

## Topic 12b – Redox Revision Notes

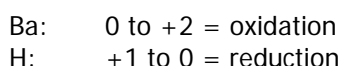
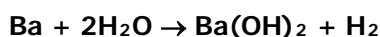
### 1) Introduction

- Oxidation is loss of electrons
- Reduction is gain of electrons
- An oxidising agent accepts electrons (and is reduced in the process)
- A reducing agent donates electrons (and is oxidised in the process)

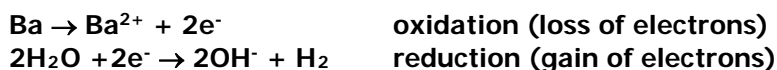
### 2) Oxidation states

- Oxidation states are “charges” assigned to each element in a reaction. The rules for assigning oxidation states are:
  - Elements are zero
  - In compounds, H is +1 and O is -2
  - In compounds, Group 1 elements are +1, Group 2 are +2, Group 6 are -2 and Group 7 are -1 (these are real charges)

- Be able to apply this to redox reactions of group 2 elements with water and oxygen e.g.

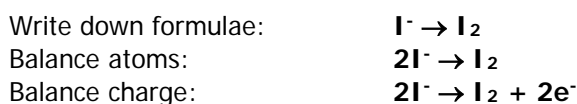


- The overall equation can be split into half-equations, one for the oxidation and one for the reduction:

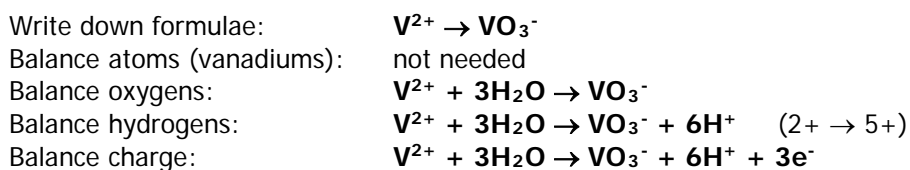


### 3) Constructing half-equations

- Constructing simple half-equations involves two steps: balancing atoms and balancing charge by adding electrons e.g. oxidation of iodide ions, I<sup>-</sup>, to iodine, I<sub>2</sub>.

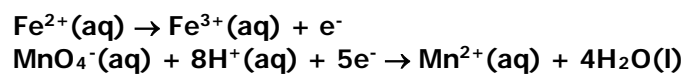


- Constructing more complicated half-equations involves 2 extra steps, namely: balancing oxygens by adding water and balancing hydrogens by adding H<sup>+</sup> e.g. oxidation of V<sup>2+</sup> to VO<sub>3</sub><sup>-</sup>



#### 4) Redox equations

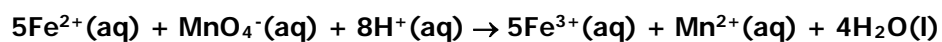
- To produce an overall redox equation, multiply one or both half-equations until the number of electrons is the same



- In this case, multiply the first half-equation by 5 to get 5 electrons in both.



- Now add the half-equations together, cancelling the electrons at the same time.



- Any species that appears on both sides of the equation needs to be cancelled  
e.g.  $\text{H}^{+}$