

Electrochemistry

Section 104 – Review of Redox Chemistry

104-1 Assign oxidation numbers to each element in the following substances:

- (a) H₂O
- (b) CO₂
- (c) NH₄Cl
- (d) Na₂SO₄
- (e) H₃PO₄

Solution:

(a) The formal oxidation state for **O is -2; for H is +1.**

(b) Since O is -2, the C can be determined to be +4 as follows:

$$C + 2(-2) = 0$$

$$C - 4 = 0$$

$$C = +4$$

Thus the oxidation states are as follows: O is -2, C is +4.

(c) Cl is -1; the ammonium ion is NH₄⁺¹ to balance. H is +1 (unless bonded to a metal), so the oxidation state of N in the ammonium ion can be determined as follows:

$$N + 4(+1) = +1$$

$$N + 4 = +1$$

$$N = +1 - 4 = -3$$

Thus the oxidation states are as follows: Cl is -1, H is +1 and N is -3.

(d) Since sulfate (SO₄⁻²) has an overall charge of -2, the two sodium ions must each be +1 (which is consistent with what would be expected for a Group 1 metal). Oxygen is normally -2, so the oxidation state of sulfur can be determined as follows:

$$S + 4(-2) = -2$$

$$S = -2 + 8 = +6.$$

Thus the oxidation numbers are as follows: Na is +1, S is +6, O is -2.

(e) The oxidation state of H is +1; of O is -2. The phosphorus in the phosphate (PO₄⁻³) can be determined as follows:

$$P + 4(-2) = -3$$

$$P - 8 = -3$$

$$P = -3 + 8 = +5.$$

Thus the oxidation numbers are as follows: H is +1, O is -2, P is +5.

104-2 Assign oxidation numbers to each element in the following substances:

- (a) NaOCl,
- (b) ClF₃
- (c) KH
- (d) I₂
- (e) Na

Solution:

(a) Na (a group 1 element) is +1; the hypochlorite ion OCl⁻ has an overall charge of -1. Since O is -2, the oxidation state of Cl can be determined as follows:

$$\text{Cl} + (-2) = -1.$$

$$\text{Cl} = -1 + 2 = +1.$$

Thus the oxidation states are Na +1; O -2; Cl +1.

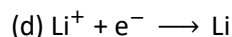
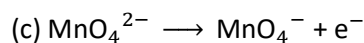
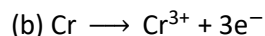
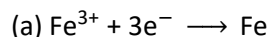
(b) In ClF₃, F is -1 since it is the most electronegative element. Thus the Cl must be +3 since the overall charge is 0. Thus the oxidation states are as follows: F is -1; Cl is +3.

(c) Since potassium is a group I metal, K has an oxidation state of +1 while H is -1.

(d) In I₂ the oxidation state of iodine is 0.

(e) In Na the oxidation state of Na is 0 (since the overall charge is 0, it is not in ionic form).

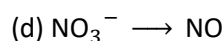
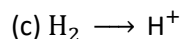
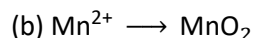
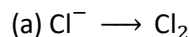
104-3 Identify each half-reaction below as either oxidation or reduction.



Solution

- (a) reduction
- (b) oxidation
- (c) oxidation
- (d) reduction

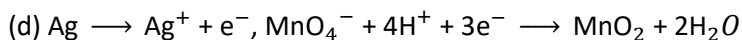
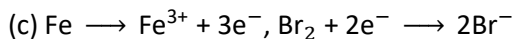
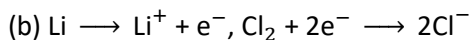
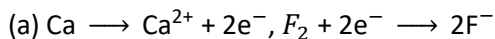
104-4 Identify each half-reaction below as either oxidation or reduction.



Solution

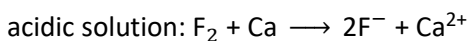
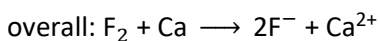
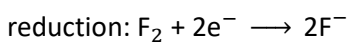
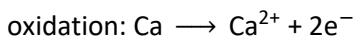
- (a) oxidation
- (b) oxidation
- (c) oxidation
- (d) reduction

104-5 Assuming each pair of half-reactions below takes place in an acidic solution, write a balanced equation for the overall reaction.

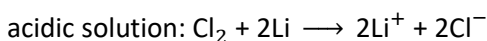
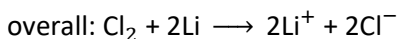
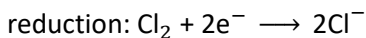
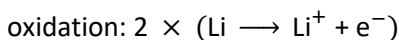


Solution

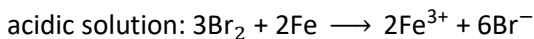
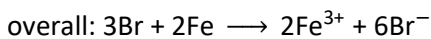
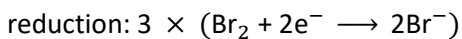
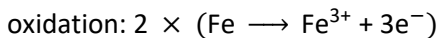
(a)



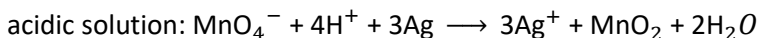
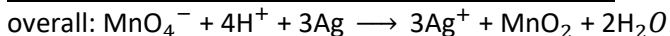
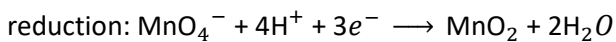
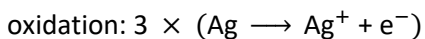
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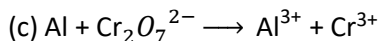
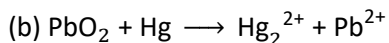
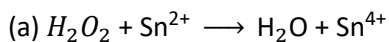
(c)



(d)

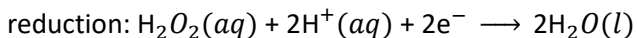
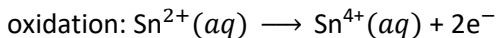


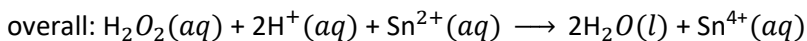
104-6 Balance the equations below assuming they occur in an acidic solution.



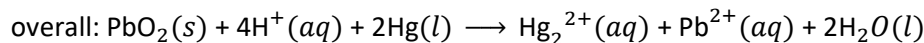
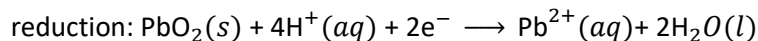
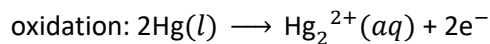
Solution

(a)

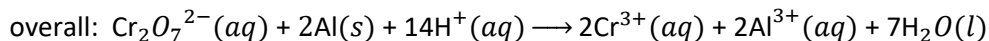
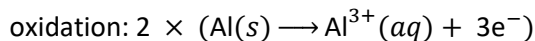
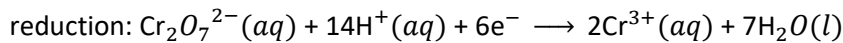




(b)



(c)



104-7 Identify the oxidant, reductant, oxidizing agent and reducing agent in each reaction of the previous exercise.

Solution

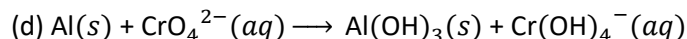
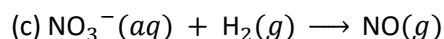
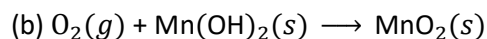
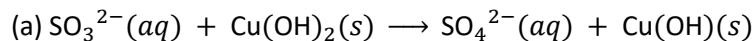
Oxidized: (a) Sn^{2+} ; (b) Hg; (c) Al,

reduced: (b) H_2O_2 ; (b) PbO_2 ; (c) $\text{Cr}_2\text{O}_7^{2-}$,

oxidizing agent: (a) H_2O_2 ; (b) PbO_2 ; (c) $\text{Cr}_2\text{O}_7^{2-}$,

reducing agent: (a) Sn^{2+} ; (b) Hg; (c) Al.

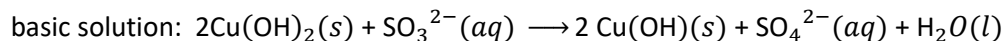
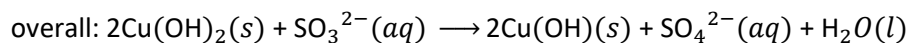
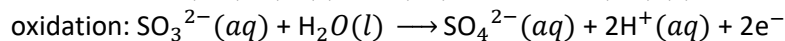
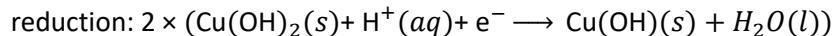
104-8 Balance the equations below assuming they occur in a basic solution.



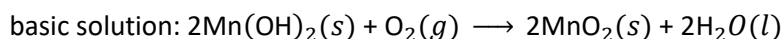
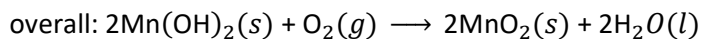
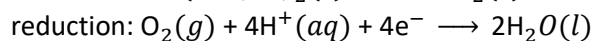
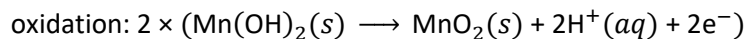
Solution

First balance in acidic solution, then “convert” to basic solution:

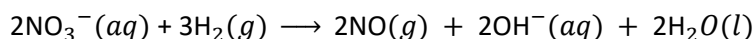
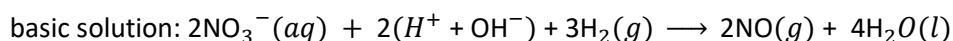
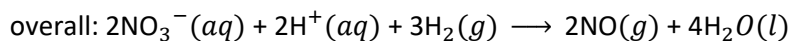
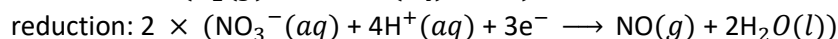
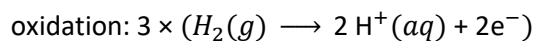
(a)



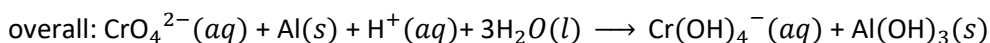
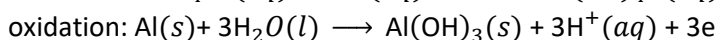
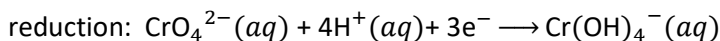
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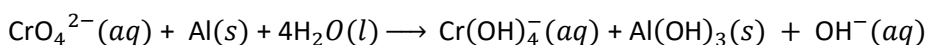
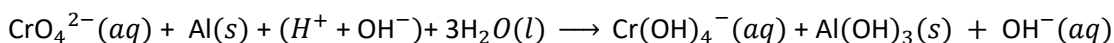
(c)



(d)



basic solution:



104-9 Identify the oxidant and reductant of each reaction of the previous exercise.

Solution

Oxidized = reducing agent: (a) SO_3^{2-} ; (b) $Mn(OH)_2$; (c) H_2 ; (d) Al ;

reduced = oxidizing agent: (a) $Cu(OH)_2$; (b) O_2 ; (c) NO_3^- ; (d) CrO_4^{2-}

104-10 Why don't hydroxide ions appear in equations for half-reactions occurring in acidic solution?

Solution

In acidic solution, $[H^+] > 1 \times 10^{-7}M > [OH^-]$. Hydroxide ion cannot appear as a reactant because its concentration is essentially zero. If it were produced, it would instantly react with the excess hydrogen ion to produce water. Thus, hydroxide should *not* appear as a reactant or product in acidic solution.

104-11 Why don't hydrogen ions appear in equations for half-reactions occurring in basic solution?

Solution

In basic solution, $[OH^-] > 1 \times 10^{-7}M > [H^+]$. Hydrogen ion cannot appear as a reactant because its concentration is essentially zero. If it were produced, it would instantly react with the excess hydroxide ion to produce water. Thus, hydrogen ion should *not* appear as a reactant or product in basic solution.

104-12 Why must the charge balance in oxidation-reduction reactions?

Solution

If the charge does not balance, then there would be a different number of electrons on each side of the equation. This would violate the law of conservation of charge.