac{0.385\ J}{g\,°C} \times \left(31.1°C - 143.2°C\right) = 165\ mL \times \frac{1.26\ g}{1\ mL} \times sp.ht. \times \left(31.1°C - 24.8°C\right)$$

$$3.23 \times 10^3 = 1.3 \times 10^3 \times (sp.ht.) \quad sp.ht. = \frac{3.23 \times 10^3}{1.3 \times 10^3} = 2.5\ J\,g^{-1}\,°C^{-1}$$

$$molar\ heat\ capacity = 2.5\ J\,g^{-1}\,°C^{-1} \times \frac{92.1\ g}{1\ mol\ C_3H_8O_3} = 2.3 \times 10^2\ J\,mol^{-1}\,°C^{-1}$$

13. How much heat, in kilojoules, is associated with the production of 283 kg of slaked lime, $Ca(OH)_2$?

$$CaO(s) + H_2O(l) \longrightarrow$$
$$Ca(OH)_2(s) \qquad \Delta H° = -65.2\,kJ$$
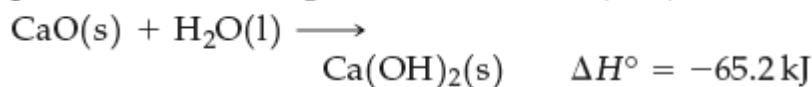
Ans:

(E) $heat = 283\ kg \times \dfrac{1000\ g}{1\ kg} \times \dfrac{1\ mol\ Ca(OH)_2}{74.09\ g\ Ca(OH)_2} \times \dfrac{65.2\ kJ}{1\ mol\ Ca(OH)_2} = 2.49 \times 10^5$ kJ of heat evolved.

17. The combustion of methane gas, the principal constituent of natural gas, is represented by the equation

$$CH_4(g) + 2O_2(g) \longrightarrow CO_2(g) + 2H_2O(l)$$
$$\Delta H^\circ = -890.3 \text{ kJ}$$
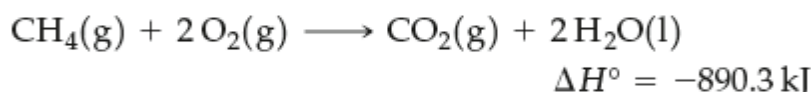
(a) What mass of methane, in kilograms, must be burned to liberate 2.80×10^7 kJ of heat?

Ans:

(M)

(a) $\text{mass} = 2.80 \times 10^7 \text{ kJ} \times \dfrac{1 \text{ mol CH}_4}{890.3 \text{ kJ}} \times \dfrac{16.04 \text{ g CH}_4}{1 \text{ mol CH}_4} \times \dfrac{1 \text{ kg}}{1000 \text{ g}} = 504 \text{ kg CH}_4.$

(b) First determine the moles of CH_4 present, with the ideal gas law.

$$\text{mol CH}_4 = \dfrac{\left(768 \text{ mmHg} \times \dfrac{1 \text{ atm}}{760 \text{ mmHg}}\right) 1.65 \times 10^4 \text{ L}}{0.08206 \dfrac{\text{L atm}}{\text{mol K}} \times (18.6 + 273.2) \text{ K}} = 696 \text{ mol CH}_4$$

$$\text{heat energy} = 696 \text{ mol CH}_4 \times \dfrac{-890.3 \text{ kJ}}{1 \text{ mol CH}_4} = -6.20 \times 10^5 \text{ kJ of heat energy}$$

(c) $V_{H_2O} = \dfrac{6.21 \times 10^5 \text{ kJ} \times \dfrac{1000 \text{ J}}{1 \text{ kJ}}}{4.18 \dfrac{\text{J}}{\text{g} \,^\circ\text{C}} (60.0 - 8.8)\,^\circ\text{C}} \times \dfrac{1 \text{ mL H}_2\text{O}}{1 \text{ g}} = 2.90 \times 10^6 \text{ mL} = 2.90 \times 10^3 \text{ L H}_2\text{O}$

29. What mass of ice can be melted with the same quantity of heat as required to raise the temperature of 3.50 mol $H_2O(l)$ by 50.0 °C? $[\Delta H^\circ_{fusion} = 6.01 \text{ kJ/mol } H_2O(s)]$

Ans:

(M) $q_{H_2O(l)} = q_{H_2O(s)}$ $m(\text{sp. ht.})_{H_2O(l)} \Delta T_{H_2O(l)} = \text{mol}_{H_2O(s)} \Delta H_{fus\ H_2O(s)}$

$(3.50 \text{ mol H}_2\text{O} \times \dfrac{18.015 \text{ g H}_2\text{O}}{1 \text{ mol H}_2\text{O}})(4.184 \dfrac{\text{J}}{\text{g }^\circ\text{C}})(50.0\,^\circ\text{C}) = (\dfrac{m}{\dfrac{18.015 \text{ g H}_2\text{O}}{1 \text{ mol H}_2\text{O}}} \times 6.01 \times 10^3 \dfrac{\text{J}}{\text{mol}})$

$13.2 \times 10^3 \text{ J} = m(333.\underline{6} \text{ J g}^{-1})$ Hence, $m = 39.6 \text{ g}$

41. A 1.620 g sample of naphthalene, $C_{10}H_8(s)$, is completely burned in a bomb calorimeter assembly and a temperature increase of 8.44 °C is noted. If the heat of combustion of naphthalene is −5156 kJ/mol $C_{10}H_8$, what is the heat capacity of the bomb calorimeter?

Ans:

(M) To determine the heat capacity of the calorimeter, recognize that the heat evolved by the reaction is the negative of the heat of combustion.

$$\text{heat capacity} = \frac{\text{heat evolved}}{\Delta T} = \frac{1.620\,\text{g}\,C_{10}H_8 \times \dfrac{1\,\text{mol}\,C_{10}H_8}{128.2\,\text{g}\,C_{10}H_8} \times \dfrac{5156.1\,\text{kJ}}{1\,\text{mol}\,C_{10}H_8}}{8.44\,°C} = 7.72\,\text{kJ/°C}$$

53. If 325 J of work is done by a system at a pressure of 1.0 atm and 298 K, what is the change in the volume of the system?

Ans:

(E) We can either convert pressure from atm to Pascals, or convert work from Joules to L·atm. We opt for the latter. Since the conversion between J and L·atm is 101.33 J/(L·atm), the amount of work is 325 J × (1 L·atm/101.33 J) = 3.207 L·atm. Therefore,
W = -PΔV
3.207 L·atm = (1.0 atm) ΔV. Solving for ΔV, we get a volume of 3.21 L.

55. What is the change in internal energy of a system if the system **(a)** absorbs 58 J of heat and does 58 J of work; **(b)** absorbs 125 J of heat and does 687 J of work;
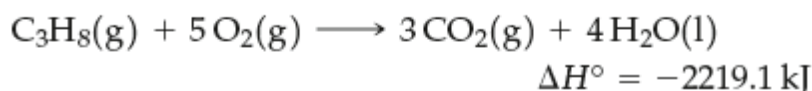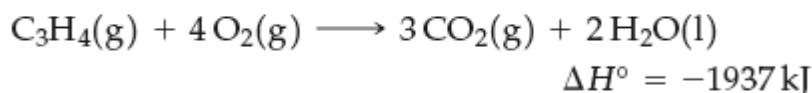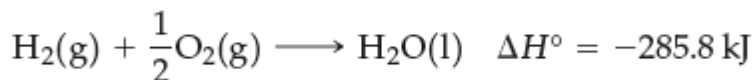
Ans:

(E)
(a) $\Delta U = q + w = +58\,\text{J} + (−58\,\text{J}) = 0$
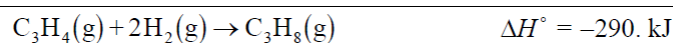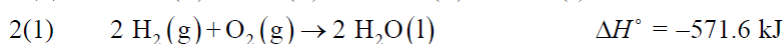(b) $\Delta U = q + w = +125\,\text{J} + (−687\,\text{J}) = −562\,\text{J}$
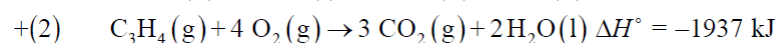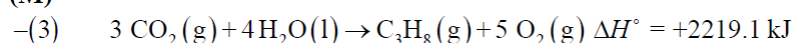(c)   280 cal × $(4.184\,\dfrac{\text{J}}{\text{cal}})$ = 1171.52 J = 1.17 kJ $\Delta U = q + w = −1.17\,\text{kJ} + 1.25\,\text{kJ} = 0.08\,\text{kJ}$

69. Use Hess's law to determine $\Delta H°$ for the reaction
$C_3H_4(g) + 2\,H_2(g) \longrightarrow C_3H_8(g)$, given that

$$H_2(g) + \frac{1}{2}O_2(g) \longrightarrow H_2O(l) \quad \Delta H° = -285.8\text{ kJ}$$

$$C_3H_4(g) + 4\,O_2(g) \longrightarrow 3\,CO_2(g) + 2\,H_2O(l)$$
$$\Delta H° = -1937\text{ kJ}$$

$$C_3H_8(g) + 5\,O_2(g) \longrightarrow 3\,CO_2(g) + 4\,H_2O(l)$$
$$\Delta H° = -2219.1\text{ kJ}$$

Ans:

(M)

$-(3)$	$3\,CO_2(g)+4\,H_2O(l)\rightarrow C_3H_8(g)+5\,O_2(g)$	$\Delta H° = +2219.1\text{ kJ}$
$+(2)$	$C_3H_4(g)+4\,O_2(g)\rightarrow 3\,CO_2(g)+2\,H_2O(l)$	$\Delta H° = -1937\text{ kJ}$
$2(1)$	$2\,H_2(g)+O_2(g)\rightarrow 2\,H_2O(l)$	$\Delta H° = -571.6\text{ kJ}$
	$C_3H_4(g)+2\,H_2(g)\rightarrow C_3H_8(g)$	$\Delta H° = -290.\text{ kJ}$

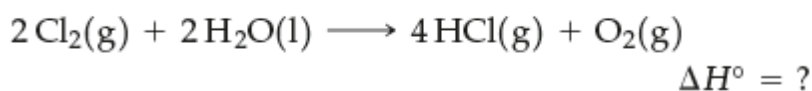**83.** Use standard enthalpies of formation from Table 7.2 to determine the enthalpy change at 25 °C for the following reaction.

$$2\,Cl_2(g) + 2\,H_2O(l) \longrightarrow 4\,HCl(g) + O_2(g)$$
$$\Delta H° = ?$$

Ans:

(E) $\Delta H° = 4\Delta H_f°[HCl(g)]+\Delta H_f°[O_2(g)]-2\Delta H_f°[Cl_2(g)]-2\Delta H_f°[H_2O(l)]$
$= 4\,(-92.31) + (0.00) - 2\,(0.00) -2\,(-285.8) = +202.4\text{ kJ}$