Chapter 4 - Aqueous Reactions and Solution Stoichiometry

Section 3 - Redox Reactions

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Introduction - Types of Reactions

You have already learned previously about this type of classification of reactions: how **atoms are rearrangement**. Just to review there are four classifications which we will bring up again in this chapter.

- **Synthesis** (combination): two substances combine to form one.

  $2Na_{(s)} + Cl_{2(g)} \rightarrow 2NaCl_{(s)}$

- **Double Displacement**: A reaction in which two elements displaces two elements.

  $AgNO_{3(aq)} + NaCl_{(aq)} \rightarrow AgCl_{(s)} + NaNO_{3(aq)}$

- **Single displacement**: A reaction where one element displaces one other element.

  $Zn_{(s)} + CuSO_{4\,(aq)} \rightarrow ZnSO_{4(aq)} + Cu_{(s)}$

- **Decomposition**: A reaction in which a single compound reacts to give two or more substances.

  $2HgO_{(s)} \rightarrow 2Hg_{(l)} + O_{2(g)}$

Types of Reactions – Chemical Type

Now we will focus on reactions in water/aqueous system. Dissociation occurs in ionic substances in water and the ions produced can react with each other as they are all in solution now. Below are the reactions we will cover.

Precipitation Reactions: In this a solid is formed when two solutions are mixed. This means that one substance combination in the solution is not soluble in water.

Neutralization Reactions: Reaction between an acid and base. This usually results in forming salt and water.

Oxidation–Reduction Reactions: One can recognize these reactions by seeing if oxygen or hydrogen is being added or removed. OR check if there is a transfer of electrons from one element/ion to another.

All these reactions have water as a solvent.

Oxidation Reduction (*Redox*) Reactions

Oxidation	Reduction
Addition of oxygen	Removal of oxygen
Removal of hydrogen	Addition of hydrogen
Loss of electrons (LEO)	Gain of electrons (GER)
Metals lose electrons hence undergo oxidation	Nonmetals gain electrons hence undergo reduction
Reducing agents – an element that causes reduction of another element and gets oxidized (loses electrons) itself	Oxidizing agent – an element that causes oxidation of another element and gets reduced (gains electrons) itself

Writing Redox Reactions

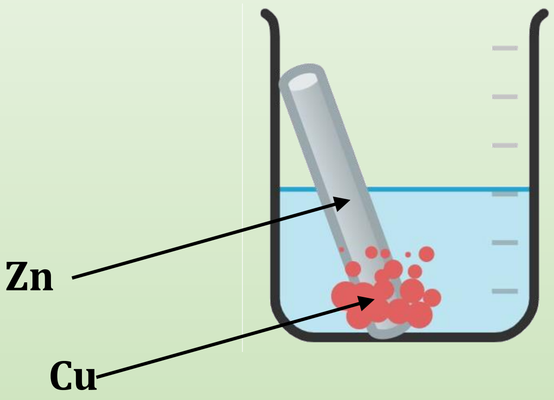
Example

$$Zn_{(s)} + Cu^{2+}_{(aq)} \rightarrow Zn^{2+}_{(aq)} + Cu_{(s)}$$

gains 2 electrons, *reduction*

$$Zn_{(s)} + Cu^{2+}_{(aq)} \rightarrow Zn^{2+}_{(aq)} + Cu_{(s)}$$

loses 2 electrons, *oxidation*



Another Example:

$$Mg_{(s)} + O_{2\ (g)} \rightarrow MgO_{(s)}$$

Mg is adding oxygen hence getting oxidized; so oxygen must be getting reduced.

But can you do this with electrons?? Yes…but we have to learn the rules to finding the oxidation numbers (valencies or charges on any element or ion).

Rules for Assigning Oxidation Numbers

1. **Elements:** The oxidation number of an element in their natural state is zero e.g., Mg, Zn, Na, Cr etc. for metals. All diatomic gases (H_2, O_2, N_2 and halogens) are zero.
2. **Monatomic ions:** The oxidation number of a monatomic ion equals the charge on the ion e.g., Na^+ is +1, Mg^{2+} is +2, Fe^{3+} is +3 etc.
3. **Oxygen:** The oxidation number of oxygen is always −2 except in H_2O_2 and other peroxides, where the oxidation number of oxygen is −1.
4. **Hydrogen:** The oxidation number of hydrogen is +1 in most compounds. The oxidation number of hydrogen is −1 when combined with metals such as CaH_2.
5. **Halogens:** The oxidation number of fluorine is −1. Each of the other halogens (Cl, Br, I) has an oxidation number of −1 in binary compounds e.g., HCl, NaBr. When the other element is another halogen above it in the periodic table or the other element is oxygen e.g., $HClO_3$.
6. **Compounds and ions:** The sum of the oxidation numbers of a compound is *zero*. The sum of the oxidation numbers of the atoms in a polyatomic ion equals the charge on the ion.

Oxidation Numbers on the Periodic Table

IA	IIA	IIIB	IVB	VB	VIB	VIIB		VIIIB		IB	IIB	IIIA	IVA	VA	VIA	VIIA
1 **H** +1, -1																
3 **Li** +1	4 **Be** +2											5 **B** +3	6 **C** +4 +2 -4	7 **N** +5 +4 +3 +2 +1 -3	8 **O** +2 +1/2 -1 -2	9 **F** -1
11 **Na** +1	12 **Mg** +2											13 **Al** +3	14 **Si** +4 -4	15 **P** +5 +3 -3	16 **S** +6 +4 +2 -2	17 **Cl** +7 +6 +6 +5 +4 +3 +1 -1
19 **K** +1	20 **Ca** +2	21 **Sc** +3	22 **Ti** +4 +3 +2	23 **V** +5 +4 +3 +2	24 **Cr** +6 +5 +4 +3 +2	25 **Mn** +7 +6 +4 +3 +2	26 **Fe** +3 +2	27 **Co** +3 +2	28 **Ni** +2	29 **Cu** +2 +1	30 **Zn** +2	31 **Ga** +3	32 **Ge** +4 -4	33 **As** +5 +3 -3	34 **Se** +6 +4 -2	35 **Br** +5 +3 +1 -1
37 **Rb** +1	38 **Sr** +2	39 **Y**	40 **Zr** +4	41 **Nb** +5 +4	42 **Mo** +6 +4 +3	43 **Tc** +7 +6 +4	44 **Ru** +8 +6 +4 +3	45 **Rh** +4 +3 +2	46 **Pd** +4 +2	47 **Ag** +1	48 **Cd** +2	49 **In** +3	50 **Sn** +4 +2	51 **Sb** +5 +3 -3	52 **Te** +6 +4 -2	53 **I** +7 +5 +1 -1
55 **Cs** +1	56 **Ba** +2	57 **La**	72 **Hf** +4	73 **Ta** +5	74 **W** +6 +4	75 **Re** +7 +6 +4	76 **Os** +8 +4	77 **Ir** +4 +3	78 **Pt** +4 +3	79 **Au** +3 +1	80 **Hg** +2 +1	81 **Tl** +3 +1	82 **Pb** +4 +2	83 **Bi** +5 +3	84 **Po** +2	85 **At** -1

Solved Problem: Assigning oxidation numbers

Assign oxidation numbers for all elements in the compounds below.

a) $MgBr_2$ and b) $KMnO_4$.

a) $MgBr_2$:

Mg +2,

Br –1 x 2= -2;

+2 +(-2) = total charge of 0

Oxidation number of Mg = +2 and Br = -1

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b) $KMnO_4$

K		Mn		O_4	
1 x	**+**	**1 y**	**+**	**4 z**	**=0**
1(+1)	+	y	+	4(−2)	= 0
+1	+	y	+	(−8)	= 0
+1	+	+7	+	(-8)	= 0

Oxidation number of K = +1, O = -2 and Mn = +7

Solved Problem: Assigning oxidation numbers in Ions

Assign oxidation numbers for elements underlined in the ions below.

a) **Cl**O_2^- b) **Cr**$_2O_7^{2-}$.

a) **Cl**O_2^-

Cl   2O = -1

x + 2y = -1

x + 2 (-2) = -1

x + (-4) = -1

+3

Oxidation number for Cl = + 3

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b) **Cr**$_2O_7^{2-}$?

2Cr + 7O = -2

$2(x) + 7(-2) = -2$

$2x + (-14) = -2$

Oxidation number of Cr = +6

2x = + 12; and x = 12/2 = +6

Solved Problem: Identifying oxidation reduction in reactions

Identify what is getting oxidized and reduced in the following reactions.

a) $FeI_{3(aq)} + Mg_{(s)} \rightarrow Fe_{(s)} + MgI_{2(aq)}$

b) $6FeSO_{4\,(aq)} + K_2Cr_2O_{7\,(aq)} + 7H_2SO_{4\,(aq)} \rightarrow$
$Cr_2(SO_4)_{3\,(aq)} + 3Fe_2(SO_4)_{3\,(aq)} + K_2SO_{4\,(aq)} + 7H_2O_{(l)}$

Calculate the oxidation numbers for all elements.

The oxidation numbers are given below the reaction.

a) $FeI_{3\,(aq)} + Mg_{(s)} \rightarrow Fe_{(s)} + MgI_{2\,(aq)}$

+3 −1 0 0 +2 −1

reduction oxidation

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b) $6\underline{\mathbf{Fe}}SO_{4\,(aq)} + K_2\underline{\mathbf{Cr_2}}O_{7\,(aq)} + 7H_2SO_{4\,(aq)} \rightarrow$
$\underline{\mathbf{Cr_2}}(SO_4)_{3\,(aq)} + 3\underline{\mathbf{Fe_2}}(SO_4)_{3\,(aq)} + K_2SO_{4\,(aq)} + 7H_2O_{(l)}$

Ignore H_2O and H_2SO_4 – focus on the transition metals.

Reactant Fe = +2 and Cr = +6; Product Fe = +3 and Cr = + 3

Fe is getting oxidized and Cr is getting reduced.

Types of Redox Reactions

Redox reactions can be of different types in terms of atom rearrangement type.

Single Displacement reactions

A common reaction: active metal replaces (displaces) a metal ion from a solution (use the activity series to predict if reaction will take place)

$$Mg_{(s)} + CuCl_{2(aq)} \rightarrow Cu_{(s)} + MgCl_{2(aq)}$$

Decomposition reactions

$$2KClO_{3\,(s)} \rightarrow 2KCl_{(s)} + 3O_{2\,(s)}$$

Combination Reactions

$$2H_{2\,(g)} + O_{2\,(g)} \rightarrow 2H_2O_{(l)}$$

Combustion reactions

Common example, hydrocarbon fuel reacts with oxygen to produce carbon dioxide and water

$$CH_{4\,(g)} + O_{2\,(g)} \rightarrow H_2O_{(l)} + CO_{2\,(g)}$$

Solved Example: Chemical reaction classification

Classify the following reactions as precipitation; acid-base; or redox reaction and any other classification that can describe the reaction.

$2H_{2\,(g)} + O_{2\,(g)} \rightarrow 2H_2O_{\,(l)}$

Redox (combustion, combination)

$Zn_{\,(s)} + H_2SO_{4\,(aq)} \rightarrow ZnSO_{4\,(aq)} + H_{2\,(g)}$

Redox (single displacement)

$H_2O_{\,(l)} + NH_3 \rightarrow NH_4^{+}{}_{(aq)} + OH^{-}{}_{(aq)}$

Acid-base (double displacement)

$2NaCl_{\,(aq)} + Pb(NO_3)_{2\,(aq)} \rightarrow PbCl_{2\,(s)} + 3NaNO_{3\,(aq)}$

Precipitation (double displacement)

Activity Series

Loses electrons easily

Activity series is used to predict which element will lose electrons (oxidize) and which will gain (reduce).
Group I metals give electrons easily whereas group VII elements gain easily.
Heavy metals are very stable and don't give oxidize easily.

Reducing agents – cause reduction, i.e., cause other elements to gain electrons. Metals are reducing agents.
Oxidizing agents – cause oxidation e.ie., cause other elements to lose electrons. Nonmetals are oxidizing agents.

Does not lose electrons easily

Element	Oxidation Half Reaction
Lithium	$Li \longrightarrow Li^{+} + 1e-$
Potassium	$K \longrightarrow K^{+} + 1e-$
Barium	$Ba \longrightarrow Ba^{+2} + 2e-$
Calcium	$Ca \longrightarrow Ca^{+2} + 2e-$
Sodium	$Na \longrightarrow Na^{+} + 2e-$
Magnesium	$Mg \longrightarrow Mg^{+2} + 2e-$
Zinc	$Zn \longrightarrow Zn^{+2} + 2e-$
Chromium	$Cr \longrightarrow Cr^{+3} + 3e-$
Iron	$Fe \longrightarrow Fe^{+2} + 2e-$
Cadmium	$Cd \longrightarrow Cd^{+2} + 2e-$
Cobalt	$Co \longrightarrow Co^{+2} + 2e-$
Nickel	$Ni \longrightarrow Ni^{+2} + 2e-$
Tin	$Sn \longrightarrow Sn^{+2} + 2e-$
Lead	$Pb \longrightarrow Pb^{+2} + 2e-$
Hydrogen	$H \longrightarrow H^{+} + 1e-$
Copper	$Cu \longrightarrow Cu^{+2} + 2e-$
Silver	$Ag \longrightarrow Ag^{+} + 1e-$
Mercury	$Hg \longrightarrow Hg^{+2} + 2e-$
Platinum	$Pt \longrightarrow Pt^{+2} + 2e-$
Gold	$Au \longrightarrow Au^{+2} + 2e-$

Key Words and Concepts

- Oxidation–Reduction Reactions
  - Oxidation
  - Reduction
  - Reducing agent
  - Oxidizing agent
  - Half reactions
  - Activity series