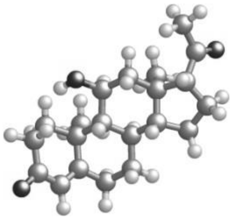

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<u>Oxidation–Reduction Reactions: Redox</u>

Part A – Oxidation Numbers:

Oxidation numbers (also called oxidation states) provide an easy way to identify oxidation-reduction reactions, and to know which is the oxidizing agent, and which is the reducing agent. The *oxidation number* of an atom in a molecule or ion is defined as the charge an atom has, or appears to have – that is, the hypothetical charged assigned to the atom, assuming that the electrons are completely held by one atom or the other. Oxidation occurs when there is an increase in oxidation number, whereas reduction occurs when there is a decrease in oxidation number.

We use the following rules for assigning oxidation numbers. Following each of the guidelines below, indicate the oxidation number of the italicized element in the box next to the compound.

1) Each atom in a pure element has an oxidation number of zero:

S_8 ☐ P_4 ☐ H_2 ☐

2) For monoatomic ions, the oxidation number is equal to the charge on the ion. (Note that in writing oxidation numbers, we usually write the sign before the number.):

Fe^{3+} ☐ Na^{+} ☐ S^{2-} ☐

3) The oxidation number of H is +1 when bonded to non-metals, and –1 when bonded to metals (hydrides.):

HCl ☐ CaH_2 ☐ H_2S ☐

4) The oxidation number of O is usually –2 in both ionic and molecular compounds. The major exception is in peroxides, which contain the O_2^{2-} ion:

H_2O ☐ H_2O_2 ☐ OH^- ☐

5) F always has an oxidation number of –1 in compounds.

6) C*l*, Br and I always have oxidation number of –1 in compounds, unless combined with O or F:

Na*Cl* ☐ *Cl*O$^-$ ☐ K*I* ☐

7) In their compounds, Group 1 metals have an oxidation number of +1, Group 2 metals have an oxidation number of +2 and Group 13 metals have an oxidation number of +3:

*K*Br ☐ *Mg*SO_4 ☐ *Al*$_2O_3$ ☐

8) The sum of oxidation numbers in a neutral compound is 0, and in a polyatomic ion it is equal to the charge on the ion:

SO_4^{2-} ☐ $H_3\mathit{PO}_4$ ☐ $Na_2\mathit{SO}_3$ ☐

• Now, practice the rules by finding the oxidation state of the boldfaced element in each of the following:

$\mathbf{P}_2O_5$ ☐ Na**H** ☐ $\mathbf{Cr}_2O_7^{2-}$ ☐

SnBr_4 ☐ Ba$\mathbf{O}_2$ ☐ $\mathbf{N}O_2^+$ ☐

Part B – Recognising Redox Reactions:

For each reaction below, determine which atoms are undergoing a change in oxidation number, identify the oxidising and reducing agents, and *name* the compounds. Remember, an *oxidizing agent* is itself *reduced*, and a *reducing agent* is itself *oxidized*.

$$5Fe^{2+}(aq) + MnO_4^-(aq) + 8H^+(aq) \rightarrow 5Fe^{3+}(aq) + Mn^{2+}(aq) + 4H_2O(l)$$

$$Cu(s) + 2NO_3^-(aq) + 4H^+(aq) \rightarrow Cu^{2+}(aq) + 2NO_2(g) + 2H_2O(l)$$

$$Fe_2O_3(s) + 2Al(s) \rightarrow 2Fe(l) + Al_2O_3(s)$$

$$Cu(s) + 2Ag^+(aq) \rightarrow Cu^{2+}(aq) + 2Ag(s)$$

Part C – Balancing Redox Reactions:

When writing redox reactions, all equations must be balanced for both mass and charge. The number of electrons lost (or produced) in oxidation must equal the number of electrons gained (or consumed) in reduction. The half-reaction method is a systematic way to balance complex redox reactions.

Balance the following redox reaction using the half-reaction method, following each step *carefully*.

$$Cr_2O_7^{2-}(aq) + I^-(aq) \rightarrow Cr^{3+}(aq) + I_2(aq)$$

Step 1 – Is this a redox reaction? Determine which atoms are undergoing a change in oxidation number and identify the oxidising and reducing agents.

Oxidising agent (is reduced):

Reducing agent (is oxidized):

Step 2 – Separate the reaction into an oxidation half-reaction, and a reduction half-reaction.

Reduction:

Oxidation:

Step 3 – Balance each half-reaction for mass (same number of elements on each side, besides O and H). Balance the number of O atoms by adding H_2O. Balance the H atoms by adding H^+.

Reduction:

Oxidation:

Step 4 – Balance each half-reaction for charge by adding electrons either on the reactant or the product side.

Reduction:

Oxidation:

Step 5 – Multiply each half-reaction by an appropriate factor to obtain the same number of electrons in both half-reactions.

Reduction:

Oxidation:

Step 6 – Add the two half-reactions and cancel any substances that appear on both sides. <u>All the electrons must cancel out by this step</u>.

Step 7 – Check to make sure all atoms and charges are balance.

Important reminders:

Notice that this redox reaction occurs in an acidic medium because you added H^+ ions to balance the hydrogen atoms. This applies to the majority of redox reactions you will encounter. Sometimes, redox reactions occur in a *basic medium*. In such cases, you must add OH^- to balance the hydrogen atoms. Never add H atoms or H_2 to balance the equation.

Never add O atoms, O^{2-} ions or O_2 to balance the oxygen atoms. Add only H_2O (acidic medium) or OH^- (basic medium).

By the way, here's a clue. A redox reaction can occur in an acidic *OR* a basic medium, *never both* at the same time. Also, if one of the reactants or products is an acid, the reaction must occur in an acidic medium (add H^+ / H_2O). Likewise, if one of the reactants or products is a base (e.g. hydroxides), the reaction must occur in a basic medium (add H_2O / OH^-).

The best way to become competent in balancing redox reactions is to practice, practice, practice!

Part D – Practice Questions

Balance the following redox reactions using the half-reaction method. Unless otherwise noted, all reactions occur in an acidic medium.

1) $CuS(s) + HNO_3(aq) \rightarrow NO(g) + CuSO_4(aq)$

2) $MnO_4^-(aq) + HSO_3^-(aq) \rightarrow Mn^{2+}(aq) + SO_4^{2-}(aq)$

3) $MnO_2(s) + Cl^-(aq) \rightarrow Mn^{2+}(aq) + Cl_2(g)$

4) $CH_2O(aq) + Ag^+(aq) \rightarrow HCO_2H(aq) + Ag(s)$

5) $Al(s) + Cu^{2+}(aq) \rightarrow Al^{3+}(aq) + Cu(s)$

6) $Al(s) + OH^-(aq) \rightarrow Al(OH)_4^-(aq) + H_2(g)$ in basic medium

7) $Fe(OH)_3(s) + Cr(s) \rightarrow Cr(OH)_3(s) + Fe(OH)_2(s)$ in basic medium

8) $Sn(s) + H^+(aq) \rightarrow Sn^{2+}(aq) + H_2(g)$

9) $Zn(s) + NO_3^-(aq) \rightarrow Zn^{2+}(aq) + N_2O(g)$

10 a) $VO_2^+(aq) + Zn(s) \rightarrow VO^{2+}(aq) + Zn^{2+}(aq)$

10 b) $VO^{2+}(aq) + Zn(s) \rightarrow V^{3+}(aq) + Zn^{2+}(aq)$

10 c) $V^{3+}(aq) + Zn(s) \rightarrow V^{2+}(aq) + Zn^{2+}(aq)$

Equations **10a**, **10b** and **10c** demonstrate the multiple oxidation states of vanadium, a transition metal. Each oxidation state has its own distinctive colour, which starts from yellow (VO_2^+) to blue (VO^{2+}), to green (V^{3+}) and finally to violet (V^{2+}). What are the oxidation states of vanadium at each stage?

• Scan the QR code below for the answers to this assignment.

http://www.chemist.sg/redox/redox_worksheet_alternative_ans.pdf