CHEM 9.2.4 ELECTROCHEMISTRY

Oxidation-reduction reactions are increasingly important as a source of energy

4.1 Explain the **displacement of metals** from solution in terms of **transfer of electrons**

- **Displacement reaction**: reaction where **a metal converts the ion** of another metal to the **neutral** atom

$$Cu_{(s)} + 2Ag^{+}_{(aq)} \rightarrow 2Ag_{(s)} + Cu^{2+}_{(aq)}$$

Copper + silver nitrate → silver + copper nitrate (net ionic equation)

- **OIL RIG** – oxidation is losing e-, reduction is gaining e-
    - **Oxidation** half-equation: $Cu_{(s)} \rightarrow Cu^{2+}_{(aq)} + 2e^{-}$
    - **Reduction** half-equation: $2Ag^{+}_{(aq)} + 2e^{-} \rightarrow 2Ag_{(s)}$
- Also called **redox/electron transfer** reactions
- More active metals displace less active metal ions

4.2 Identify the relationship between **displacement of metal ions** in solution by other metals to the **relative activity of metals**

- Higher in activity series = **displaces other metal** (which is in solution)
    - Copper becomes copper nitrate as it is more reactive
- Higher in activity series = more active – loses electrons to lower metals = oxidises
- Hydrogen – above reacts with dilute acid to form H_2

oxidises (best reductant)

reverse this metal's half-equation

reduces (best oxidant)

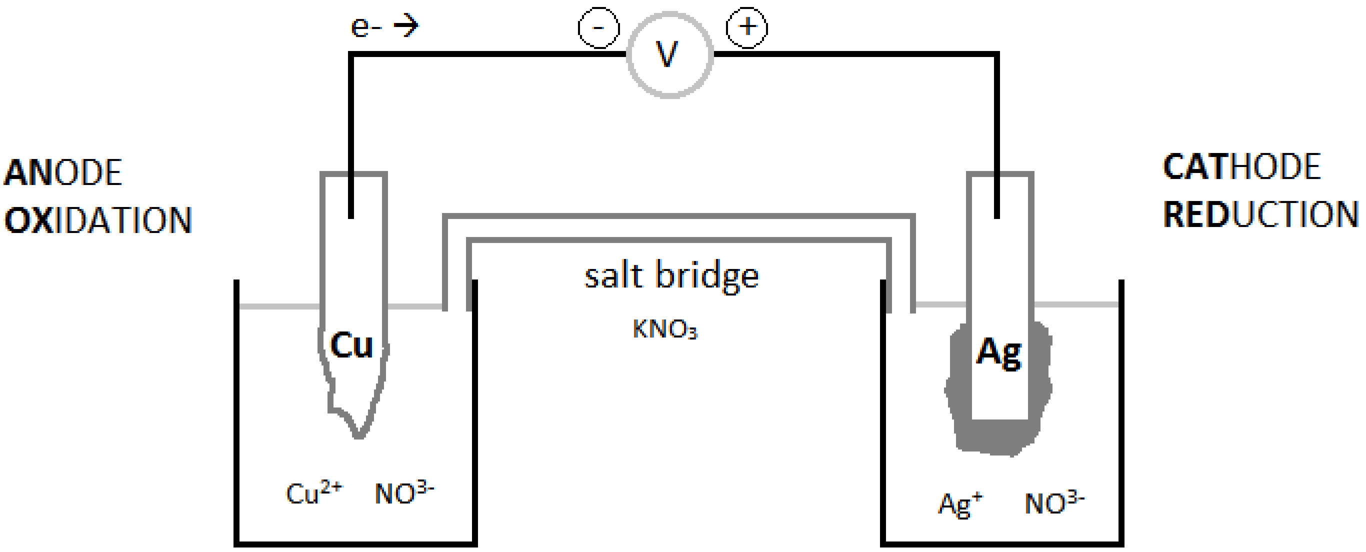
4.3 Account for changes in the **oxidation state** of species in terms of their **loss or gain** of electrons

- Oxidation state: **charge of ion** (incl. the sign)
- **Oxygen is always -2, H is always +1**

Type	What to do	Type
Element	0	Cl_2 = 0
Ion	Charge on ion	Cl^- = -1
Polyatomic	Sum = charge	SO_4^{2-} – since O is -2 O = -8, S = 6
Molecular Compound	Sum is 0	SO_2 – since O is 2 O = -4, S = +4

- Used to find if **oxidation** or **reduction:**
    - $Mg \rightarrow Mg^{2+} + e^-$
      0 → +2 - +'ve change means losing electrons – therefore Mg is reductant.
    - $2H^+ + 2e^- \rightarrow H_2$
      2+ → 0 - -'ve change means gaining electrons – therefore H^+ is oxidant

GALVANIC CELL (4.4, 4.5, 4.6)



4.4 Describe and explain **galvanic cells** in terms of **oxidation/reduction reactions**

- Both reactions **occur in different locations**
- Normally drawn with **left more active**
  - Anode (left) is the oxidation of **copper** $Cu_{(s)} \rightarrow Cu^{2+}_{(aq)} + 2e^-$
  - Cathode (right) is the reduction of **silver** $2Ag^+_{(aq)} + 2e^- \rightarrow 2Ag_{(s)}$

4.5 Outline the **construction of galvanic cells** and trace the **direction of electron flow**

- See above diagram for construction
- Salt bridge used to balance imbalance of + and – ions, normally KNO_3
  - Allows migration to occur
  - NO_3 ions **pushed towards anode**
  - K ions **pushed towards cathode**
- Electrons pass through wire from left to right – voltage is always +'ve

4.6 Define the terms **anode, cathode, electrode and electrolyte** to describe galvanic cells

- **Anode**: The electrode in which **oxidation** occurs — Cu electrode — **negative**
- **Cathode**: The electrode in which **reduction** occurs — Ag electrode — **positive**
- **Electrode**: **conductors of a cell** that is **connected to an external circuit**
- **Electrolyte**: substance which, in solution or molten, **conducts electricity**