Chemistry 30 (AB) Errata (2008 Edition)

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Please see the attached pages for the complete text.

NOTES

Example 3

When 1.00 g of diphenylmethanone, $(C_6H_5)_2CO_{(s)}$, burns to produce $CO_{2(g)}$ and $H_2O_{(l)}$, 35.74 kJ of heat energy are released in a bomb calorimeter. Calculating a molar enthalpy of combustion of diphenylmethanone, what is a molar enthalpy of formation?

Solution

The balanced reaction equation for diphenylmethanone combustion is

$$\begin{array}{ccccccc} (C_6H_5)_2CO_{(s)} & + & 15O_{2(g)} & \rightarrow & 13CO_{2(g)} & + & 5H_2O_{(l)} \\ ? & & 0 & & -393.5\ \text{kJ/mol} & & -285.8\ \text{kJ/mol} \end{array}$$

Diphenylmethanone releases 35.74 kJ of energy per gram burned. Diphenylmethanone's molar mass can be calculated, straight from its chemical formula, to be 182.23 g/mol:

$$\frac{-35.74\ \text{kJ}}{1\ \text{g}} = \frac{\Delta_c H_m{}^\circ}{1\ \text{mol}} = \frac{\Delta_c H^\circ}{182.23\ \text{g}}$$

$$\Delta_c H_m{}^\circ = -6\,512.9\ \text{kJ/mol}$$

$$\sum_{\text{reactants}} (n\Delta_f H_m{}^\circ) = \sum_{\text{products}} (n\Delta_f H_m{}^\circ) - \Delta_r H^\circ$$

$$n\Delta_f H_m{}^\circ((C_6H_5)_2CO_{(s)}) = (13\ \text{mol})(-393.5\ \text{kJ/mol}) + (5\ \text{mol})(-285.8\ \text{kJ/mol}) - (-6\,512.9\ \text{kJ})$$

$$n\Delta_f H_m{}^\circ((C_6H_5)_2CO_{(s)}) = -31.6\ \text{kJ}$$

In this case $n = 1$ $\therefore \Delta_f H_m{}^\circ\left((C_6H_5)_2CO_{(s)}\right) = -31.6\ \text{kJ}$

PRACTICE EXERCISE

Formulae: $n\Delta_c H_m = Q_{H_2O_{(l)}} + Q_{can}$ $\quad n\Delta_c H_m = C\Delta T$

1. a) What quantity of water could be heated from 20.0˚C to 100˚C, without boiling, on a stove that burns 60.0 g of butane ($\Delta_c H_m = -2\ 657.4$ kJ/mol), assuming that all the heat goes into heating the water only?

b) What are three factors involved in heating the water in question 1 that make the calculated mass of water unrealistic?

2. What quantity of methane would have to burn on an efficient stove top to heat 1.00 kg of water, in a 300 g aluminum pot, from 10.0°C to 90.0°C, without boiling? $(\Delta_c H_m = -802.5 \text{ kJ/mol})$

3. a) Which automotive fuel, natural gas (assume methane) or propane, releases more energy per kilogram of fuel burned? $(\Delta_c H_m)_{CH_{4(g)}} = -802.5$ kJ/mol, $(\Delta_c H_m)_{C_3H_{8(g)}} = -2\ 043.9$ kJ/mol.

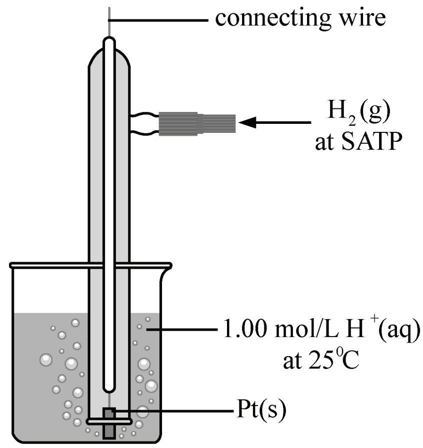
NOTES

oxidizing agent present in a cell is found higher on a redox table than the strongest reducing agent present, the reaction will proceed spontaneously. If the strongest oxidizing agent is lower than the strongest reducing agent, the reaction is non-spontaneous—this is found with electrolytic cells.

THE STANDARD HYDROGEN HALF-CELL

The standard reduction half-potential of a cell is determined by comparing it to a standard hydrogen half-cell (called a reference half-cell) operating at SATP conditions of 25˚C and 100 kPa. The standard hydrogen half-cell uses an inert metal electrode, such as platinum, in a solution of 1.00 mol/L $H^+_{(aq)}$, with $H_{2(g)}$ passing over the surface of the electrode. As the H^+ is reduced to $H_{2(g)}$, the potential that results is arbitrarily said to be 0 volts.

$$2H^+_{(aq)} + 2e^- \rightarrow H_{2(g)} \qquad E^\circ_r = 0.00V$$



Pt(s)| H_2(g)| H^+(aq) E^0_r= 0.00 V

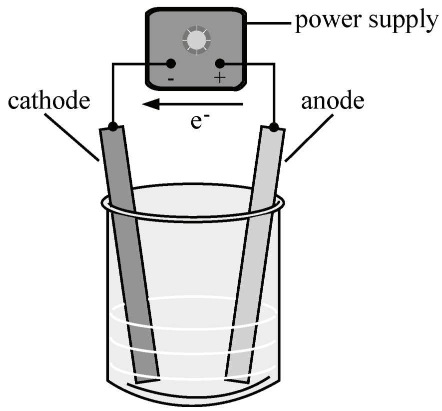
Because the standard reduction potential (E°_r) measures the tendency to undergo a reduction, a half-cell with a greater tendency to be reduced (a stronger oxidizing agent) than hydrogen ions is assigned a positive E°_r. A half-cell with a lesser tendency to be reduced (a weaker oxidizing agent) than hydrogen ions is assigned a negative E°_r. A voltmeter is used to measure the flow of electrons between half-cells and to find the cell potential.

Lesson 5 ELECTROLYTIC CELLS

NOTES

Electrolytic cells require an input of electricity to force a non-spontaneous reaction to occur. The cell potential (E°cell) is negative. An electrolytic cell consists of an electrolyte and two electrodes attached to an external power supply. The process of supplying energy to force a non-spontaneous reaction is called *electrolysis*. The power supply forces an electron transfer inside the cell, causing electrons to move away from the entity they are most attracted to.

As in the voltaic cell, the strongest oxidizing agent undergoes a reduction at the cathode, and the strongest reducing agent undergoes an oxidation at the anode. In a voltaic cell, the flow of electrons originates at an electrode (the anode), while in an electrolytic cell, the electron transfer is forced by an external power supply. In electrolytic cells, the cathode is the negative electrode and the anode is the positive electrode.



The minimum voltage required to initiate the reaction can be determined by writing the half-reaction equations of the cell and calculating the cell potential. The strongest oxidizing and reducing agents present are found using a redox table. If the strongest oxidizing agent is positioned lower on the redox table than the strongest reducing agent, the cell potential is negative and the reaction is non-spontaneous.

The minimum voltage required for the reaction to occur is the absolute value of $E°_{cell}$. This means that a greater input of energy (a higher voltage) is required as the value of $E°_{cell}$ becomes more negative.

ELECTROLYSIS

In the electrolysis of water, only the entities present in the electrolyte react to form products. Electrodes of an inert metal, such as platinum, are used because they do not react with the other components of the cell, but provide a surface to attach the power source to. The electrodes are immersed in water, and $Na_2SO_{4(aq)}$ is added to the solution to provide a sufficient amount of $Na_2SO_{4(aq)}$.
Using $H_2SO_{4(aq)}$ as the electrolyte is even more effective due to the presence of H^+ ions

