Chem 111

Lecture 22

Announcements

4:00-6:00 PM, ISB 135
6:45-8:00 PM, ISB 135
8:15-9:30 PM, ISB 135

Let's Practice

Determine the oxidation state of sulfur in each of the following.

a) H_2S $(+1)x2 + x = 0$ $x = -2$

b) S_8 $\emptyset$

c) SCl_2 $(-1)x2 + x = 0$ $x = +2$

d) Na_2SO_3 $(+1)2 + y + (-2)x3 = y = +4$

e) SO_4^{2-} $y + (-2)(4) = -2$ $y = +6$

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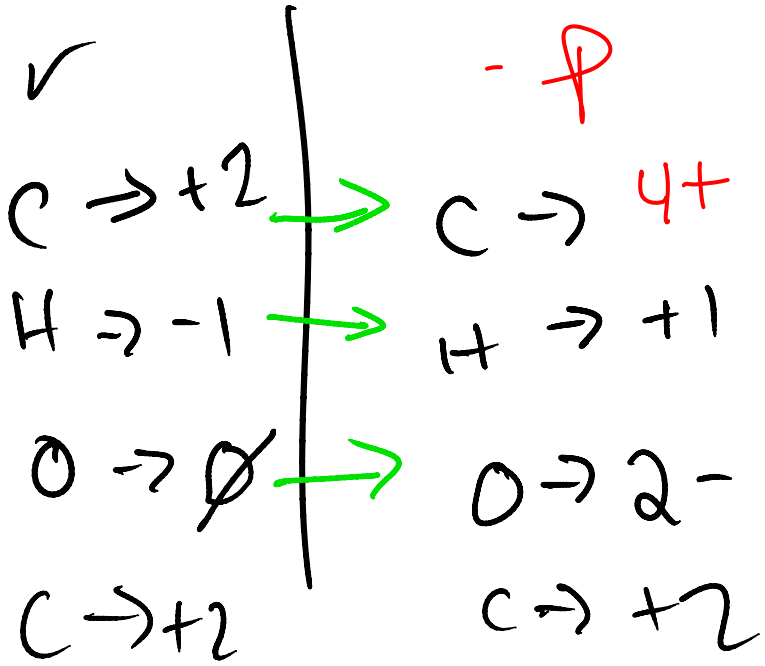
Let Us Practice

Which of the following represents a redox reaction?

$2\ Cu(s) + 2\ HCl(aq) \rightarrow 2\ CuCl(aq) + H_2(g)$

~~$CuSO_4(aq) + 2\ KOH(aq) \rightarrow Cu(OH)_2(s) + K_2SO_4(aq)$~~

$C_2H_4(g) + 3\ O_2(g) \rightarrow 2\ CO_2(g) + 2\ H_2O(l)$



Practice, Let's

Calculate ΔH_r for the combustion of gaseous dimethyl ether, $CH_3OCH_3(g)$.

$$CH_3OCH_3(g) + 3O2(g) \rightarrow 2CO_2(g) + 3H_2O(l)$$

	ΔH°_f kJ/mol
$C_3H_8(g)$	-103.85
$CO_2(g)$	-393.5
$H_2O(l)$	-285.5
$C_6H_6(l)$	49.04
$C_2H_6O(g)$	–184.1

$$\Delta H_{rxn} = \sum \Delta H_f(\text{products}) - \sum \Delta H_f(\text{reactants})$$

$$= \left(-393.5 \times 2 + (-285.5) \times 3\right) - \left(-184.1 + 0\right)$$

$$= -1460 \text{ kJ/mol}$$

Practice, Let us

Excited hydrogen atoms emit light in the infrared at 1.87 × 10^{-6} m. What is the energy of a single photon with this wavelength?

$c = \lambda\nu$ $\quad E = h\nu$

$\frac{c}{\nu} = \lambda$ $\quad E = \frac{hc}{\lambda}$ combine

$$= \frac{(6.63\times10^{-34})(3\times10^{8})}{1.87\times10^{-6}}$$

$$= 1.06\times10^{-19}\,J$$

Practice

How many grams of water can be cooled from 35 °C to 22 °C by the evaporation of 10g water. The heat of vaporization of water is 2.4 kJ/g.

① Evaporation Heat

$q = \Delta H_{vap}$ (amount)

$= 2.4 \frac{kJ}{g} \times 10\,g$

$= 24\,kJ$ or

$= 24000\,J$

Check units! That'll help you decide if you did it correctly

② $q_{gained} + q_{lost} = 0$

q_{gained} → H_2O evap.

q_{lost} → cooled water

$q_{gained} = -q_{lost}$

$24000\,J = -m\,C_{sp}\,\Delta T$

$= -x\ 4.184 \frac{J}{°C}(22-35)$

$441\,g = x$

Let's

Under constant-volume conditions the heat of combustion of glucose is 15.57 kJ/g. A 2.500 g sample of glucose is burned in a bomb calorimeter. The temperature of the calorimeter increased from 20.55 °C to 23.23 °C. What is heat capacity of the calorimeter (water included)? If the calorimeter contained 2.700 kg of water, what is the heat capacity of the "dry calorimeter"?

heat given off by rxn

$q_r = 15.57 \frac{kJ}{g}\ 2.500\,g$

$q_r = -38.93\ kJ$ or

$q_r = -38930\ J$

$q_r + q_{cal} = 0$

$-q_r = q_{cal}$

$-q_r = \Delta T\, C_{cal}$ (includes H_2O)

$38930\ J = \Delta T\, C_{cal}$

$38930\ J = (23.23 - 20.55)\, C_{cal}$

$14530\ J = C_{cal}$

Dry: $q_r + q_{cal} + q_w = 0$

$-q_r = q_{cal} + q_w$

$38930 = q_{cal} + q_w$

$= (23.23 - 20.55)\, C_{cal} - 2700\,(4.184)\,(23.23 - 2055)$

$C_{cal} = 3229\ J$

Let us

$$H_2S(g) + 2NaOH(aq) \rightarrow 2H_2O(l) + Na_2S$$

How many grams of sodium sulfide are formed if 1.50 g of hydrogen sulfide is bubbled into a solution of 1.65 g sodium hydroxide?

$1.5\,g\,H_2S \times \frac{mol}{(1\times2+32)\,g} = 0.0441\ mol\ H_2S \rightarrow 1:1\ w/\ Na_2S \Rightarrow 0.0441\ mol\ Na_2S$

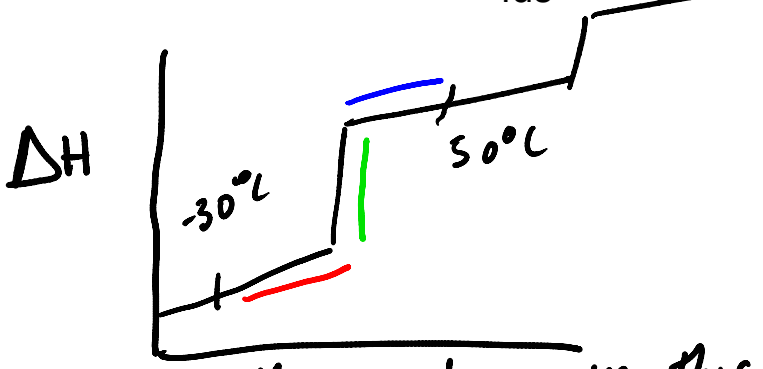
$1.6\,g\,NaOH \times \frac{mol}{(23+16+1)} = 0.041\ mol\ NaOH \rightarrow 2:1\ w/\ Na_2S \Rightarrow 0.0206\ mol\ Na_2S$

Limiting →

$0.0206\ mol\ Na_2S \times \frac{(23\times2+32)}{mol} = \boxed{1.6\,g}$

Let's Practice

What is the enthalpy change during the process in which 100.0 g of water at 50 °C is cooled to ice at -30 °C? Specific heat of ice 2.09 J/g-K and ΔH_{fus} = -6.01 kJ/mol.



you have three steps in this process

$q = m \Delta T C_{sp}$

$= (100)(0-50)(4.184)$

$= -20740\ J$

$q = \Delta H_{fus}$ (amount)

$100\ g\ H_2O \times \frac{mol}{(1\times2+18)\ g} = 5.55\ mol$

$q = -6.01 \frac{kJ}{mol} \times 5.55\ mol$

$q = -33380\ J$

$q = m \Delta T C_{sp}$

$= (100)(-30-0)(2.09)$

$= -6270$

$q_{total} = q_1 + q_2 + q_3$

$= -20740\ J + -33380\ J + -6270$

$= -60390\ J$