TYPES OF CHEMICAL REACTIONS

Direct Combination (Synthesis) Reactions

Two or more elements react to form a single product. Direct combination reactions can be represented by the general formula:

$$A + B \rightarrow AB$$

where A and B are elements.

The following are examples of direct combination reactions:

1. The reaction of a metal with oxygen to produce the oxide of the metal.
   Metal + Oxygen → Metal oxide
   $$2Mg_{(s)} + O_{2\,(g)} \rightarrow 2MgO_{(s)}$$
   Some metals combine with the oxygen in the air at room temperature; others have to be heated to form the oxides.

2. Heating a non-metal in air or oxygen to produce the non-metal oxide:
   Non-metal + Oxygen → Non-metal oxide
   $$S_{(s)} + O_{2\,(g)} \xrightarrow{heat} SO_{2\,(g)}$$
   $$C_{(s)} + O_{2\,(g)} \xrightarrow{heat} CO_{2\,(g)}$$

3. The reaction of a metal with a non-metal to produce a salt:
   Metal + Non-metal → Salt
   $$2Na_{(s)} + Cl_{2\,(g)} \rightarrow 2NaCl_{(s)}$$
   $$Fe_{(s)} + S_{(s)} \rightarrow FeS_{(s)}$$

Decomposition Reactions

In decomposition reactions, one substance undergoes a reaction to form two or more substances. The substance broken down is always a compound, and the products may be compounds or elements. Heat is often necessary for this process, in which case, the reaction may be described as thermal decomposition. The thermal decomposition of some metallic hydroxides, carbonates and nitrates are examples. Others include:

Loss of some or all of their water of crystallization by some hydrated salts on heating:

$$CuSO_4.5H_2O_{(s)} \xrightarrow{heat} CuSO_{4\,(s)} + 5H_2O_{(l)}$$

$$MgSO_4.7H_2O_{(s)} \xrightarrow{heat} MgSO_{4\,(s)} + 7H_2O_{(l)}$$

The decomposition of water by electrolysis is an example of decomposition which does not require heat.

Substitution (Displacement) Reactions

In a substitution reaction, one element displaces another element from a compound. They can be represented by the general equation:

$$A + BX \rightarrow AX + B$$

A and B may be metals or non-metals.

Displacement reactions can generally be divided into one of three types:

1. A more reactive metal displaces a less reactive metal from an aqueous solution of its salt.

$$Zn_{(s)} + CuSO_{4\,(aq)} \rightarrow ZnSO_{4\,(aq)} + Cu_{(s)}$$
$$Cu_{(s)} + 2AgNO_{3\,(aq)} \rightarrow Cu(NO_3)_{2\,(aq)} + 2Ag_{(s)}$$

2. A reactive metal displaces hydrogen from an acid or water.

$$Mg_{(s)} + 2HCl_{(aq)} \rightarrow MgCl_{2\,(aq)} + H_{2\,(g)}$$
$$Zn_{(s)} + 2HCl_{(aq)} \rightarrow ZnCl_{2\,(aq)} + H_{2\,(g)}$$
$$2Na_{(s)} + 2H_2O_{(l)} \rightarrow 2NaOH_{(aq)} + H_{2\,(g)}$$

For a metal to displace hydrogen ions from an acid, the metal must be above hydrogen in the reactivity series. Thus, both magnesium and zinc will react with hydrochloric acid to produce hydrogen gas, but copper with not react with dilute hydrochloric acid.

3. A more reactive non-metal displaces a less reactive non-metal

$$F_{2\,(g)} + 2NaCl_{(aq)} \rightarrow 2NaF_{(aq)} + Cl_{2\,(g)}$$
$$Cl_{2\,(g)} + 2KI_{(aq)} \rightarrow 2KCl_{(aq)} + I_{2\,(g)}$$

<u>Activity</u>

1. Complete and balance the following:
   a) $Fe_{(s)} + Cu(NO_3)_{2\,(aq)} \rightarrow$
   b) $Al_{(s)} + H_2SO_{4\,(aq)} \rightarrow$
   c) $Br_{2(l)} + 2KI_{(aq)} \rightarrow$

2. Complete and balance the following:
   a) $Ca_{(s)} + O_{2(g)} \rightarrow$
   b) $Mg_{(s)} + S_{(s)} \rightarrow$
   c) $2Al_{(s)} + 3Cl_2\,(g) \rightarrow$

3. Complete and balance the following decomposition reactions:
   a) $H_2O_{2\,(l)} \rightarrow$
   b) $2MgCO_{3(s)} \rightarrow$
   c) $BaCl_2.2H_2O_{(s)} \rightarrow$

Ionic Precipitation Reactions

Ionic precipitation reactions are sometimes referred to as double decomposition reactions. In these reactions, two soluble ionic compounds react to form one soluble and one insoluble compound. Ionic precipitation reactions are generally represented by:

$$AB + CD \rightarrow AD + CB$$

In these reactions, two compounds "exchange" ions/radicals.

You can predict whether or not a precipitate will form by studying the **solubility characteristics** of compounds. Examples of ionic precipitation reactions are shown below:

$$AgNO_{3\,(aq)} + NaCl_{(aq)} \rightarrow AgCl_{(s)} + NaNO_{3\,(aq)}$$
$$Ag^{+}_{(aq)} + Cl^{-}_{(aq)} \rightarrow AgCl_{(s)}$$

$$BaCl_{2\,(aq)} + H_2SO_{4\,(aq)} \rightarrow BaSO_{4\,(s)} + 2HCl_{(aq)}$$
$$Ba^{2+}_{(aq)} + SO^{2-}_{4\,(aq)} \rightarrow BaSO_{4\,(s)}$$

$$FeSO_{4\,(aq)} + 2NaOH_{(aq)} \rightarrow Fe(OH)_{2\,(s)} + Na_2SO_{4\,(aq)}$$

$$Fe^{2+}_{(aq)} + 2OH^{-}_{(aq)} \rightarrow Fe(OH)_{2\,(s)}$$

Neutralization Reactions

All neutralization reactions involve reaction between an acid (or acidic oxide) and a base. Water is produced in neutralization reactions; heat is also given out (Neutralization reactions are exothermic). You can investigate neutralization reactions in the laboratory by following the change in pH or in temperature as the acid is added to the alkali.

Reversible Reactions

Reversible reactions can proceed in both directions, that is, from reactants to products and vice versa. An important reversible reaction is the one between nitrogen and hydrogen, in a closed container, to produce ammonia.

$$N_{2(g)} + 3H_{2\,(g)} \leftrightarrow 2NH_{3\,(g)}$$

At first, the reaction proceeds in the forward direction only (from left to right). As the quantity of product increases, some of the ammonia dissociates and a backward reaction begins as well. The reaction proceeds in both directions until the rate of the forward reaction equals the rate of the backward reaction. At this point, the reaction is said to be in dynamic equilibrium. The reaction vessel will contain both products and reactants. This reaction is used in the industrial production of ammonia. In a reversible reaction, the conditions must be carefully controlled.

Reduction-Oxidation Reactions (Redox Reactions)

Redox reactions are chemical reactions in which oxidation and reduction take place at the same time. Oxidation and reduction are reactions that involve the loss and gain of electrons (respectively). Sometimes there is a total transfer of electrons (from metal to non-metal) but in some cases (between non-metals) there is only partial transfer.

Total Transfer of Electrons

Electron loss by an atom or ion is oxidation (OIL = oxidation is loss).

- When a sodium atom loses its valence electron, we say the sodium atom is oxidized to a sodium ion:

$$Na \rightarrow Na^{+} + e^{-}$$

- Similarly, when an iron (II) ion loses an electron, it is oxidized to an iron (III) ion:

$$Fe^{2+} \rightarrow Fe^{3+} + e^{-}$$

- Non-metal ions can also be oxidized to atoms:

$$2Cl^{-} \rightarrow Cl_2 + 2\,e^{-}$$

Each chloride ion loses an electron forming a chlorine atom. The chlorine atoms then pair up to form chlorine molecules.

Non-metals gain electrons when combining with metals. Electron gain by an atom or ion is reduction (RIG = reduction is gain).

When a chlorine atom gains an electron, we say the chlorine atom is reduced to the chloride ion:

$$Cl_2 + 2e^- \rightarrow 2Cl^-$$
$$Cl + e^- \rightarrow Cl^-$$

Each chlorine atom in the molecule gains an electron, forming a chloride ion.

Metal ions can also be reduced. Copper (II) ions can be reduced to copper atoms by gaining electrons:

$$Cu^{2+} + 2e^- \rightarrow Cu$$

Similarly, Fe^{3+} ions can be reduced to Fe^{2+} ions:

$$Fe^{3+} + e^- \rightarrow Fe^{2+}$$

In order for electron loss to occur, there must be an atom or ion to receive the electrons. Similarly, for electron gain to occur, there must be an electron donor. Therefore, reduction and oxidation occur simultaneously; this is where the name ‘redox’ comes from.

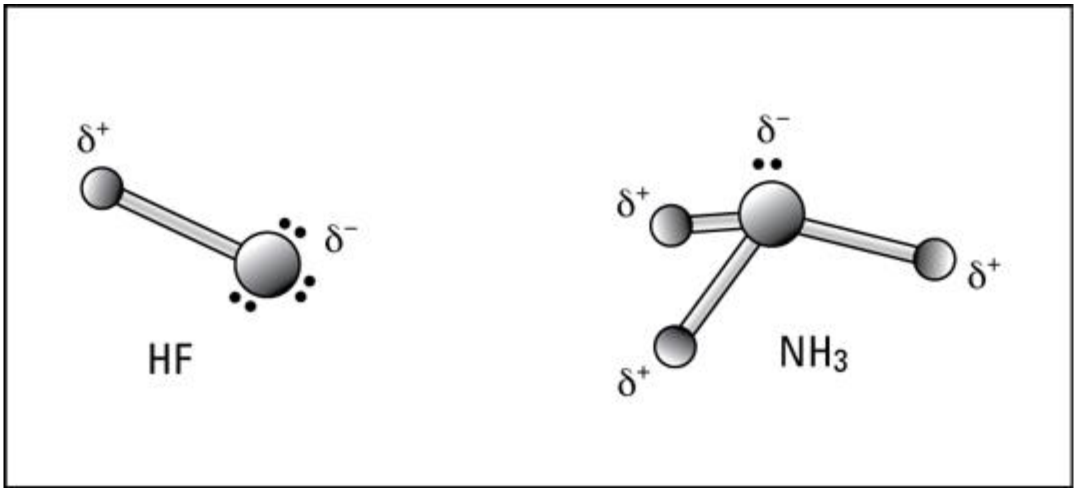
In the reaction between sodium metal and chlorine gas to form the ionic compound sodium chloride, sodium atoms lose electrons and chlorine atoms gain electrons:

$$2Na_{(s)} + Cl_{2\,(g)} \rightarrow 2Na^+Cl^-$$

The sodium atoms are oxidized to sodium ions, and, at the same time, the chlorine atoms are reduced to chloride ions.

Partial Transfer of Electrons

When two different non-metal atoms share electrons, the molecules may be polar, as the electrons are pulled closer to the more electronegative atom, which is thus reduced. The less electronegative atom in the compound partially loses the bonding electrons to the more electronegative atom and is oxidized.



In hydrogen fluorine, the more electronegative fluorine atom is reduced and the less electronegative hydrogen atom is oxidized.
In ammonia, the more electronegative nitrogen atom is reduced and the less electronegative hydrogen atoms are oxidized.

OXIDATION STATES OR OXIDATION NUMBERS

An oxidation state or oxidation number is a number assigned to an ionic compound or an atom in a molecule. The oxidation number is positive if electrons are totally or partially lost and is negative if electrons are partially or totally gained. The oxidation number of uncombined elements is zero.

- For ions, the oxidation state is the number of electrons lost or gained when the ion is formed from its element.
- For atoms in covalent compounds, oxidation state is the number of electrons that an atom partially gains or loses in a molecule.

When considering oxidation state, note the following:

- Oxidation state does not represent an actual charge on the atom and should not be confused, or used interchangeably with formal ionic charges.
- In describing oxidation states, the sign always precedes the number, e.g. the oxidation state of oxygen in the oxide ion –2, while the formal charge is written as 2– and the oxidation state of Fe in Fe^{3+} is +3 while the charge is 3+.

The chemical notation indicating the charge and the oxidation number is as follows:

$$X^C \rightarrow\ C\ is\ the\ charge\ or\ oxidation\ number$$

e.g. Fe^{3+} (charge) or Fe^{+3} (oxidation number)

Rules for Determining Oxidation Numbers

Oxidation numbers are assigned to elements using these rules:

Rule 1: The oxidation number of an element in its free (uncombined) state is zero — for example, Al(s) or Zn(s). This is also true for elements found in nature as diatomic (two-atom) elements (e.g. H_2, O_2, N_2, F_2, Cl_2, Br_2, I_2) and for sulphur, found as S_8

Rule 2: The oxidation number of a monatomic (one-atom) ion is the same as the charge on the ion, for example: $Na^+ = +1$, $S^{2-} = -2$

Rule 3: The sum of all oxidation numbers in a neutral compound is zero. The sum of all oxidation numbers in a polyatomic (many-atom) ion is equal to the charge on the ion. This rule often allows chemists to calculate the oxidation number of an atom that may have multiple oxidation states, if the other atoms in the ion have known oxidation numbers.

Rule 4: The oxidation number of an alkali metal (Group I) in a compound is +1; the oxidation number of an alkaline earth metal (Group II) in a compound is +2.

Rule 5: The oxidation number of oxygen in a compound is usually –2. If, however, the oxygen is in a class of compounds called peroxides (for example, hydrogen peroxide), then the oxygen has an oxidation number of –1. If the oxygen is bonded to fluorine, the number is +1.

Rule 6: The oxidation state of hydrogen in a compound is usually +1. If the hydrogen is part of a binary metal hydride (compound of hydrogen and a metal), then the oxidation state of hydrogen is –1.

Rule 7: The oxidation number of fluorine is always –1. Chlorine, bromine, and iodine usually have an oxidation number of –1, unless they're in combination with an oxygen or fluorine.

$$Zn(s) \rightarrow Zn^{2+} + 2e^-$$

Notice that the zinc metal (the reactant) has an oxidation number of zero (rule 1), and the zinc cation (the product) has an oxidation number of +2 (rule 2). In general, you can say that a substance is oxidized when there's an increase in its oxidation number.

Reduction works the same way. Consider this reaction:

$$Cu^{2+} + 2e^- \rightarrow Cu(s)$$

The copper is going from an oxidation number of +2 to zero. A substance is reduced if there's a decrease in its oxidation number.

Example:

Find the oxidation state of (a) sulphur in H_2SO_4 and (b) nitrogen in $Mg(NO_3)_2$,

Solution:

(a) The oxidation states of H and O are +1 and – 2 respectively. The sum of the oxidation states of the elements in H_2SO_4 = 0.

2 × oxidation state of H + oxidation state of S + 4 × oxidation state of O = 0
(2 × +1) + (oxidation state of S) + (4 × – 2) = 0

The oxidation state of S in H_2SO_4 = +6

(b) The oxidation states of Mg and O are +2 and – 2, respectively. The sum of the oxidation states of the elements in $Mg(NO_3)_2$ = 0

Oxidation state of Mg + 2 × oxidation state of N + 6 × oxidation of O = 0
(+2) + (2 × oxidation of N) + (6 × – 2) = 0

The oxidation state of N in $Mg(NO_3)_2$ = +5

Activity

Find the oxidation state of the particular elements in the following:

1. Cr in CrO_4^{2-}
2. Mn in Mn_2O_7
3. I in $Ca(IO_3)_2$
4. V in VO_2^+
5. N in NH_4^+
6. P in H_3PO_3

Naming Compounds Using Oxidation Numbers

Sodium, chlorine and oxygen form four compounds that are all called sodium chlorate. In each of these compounds the chlorine has a different oxidation number. The oxidation number is written in brackets after the chemical formula to distinguish them from each other:

	$NaClO$	$NaClO_2$	$NaClO_3$	$NaClO_4$
Oxidation state of chlorine	+1	+3	+5	+7
Name	Sodium chlorate (I)	Sodium chlorate (III)	Sodium chlorate (V)	Sodium chlorate (VII)

Oxidizing and Reducing Agents

- An oxidizing agent brings about the oxidation of another substance. In the process it is reduced (it accepts electrons).
- A reducing agent brings about the reduction of another substance. In the process it is oxidized (it donates electrons).

Consider the redox reaction between zinc metal and aqueous copper (II) sulphate. The equation for the reaction is:

$$Zn_{(s)} + CuSO_{4\,(aq)} \rightarrow ZnSO_{4\,(aq)} + Cu_{(s)}$$
$$Zn_{(s)} + Cu^{2+} \rightarrow Zn^{2+} + Cu_{(s)}$$

Ionic half-equations:

$$Zn_{(s)} - 2e^{-} \rightarrow Zn^{2+}$$
$$Cu^{2+} + 2e^{-} \rightarrow Cu_{(s)}$$

Common Oxidizing and Reducing Agents

Oxidizing Agent	**Colour Change Or Other Observable Sign**	**Products**
Nitric acid (conc.)	Brown gas evolved	NO_2(g), water, nitrate(aq)
Hot concentrated sulphuric acid	Gas produced, has characteristic smell	SO_2(g), sulphate (aq), formed with metals or their compounds
Potassium manganate (VII)/dilute sulphuric acid	Colour changes from purple to colourless	Mn^{2+} (aq)
Potassium dichromate (VI)/acid	Colour changes from orange to green	Cr^{3+} (aq)

Iron (III) salts: Fe^{3+} (aq)	Yellow to pale green	Fe^{2+} (aq)
Hydrogen peroxide	Effervescence, colourless oxygen gas evolved	O_2
I_2 (aq)	From brown to colourless	I^- (aq)

Reducing Agent	**Colour Change or Other Observable Sign**	**Products**
Hydrogen sulphide	Yellow colloidal suspension formed	Sulphur and water
Sulphur dioxide	No significant observable change	H_2SO_4 or SO_4^{2-} (aq)
Sulphite	No observable change	H_2SO_4 or SO_4^{2-} (aq)
Conc. HCl	Yellow-green gas	Chlorine
KI/H^+ (aq)	Brown solution or black precipitate of I_2	Iodine, water
Fe^{2+} (aq)	Turns yellow or brown	Fe^{3+}
Hydrogen peroxide	Effervescence, colourless hydrogen gas evolved	H_2 (g)

A few substances can act as both oxidizing and reducing agents. Examples include the following:

- Hydrogen peroxide (H_2O_2) oxidizes iron (II) to iron (III), but reduces acidified potassium manganate (VII) solution
- Sodium nitrate ($NaNO_3$) oxidizes the iodide ion to iodine, but reduces purple potassium manganate (VII) to the colourless manganese (II) ions.
- Sulphur dioxide (SO_2) oxidizes hydrogen sulphide to sulphur, but reduces many other substances.

Tests for Oxidizing and Reducing Agents

In some experiments, it may help to identify a substance by testing to see whether it is an oxidizing or a reducing agent. The following tests are chosen since they are easy to carry out and give an observable change.

Tests for Oxidizing Agents

Reactant	Observable Change	Equation
KI(aq)	Solution changes from colourless to brown (iodine)	$2I^-(aq) - 2e^- \rightarrow I_2(aq)$
Fe^{2+} (aq)	Colour change from pale green to yellow-brown	$Fe^{2+}(aq) - e^- \rightarrow Fe^{3+}(aq)$
Na_2S or H_2S	Colloidal suspension of	$S^{2-}(aq) - 2e^- \rightarrow S$

	sulphur	

Tests for Reducing Agents

A reducing agent should decolourize acidified potassium manganate (VII) ion:

$$MnO_{4\,(aq)}^{-} + 5e^{-} \rightarrow Mn^{2+}_{(aq)}$$

A reducing agent should also change acidified potassium dichromate (VI) (orange) to the green chromium (III) ion:

$$CrO_{7\,(aq)}^{2-} + 3e^{-} \rightarrow 2Cr^{3+}_{(aq)}$$

Solutions should be acidified, usually with dilute sulphuric acid, if oxidizing and reducing agents are to function efficiently.

Reactions Between Oxidizing and Reducing Agents

The reactions between some common oxidizing and reducing agents are shown in the table below. Note the behaviour of H_2O_2.

	SO_2(g)/SO_3^{2-}(aq)	**Conc. HCl**	**KI/H^+(aq)**	**Fe^{2+}(aq)**	**H_2O_2**
Conc. H_2SO_4		Complex reaction; Cl_2 gas evolved; SO_4^{2-} ions converted to SO_2(g), S(s), H_2S(g)	I_2 liberated; SO_2(g) evolved	Fe^{3+}(aq) formed	Effervescence; SO_2(g) evolved
$KMnO_4/H^+$	H_2SO_4 and Mn^{2+} formed	Cl_2 evolved; purple MnO_4^- converted to colourless Mn^{2+}	I_2 liberated; Mn^{2+} (aq) formed	Fe^{3+}(aq) formed	MnO_4^- (aq) decolourized
$K_2Cr_2O_7/H^+$	H_2SO_4 and Cr^{3+}(aq) formed; colour change from orange to	Cl_2 evolved; orange $Cr_2O_7^{2-}$ converted to green Cr^{3+}	I_2 liberated; Cr^{3+}(aq) formed	Fe^{3+}(aq) formed	$Cr_2O_7^{2-}$ converted to Cr^{3+}

	green				
H_2O_2 (aq)	H_2SO_4 and H_2O formed	I_2 liberated		Fe^{3+}(aq) formed	
Fe^{3+} (aq)	Green Fe^{2+} (aq) formed; also H_2SO_4	I_2 liberated; Fe^{2+}(aq) formed			