CHEMISTRY – UNIT 5B STUDY GUIDE

When is it happening?	Wednesday May 6th
How long will it take?	35 minutes
What is the format?	multiple choice questions (5- 10 points), short answer questions (15 - 20 points)
What is it worth?	9.4% of your Q4 grade
What will it cover?	See below
What resources will be useful?	Unit 5B Lesson Helpsheets, Lab Reports 5.6 – 5.7 Class Worksheets 5.6 – 5.10

Introduction to Oxidation and Reduction

- during chemical reactions, atoms can gain or lose electrons
- when atoms lose electrons, their charge increases; we call this **oxidation** (OIL); we can represent oxidation with a half-equation, eg $Zn \rightarrow Zn^{2+} + 2e^-$
- when atoms gain electrons, their charge decreases; we call this **reduction** (RIG); we can represent reduction with a half-equation, eg $Cu^{2+} + 2e^- \rightarrow Cu$
- a reaction in which electrons are transferred from one atom to another is called a redox reaction
- during redox reactions, one atom is oxidized and one atom is reduced
- the substance containing the atom being reduced is the **oxidizing agent (OA)** and the substance containing the atom being oxidized is called the **reducing agent (RA)**
- you can identify the atom being oxidized and reduced in a chemical reaction if you know the charge which atoms have when they are in elements and compounds:

Charge	-2	-1	0	+1	+2
Which atoms?	Oxygen eg O^{2-} when it is in a compound	All atoms in Group 7 eg Cl^-, Br^- when they are in compounds	All atoms when they are on their own or in elements Eg Na, Cu, H_2, Cl_2	all atoms in Group 1 eg Na^+, K^+ and hydrogen (H^+) when they are in compounds	all atoms in Group 2 eg Ca^{2+}, Mg^{2+} and most common metals eg Zn^{2+}, Cu^{2+} when they are in compounds
Examples	O in MgO O in H_2O O in CuO	Cl in NaCl Cl in $MgCl_2$ Br in $CaBr_2$	Cu in Cu H in H_2 O in O_2	Na in NaCl K in K_2O H in HBr	Cu in CuO Mg in $MgCl_2$ Ca in $CaBr_2$
	⇒⇒⇒⇒⇒⇒⇒⇒⇒⇒⇒⇒⇒⇒⇒⇒⇒⇒⇒⇒⇒⇒⇒⇒⇒⇒⇒⇒ oxidation (charge increases)				
	⇐⇐⇐⇐⇐⇐⇐⇐⇐⇐⇐⇐⇐⇐⇐⇐⇐⇐⇐⇐⇐⇐⇐⇐⇐⇐⇐⇐⇐⇐ reduction (charge decreases)				

Example: $Zn + CuO \rightarrow ZnO + Cu$ Charges: 0 +2,-2 → +2,-2 0 Zn is oxidized from 0 to +2 ($Zn \rightarrow Zn^{2+} + 2e^-$) Cu is reduced from +2 to 0 ($Cu^{2+} + 2e^- \rightarrow Cu$) Zn is the RA, CuO is the OA	Example: $2Na + Cl_2 \rightarrow 2NaCl$ Charges: 0 0 → +1,-1 Na is oxidized from 0 to +1 ($Na \rightarrow Zn + 2e^-$) Cl is reduced from 0 to -1 ($Cl_2 + 2e^- \rightarrow 2Cl^-$) Na is the RA, Cl_2 is the OA

Reactivity of Metals and Displacement Reactions

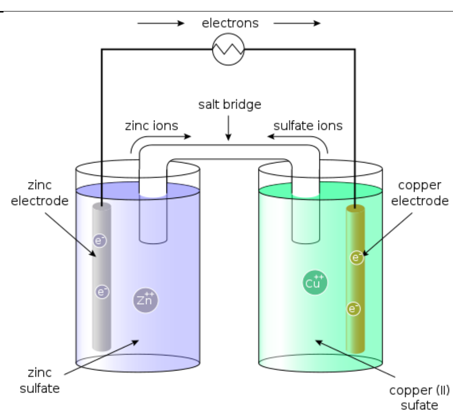
- the reactivity of a metal depends on how easily it loses electrons; the more easily it loses electrons, the more reactive it is
- the most reactive metals are those in Group 1, followed by those in Group 2; these metals get more reactive as you go down the group
- when metals react with water and acids, the metal is oxidized and H is reduced:
  eg $2Na + 2H_2O \rightarrow 2NaOH + H_2$
  eg $Mg + 2HCl \rightarrow MgCl_2 + H_2$
- metals can displace less reactive metals from their compounds; the more reactive metal is oxidized and the less reactive metal is reduced
  eg $Zn + CuCl_2 \rightarrow ZnCl_2 + Cu$
  (Zn (oxidized) is more reactive than Cu (reduced))
- metals which are less reactive than carbon can be extracted from their oxides by heating with carbon
  eg $ZnO + C \rightarrow Zn + CO$ (C is more reactive than Zn)
- metals which are more reactive than carbon cannot be extracted from their oxides like this

Reactivity Series of Metals (with C and H added)

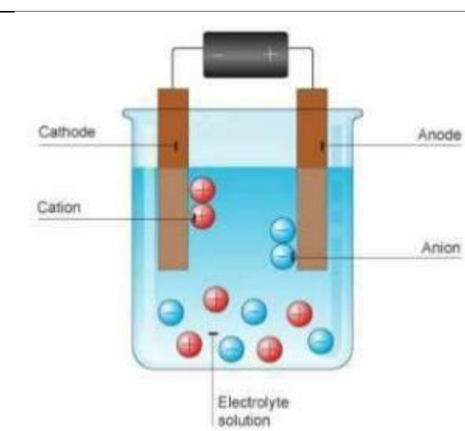
K (explosive reaction with water)
Na (explosive reaction with water)
Ca (steady reaction with water)
Mg (fast reaction with acids)
Al (anomalously slow - forms a stable oxide layer)
C
Zn (steady reaction with acids)
Fe (slow reaction with acids)
Sn (very slow reaction with acids)
Pb (very slow reaction with acids)
H
Cu (no reaction with acids)
Ag (no reaction with acids)
Au (no reaction with acids)

Electrochemical Cells

- there are two types of electrolytic cell: **galvanic cells** and **electrolytic cells**
- a **galvanic cell** is a device which uses a redox reaction to create electricity
- an **electrolytic cell** is a device which uses electricity to force a redox reaction to take place
- both cell types contain conducting rods (electrodes) immersed in a liquid which contains ions (electrolyte)
- reduction always takes place at the **cathode** and oxidation takes place at the **anode**



a typical galvanic cell
reduction takes place in the cathodic half-cell
oxidation takes place in the anodic half-cell
electrons move from the (-) anode to the (+) cathode



a typical electrolytic cell
cations move to the (-) cathode and are reduced
anions move to the (+) anode and are oxidised

Type	Galvanic	electrolytic
purpose	all batteries are made from galvanic cells	- to extract reactive metals from their ores - to make various chemicals - electroplating
energy change	chemical potential → electrical	electrical → chemical potential
terminals	cathode (reduction) is +ve anode (oxidation) is -ve	cathode (reduction) is -ve anode (oxidation) is +ve
chemical reaction	- can be any redox reaction - the simplest cells involve metal displacement reactions; the more reactive metal is oxidised and the less reactive metal is reduced	- in molten electrolytes, the compound is separated into its elements - in aqueous solutions, either the ionic compound or water can be separated into its elements
visible changes	(simple cells only) - anode gets smaller; anodic solution gets more concentrated - cathode gets bigger; cathodic solution gets less concentrated	- cations move to the cathode and are reduced; a metal or hydrogen is produced - anions move to the anode and are oxidised; a halogen or oxygen is produced

- Useful galvanic cells include: lead-acid cell, alkali cell, lithium-ion cell
- Some galvanic cells can be recharged
  - electricity is forced through the cell in the opposite direction, reversing the chemical reaction
  - during this process the galvanic cell is being converted into an electrolytic cell
  - if the reverse reaction can't take place the cell cannot be recharged