Chemistry Honors Study Guide

Test 3 S2

Test date: May 2, 2025

1 Gasses and Heat in Stoichiometry

Heat of Reaction Stoichiometry

$$7\mathrm{H}_{2(g)} + 2\mathrm{NO}_{2(g)} \longrightarrow 2\mathrm{NH}_{3(g)} + 4\mathrm{H_2O}_{(l)} \quad \Delta H = 142.5\,kJ$$

In the above equation, what would ΔH be if 4.8 mol of $\mathrm{H_2}$ were reacted? Is this reaction endothermic or exothermic? Where does heat belong in this equation: with the reactants or products?

This reaction is **endothermic** because the ΔH is positive, meaning it absorbs heat. Heat belongs on the **reactants** side.

$$7\mathrm{H}_{2(g)} + 2\mathrm{NO}_{2(g)} + \text{heat} \longrightarrow 2\mathrm{NH}_{3(g)} + 4\mathrm{H_2O}_{(l)} \quad \Delta H = 142.5\,kJ$$

Step 1: Find the heat per mole of $\mathrm{H_2}$.

$$\frac{142.5\,kJ}{7\,mol\ \mathrm{H_2}} = 20.36\,\frac{kJ}{mol}$$

Step 2: Find the total heat by multiplying by 4.8 mol.

$$20.36\,\frac{kJ}{mol} \times 4.8\,mol = \boxed{97.73\,kJ}$$

If 4.8 mol of $\mathrm{H_2}$ were reacted, the ΔH would be 97.73 kJ.

Gasses

$$6\mathrm{HCl} + 2\mathrm{Al} \longrightarrow 2\mathrm{AlCl_3} + 3\mathrm{H_2}$$

At STP, how many ml of $\mathrm{H_2}$ gas are produced from 12 g of solid Al? (1 mol = 22.4 L at STP)

Using stoichiometry:

$$(12\,g\ \mathrm{Al}) \times \left(\frac{1\,mol\ \mathrm{Al}}{26.98\,g\ \mathrm{Al}}\right) \times \left(\frac{3\,mol\ \mathrm{H_2}}{2\,mol\ \mathrm{Al}}\right) \times \left(\frac{22.4\,L\ \mathrm{H_2}}{1\,mol\ \mathrm{H_2}}\right) \times \left(\frac{1000\,ml}{1\,L}\right)$$

$$= \boxed{14944.4\,ml}$$

Heats of Formation

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

Heats of formation equation:

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

$$\mathrm{CS_2 + 3O_2 \longrightarrow CO_2 + 2SO_2}$$

Find the heat of formation given the following:

$$\Delta H_\mathrm{f}\,(\mathrm{CO_2}) = -393.5\,\frac{kJ}{mol}$$
$$\Delta H_\mathrm{f}\,(\mathrm{SO_2}) = -296.8\,\frac{kJ}{mol}$$
$$\Delta H_\mathrm{f}\,(\mathrm{CS_2}) = 87.9\,\frac{kJ}{mol}$$

Solution: Using 1:

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

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)

$$\begin{aligned} \mathrm{Na} &\longrightarrow \mathrm{Na^+} + e^- && \text{oxidation } 1/2 \text{ reaction} \\ \mathrm{Cl_2} + 2e^- &\longrightarrow \mathrm{2Cl^-} && \text{reduction } 1/2 \text{ reaction} \\ \mathrm{2Na + Cl_2} &\longrightarrow \mathrm{2NaCl} && \text{overall reaction} \end{aligned}$$

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 H_2O_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 $Mg(IO_3)_2$?

Solution: By looking at the charge on Mg, the charge on $(IO_3)_2$ can be found to be 2-, so the charge on one ion is 1-, making the ion $(IO_3)^-$. The oxidation number of O in this compound is 2- by rule 3. There are 3 oxygen atoms, so:

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

Solving for I:

$$-6 + \mathrm{I} = -1$$
$$\mathrm{I} = -1 + 6$$
$$\mathrm{I} = 5$$

The charge on I is $\boxed{+5}$.

Writing Half-Reactions

$$2\mathrm{Ag} + \mathrm{H_2S} \longrightarrow \mathrm{Ag_2S} + \mathrm{H_2}$$

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 ½ reaction:

$$2\mathrm{Ag} \longrightarrow \text{(break the ionic bond) } 2\mathrm{Ag}^+ + 2e^-$$

Reduction ½ reaction:

$$2\mathrm{H}^+ + 2e^- \longrightarrow \mathrm{H_2}$$

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

$$\mathrm{Al_2S_3} + 6\mathrm{H_2O} \longrightarrow 2\mathrm{Al(OH)_3} + 3\mathrm{H_2S}$$

is not a redox reaction.

Balancing Redox Reactions

Assume an acidic, aqueous environment:

$$PbO_2 + I_2 \longrightarrow Pb^{2+} + IO_3^-$$

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

$$\text{Oxidation half-reaction: } I_2 \longrightarrow IO_3^-$$
$$\text{Reduction half-reaction: } PbO_2 \longrightarrow Pb^{2+}$$

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

$$I_2 \longrightarrow 2IO_3^-$$
$$PbO_2 \longrightarrow Pb^{2+}$$

3. Balance O by adding H_2O.

$$6H_2O + I_2 \longrightarrow 2IO_3^-$$
$$PbO_2 \longrightarrow Pb^{2+} + 2H_2O$$

4. Balance H by adding H^+.

$$6H_2O + I_2 \longrightarrow 2IO_3^- + 12H^+$$
$$4H^+ + PbO_2 \longrightarrow Pb^{2+} + 2H_2O$$

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

$$6H_2O + I_2 \longrightarrow 2IO_3^- + 12H^+ + 10e^-$$
$$2e^- + 4H^+ + PbO_2 \longrightarrow Pb^{2+} + 2H_2O$$

6. Multiply to cancel out the electrons.

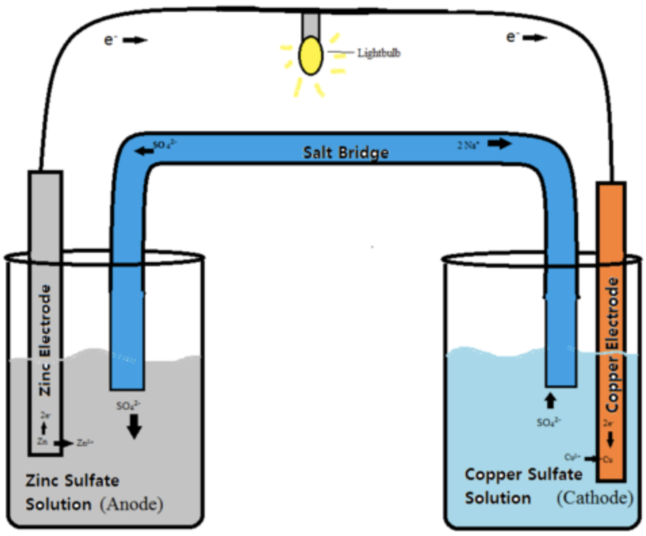
$$6H_2O + I_2 \longrightarrow 2IO_3^- + 12H^+ + 10e^-$$
$$10e^- + 20H^+ + 5PbO_2 \longrightarrow 5Pb^{2+} + 10H_2O$$

7. Combine and cancel.

$$\cancel{6H_2O} + I_2 \longrightarrow 2IO_3^- + \cancel{12H^+} + \cancel{10e^-}$$
$$\cancel{10e^-} + \cancel{20H^+} 8H^+ + 5PbO_2 \longrightarrow 5Pb^{2+} + \cancel{10H_2O} 4H_2O$$
$$\boxed{I_2 + 8H^+ + 5PbO_2 \longrightarrow 5Pb^{2+} + 4H_2O + 2IO_3^-}$$

3 Batteries

Below is a diagram of a Voltaic/Galvanic cell[1]:



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:

$$Cu^{2+} + 2e^- \longrightarrow Cu \; (0.34 \text{ V})$$
$$Zn^{2+} + 2e^- \longrightarrow Zn \; (-0.726 \text{ V})$$

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

[1] bit.ly/4jMzl6a

$$\mathrm{Cu}^{2+} + 2e^- \longrightarrow \mathrm{Cu}\ (0.34\ \mathrm{V},\ \text{reduction reaction})$$
$$\mathrm{Zn} \longrightarrow \mathrm{Zn}^{2+} + 2e^-\ (0.726\ \mathrm{V},\ \text{oxidation reaction})$$

Combine:

$$\mathrm{Cu}^{2+} + \mathrm{Zn} \longrightarrow \mathrm{Cu} + \mathrm{Zn}^{2+}\ \boxed{(1.1\ \mathrm{V})}$$

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.