Chemistry Honors Study Guide

Test 3 S2

Test date: TBD

1 Gasses and Heat in Stoichiometry

Gasses

$$6\text{HCl} + 2\text{Al} \longrightarrow 2\text{AlCl}_3 + 3\text{H}_2$$

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

Using stoichiometry:

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

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:

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

$$\text{CS}_2 + 3\text{O}_2 \longrightarrow \text{CO}_2 + 2\text{SO}_2$$

Find the heat of formation given the following:

$$\Delta H_\text{f}\,(\text{CO}_2) = -393.5\ \frac{kJ}{mol}$$
$$\Delta H_\text{f}\,(\text{SO}_2) = -296.8\ \frac{kJ}{mol}$$
$$\Delta H_\text{f}\,(\text{CS}_2) = 87.9\ \frac{kJ}{mol}$$

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).

Reactions

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

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

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 $^1/_2$ reaction:

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

Reduction $^1/_2$ 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.