Solutions
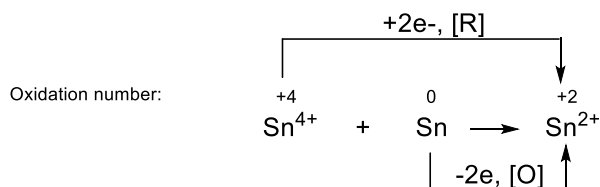
Problem 1

a.

i.   Cesium(I) fluoride (Incorrect)

ii.   Lithium(I) fluoride (Incorrect)

iii.   Tin fluoride (Incorrect)

iv.   Silver(I) fluoride (Incorrect)

v.   Zinc(II) fluoride (Incorrect)

b.

i. Oxidation number: +2 -1
 CaH_2 Ca: +2
                                              H: -1

ii. Oxidation number: ? -2 $2(?) + 5(-2) = 0$ gives $? = +5$ I: +5
    I_2O_5 O: -2

iii. Oxidation number: ?-2 $? + 3(-2) = -3$ gives $? = +3$ B: +3
    $BO_3{}^{3-}$ O: -2

iv. Oxidation number: ?-2 $? - 2 = -1$ gives $? = +1$ I: +1
    IO^- O: -2

v. Oxidation number: ? -2 $2(?) + 4(-2) = 0$ gives $? = +4$ N: +4
    N_2O_4 O: -2

c.

i.   **Step 1.** Balance the number of carbon atoms:
        $C_3H_6O + O_2 \rightarrow 3CO_2 + H_2O$
        Step 2. Balance the number of hydrogen atoms:
        $C_3H_6O + O_2 \rightarrow 3CO_2 + 3H_2O$
        Step 3. Balance the number of oxygen atoms:
        $C_3H_6O + 4O_2 \rightarrow 3CO_2 + 3H_2O$

ii.   **Step 1.** Decide which species that are oxidized ([O]) and reduced ([R]):
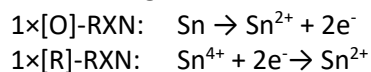
                              +2e-, [R]

        Oxidation number: +4   0   +2
                              Sn^{4+} + Sn $\longrightarrow$ Sn^{2+}
                                          -2e, [O]

        Step 2. Divide the reaction into an oxidation reaction ([O]-RXN) and reduction
        reaction ([R]-RXN):
         [O]-RXN: $Sn \rightarrow Sn^{2+}$
         [R]-RXN $Sn^{4+} \rightarrow Sn^{2+}$
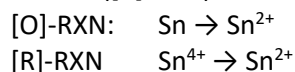
Step 3. Balance the charge by adding electrons (e^-):
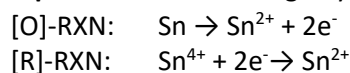
[O]-RXN: $Sn \rightarrow Sn^{2+} + 2e^-$

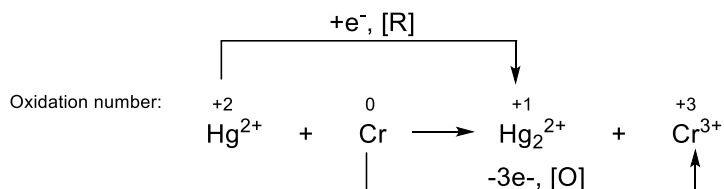[R]-RXN: $Sn^{4+} + 2e^- \rightarrow Sn^{2+}$

Step 4. Equalize the number of electrons lost in the oxidation reaction by the number of electrons gained in the reduction reaction by multiplying each reaction by a suitable integer:
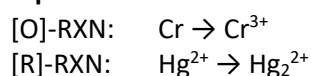
$1\times$[O]-RXN: $Sn \rightarrow Sn^{2+} + 2e^-$

$1\times$[R]-RXN: $Sn^{4+} + 2e^- \rightarrow Sn^{2+}$

Step 5. Add the two half-reaction together to get the balanced reaction equation:

$1\times$[O]-RXN: $Sn \rightarrow Sn^{2+} + 2e^-$

$1\times$[R]-RXN: $Sn^{4+} + 2e^- \rightarrow Sn^{2+}$

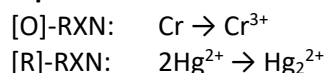Balanced reaction: $Sn^{4+} + Sn \rightarrow 2Sn^{2+}$
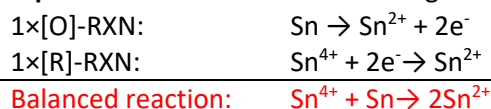
iii.   **Step 1.** Decide which species that are oxidized ([O]) and reduced ([R]):

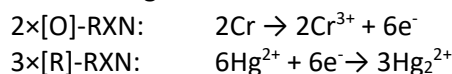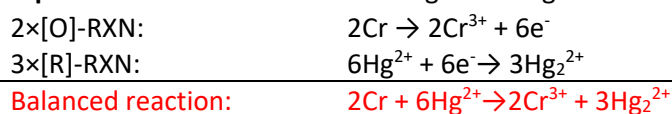**Step 2.** Divide the reaction into an oxidation reaction ([O]-RXN) and reduction:

[O]-RXN: $Cr \rightarrow Cr^{3+}$

[R]-RXN: $Hg^{2+} \rightarrow Hg_2^{2+}$

Step 3. Balance the number of atoms:
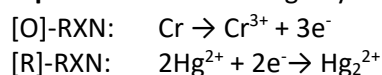
[O]-RXN: $Cr \rightarrow Cr^{3+}$

[R]-RXN: $2Hg^{2+} \rightarrow Hg_2^{2+}$

Step 4. Balance the charge by adding electrons (e^-):

[O]-RXN: $Cr \rightarrow Cr^{3+} + 3e^-$

[R]-RXN: $2Hg^{2+} + 2e^- \rightarrow Hg_2^{2+}$

Step 5. Equalize the number of electrons lost in the oxidation reaction by the number of electrons gained in the reduction reaction by multiplying each reaction by a suitable integer:

$2\times$[O]-RXN: $2Cr \rightarrow 2Cr^{3+} + 6e^-$

$3\times$[R]-RXN: $6Hg^{2+} + 6e^- \rightarrow 3Hg_2^{2+}$

Step 6. Add the two half-reaction together to get the balanced reaction equation:

$2\times$[O]-RXN: $2Cr \rightarrow 2Cr^{3+} + 6e^-$

$3\times$[R]-RXN: $6Hg^{2+} + 6e^- \rightarrow 3Hg_2^{2+}$

Balanced reaction: $2Cr + 6Hg^{2+} \rightarrow 2Cr^{3+} + 3Hg_2^{2+}$

d.

C_{conc} = 14.5 M $\qquad\qquad\qquad\qquad$ C_{dil} = 4.50 M

V_{conc} = ? $\qquad\qquad\qquad\qquad$ $V_{dil} = V_{conc} + V_{H2O}$ = 1.50 L **(1)**

Insert real values into equation **2** to calculate V_{conc}:

$V_{conc} \times C_{conc} = V_{dil} \times C_{dil}$ **(2)**

$V_{conc} \times$ (14.5 M) = (1.50 L) × (4.50 M) gives V_{conc} = 0.466 L = 466 mL

Insert real values into equation **1** to find V_{H2O}:
1.50 L = 0.466 L + V_{H2O} gives V_{H2O} = 1.034 L = 1034 mL

Problem 2

a.

C = 3.00 mol/L $\qquad\qquad\qquad\qquad$ M = 58.44 g/mol

V = 0.50 L $\qquad\qquad\qquad\qquad$ m = ?

$$C = \frac{n}{V} \text{ gives } n = C \times V \textbf{ (1)}$$
$$n = \frac{m}{M} \text{ gives } m = n \times M \textbf{ (2)}$$

Insert real values into equation **1** to find number of moles NaCl in the solution:
$$n = \left(3.00\frac{mol}{L}\right) \times (0.50 \text{ L}) = 1.5 \text{ mol}$$

Insert real values into equation **2** to find the mass of NaCl in the solution:
$$m = (1.5 \text{ mol}) \times \left(58.44\frac{g}{mol}\right) = 88 \text{ g}$$

b.

Assume $V_{solution}$ = 1.0 L which corresponds to m(solution) = 1.25 kg NaCl solution = 1250 g NaCl solution

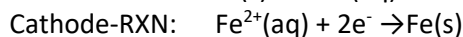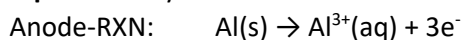$$m(NaCl) = n(NaCl) \times M(NaCl) \textbf{ (1)}$$
$$n(NaCl) = C \times V(solution) \textbf{ (2)}$$

Mass-percent NaCl in the solution is:
$$\text{Mass\% NaCl} = \frac{m(NaCl)}{m(solution)} \times 100\% = \frac{n(NaCl) \times M(NaCl)}{m(solution)} \times 100\%$$
$$= \frac{C \times V(solution) \times M(NaCl)}{m(solution)} \times 100\%$$
$$= \frac{\left(3.00\frac{mol}{L}\right)(1.0 \text{ L}) \times (58.44 g/mol)}{1250 \text{ g}} \times 100\% = 14 \text{ mass\% NaCl}$$

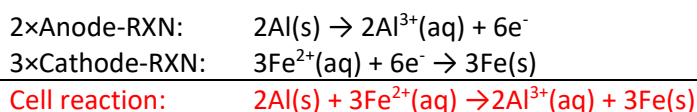c. **Step 1.** Identify anode and cathode reaction:

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

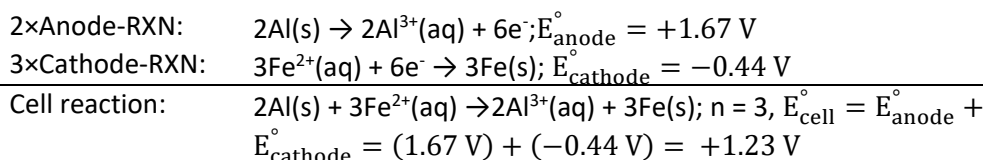Cathode-RXN: $Fe^{2+}(aq) + 2e^- \rightarrow Fe(s)$

Step 2. Multiply the half reactions with such factor that the number of electrons lost from the anode becomes equal to the number of electrons gained by the cathode:

2×Anode-RXN: $2Al(s) \rightarrow 2Al^{3+}(aq) + 6e^-$
3×Cathode-RXN: $3Fe^{2+}(aq) + 6e^- \rightarrow 3Fe(s)$
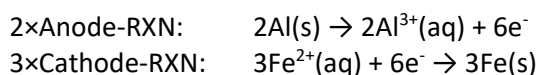
Step 3. Add the half-cell reactions together to obtain the cell reaction:

2×Anode-RXN: $2Al(s) \rightarrow 2Al^{3+}(aq) + 6e^-$
3×Cathode-RXN: $3Fe^{2+}(aq) + 6e^- \rightarrow 3Fe(s)$
Cell reaction: $2Al(s) + 3Fe^{2+}(aq) \rightarrow 2Al^{3+}(aq) + 3Fe(s)$

d.

2×Anode-RXN: $2Al(s) \rightarrow 2Al^{3+}(aq) + 6e^-$; $E^{\circ}_{anode} = +1.67\ V$
3×Cathode-RXN: $3Fe^{2+}(aq) + 6e^- \rightarrow 3Fe(s)$; $E^{\circ}_{cathode} = -0.44\ V$
Cell reaction: $2Al(s) + 3Fe^{2+}(aq) \rightarrow 2Al^{3+}(aq) + 3Fe(s)$; n = 3, $E^{\circ}_{cell} = E^{\circ}_{anode} +$
$E^{\circ}_{cathode} = (1.67\ V) + (-0.44\ V) = +1.23\ V$

$$E_{cell} = E^{\circ}_{Cell} - \frac{0.059}{6}\log Q = 1.23 - \frac{0.059}{6}\log\frac{0.300^2}{0.100^3} = 1.21\ V$$

Problem 3

a. Chemical equilibrium is achieved when the reaction rate in the both directions are equal. Therefore, a reaction in chemical equilibrium is described in figures C, E, H, and J.

b. The reaction can be forced away from its equilibrium by changing the concertation (or pressure) for any of the components in the reaction equation. (Note; the reaction is not forced away from its chemical equilibrium upon volume change as the number of moles of gas molecules are the same at both side of the reaction equation).

c.

$$PV = nRT\ gives\ [x] = \frac{n}{V} = \frac{P}{RT}\ (1)$$

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

Calculate the equilibrium molar concentrations for NO_2 and N_2O_4 by using equation **1**:

$$[NO_2] = \frac{0.337\ atm}{(0.082atm/Kmol) \times (298.15\ K)} = 0.0138\ M$$

$$[N_2O_4] = \frac{1.00\ atm}{(0.082atm/Kmol) \times (298.15\ K)} = 0.0409\ M$$

Calculate the equilibrium constant by using equation **2**:

$$K_c = \frac{0.0138^2}{0.0409} = 4.66 \times 10^{-3}$$

d.

	N₂O₄(g)	⇌	2NO₂(g)
Initial:	$\dfrac{0.0409\ M}{2} = 0.02045\ M$		$\dfrac{0.0138\ M}{2} = 0.00690\ M$
Change:	-x		+2x
Equilibrium:	(0.02045 - x) M		(0.00690 + 2x) M

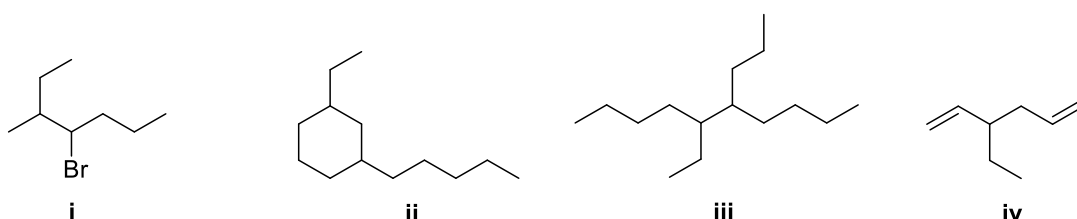Insert the molar concentrations into equation **2** in **Problem 3c** and solve it for x:

$$K_c = \frac{(0.00690 + 2x)^2}{0.02045} = 4.66 \times 10^{-3} \text{gives x} = 0.00127625\ M$$

$$[NO_2] = 0.00690 + 2 \times 0.00127625\ M = 9.45 \times 10^{-3} M$$
$$[N_2O_4] = 0.02045 - 0.00127625\ M = 1.92 \times 10^{-2}\ M$$
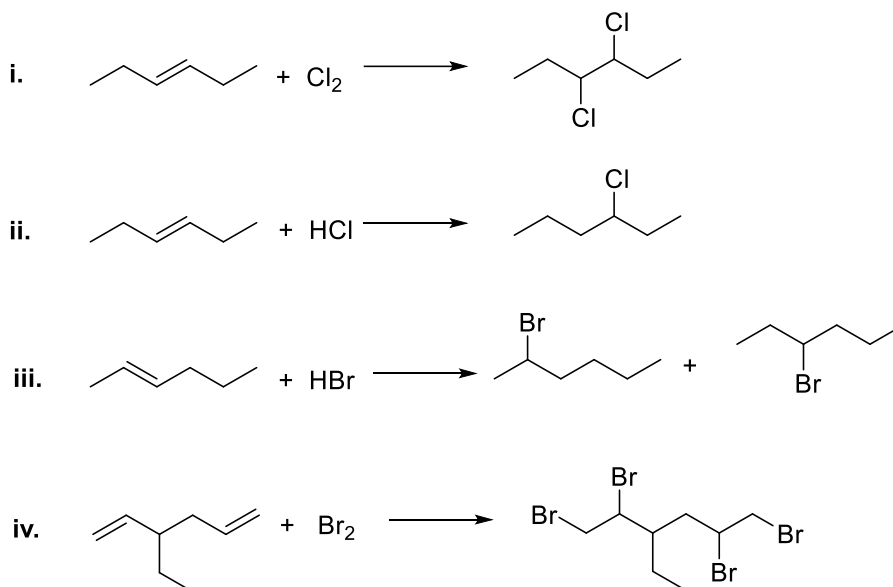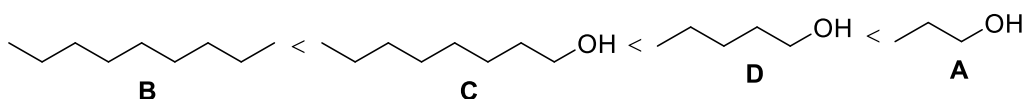
Problem 4

a.

i	ii	iii	iv
4-bromo-3-methylheptane (4-brom-3-metylheptan)	1-ethyl-3-pentylcyclohexane (1-etyl-3-pentylsykloheksan)	5-ethyl-6-propyldecane (5-etyl-6-propyldekan)	3-ethylhexa-1,5-diene (3-etylheksan-1,5-dien)

b.

i.

ii.

iii.

iv.

c. The boiling point for isopropanol is higher because the interactions between isopropanol molecules are stronger than the interactions between butane molecules.

d.

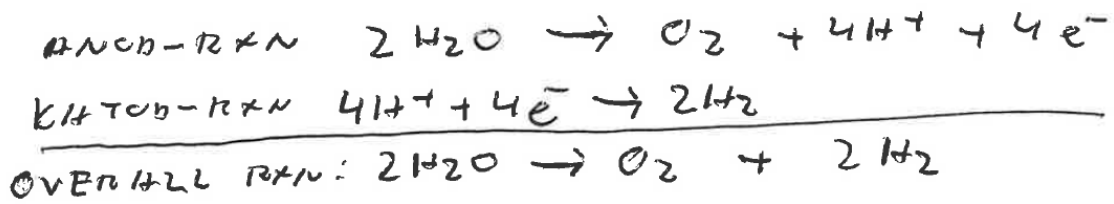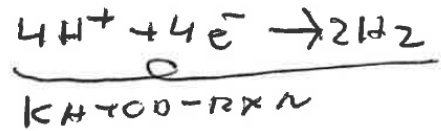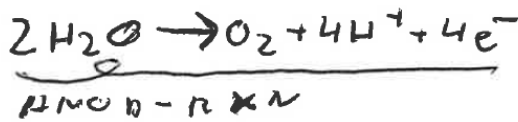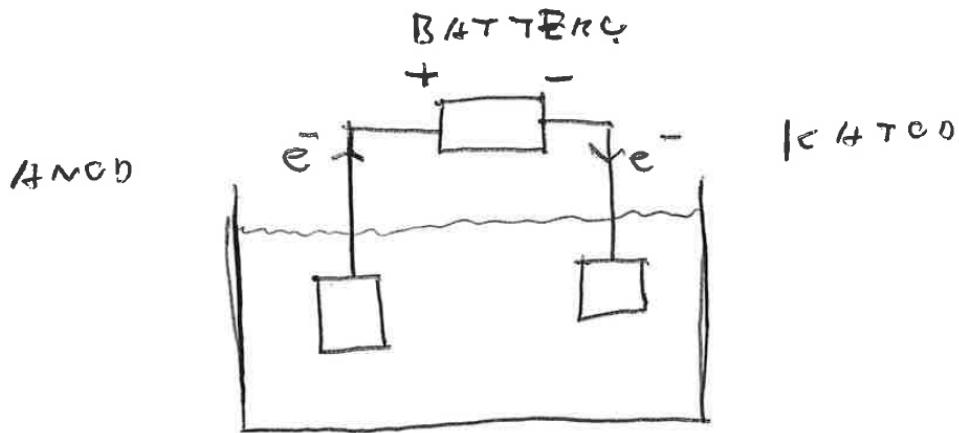**B** < **C** OH < **D** OH < **A** OH

Problem 5

a. The source of O_2 in the Chapman cycle is the photosynthesis. The production of ozone in the chapman cycle is presented below:

    i. O_2 + UV-light → 2O

    ii. $O + O_2 → O_3$

b. Because the amount of incoming solar light is largest in the stratosphere above the tropics. The UV-light from the sun is necessary for ozone to formed (see step i in **Problem 5a**)

c.

    i. $CF_2Cl_2 → CF_2Cl + Cl$

    ii. $O_2 + Cl → O + ClO$

    iii. $ClO + O_3 → Cl + 2O_2$

d. Chlorofluorocarbons (CFCs)

Problem 6

a. An endotherm reaction absorbs heat when it occurs and an exothermic reaction releases heat.

b. For example, when a heating pump is used to heat a house, the refrigerant is vaporized (**endothermic process**) in the part of the coil outside the house at low pressure. Then, the vaporized refrigerant is moved to the part of the coil inside the house where it undergoes condensation (exothermic process). The heat released from the condensation is used to heat the house.
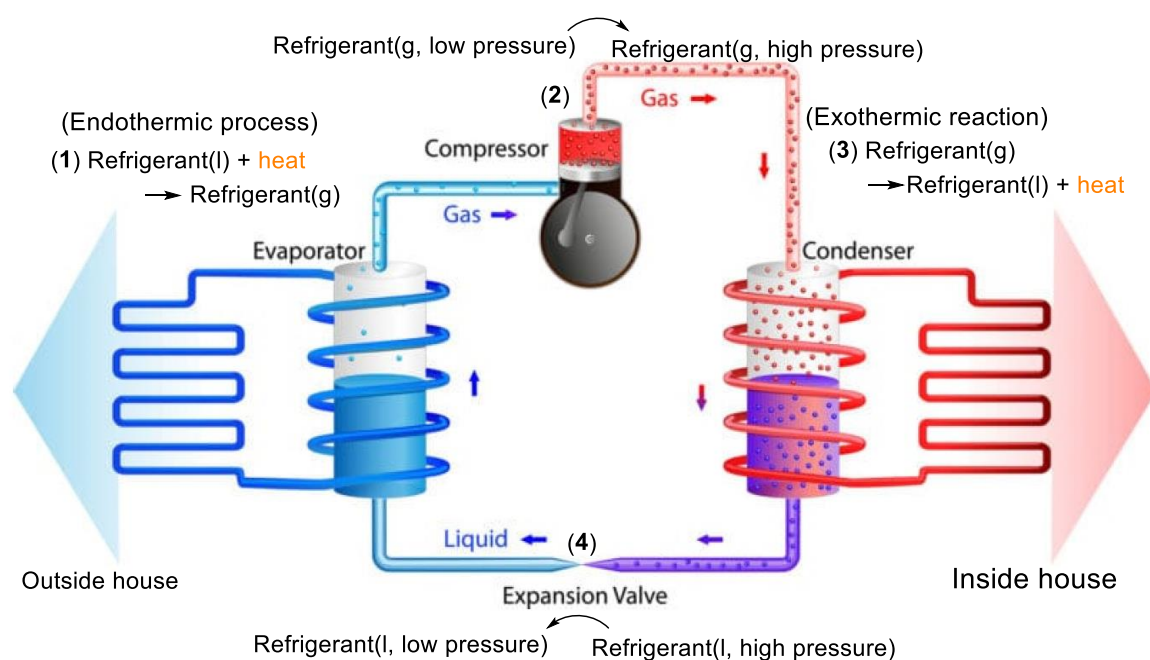
c. The wind energy that provides kinetic energy to the wind turbines is a converted form of solar energy.

d. $H_2(g)$ can be produced by the electrolysis of water in an electrolytic cell.

BATTERY

$+$ $-$

 e⁻   e⁻

ANOD

CATOD

$$2H_2O \rightarrow O_2 + 4H^+ + 4e^-$$
ANOD-RXN

$$4H^+ + 4e^- \rightarrow 2H_2$$
KATOD-RXN

ANOD-RXN $2H_2O \rightarrow O_2 + 4H^+ + 4e^-$

KATOD-RXN $4H^+ + 4e^- \rightarrow 2H_2$

OVERALL RXN: $2H_2O \rightarrow O_2 + 2H_2$

Problem 7

a. Primary pollutants are formed directly from the source whereas secondary pollutants are formed from primary pollutants.

b. Nitrogen oxides (NOx) and sulfur dioxide (SO_2)

c. Ozone is harmful to living organisms when it is present in the troposphere

d. Volatile organic compounds (VOCs) and nitrogen oxides (NOx).