Oxidation Reduction Reactions- Answer Key

4.43 Even though the problem doesn't ask you to assign oxidation numbers, you need to be able to do so in order to determine what is being oxidized or reduced.

	(i) Half Reactions	(ii) Oxidizing Agent	(iii) Reducing Agent
(a)	$Sr \rightarrow Sr^{2+} + 2e^-$ $O_2 + 4e^- \rightarrow 2O^{2-}$	O_2	Sr
(b)	$Li \rightarrow Li^+ + e^-$ $H_2 + 2e^- \rightarrow 2H^-$	H_2	Li
(c)	$Cs \rightarrow Cs^+ + e^-$ $Br_2 + 2e^- \rightarrow 2Br^-$	Br_2	Cs
(d)	$Mg \rightarrow Mg^{2+} + 2e^-$ $N_2 + 6e^- \rightarrow 2N^{3-}$	N_2	Mg

4.45 The oxidation number for hydrogen is +1 (rule 4), and for oxygen is −2 (rule 3). The oxidation number for sulfur in S_8 is zero (rule 1). Remember that in a neutral molecule, the sum of the oxidation numbers of all the atoms must be zero, and in an ion the sum of oxidation numbers of all elements in the ion must equal the net charge of the ion (rule 6).

$$H_2S\ (-2),\ S^{2-}\ (-2),\ HS^-\ (-2) < S_8\ (0) < SO_2\ (+4) < SO_3\ (+6),\ H_2SO_4\ (+6)$$

The number in parentheses denotes the oxidation number of sulfur.

4.47 See Section 4.4 of the text.

(a) $\underline{Cl}F$: F −1 (rule 5), **Cl +1** (rule 6)

(b) $\underline{I}F_7$: F −1 (rule 5), **I +7** (rules 5 and 6)

(c) $\underline{C}H_4$: H +1 (rule 4), **C −4** (rule 6)

(d) $\underline{C}_2H_2$: H +1 (rule 4), **C −1** (rule 6)

(e) $\underline{C}_2H_4$: H +1 (rule 4), **C −2** (rule 6),

(f) $K_2\underline{Cr}O_4$: K +1 (rule 2), O −2 (rule 3), **Cr +6** (rule 6)

(g) $K_2\underline{Cr}_2O_7$: K +1 (rule 2), O −2 (rule 3), **Cr +6** (rule 6)

(h) $K\underline{Mn}O_4$: K +1 (rule 2), O −2 (rule 3), **Mn +7** (rule 6)

(i) $NaH\underline{C}O_3$: Na +1 (rule 2), H +1 (rule 4), O −2 (rule 3), **C +4** (rule 6)

(j) $\underline{Li}_2$: **Li 0** (rule 1)

(k) $Na\underline{I}O_3$: Na +1 (rule 2), O −2 (rule 3), **I +5** (rule 6)

(l) $K\underline{O}_2$: K +1 (rule 2), **O −1/2** (rule 6)

(m) $\underline{P}F_6^-$: F −1 (rule 5), **P +5** (rule 6)

(n) $K\underline{Au}Cl_4$: K +1 (rule 2), Cl −1 (rule 5), **Au +3** (rule 6)

4.49 **(a)** $\underline{Cs}_2O$, **+1** **(b)** $Ca\underline{I}_2$, **−1** **(c)** $\underline{Al}_2O_3$, **+3** **(d)** $H_3\underline{As}O_3$, **+3** **(e)** $\underline{Ti}O_2$, **+4**
(f) $\underline{Mo}O_4^{2-}$, **+6** **(g)** $\underline{Pt}Cl_4^{2-}$, **+2** **(h)** $\underline{Pt}Cl_6^{2-}$, **+4** **(i)** $\underline{Sn}F_2$, **+2** **(j)** $\underline{Cl}F_3$, **+3**
(k) $\underline{Sb}F_6^-$, **+5**

4.51 If nitric acid is a strong oxidizing agent and zinc is a strong reducing agent, then zinc metal will probably reduce nitric acid when the two react; that is, N will gain electrons and the oxidation number of N must decrease. Since the oxidation number of nitrogen in nitric acid is +5 (verify!), then the nitrogen-containing product must have a smaller oxidation number for nitrogen. The only compound in the list that doesn't have a nitrogen oxidation number less than +5 is N_2O_5, (what is the oxidation number of N in N_2O_5?). This is never a product of the reduction of nitric acid.

4.53 In order to work this problem, you need to assign the oxidation numbers to all the elements in the compounds. In each case oxygen has an oxidation number of −2 (rule 3). These oxidation numbers should then be compared to the range of possible oxidation numbers that each element can have. Molecular oxygen is a powerful oxidizing agent. In $\mathbf{SO_3}$ alone, the oxidation number of the element bound to oxygen (S) is at its maximum value (+6); the sulfur cannot be oxidized further. The other elements bound to oxygen in this problem have less than their maximum oxidation number and can undergo further oxidation.

4.55 **(a)** Disproportionation reaction **(b)** Displacement reaction
(c) Decomposition reaction **(d)** Combination reaction

4.57 **(b)** $\overset{+6}{V}O_2 \rightarrow \overset{+4}{V}O_2$

(d) $\overset{+3}{N}O_2^- \rightarrow \overset{+5}{N}O_3^-$

(e) $Cr^{3+} \rightarrow \overset{+6}{Cr}O_4^{2-}$