

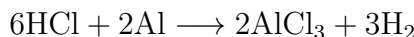
# Chemistry Honors Study Guide

## Test 3 S2

Test date: May 2, 2025

## 1 Gasses and Heat in Stoichiometry

### Gasses



At STP, how many *ml* of  $\text{H}_2$  gas are produced from 12 *g* of solid Al? (1 *mol* = 22.4 *L* at STP)

Using stoichiometry:

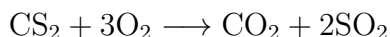
$$\begin{aligned} (12 \text{ g Al}) \times \left( \frac{1 \text{ mol Al}}{26.98 \text{ g Al}} \right) \times \left( \frac{3 \text{ mol H}_2}{2 \text{ mol Al}} \right) \times \left( \frac{22.4 \text{ L H}_2}{1 \text{ mol H}_2} \right) \times \left( \frac{1000 \text{ ml}}{1 \text{ L}} \right) \\ = \boxed{14944.4 \text{ ml}} \end{aligned}$$

### Heats of Formation

$\Delta H_f$  is the heat absorbed/released when compounds are formed from elemental units. The  $\Delta H_f$  of elements, including diatomic elements, is always 0.

Heats of formation equation:

$$\Delta H_{\text{rxn}} = \sum \Delta H_{\text{f(products)}} - \sum \Delta H_{\text{f(reactants)}} \quad (1)$$



Find the heat of formation given the following:

$$\begin{aligned} \Delta H_f(\text{CO}_2) &= -393.5 \frac{\text{kJ}}{\text{mol}} \\ \Delta H_f(\text{SO}_2) &= -296.8 \frac{\text{kJ}}{\text{mol}} \\ \Delta H_f(\text{CS}_2) &= 87.9 \frac{\text{kJ}}{\text{mol}} \end{aligned}$$

Solution: Using 1:

$$\begin{aligned} & [-393.5 + 2(-296.8)] - [3(0) + 87.9] \\ &= \boxed{1075 \frac{kJ}{mol}} \end{aligned}$$

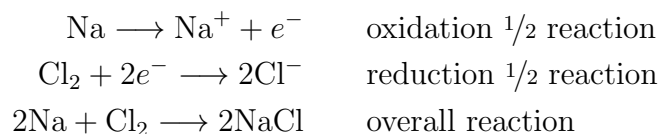
## 2 Redox

### Definitions

- **redox reaction**: Short for *oxidation/reduction reaction*.
- **oxidation**: Losing electrons (becoming more positive).
- **reduction**: Gaining electrons (becoming more negative).
- **oxidizing agent**: Oxidizing another compound and is being reduced.
- **reducing agent**: Reducing another compound and is being oxidized.

### Reactions

LEO the lion says GER (Losing Electrons = Oxidation, Gaining Electrons = Reduction)



### Assigning Oxidation Numbers

#### Rules

1. An element in its elemental state is neutral.
2. H in a compound is always +1.
3. O in a compound is -2 except for  $\text{H}_2\text{O}_2$ ; in that case, the oxidation number of O is -1.
4. Monatomic ions are whatever ionic charge they would normally form.
5. All oxidation numbers must add to the overall charge of the molecule or ion.

Example: What is the oxidation number of I in  $\text{Mg}(\text{IO}_3)_2$ ?

Solution: By looking at the charge on Mg, the charge on  $(\text{IO}_3)_2$  can be found to be 2-, so the charge on one ion is 1-, making the ion  $(\text{IO}_3)^-$ . The oxidation number of O in this compound is 2- by rule 3. There are 3 oxygen atoms, so:

$$3(-2) + \text{I} = -1$$

Solving for I:

$$-6 + I = -1$$

$$I = -1 + 6$$

$$I = 5$$

The charge on I is +5.

## Writing Half-Reactions

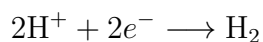


To write half-reactions, look at the oxidation numbers on each element. If the oxidation number has increased, it has been oxidised. If the number has decreased, it has been reduced.

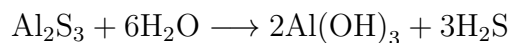
Oxidation  $1/2$  reaction:



Reduction  $1/2$  reaction:



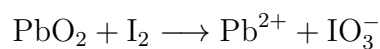
If *none* of the oxidation numbers change, the reaction is *not* a redox reaction. For example,



is not a redox reaction.

## Balancing Redox Reactions

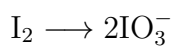
Assume an acidic, aqueous environment:



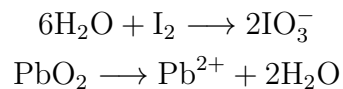
1. Separate into oxidation and reduction half-reactions by looking at oxidation numbers.



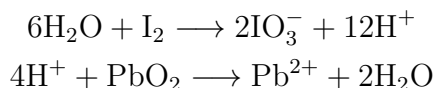
2. Balance elements that are *not* H or O.



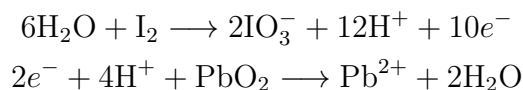
3. Balance O by adding H<sub>2</sub>O.



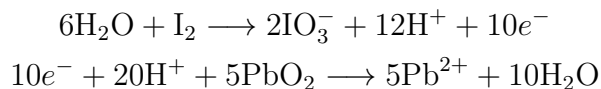
4. Balance H by adding H<sup>+</sup>.



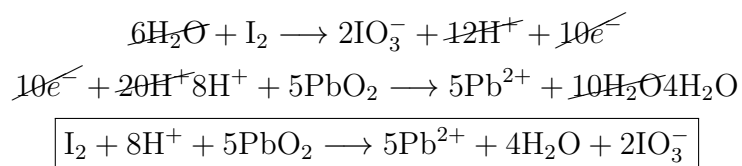
5. Add electrons to balance the charges. When balanced correctly, electrons should be on opposite sides (one on reactants, one on products).



6. Multiply to cancel out the electrons.



7. Combine and cancel.

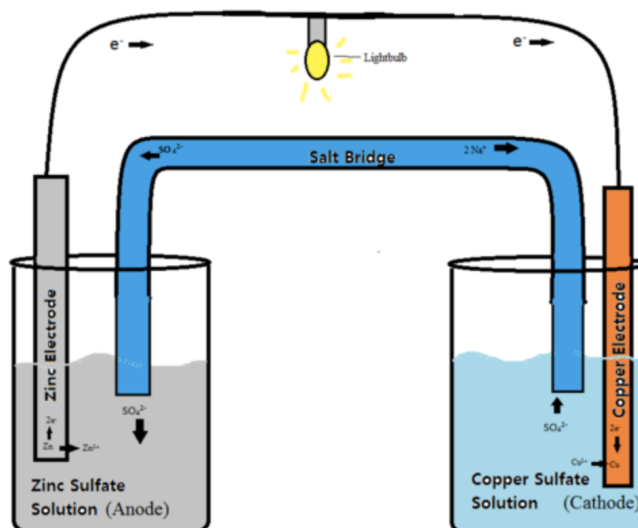


### 3 Batteries

Below is a diagram of a Voltaic/Galvanic cell<sup>1</sup>:

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<sup>1</sup>[bit.ly/4jMzl6a](https://bit.ly/4jMzl6a)



containing:

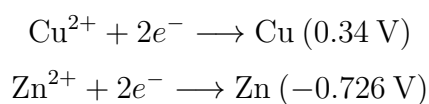
**electrodes** to carry electricity. The **anode** is oxidised, losing mass as the material is released into solution, and the **cathode** is reduced, gaining mass as solid metal deposits onto it.

**salt bridge** to close the circuit.

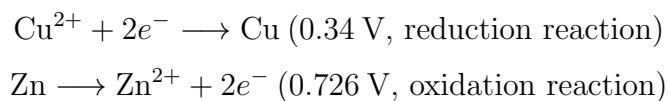
**solutions** that the electrodes are suspended in.

Electrons flow from the anode to the cathode.

To find the voltage of this battery:



Flip the equation that is higher on the standard reduction potential table:



Combine:



N.B. When **multiplying** an equation to balance out the electrons, **the voltage does not change**. It only changes in sign if the equation is flipped.