Chemistry BC2001x: General Chemistry I

Lecture 18: Tuesday November 17, 2009

Topic: Oxidation Reduction Chemistry: Oxidation states

Pick up:

- **Oxidation-Reduction:**
  states and balancing equations
  representative states
- **Graded problem set 8**

Determining oxidation states

1) **All atoms in a pure element have oxidation state 0.**
2) **The oxidation state of the element in a monatomic ion is equal to the charge on the ion.**
3) **The oxidation state of hydrogen, H, in any ion or molecule (other than H_2) is +1,** with two exceptions:
   (i) The oxidation state of hydrogen is –1 in metal hydrides.
   (ii) The oxidation state of hydrogen is –1 when covalently bound to B in various compounds.
4) **The oxidation state of oxygen, O, in any ion or molecule (other than O_2 or O_3) is –2,** with three exceptions:
   (i) The covalent molecule OF_2, in which the oxidation state of O is +2 because the oxidation state of F is always –1 (except in F_2). F is the only element more electronegative than O.
   (ii) Peroxides, in which the oxidation state of O is –1.
   (iii) Superoxides, in which the oxidation state of O is –½.
5) **a. In all covalently bonded molecules or ions that do not contain hydrogen or oxygen, the more electronegative element is assigned its common negative oxidation state: –3 for N or P; –2 for S, Se, or Te; –1 for F, Cl, Br, I.**
   b. If the molecule or ion does contain H and/or O as well as several other elements, assign +1 to H, assign –2 to O, and assign the most common negative oxidation state to the most electronegative other element, in order to assign the remaining oxidation state.
6) **The sum of all the oxidation numbers must equal the total charge on an ion or molecule,** namely, the ionic charge for an ion, or zero for a neutral molecule.

When working with ionic substances, determine oxidation states separately in the two ions.

Representative oxidation states

NITROGEN (Group VA or 15) Maximum (highest) oxidation state: +5 Minimum: –3

+5	NO_3^-	Nitrate ion. A strong oxidizing agent in acid solution.
+4	NO_2	Nitrogen dioxide, a reddish-brown gas. (Dimerizes to colorless N_2O_4.) NO_2 is the principal product of the reduction of *concentrated* HNO_3.
+3	NO_2^-	Nitrite ion (in basic solution).
	HNO_2	Nitrous acid (in acidic solution).
+2	NO	Nitric oxide, a colorless gas. Reacts rapidly with O_2 to form NO_2. NO is the principal product of the reduction of *dilute* HNO_3.
+1	N_2O	Nitrous oxide (laughing gas). Seldom formed in oxidation-reduction reactions.
0	N_2	Nitrogen gas. (Also called dinitrogen.) Very stable and unreactive.
–1	$HONH_2$	Hydroxylamine. A weak base.
–2	N_2H_4	Hydrazine. A weak base.
–3	NH_3	Ammonia (in basic solution). A colorless gas, very soluble in water.
	NH_4^+	Ammonium ion (in acidic solution).

SULFUR (Group VIA or 16) Maximum (highest) oxidation state: +6 Minimum: –2

+6	SO_3	Sulfur trioxide. Dissolves in water to form the strong acid H_2SO_4.
	SO_4^{2-}	Sulfate ion.
+4	SO_3^{2-}	Sulfite ion. Easily oxidized to sulfate ion.
	SO_2	Sulfur dioxide. Dissolves in water to form the weak acid H_2SO_3.
0	S	Elemental sulfur. Actually exists as 8-membered rings $S_8(s)$.
–2	S^{2-}	Sulfide ion (in basic solution).
	H_2S	Hydrogen sulfide (in acidic solution). A gas, soluble 0.1 M in H_2O.

Balancing redox equations in aqueous solution

Recommended single procedure that works for both acidic and basic solutions

I. Write balanced **half-reactions** for the oxidation and the reduction, following the steps below.

For each half-reaction, initially written in skeletal form:

1. Determine the **oxidation states** of the element that is being oxidized or reduced, in the form in which it appears as a reactant and in the form in which it appears as a product.
2. Balance the **number of atoms** of the element being oxidized or reduced.
3. Calculate the total **number of electrons** lost or gained in the half-reaction, and add that number of electrons to the correct side of the half-reaction.
4. Calculate the net **charge**, including the added electrons, on each side of the half-reaction. At this point, in general, the net charge on each side of the half-reaction will be different.
   (a) If the reaction is occurring in an **acidic** or in a **neutral** solution, balance the net charge, including electrons, by adding $\mathbf{H^+}$(aq) ions [or H_3O^+(aq) ions, if you prefer] to one side.
   (b) If the oxidation-reduction reaction is occurring in **basic** solution balance the net charge, including electrons, by adding $\mathbf{OH^-}$(aq) ions to one side.
5. Balance the total number of **O atoms** on each side of the half-reaction by adding $\mathbf{H_2O}$ to one side. (This step will simultaneously balance the number of H atoms on each side.)
6. The half-reaction should now be complete and balanced. Before proceeding any further:
   (a) Check that the numbers of atoms of every element are the same on each side.
   (b) Check that the total net charge is the same on each side of the half-reaction.
   If everything is not in balance, go back to the beginning and find the error.

II. Combine the two half-reactions to yield the net overall oxidation-reduction reaction equation.

1. Multiply each half-reaction by a integer such that the total number of electrons **lost** in the oxidation equals the total number of electrons **gained** in the reduction.
2. Add the half-reactions; the electrons cancel. **No electrons should appear in any overall balanced oxidation-reduction equation.**
3. Check to see if H^+, or OH^-, or H_2O appears on both sides of the equation. If so, **simplify**.
4. The overall oxidation-reduction reaction should now be complete and balanced. Test this:
   (a) Check that the numbers of atoms of every element are the same on each side.
   (b) Check that the total net charge is the same on each side of the overall equation.