CHEMISTRY LEVEL 4C

OXIDATION REDUCTION (REDOX)

THEORY SUMMARY

&

REVISION QUESTIONS

INDEX:

CHEMISTRY (CHM 415114)
# OXIDATION ~ REDUCTION CHEMISTRY
INTRODUCTORY THEORY (CRITERION 5)

1. INTRODUCTION:

A significant proportion of the chemical reactions that we have studied so far are ones in which electrons are transferred from one chemical species to another.
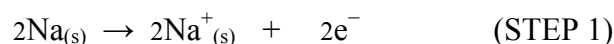
These so called 'electron transfer reactions' are more commonly referred to as "oxidation~reduction" reactions or "**REDOX**" reactions.

e.g. Consider the reaction between sodium metal and chlorine gas:

The equation could be simply written as:

$$2Na_{(s)} \quad + \quad Cl_{2(g)} \quad \rightarrow \quad 2NaCl_{(s)}$$
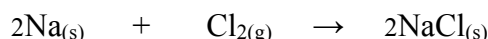
A closer inspection of the chemistry involved, reveals that in fact this reaction process involves two separate steps, one in which sodium atoms lose electrons and the other where chlorine atoms gain those electrons; i.e.

$$2Na_{(s)} \rightarrow 2Na^{+}_{(s)} \; + \; 2e^{-} \qquad \text{(STEP 1)}$$
$$Cl_{2(g)} \; + \; 2e^{-} \rightarrow \quad 2Cl^{-}_{(s)} \qquad \text{(STEP 2)}$$

These two 'sub-reactions' or more correctly **HALF-EQUATIONS,** show one process in which electrons are lost (STEP 1) and the other where electrons are gained (STEP 2)

Sodium atoms change their electron arrangement from $Na = 2)8)1$ to $Na^{+} = 2)8$
Chlorine atoms change their electron arrangement from $Cl = 2)8)7$ to $Cl^{-} = 2)8)8$

A vast number of chemical reactions are ones that involve this loss and gain of electrons (redox) and thus this topic deservedly becomes a major study unit for our chemistry course. Examples of oxidation/reduction reactions include:

- Combustion of fuels
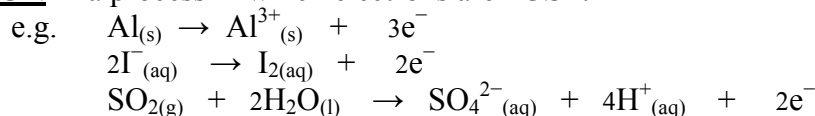- Reactions of metals with oxygen, acids, ...
- Photographic reactions
- Biochemical reactions
- Manufacture of plastics, dyes, paints, ...
- Electrochemical processes that occur in batteries
- Electrolytic processes in industry

Any reaction where an element changes its VALENCY is an oxidation/reduction reaction or conveniently abbreviated to a:
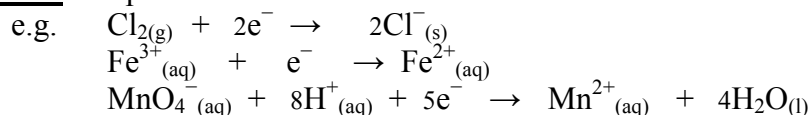
"REDOX" REACTION

2. DEFINITIONS:

(i) **OXIDATION** = a process in which electrons are **_LOST._**

e.g.
$$Al_{(s)} \rightarrow Al^{3+}_{(s)} + 3e^-$$
$$2I^-_{(aq)} \rightarrow I_{2(aq)} + 2e^-$$
$$SO_{2(g)} + 2H_2O_{(l)} \rightarrow SO_4^{2-}_{(aq)} + 4H^+_{(aq)} + 2e^-$$

(oxidation half-equations always have electrons appearing on the 'product' or right side of the arrow.)

(ii) **REDUCTION** = a process in which electrons are **_GAINED._**

e.g.
$$Cl_{2(g)} + 2e^- \rightarrow 2Cl^-_{(s)}$$
$$Fe^{3+}_{(aq)} + e^- \rightarrow Fe^{2+}_{(aq)}$$
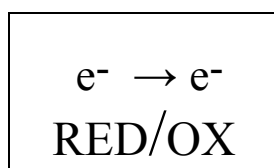$$MnO_4^-_{(aq)} + 8H^+_{(aq)} + 5e^- \rightarrow Mn^{2+}_{(aq)} + 4H_2O_{(l)}$$

(reduction half-equations always have electrons appearing on the 'reactant' or left side of the arrow.)

(iii) **MEMORY AID**

The two terms of oxidation and reduction are frequently condensed into an abbreviated form expressed as REDOX.
Take the word "redox" and split it into its two parts; i.e.

$$\boxed{\begin{array}{c} e^- \rightarrow e^- \\ RED/OX \end{array}}$$

If the electrons are on the left-side of the arrow it's a REDUCTION half-equation.
If the electrons are on the right-side of the arrow it's an OXIDATION half-equation.

(iv) Oxidation and reduction **_must always occur together_** with one chemical species losing electrons and another gaining those same electrons.

(v) In all redox systems:

> **THE NUMBER OF ELECTRONS LOST = THE NUMBER OF ELECTRONS GAINED**

(vi)   The fact that the number of electrons lost is the same as the number of electrons gained will turn out to provide us with a very useful and comparatively easy way to balance redox chemical equations!

(vii) Before we get heavily into redox chemistry, it will pay us to review some chemical ideas we met in year 11 concerning atomic structure and the electron/proton make-up of various atoms and ions.

(viii) The word 'oxidised' seems to suggest a chemical change involving oxygen. This was an early usage of the term but now we use the term 'oxidation' in a much more general sense where there doesn't need to be oxygen involved at all.

3. PROTON ~ ELECTRON NUMBERS:

For each of the following chemical species, give the name, total number of protons and electrons present. For ionic species, indicate whether the ion is a cation or an anion.
The first example has been done for you.

SPECIES	NAME	PROTONS	ELECTRONS
Al^{3+}	aluminium 3+ cation	13	10
Na			
Ca^{2+}			
S			
Bi			
Sn			
Br			
Xe			
Mg^{2+}			
Cu^{2+}			
O^{2-}			
P^{3-}			
SO_3			
HCl			
O_2			
$PO_4{}^{3-}$			
$S_2O_3{}^{2-}$			
PCl_3			
H_2SO_4			

4. INTRODUCTION TO HALF-EQUATIONS:

Complete the following half-equations by indicating the number of electrons transferred and whether the electrons are gained or lost. State whether the process is oxidation or reduction.

The first two examples have been completed for you.

Ca	$\rightarrow$	Ca^{2+}	+	$2e^-$	**(OXIDATION)**
P + $3e^-$	$\rightarrow$	P^{3-}			**(REDUCTION)**
Al^{3+}	$\rightarrow$	Al			
S	$\rightarrow$	S^{2-}			
Cu	$\rightarrow$	Cu^{2+}			
Br	$\rightarrow$	Br^-			
H	$\rightarrow$	H^+			
H	$\rightarrow$	H^-			
As	$\rightarrow$	As^{3-}			
Fe^{2+}	$\rightarrow$	Fe^{3+}			
O_2	$\rightarrow$	O^{2-}			
Cr^{3+}	$\rightarrow$	Cr^{2+}			
N^{3-}	$\rightarrow$	N_2			
Mg^{2+}	$\rightarrow$	Mg			
Au^{3+}	$\rightarrow$	Au			
Cl_2	$\rightarrow$	Cl^-			
S^{2-}	$\rightarrow$	S_8			
P_4	$\rightarrow$	P^{3-}			
O^{2-}	$\rightarrow$	O_3			

5. BALANCING HALF-EQUATIONS:

* Balance the following half-equations by using the steps:
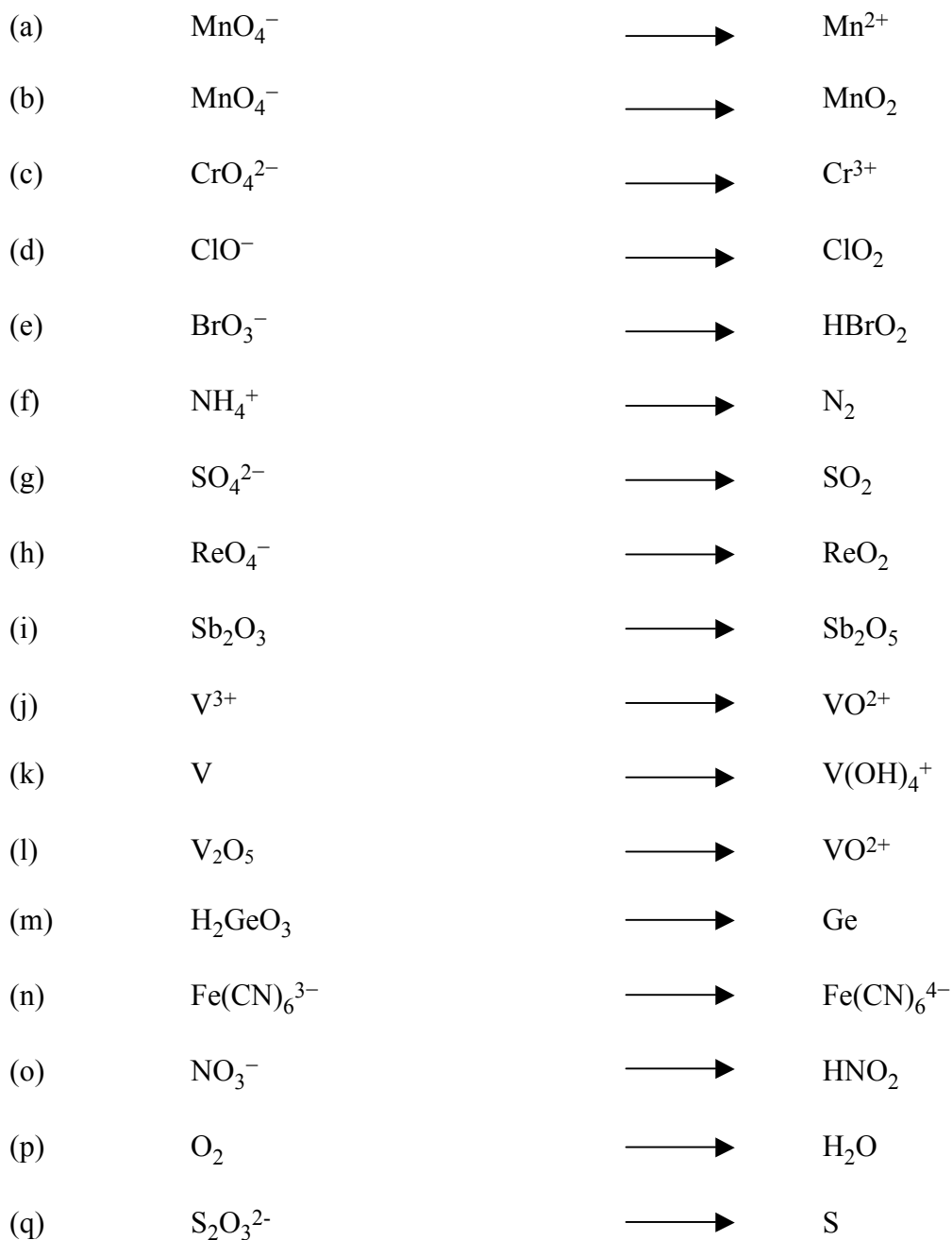   (i) balance the element other than oxygen or hydrogen.

(ii) balance oxygen (if necessary) by using water H_2O.
(iii) balance hydrogen (if necessary) by using H^+ ions.
(iv) balance electrical charge by using electrons e^-

(assume all ions are in aqueous solution)

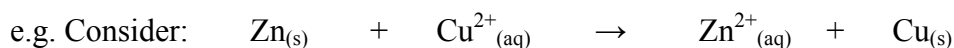* Indicate whether the half-equation is either **OXIDATION** or **REDUCTION**.

(a) MnO_4^- $\longrightarrow$ Mn^{2+}

(b) MnO_4^- $\longrightarrow$ MnO_2

(c) CrO_4^{2-} $\longrightarrow$ Cr^{3+}

(d) ClO^- $\longrightarrow$ ClO_2

(e) BrO_3^- $\longrightarrow$ $HBrO_2$

(f) NH_4^+ $\longrightarrow$ N_2

(g) SO_4^{2-} $\longrightarrow$ SO_2

(h) ReO_4^- $\longrightarrow$ ReO_2

(i) Sb_2O_3 $\longrightarrow$ Sb_2O_5

(j) V^{3+} $\longrightarrow$ VO^{2+}

(k) V $\longrightarrow$ $V(OH)_4^+$

(l) V_2O_5 $\longrightarrow$ VO^{2+}

(m) H_2GeO_3 $\longrightarrow$ Ge

(n) $Fe(CN)_6^{3-}$ $\longrightarrow$ $Fe(CN)_6^{4-}$

(o) NO_3^- $\longrightarrow$ HNO_2

(p) O_2 $\longrightarrow$ H_2O

(q) $S_2O_3^{2-}$ $\longrightarrow$ S

6. OXIDISING AGENTS & REDUCING AGENTS:

In order that a chemical species be able to undergo oxidation, (i.e. lose electrons) there has to be another chemical species that will accept these electrons. The species that accepts the electrons is called an OXIDISING AGENT (or an oxidant or oxidiser)
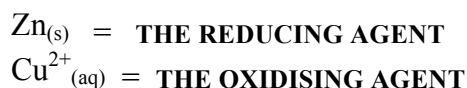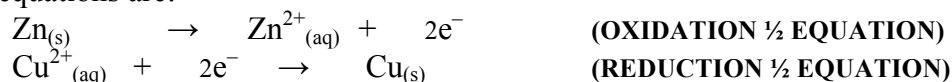
In order that a chemical species be able to undergo reduction, (i.e. gain electrons) there has to be another chemical species that will donate these electrons. The species that donates the electrons is called a REDUCING AGENT (or a reductant or reducer)

i.e.	**AN OXIDISING AGENT**	=	**AN ELECTRON ACCEPTOR**
	A REDUCING AGENT	=	**AN ELECTRON DONOR**

Although initially it may sound a little confusing, the oxidising agent will always be the chemical species undergoing the reduction process and the reducing agent will always be the chemical species undergoing the oxidation process.

e.g. Consider: $\qquad Zn_{(s)} \quad + \quad Cu^{2+}_{(aq)} \quad \rightarrow \quad Zn^{2+}_{(aq)} \quad + \quad Cu_{(s)}$

The two half-equations are:

$\qquad Zn_{(s)} \quad \rightarrow \quad Zn^{2+}_{(aq)} \quad + \quad 2e^{-} \qquad$ **(OXIDATION ½ EQUATION)**
$\qquad Cu^{2+}_{(aq)} \quad + \quad 2e^{-} \quad \rightarrow \quad Cu_{(s)} \qquad$ **(REDUCTION ½ EQUATION)**

$\qquad Zn_{(s)} =$ **THE REDUCING AGENT**
$\qquad Cu^{2+}_{(aq)} =$ **THE OXIDISING AGENT**

(i) The zinc metal undergoes oxidation and in so doing, loses two electrons.
(ii) These two electrons that are lost by Zn metal are now available to bring about the reduction of the Cu^{2+} ions.
(iii) As the Zn metal provides the electrons that cause the ***reduction*** of the Cu^{2+} ions, we say that in this reaction, Zn is the ***reducing*** agent.
(iv) Similarly, the Cu^{2+} ions take electrons away from the Zn metal, thus causing the Zn metal to be oxidised. Consequently, we describe the Cu^{2+} ions in this case as being ***oxidising*** agents.

SUMMARY:

(i) THE OXIDISING AGENT IS THE SPECIES UNDERGOING REDUCTION

(ii) THE REDUCING AGENT IS THE SPECIES UNDERGOING OXIDATION

Q. In the net chemical reaction:

$\qquad Al_{(s)} \quad + \quad 3Ag^{+}_{(aq)} \quad \rightarrow \quad Al^{3+}_{(aq)} \quad + \quad 3Ag_{(s)}$
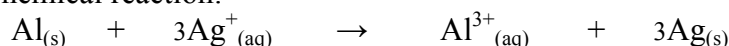
(i) Which chemical species is the oxidising agent? $\qquad\qquad$ (silver ions; Ag^{+})
(ii) Which chemical species is the reducing agent? $\qquad\qquad$ (aluminium metal Al)

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7. BALANCING REDOX REACTIONS: (THE HALF-EQUATION METHOD)

The process we use to balance redox equations is quite straight forward and relies upon the fact that the electrons lost in the oxidation half-equation must be identical in number to the electrons gained in the reduction half-equation.
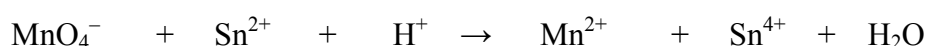*The half equations are balanced in the procedure described earlier at the top of page 7.
* Spectator ions are deliberately omitted to make the process a little easier to follow.

EXAMPLE:
(i) Balance the following net ionic equation involving aqueous ionic species.
(ii) Identify both the oxidising agent and the reducing agent

$$MnO_4^- \quad + \quad Sn^{2+} \quad + \quad H^+ \quad \rightarrow \quad Mn^{2+} \quad + \quad Sn^{4+} \quad + \quad H_2O$$

STEP 1. Identify the redox *conjugate pairs* on either side of the equation as these will form the bases for constructing your two half-equations.
In this case, the two related conjugate pairs are: $\{MnO_4^- \ \& \ Mn^{2+}\}$ and $\{Sn^{2+} \ \& \ Sn^{4+}\}$

Note that the H^+ ions and H_2O are not a redox conjugate pair because in water, the hydrogen is still in the H^+ oxidation state. (see later) .

STEP 2. Construct half-equations based upon the redox pairs you've identified.
This is done by: (i) balancing the element other than oxygen or hydrogen.
                 (ii) balancing oxygen (if necessary) by using water H_2O.
                 (iii) balancing hydrogen (if necessary) by using H^+ ions.
                 (iv) balancing electrical charge by using electrons (e^-)
Thus, one of our redox pairs leads to:
$$MnO_4^- \quad + \quad 8H^+ \quad + \quad 5e^- \quad \rightarrow \quad Mn^{2+} \quad + \quad 4H_2O \quad \text{(a reduction ½ equation)}$$
The other redox pair gives: $Sn^{2+} \quad \rightarrow \quad Sn^{4+} \quad + \quad 2e^- \quad$ (an oxidation ½ equation)
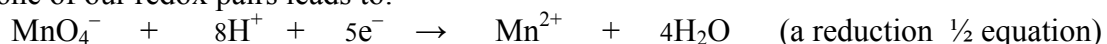
STEP 3. Now adjust the two half-equations so as to have the *same* number of electrons lost as are gained. i.e. multipy right through the first half-equation by 2 and multipy right through the second half-equation by 5, thus giving both equations $10e^-$.

i.e. $2MnO_4^- \quad + \quad 16H^+ \quad + \quad 10e^- \quad \rightarrow \quad 2Mn^{2+} \quad + \quad 8H_2O$
                           $5Sn^{2+} \quad \rightarrow \quad 5Sn^{4+} \quad + \quad 10e^-$
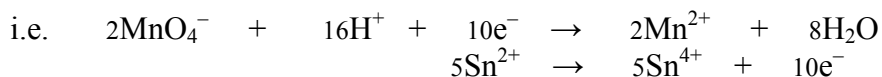
STEP 4. Add together the two half-equations and if you've done it correctly, the electrons won't appear in the final balanced net equation; i.e.
i.e. $2MnO_4^- \quad + \quad 16H^+ \quad + \quad 5Sn^{2+} \quad \rightarrow \quad 2Mn^{2+} \quad + \quad 5Sn^{4+} \quad + \quad 8H_2O$
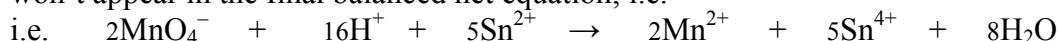
NOTE: The oxidising agent was acidified MnO_4^- and the reducing agent was Sn^{2+}.

STEP 5. Check your answer. A quick check is to see that the total electrical charge on the reactant side is the same as on the product side. In this case: $24+ = 24+$ (Checked OK!)

BALANCE THE FOLLOWING NET REDOX EQUATIONS:

(a) $Al + Sn^{2+} \rightarrow Al^{3+} + Sn$

(b) $MnO_4^- + Cl^- + H^+ \rightarrow Mn^{2+} + Cl_2 + H_2O$

(c) $HNO_3 + SO_2 \rightarrow SO_4^{2-} + H^+ + NO_2$

(d) $Cr_2O_7^{2-} + Fe^{2+} + H^+ \rightarrow Cr^{3+} + Fe^{3+} + H_2O$

(e) $H_2O_2 + N_2O \rightarrow NO_2 + H_2O$

(f) $Br_2 + Cr^{3+} + H_2O \rightarrow Br^- + CrO_4^{2-} + H^+$

(g) $SO_2 + Au^+ + H_2O \rightarrow SO_3 + Au + H^+$

(h) $GeO_2 + Cl_2 + H_2O \rightarrow GeO_4^{2-} + Cl^- + H^+$

(i) $Fe^{3+} + Fe \rightarrow Fe^{2+}$

(j) $MnO_4^{2-} + Sn(OH)_4^{2-} + H^+ \rightarrow MnO_2 + Sn^{4+} + H_2O$

(k) $S_2O_3^{2-} + NH_4^+ + H^+ \rightarrow N_2 + S + H_2O$

(l) $NO_3^- + V + H^+ \rightarrow VO_2^+ + HNO_2 + H_2O$

(m) $V_2O_5 + ReO_2 + H^+ \rightarrow VO^{2+} + ReO_4^- + H_2O$

8. AQUEOUS SPECTATOR IONS:

The lists below show ions that are *spectator ions* and as such will never appear in a net ionic redox equation!

THESE ANIONS CAN'T BE OXIDISED IN AQUEOUS SOLUTION

nitrate ions	$NO_3^-{}_{(aq)}$	sulfate ions	$SO_4^{2-}{}_{(aq)}$
carbonate ions	$CO_3^{2-}{}_{(aq)}$	phosphate ions	$PO_4^{3-}{}_{(aq)}$
hydrogensulfate ions	$HSO_4^-{}_{(aq)}$		

THESE CATIONS CAN'T BE REDUCED IN AQUEOUS SOLUTION

|			
potassium ions	$K^+{}_{(aq)}$	sodium ions	$Na^+{}_{(aq)}$
magnesium ions	$Mg^{2+}{}_{(aq)}$	calcium ions	$Ca^{2+}{}_{(aq)}$
barium ions	$Ba^{2+}{}_{(aq)}$	aluminium ions	$Al^{3+}{}_{(aq)}$
lithium ions	$Li^+{}_{(aq)}$	strontium ions	$Sr^{2+}{}_{(aq)}$

9. OXIDATION NUMBERS:

These are numbers that apply to an ***element*** and relate to the electrons involved in bonding whether the bonding is ionic or covalent.

The concept of oxidation number can serve as a very useful electron "accounting" device and has applications in balancing redox equations.

To determine the value of an element's oxidation number, there are a few simple rules you need to apply. These are:
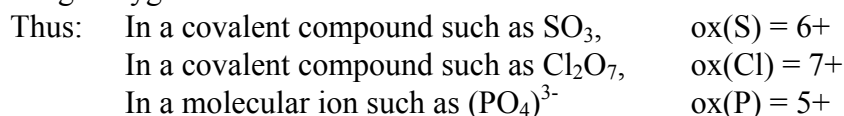
RULE 1. Any uncombined neutral atom has an oxidation number of zero.

e.g. $Al_{(s)}$ ox(Al) = 0

           $Cl_{2(s)}$ ox(Cl) = 0
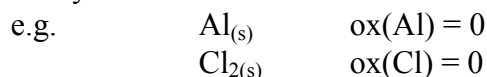
RULE 2. For metallic ions and non-metal elemental ions, the oxidation number is the valency(ionic charge)

e.g. Cu^{2+} ox(Cu) = 2+

           S^{2-} ox(S) = 2–
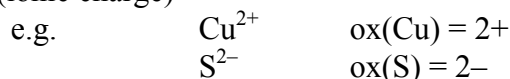
RULE 3. The sum of the oxidation numbers in a neutral compound (e.g. H_2SO_4) is always equal to zero and for polyatomic ions, (e.g. SO_4^{2-}) the sum of the oxidation numbers is the overall charge on the ion. This rule is now applied in rules 4 and 5.

RULE 4. In all compounds or molecular ions involving oxygen combined with other elements, assign oxygen an oxidation number of 2–.

Thus: In a covalent compound such as SO_3,   ox(S) = 6+

        In a covalent compound such as Cl_2O_7,   ox(Cl) = 7+

        In a molecular ion such as $(PO_4)^{3-}$ ox(P) = 5+

RULE 5. In all compounds or molecular ions involving hydrogen combined with other elements, assign hydrogen an oxidation number of 1+.

Thus: In a covalent compound such as NH_3,   ox(N) = 3–

        In a covalent compound such as H_2SO_3,   ox(S) = 4+

        In a molecular ion such as $(HPO_2)^{2-}$ ox(P) = 1+
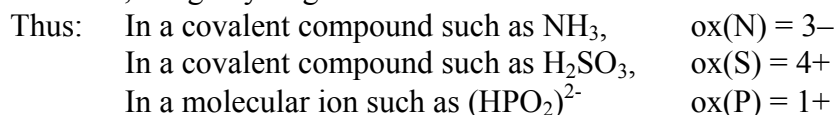
Examples: Determine the oxidation number (sometimes called the oxidation *state*) of the elements shown in bold below;

(i) $Na_2\mathbf{C}O_3$	(ii) $Ca\mathbf{Cr}O_4$	(iii) $H\mathbf{Mn}O_4$
(iv) $\mathbf{N}_2O_5$	(v) $\mathbf{As}O_3^{3-}$	(vi) $\mathbf{S}_8$
(vii) $\mathbf{Cl}O_3^{-}$	(viii) $K\mathbf{Br}O_4$	(ix) $\mathbf{V}O_2^{+}$
(x) $\mathbf{Cl}O^{-}$	(xi) $K\mathbf{Br}O_2$	(xii) $\mathbf{V}O^{2+}$
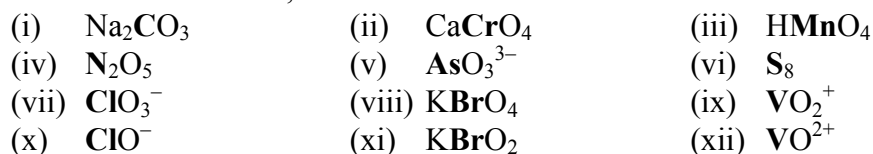
* EXCEPTIONS TO THE RULES.

     Peroxides such as H_2O_2 and Na_2O_2 are exceptions. In peroxides, ox(O) = 1–

     Metal hydrides such as LiH and NaH are exceptions. In metal hydrides, ox(H) = 1–

10. USING OXIDATION NUMBERS:

(i) BALANCING EQUATIONS
Rather than using the "half-equation" method for balancing redox equations, it is possible to use the oxidation numbers of the elements undergoing loss and gain of electrons. Students may prefer to use the ½-equation method but this alternate procedure is shown below for comparison.
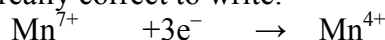e.g. Balance

$$MnO_4^- + SO_2 + H_2O \rightarrow MnO_2 + SO_4^{2-} + H^+$$

In the permanganage ion MnO_4^-, the ox(Mn) = +7 and in this reaction, the manganese is reduced to MnO_2 where the ox(Mn) = +4.
i.e. This change can be represented by :

$$\{ ox(Mn) = +7 \} \quad +3e^- \quad \rightarrow \quad \{ ox(Mn) = +4 \}$$
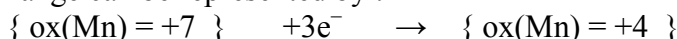
Note that this is not really correct to write:

$$Mn^{7+} \quad +3e^- \quad \rightarrow \quad Mn^{4+}$$

because this infers that there are manganese 7+ ions in MnO_4^- which is not the case.
In sulfur dioxide SO_2 , the ox(S) = +4 and in this reaction, the sulfur is oxidised to SO_4^{2-} where the ox(S) = +6.
i.e. This change can be represented by :

$$\{ ox(S) = +4 \} \quad \rightarrow \quad \{ ox(S) = +6 \} \quad + \quad 2e^-$$

Combining the two we get:

$\{ ox(Mn) = +7 \} + \mathbf{3e^-} \rightarrow \{ ox(Mn) = +4 \}$	**REDUCTION**

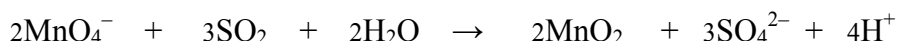|	
$\{ ox(S) = +4 \} \rightarrow \{ ox(S) = +6 \} + \mathbf{2e^-}$	**OXIDATION**
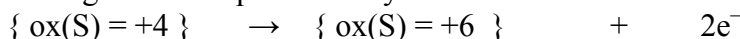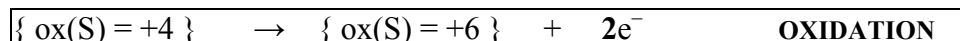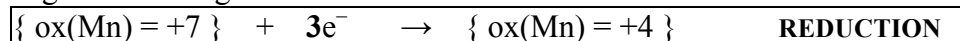
In order to have the same number of elecrons lost as are gained, we need to multiply the first by 3 and the second by 2 giving $6e^-$ in both changes.
This means that our balanced equation must have manganese and sulfur present in the ratio

$$Mn : S = 2 : 3$$
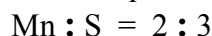
So the balanced net equation is:

$$2MnO_4^- + 3SO_2 + 2H_2O \rightarrow 2MnO_2 + 3SO_4^{2-} + 4H^+$$

The water and hydrogen ions are adjusted at the end to achieve the balancing.

(ii) COMPARING OXIDISING & REDUCING STRENGTH
Where the *SAME* element exists in a number of different oxidation states (i.e. it has several different oxidation numbers) then usually,

> **THE HIGHEST OXIDATION STATE IDENTIFIES THE MOST POWERFUL OXIDISER**
> **THE LOWEST OXIDATION STATE IDENTIFIES THE MOST POWERFUL REDUCER**

Q1. Which is the most powerful oxiliser amongst Fe, Fe^{2+} and Fe^{3+}?   (Fe^{3+})
Q2. Which is the most powerful reducer amongst Fe, Fe^{2+} and Fe^{3+}?   (Fe)
Q3. By considering he oxidation state of chlorine, which is the most powerful oxidising agent amongst HClO, $HClO_2$, $HClO_3$ and $HClO_4$?   ($HClO_4$)
Q4. By considering he oxidation state of sulfur, which is the most powerful reducing agent amongst H_2S, H_2SO_3, S_8, $Na_2S_2O_3$, H_2SO_4 and $H_2S_2O_7$?   (H_2S)

11. THE ELECTROCHEMICAL SERIES: "E.C.S.":

(a) The tendency for different species to lose or gain electrons varies significantly. For example, fluorine molecules (F_2) have a very high tendency to gain electrons and thus fluorine is described as a very powerful oxidiser (oxidising agent).

Consider: $$F_2 \quad + \quad 2e^- \quad \rightleftharpoons \quad 2F^-$$

(i) The forward reaction, as described above has a very high tendency to occur making F_2 a very powerful electron acceptor; i.e. F_2 is a very powerful oxidising agent.
(ii) Conversely, the reverse reaction (as shown by the 'back' arrow) will have an extremely low tendency to occur meaning that fluoride ions (F^-) are very weak electron donors. This means that fluoride ions are very weak reducing agents.

F_2/F^- is called a REDOX CONJUGATE PAIR.

(b) Sodium ions (Na^+) have an extremely low tendency to gain electrons and are thus, very weak oxidisers.

Consider: $$Na^+ \quad + \quad e^- \quad \rightleftharpoons \quad Na$$

(i) The forward reaction, as described above has a very low tendency to occur making Na^+ a very weak electron acceptor; i.e. Na^+ is a very weak oxidising agent.
(ii) Conversely, the reverse reaction (as shown by the 'back' arrow) will have an extremely high tendency to occur meaning that sodium atoms (Na) are very strong electron donors. This means that sodium atoms are very strong reducing agents.

Na^+/Na is another REDOX CONJUGATE PAIR.

(c) Lead ions (Pb^{2+}) are somewhat intermediate in strength as oxidisers (i.e. they would be described as neither strong nor weak).

Consider: $$Pb^{2+} \quad + \quad 2e^- \quad \rightleftharpoons \quad Pb$$

(i) The forward reaction, as described above has a moderate tendency to occur making Pb^{2+} ions a middle strength oxidising agent.
(ii) Similarly, the reverse reaction (as shown by the 'back' arrow) has a moderate tendency to occur making Pb metal a middle strength reducing agent.

Pb^{2+}/Pb is another REDOX CONJUGATE PAIR.

(d) All the commonly encountered redox conjugate pairs can be placed in order with the oxidising agents listed with decreasing tendency to gain electrons. This gives rise to a very useful table that we call **THE ELECTROCHEMICAL SERIES.** The tendency for each oxidiser to gain electrons and undergo reduction is measured as an electric potential or voltage.

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12. STANDARD REDUCTION POTENTIALS:

AQUEOUS IONS (1.00 mol L^{-1} and 25°C)

REDUCTION HALF EQUATIONS

VERY STRONG OXIDISERS	Reduction Half Equation						VERY WEAK REDUCERS	E°(volts)
	$F_2(g)$	+	$2e^-$	⇌	$2F^-(aq)$			+2.87
	$H_2O_2(aq) + 2H^+(aq)$	+	$2e^-$	⇌	$2H_2O(l)$			+1.78
	$Au^{3+}(aq)$	+	$3e^-$	⇌	$Au(s)$			+1.50
	$MnO_4^-(aq) + 8H^+(aq)$	+	$5e^-$	⇌	$Mn^{2+}(aq) + 4H_2O(l)$			+1.49
	$Cl_2(g)$	+	$2e^-$	⇌	$2Cl^-(aq)$			+1.36
	$Cr_2O_7^{2-}(aq) + 14H^+(aq)$	+	$6e^-$	⇌	$2Cr^{3+}(aq) + 7H_2O(l)$			+1.36
	$MnO_2(s) + 4H^+(aq)$	+	$2e^-$	⇌	$Mn^{2+}(aq) + 2H_2O(l)$			+1.28
	$O_2(g) + 4H^+(aq)$	+	$4e^-$	⇌	$2H_2O(l)$			+1.23
	$Br_2(l)$	+	$2e^-$	⇌	$2Br^-(aq)$			+1.09
	$NO_3^-(aq) + 4H^+(aq)$	+	$3e^-$	⇌	$NO(g) + 2H_2O(l)$			+0.96
	$Ag^+(aq)$	+	e^-	⇌	$Ag(s)$			+0.80
	$Hg^{2+}(aq)$	+	$2e^-$	⇌	$Hg(l)$			+0.78
	$NO_3^-(aq) + 2H^+(aq)$	+	e^-	⇌	$NO_2(g) + H_2O(l)$			+0.78
	$Fe^{3+}(aq)$	+	e^-	⇌	$Fe^{2+}(aq)$			+0.77
	$O_2(g) + 2H^+(aq)$	+	$2e^-$	⇌	$H_2O_2(aq)$			+0.68
	$I_2(s)$	+	$2e^-$	⇌	$2I^-(aq)$		INCREASING STRENGTH AS REDUCERS	+0.54
	$O_2(g) + 2H_2O(l)$	+	$4e$	⇌	$4OH^-(aq)$			+0.40
	$Cu^{2+}(aq)$	+	$2e^-$	⇌	$Cu(s)$			+0.34
	$SO_4^{2-}(aq) + 4H^+(aq)$	+	$2e^-$	⇌	$SO_2(g) + 2H_2O(l)$			+0.17
INCREASING STRENGTH AS OXIDISERS	$Sn^{4+}(aq)$	+	$2e^-$	⇌	$Sn^{2+}(aq)$			+0.15
	$S(s) + 2H^+(aq)$	+	$2e^-$	⇌	$H_2S(g)$			+0.14
	$2H^+(aq)$	+	$2e^-$	⇌	$H_2(g)$			0.00
	$Pb^{2+}(aq)$	+	$2e^-$	⇌	$Pb(s)$			−0.13
	$Sn^{2+}(aq)$	+	$2e^-$	⇌	$Sn(s)$			−0.14
	$Ni^{2+}(aq)$	+	$2e^-$	⇌	$Ni(s)$			−0.25
	$Co^{2+}(aq)$	+	$2e^-$	⇌	$Co(s)$			−0.28
	$Cd^{2+}(aq)$	+	$2e^-$	⇌	$Cd(s)$			−0.40
	$Fe^{2+}(aq)$	+	$2e^-$	⇌	$Fe(s)$			−0.41
	$Cr^{3+}(aq)$	+	$3e^-$	⇌	$Cr(s)$			−0.74
	$Zn^{2+}(aq)$	+	$2e^-$	⇌	$Zn(s)$			−0.76
	$2H_2O(l)$	+	$2e^-$	⇌	$2OH^-(aq) + H_2(g)$			−0.83
	$Mn^{2+}(aq)$	+	$2e^-$	⇌	$Mn(s)$			−1.18
	$Al^{3+}(aq)$	+	$3e^-$	⇌	$Al(s)$			−1.71
	$Mg^{2+}(aq)$	+	$2e^-$	⇌	$Mg(s)$			−2.38
	$Na^+(aq)$	+	e^-	⇌	$Na(s)$			−2.71
	$Ca^{2+}(aq)$	+	$2e^-$	⇌	$Ca(s)$			−2.87
VERY WEAK OXIDISERS	$Sr^{2+}(aq)$	+	$2e^-$	⇌	$Sr(s)$		VERY STRONG REDUCERS	−2.89
	$K^+(aq)$	+	e^-	⇌	$K(s)$			−2.92
	$Li^+(aq)$	+	e^-	⇌	$Li(s)$			−3.05

OXIDATION HALF EQUATIONS

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13. STANDARD REDUCTION POTENTIALS (E^o):
**(EXPLANATORY NOTES)**

1. The half-equations are all written with the forward arrow showing a reduction process and the reverse arrow showing an oxidation process.

2. All the chemical species down the left-side of the page are oxidising agents with the order such that the strongest oxidisers are at the top (F_2, acidified H_2O_2,….) and the weakest oxidisers are at the bottom (Li^+, K^+,…)

3. All the chemical species down the right-side of the page are reducing agents with the order such that the weakest reducers are at the top (F^-, H_2O,….) and the strongest reducers are at the bottom (Li, K …)

4. The E^o value indicates the tendency for the *forward* reaction to occur measured as an electric potential (in volts) where the reduction of hydrogen ions to hydrogen gas is assigned an arbitrary zero value; i.e.

$$2H^+ \; + \; 2e^- \; \rightleftharpoons \; H_{2(g)} \qquad E^o = 0.00 \text{ volts}$$
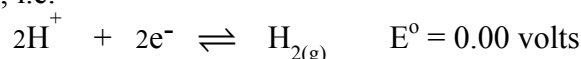
The arbitrary zero value doesn't mean that hydrogen ions have no tendency to gain electrons; it simply serves as a comparison point in just the same way that altitudes and sea depths are measured as heights above or below sea level. This doesn't mean sea level doesn't have any height but rather, it is used as a comparison point.

i.e. oxidisers with a stronger attraction for e^- than hydrogen ions will have positive E^o values whereas oxidisers with a weaker attraction for e^- than hydrogen ions will have negative E^o values.

5. Logically, if the forward half-equation has an $E^o = + x$ volts, then the reverse half-equation will have an $E^o = - x$ volts.

e.g..
$$Ag^+ \; + \; e^- \; \rightleftharpoons \; Ag \qquad\qquad E^o = +0.80 \text{ volts}$$
$$Ag \; \rightleftharpoons \; Ag^+ \; + \; e^- \qquad\qquad E^o = -0.80 \text{ volts}$$
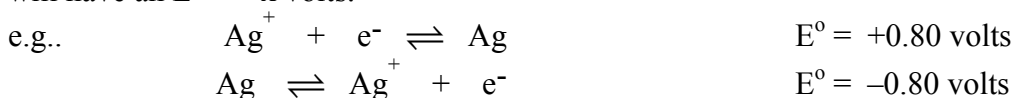
6. The E^o values are determined at 'standard state' conditions of 1.00 mol L^{-1} ionic concentrations, 1.00 atm pressure and 25^oC. For different conditions, the tendency will not be the same.

7. Changing the amount of reactant does **NOT** alter the tendency to lose or gain electrons;

i.e.
$$Cu^{2+} \; + \; 2e^- \; \rightleftharpoons \; Cu \qquad E^o = +0.34 \text{ volts}$$
$$2Cu^{2+} \; + \; 4e^- \; \rightleftharpoons \; 2Cu \qquad E^o = +0.34 \text{ volts} \quad (\text{note } E^o \text{ is } \textbf{\textit{not}} \text{ doubled})$$
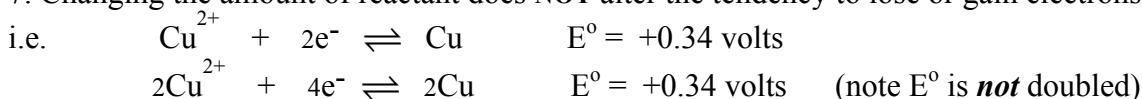
8. Any redox reaction has to have both an oxidising agent and a reducing agent. As we've seen already, this means that all redox systems must comprise (at least) a reduction half-equation and an oxidation half-equation. This translates to meaning that a net redox equation will comprise one left → right (reduction) half-equation and one right → left (oxidation) half-equation from the ECS.

9. The way the Electrochemical Series (ECS) is constructed (see page 14) enables us to make generally reliable predictions about which redox reacions will occur spontaneously (reactions that take place without an energy input) and those that will be non-spontaneous (need an energy input to occur).

e.g. Consider the three half-equations in the order as they appear in the E.C.S.

$Hg^{2+} + 2e^- \rightleftharpoons Hg$	$E^o = +0.78$ volts
$Cu^{2+} + 2e^- \rightleftharpoons Cu$	$E^o = +0.34$ volts
$Zn^{2+} + 2e^- \rightleftharpoons Zn$	$E^o = -0.76$ volts

The strongest of the three oxidisers listed is Hg^{2+} and the strongest of the three reducers listed is Zn.

(i) Hg^{2+} as an oxidiser will react with any reducing agent that is stronger than Hg i.e. any reducers that are *below* Hg on the ECS.

i.e. Hg^{2+} will react spontaneously with Cu and Zn.

(ii) Zn as a reducing agent will react with any oxidising agent that is stronger than Zn^{2+} i.e. any oxidisers that are *above* Zn^{2+} on the ECS.

i.e. Zn will react spontaneously with Cu^{2+} and Hg^{2+}.

10. **USEFUL GENERAL RULE:**

An oxidising agent (from the left side of the ECS) will react *spontaneously* with any reducing agent (from the right side of the ECS) providing that the oxidiser and reducer are positioned so that the connector is sloped as shown:
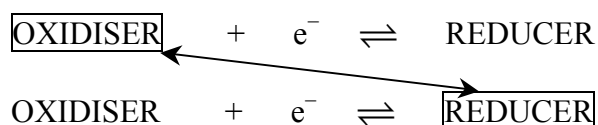
This means that the overall redox reaction that occurs spontaneously will be made up from the combination of two half-equations with the:

REDUCTION ½ EQUATION BEING THE LEFT → RIGHT ½ EQUATION NEAREST THE TOP OF THE ECS
AND THE
OXIDATION ½ EQUATION BEING THE RIGHT → LEFT ½ EQUATION NEAREST THE BOTTOM OF THE ECS

11. The reliability of predictions of spontaneity are dependent upon the conditions being 'standard state'. The fact that a reaction occurs spontaneously does not imply that the rate will be fast.

e.g. we know that iron (Fe) will be oxidised ("rust") spontaneously in an O_2/H_2O environment and yet the reaction is so slow that observing Fe + O_2/H_2O for an hour or so might, reveal no visible chemical change whatsoever.

12. Generally the more powerful the oxidiser and the more powerful the reducer (i.e. the further apart they are on the ECS) the more likely the spontaneous reaction will occur rapidly.

14. QUESTIONS BASED ON USE OF THE E.C.S:

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Q1. Identify an oxidising agent that is stronger than Cd^{2+} ions but weaker than Pb^{2+} ions.

$(Co^{2+}, Ni^{2+}, Sn^{2+})$

Q2. Use the ECS to arrange the halogen elements in order of *decreasing* strength as oxidising agents. (F_2, Cl_2, Br_2, I_2)

Q3. Identify a reducing agent that is stronger than $SO_{2(aq)}$ but weaker than $H_2S_{(g)}$.

$(Cu^+$ or $Sn^{2+})$

Q4. Complete the following statement correctly:

"Given the ½-equation: $Au^{3+} + 3e^- \rightleftharpoons Au$ and the fact that Au^{3+} is a very powerful oxidiser, it means that Au must be a ...."

(very weak reducer)

Q5. Careful inspection of the ECS on page 14 reveals that some chemical species such as Fe^{2+}, Sn^{2+} and H_2O_2 can sometimes act as oxidants and at other times act as reductants.

    (i) When Fe^{2+} acts as an oxidant (oxidiser), what is the reduction product?   (Fe)

    (ii) When Fe^{2+} acts as a reductant (reducer), what is the oxidation product? (Fe^{3+})

    (iii) For hydrogen peroxide (H_2O_2) to act as a powerful oxidising agent, what conditions are necessary and what is the reduction product?

        (the H_2O_2 must be acidified, i.e. $H^+_{(aq)}$ is needed and the reduction product is water)

Q6. Identify one oxidising agent that will oxidise zinc metal but not iron metal. (Cr^{3+})

Q7. Identify one reducing agent that will reduce Cu^+ to Cu metal but not reduce Cu^{2+} to Cu metal.

$(OH^-_{(aq)})$

Q8. For each of the following, search the ECS and identify the two appropriate half-equations ensuring that one is reduction (forward arrow) and that one is oxidation (reverse arrow). Then adjust them (if necessary) and combine these two half-equations to give a balanced net equation.

    (i) $Ag^+_{(aq)}$ reacts with $H_2O_{2(aq)}$ (ii) $Pb_{(s)}$ reacts with $MnO_{2(aq)}/H^+_{(aq)}$

    (iii) $SO_4^{2-}{}_{(aq)}/H^+_{(aq)}$ reacts with $Al_{(s)}$ (iv) $MnO_4^-{}_{(aq)}/H^+_{(aq)}$ reacts with aqueous iodide ions.

     (v) acidified hydrogen peroxide reacts with aqueous sulfur dioxide ($SO_{2(aq)}$)

Q9. "Free radicals" are molecules or ions with unpaired electrons and are thus highly reactive. Medical evidence suggests that free-radicals are the probable cause of a number of health problems. By *gaining* an additional electron from an electron donor, the free-radical can be eliminated. Suggest a reason why 'anti-oxidants' such as vitamin C may help eliminate dangerous free-radicals.

Q10. Oxalic acid has the formula $H_2C_2O_4$. In redox reactions, oxalic acid is converted to carbon dioxide (CO_2). Is oxalic acid acting as a reducing agent or an oxidising agent in this change? Explain.   (reducing agent)

Q11. Acidified potassium permanganate is a very powerful oxidising agent commonly used in the laboratory. Why is $MnO_4^-{}_{(aq)}$ usually acidified with sulfuric acid but never hydrochloric acid?

(HINT: can $MnO_4^-{}_{(aq)}$ react with the acid's anion?)

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15. COMMON OXIDISERS & REDUCERS YOU SHOULD KNOW:

The tables below show the more common oxidisers and reducers that you are likely to encouter during your chemistry course.

COMMON OXIDISERS

OXIDISER	USUAL REDUCTION PRODUCT
oxygen O_2	water H_2O
fluorine F_2	fluoride ions F^-
chlorine Cl_2	chloride ions Cl^-
bromine Br_2	bromide ions Br^-
iodine I_2	iodide ions I^-
(acidified) hydrogen peroxide H_2O_2	water H_2O
(acidified) permanganate ions MnO_4^-	manganese(II) ions Mn^{2+}
(alkaline) permanganate ions MnO_4^-	manganese dioxide MnO_2
(acidified) dichromate ions $Cr_2O_7^{2-}$	chromium(III) ions Cr^{3+}
hydrogen ions H^+	hydrogen gas H_2
(hot conc.) sulfuric acid H_2SO_4	sulfur dioxide SO_2
(acidified) nitrate ions NO_3^-	nitrogen dioxide NO_2
hypochlorite ions ClO^-	chloride ions Cl^-
iodate ions IO_3^-	iodine I_2

COMMON REDUCERS

REDUCER	USUAL OXIDATION PRODUCT
hydrogen H_2	water H_2O
carbon C	carbon monoxide CO or carbon dioxide CO_2
carbon monoxide CO	carbon dioxide CO_2
zinc Zn	zinc ions Zn^{2+}
aluminium Al	aluminium ions Al^{3+}
sulfur dioxide SO_2	sulfate ions SO_4^{2-}
sulfite ions SO_3^{2-}	sulfate ions SO_4^{2-}
hydrogen sulfide H_2S	sulfur S
iron(II) ions Fe^{2+}	iron(III) ions Fe^{3+}
oxalic acid $H_2C_2O_4$	carbon dioxide CO_2
oxalate ions $C_2O_4^{2-}$	carbon dioxide CO_2
(alkaline) hydrogen peroxide H_2O_2	oxygen O_2
thiosulfate ions $S_2O_3^{2-}$	tetrathionate ions $S_4O_6^{2-}$
iodide ions I^-	iodine I_2

16. REDOX STOICHIOMETRY PROBLEM (CRITERION 8):

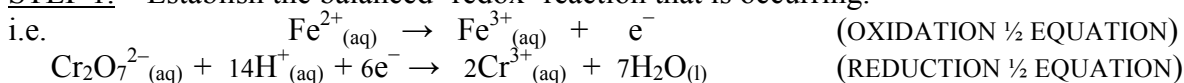The following problem is a typical example of the applications of our 'redox' theory.

Question. An aqueous solution of potassium dichromate was prepared by dissolving 2.312 g of $K_2Cr_2O_7$ crystals in 500.0 mL of acidic solution.

A piece of 'impure' iron wire of mass 0.5972 g was completely dissolved in sulfuric acid so that the iron was now ALL in the form of $Fe^{2+}_{(aq)}$ ions. The resulting solution was diluted to 100.0 mL and then 25.00 mL samples of this solution were then titrated against the acidified potassium dichromate solution. The average volume of $K_2Cr_2O_{7(aq)}$ required was 27.80 mL.

Find the % purity of the iron wire.
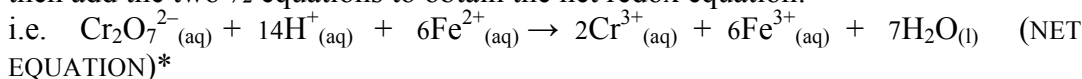ANSWER.

STEP 1.   Establish the balanced 'redox' reaction that is occurring.
i.e. $Fe^{2+}_{(aq)} \rightarrow Fe^{3+}_{(aq)} + e^-$ (OXIDATION ½ EQUATION)
    $Cr_2O_7{}^{2-}_{(aq)} + 14H^+_{(aq)} + 6e^- \rightarrow 2Cr^{3+}_{(aq)} + 7H_2O_{(l)}$ (REDUCTION ½ EQUATION)

In order to have $6e^-$ lost and as well as gained, we need to multiply the first ½ equation by 6; then add the two ½ equations to obtain the net redox equation:
i.e. $Cr_2O_7{}^{2-}_{(aq)} + 14H^+_{(aq)} + 6Fe^{2+}_{(aq)} \rightarrow 2Cr^{3+}_{(aq)} + 6Fe^{3+}_{(aq)} + 7H_2O_{(l)}$ (NET EQUATION)*

STEP 2.   Find the molarity of the $K_2Cr_2O_{7(aq)}$ solution.
                $n(K_2Cr_2O_7)$ =   mass/molar mass $M(K_2Cr_2O_7) = 294.2$ g mol^{-1}
                            =   2.312/294.2
                            = 7.859×10^{-3} mole   dissolved in 500 mL of solution
   Thus, concentration is:
              $[K_2Cr_2O_{7(aq)}]$ = moles/litres
                         = $7.859 \times 10^{-3}/0.5000$
                         =   0.01572 mol L^{-1}

STEP 3.   Find the average number of moles of $K_2Cr_2O_7$ used per titre.
              $n(K_2Cr_2O_7)$ =   molarity x litres
                         = $0.01572 \times 27.80 \times 10^{-3}$ mole
                         = 4.369×10^{-4} mole

STEP 4.   Find the average number of moles of $Fe^{2+}_{(aq)}$ that reacted per titre.
                $n(Fe^{2+})$ = $4.369 \times 10^{-4} \times (6/1)$ mole   (6:1 mole ratio from *)
                       = 2.622×10^{-3} mole in each 25 mL sample
                       = $2.622 \times 10^{-3} \times (100/25)$ mole in the original 100.0 mL
                     =   0.01049 mole in the original 100.0 mL

 STEP 5.   Find the mass of Fe^{2+} and hence Fe present and thus the % by mass.
               mass of Fe =   mole x $A_r(Fe)$ = 0.01049 x 55.85 g
                         =   0.5857 g
       Thus, percentage purity =   (0.5857/0.5972) x 100
                          =   98.07 %   (to 4 significant figures)

17. PREVIOUS TESTS & ANSWERS: CHEMISTRY (LEVEL 4C) OXIDATION ~ REDUCTION TEST

CRITERIA 5 & 8

30 marks

Q1. Complete the following statements:

(a) **OXIDATION** involves a process in which electrons are ....

(b) An **OXIDISING AGENT** (oxidiser or oxidant) is an electron ....

(c) Because Na is a powerful reducing agent, it means that Na^+ is a ....

(3 marks)

Q2. Give the oxidation number of:

(a) bromine in BrO_2^- ox(Br) = ...

(b) phosphorus in H_3PO_3 ox(P) = ...

(c) nitrogen in N_2 ox(N) = ...

(d) vanadium in VO^{2+} ox(V) = ...

(e) iodine in H_5IO_6 ox(I) = ...

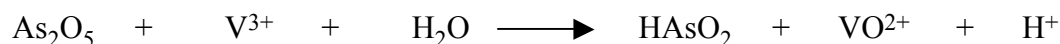(f) bismuth in BiO^+ ox(Bi) = ...

(6 marks)

Q3. Complete the balancing of the following half-equations and indicate whether they are oxidation or reduction.   OXN or REDN?

(a) $HClO_3$ $\longrightarrow$ Cl_2 (............)

(b) $Cr_2O_7^{2-}$ $\longrightarrow$ $Cr(OH)_3$ (............)

(c) HNO_3 $\longrightarrow$ N_2O (............)

(6 marks)

Q4 (a) Find the two "half-equations" and hence balance the following redox equation, :

$$As_2O_5 \; + \; V^{3+} \; + \; H_2O \; \longrightarrow \; HAsO_2 \; + \; VO^{2+} \; + \; H^+$$

(4 marks)

(b) In this reaction, which chemical species is:

(i) the reducing agent? ....

(ii) the oxidising agent? ....

(2 marks)

Q5. Use the electrochemical series (ECS) to find:

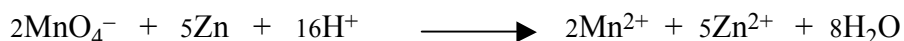(i) a reactant species that will reduce Ag^+ ions to Ag metal but will not reduce Cu^{2+} ions to Cu metal.

(1 mark)

(ii) a reactant species that will oxidise Sn metal to Sn^{2+} ions but will not oxidise Sn^{2+} ions to Sn^{4+} ions.

(1 mark)

Q6. Consider the balanced redox net ionic equation:

$$2MnO_4^- + 5Zn + 16H^+ \longrightarrow 2Mn^{2+} + 5Zn^{2+} + 8H_2O$$

A sample of *impure* zinc metal of mass 0.200 g was analysed by oxidising it with a solution of acidified $KMnO_4$ of concentration 0.0350 mol L^{-1}. The end-point was achieved after the addition of 27.6 mL of $KMnO_4$ solution. (Assume that the impurity in the metal doesn't react) Use these data to find the percentage by mass of zinc in the sample. (use 3 significant figures)

(7 marks)

~~~~~~

NAME: **ANSWERS**

26A
22B
16C

CHEMISTRY (LEVEL 5C)
OXIDATION ~ REDUCTION
TEST

CRITERIA 7 & 10

30 marks

Q1. Complete the following statements:

(a) OXIDATION involves a process in which electrons are ...... **LOST** ......

(b) An OXIDISING AGENT (oxidiser or oxidant) is an electron ...... **ACCEPTOR** ......

(c) Because Na is a powerful reductant means that Na+ is a **WEAK OXIDANT**

(3 marks)

Q2. Give the oxidation number of:

(a) bromine in BrO_2^{1-} ox(Br) = **+3**

(b) phosphorus in H_3PO_3 ox(P) = **+3**

(c) nitrogen in N_2 ox(N) = **0**

(d) vanadium in VO^{2+} ox(V) = **+4**

(e) iodine in H_5IO_6 ox(I) = **+7**

(f) bismuth in BiO^{1+} ox(Bi) = **+3**
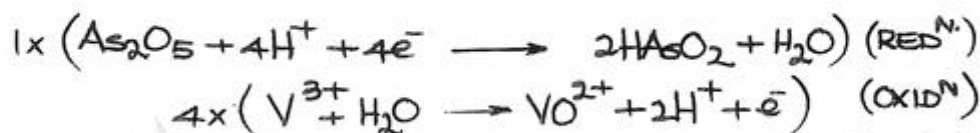
(6 marks)

Q3. Complete the balancing of the following half-equations and indicate whether they are oxidation or reduction.

(a) $2HClO_3 + 10H^+ + 10e^- \longrightarrow Cl_2 + 6H_2O$ (**RED^N**)

(b) $Cr_2O_7^{2-} + 8H^+ + 6e^- \longrightarrow 2Cr(OH)_3 + H_2O$ (**RED^N**)

(c) $2HNO_3 + 8H^+ + 8e^- \longrightarrow N_2O + 5H_2O$ (**RED^N**)

(6 marks)

Q4(a) Extract the half-equations and hence balance the net redox equation:

$$1\,As_2O_5 + 4\,V^{3+} + 3\,H_2O \longrightarrow 2\,HAsO_2 + 4\,VO^{2+} + 4\,H^+$$

$$1 \times (As_2O_5 + 4H^+ + 4e^- \longrightarrow 2HAsO_2 + H_2O)\ (RED^N)$$
$$4 \times (V^{3+} + H_2O \longrightarrow VO^{2+} + 2H^+ + e^-)\ (OXID^N)$$

(4 marks)

(b) In this reaction, which chemical species is:

(i) the reducing agent? ...... V^{3+} ......
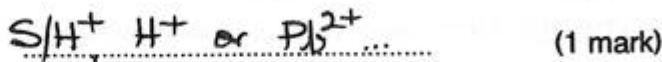
(ii) the oxidising agent? ...... As_2O_5 ......

(2 marks)

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Q5. Use the electrochemical series to find:

(i) a reactant species that will reduce Ag^{1+} ions to Ag metal but will not reduce Cu^{2+} ions to Cu metal.

Fe^{2+}, H_2O_2, I^- or OH^- ....   (1 mark)

(ii) a reactant species that will oxidise Sn metal to Sn^{2+} ions but will not oxidise Sn^{2+} ions to Sn^{4+} ions.

$S/H^+, H^+$ or Pb^{2+} ....   (1 mark)

Q6. Consider the balanced redox net ionic equation:

$$2MnO_4^{1-} + 5Zn + 16H^+ \longrightarrow 2Mn^{2+} + 5Zn^{2+} + 8H_2O$$

A sample of impure zinc metal of mass 0.200 g was analysed by oxidising it with a standardised solution of acidified $KMnO_4$ of concentration 0.0350 mol L^{-1}. The end-point was achieved after the addition of 27.6 mL of $KMnO_4$ solution. (Assume that the impurity in the metal doesn't react with the potassium permanganate solution)
Use these data to find the percentage by mass of zinc in the sample.

$$n(MnO_4^-) \text{ used} = \text{molarity} \times L$$
$$= (0.0350 \times 27.6 \times 10^{-3}) \text{ mole}$$
$$= 9.66 \times 10^{-4} \text{ mole}$$

$$\therefore n(Zn) \text{ present} = \left(9.66 \times 10^{-4} \times \frac{5}{2}\right) \text{ mole}$$
$$= 2.42 \times 10^{-3} \text{ mole}$$

$$\therefore \text{mass of } Zn \text{ present} = (\text{mole} \times M_r) \text{ g}$$
$$= (2.42 \times 10^{-3} \times 65.4) \text{ g}$$
$$= 0.158 \text{ g}$$

$$\text{Thus percentage by mass of } Zn = \left(\frac{\text{mass of } Zn}{\text{mass of sample}} \times 100\right)\%$$
$$= \left(\frac{0.158}{0.200} \times 100\right)$$
$$= 79.0\%$$

(7 marks)

OXIDATION ~ REDUCTION
REVISION QUESTIONS   **CRITERION**
5

Q1. Complete the following statements:

(a) REDUCTION involves a process in which electrons are ....

(b) An OXIDISING AGENT (oxidiser or oxidant) is an electron ....

(c) The fact that Au^{3+} is a powerful oxidising agent means that Au is a.....................

Q2. Give the oxidation number of:

(a) chlorine in ClO_4^- ox(Cl) = ...

(b) phosphorus in HPO_3^{2-} ox(P) = ...

(c) chlorine in Cl_2 ox(Cl) = ...

(d) vanadium in $V(OH)_4^+$ ox(V) = ...

(e) antimony in SbO_3^- ox(Sb) = ...

(f) chlorine in Cl^- ox(Cl) = ...

Q3. Complete the balancing of the following half-equations and indicate whether they are oxidation or reduction.

(a) H_2MoO_4 $\longrightarrow$ Mo

(b) H_2SeO_3 $\longrightarrow$ SeO_4^{2-}

(c) HNO_2 $\longrightarrow$ N_2O_4

Q4(a) Balance the redox equation:

H_2GeO_3 + Cr^{3+} + H_2O $\longrightarrow$ Ge   + CrO_4^{2-} + H^+
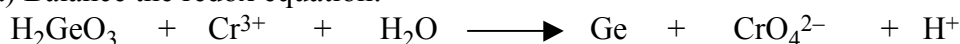
(b) In this reaction, which chemical species is the reducing agent?   ....

Q5. Which is most likely to be the most powerful oxidiser amongst the following oxy-acids of bromine? $HBrO_1$ $HBrO_2$ $HBrO_3$ $HBrO_4$ Explain briefly.

Q6. When Al metal reacts with nitric acid HNO_3 the metal is oxidised to Al^{3+} and the nitric acid is reduced to NO_2 gas.

(a) Write the balanced net ionic equation for this reaction.

(b) Write the balanced total equation for this reaction.

(c) What mass of aluminium metal reacts with 20.0 g of pure HNO_3?

(d) What mass of aluminium metal reacts with 125 mL of 8.00 mol/L HNO_3?

(e) If 25.0 g of aluminium metal was to react completely with excess nitric acid, what mass of NO_2 gas would be produced?

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OXIDATION ~ REDUCTION
REVISION QUESTIONS

CRITERION 7

Q1. Complete the following statements:

(a) REDUCTION involves a process in which electrons are ...GAINED..........

(b) An OXIDISING AGENT (oxidiser or oxidant) is an electron .ACCEPTOR.....

(c) The fact that Au^{3+} is a powerful oxidising agent means that Au is a..WEAK..REDUCER.

Q2. Give the oxidation number of:

(a) chlorine in ClO_4^{1-} ox(Cl)   = $+7$

(b) phosphorus in HPO_3^{2-} ox(P)   = $+3$

(c) chlorine in Cl_2 ox(Cl)   = O

(d) vanadium in $V(OH)_4^{1+}$ ox(V)   = $+5$

(e) antimony in SbO_3^{1-} ox(Sb)   = $+5$

(f) chlorine in Cl^{1-} ox(Cl)   = -1

Q3. Complete the balancing of the following half-equations and indicate whether they are oxidation or reduction.

(a) $H_2MoO_4 + 6H^+ + 6e^- \longrightarrow$ $Mo + 4H_2O$ (reduction)

(b) $H_2SeO_3 + H_2O \longrightarrow$ $SeO_4^{2-} + 4H^+ + 2e^-$ (oxidation)

(c) $2HNO_2 \longrightarrow$ $N_2O_4 + 2H^+ + 2e^-$ (oxidation)

Q4(a) Balance the redox equation:

$$3H_2GeO_3 + 4Cr^{3+} + 7H_2O \longrightarrow 3Ge + 4CrO_4^{2-} + 20H^{1+}$$

(b) In this reaction, which chemical species is the reducing agent? Cr^{3+}

Q5. Which is most likely to be the most powerful oxidiser amongst the following oxy-acids of bromine? $HBrO_1$ $HBrO_2$ $HBrO_3$ ($HBrO_4$)

Explain briefly.   ox(Br) in $HBrO_4$ is $7+$ ∴ STRONGEST e^- ACCEPTOR!

Q6. When Al metal reacts with nitric acid HNO_3 the metal is oxidised to Al^{3+} and the nitric acid is reduced to NO_2 gas.

(a) Write the balanced net ionic equation for this reaction. $Al + 3HNO_3 + 3H^+ \longrightarrow Al^{3+} + 3NO_2 + 3H_2O$.

(b) Write the balanced total equation for this reaction. $Al + 6HNO_3 \longrightarrow Al(NO_3)_3 + 3NO_2 + 3H_2O$

(c) What mass of aluminium metal reacts with 20.0 g of pure HNO_3? $1.43\ g$

(d) What mass of aluminium metal reacts with 125 mL of 8.00 mol/L HNO_3? $4.50\ g$

(e) If 25.0 g of aluminium metal was to react completely with excess nitric acid, what mass of NO_2 gas would be produced? $128\ g$

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