10/23/2019
"Mole Day"

last time:

calculate q_{Ag} if a 17.3g sample changes temp from 25.1°C to 137.8°C.

$C_{s,Ag} = 0.235\ J/g \cdot °C$.

$\Delta t = t_F - t_I$

$$q = m \cdot C_s \cdot \Delta t = 17.3g \times 0.235\ J/g \cdot °C \times \underbrace{(137.8°C - 25.1°C)}_{\Delta t = t_F - t_I}$$

$$= +458\ J.$$

Q: If we take the same heat, + let it be absorbed by 17.3g H_2O @ 25.1°C ... what will the final temp of H_2O be? $C_{s,H_2O} = 4.184\ J/g \cdot °C$

$$q = m \cdot C_s \cdot \Delta t$$

$$\Delta t = \frac{q}{m \cdot C_s} = \frac{+458\cancel{J}}{17.3\cancel{g} \times 4.184\ \cancel{J}/\cancel{g} \cdot °C}$$

$$= +6.33\ °C \text{ (increased)}.$$

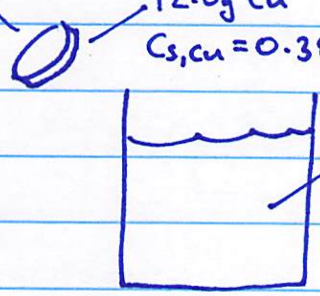
$$t_F = t_i + \Delta t$$

$$= \underset{1dp}{25.1°C} + \underset{2dp}{6.33°C} = \underset{1dp}{31.4°C}$$

Thermal E transfer

$$\boxed{\text{hot}} \xrightarrow{q} \boxed{\text{cold}}$$

$$q_{hot} = -q_{cold}$$



99.7°C

12.0g Cu

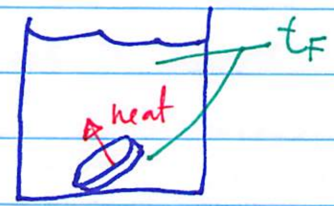
$C_{s,Cu} = 0.385\ J/g\cdot°C$

58.0g H_2O

$t = 22.0°C$

$C_{s,water} = 4.184\ J/g\cdot°C$

→



$$q_{H_2O} = -q_{Cu}$$

$$q_{H_2O} = -q_{Cu}$$

$$m_{H_2O} \cdot C_{s,H_2O} \cdot \Delta t_{H_2O} = -m_{Cu} \cdot C_{s,Cu} \cdot \Delta t_{Cu}$$

$$58.0g \times 4.184\ \tfrac{J}{g\cdot°C} \times (t_F - 22.0°C) = -12.0g \times 0.385\ \tfrac{J}{g\cdot°C} \times (t_F - 99.7°C)$$

1 2 3 4

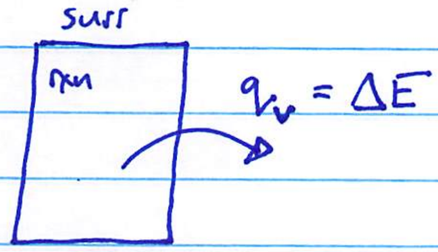
$$242.672\ J/°C \times t_F - 5338.784\ J = -4.62\ J/°C \times t_F + 460.614\ J$$

$$\frac{\cancel{247.292\ J/°C} \times t_F}{\cancel{247.292\ J/°C}} = \frac{5799.398\ J}{247.292\ J/°C}$$

$$\boxed{t_F = 23.5\ °C}$$

Enthalpy

turns out:



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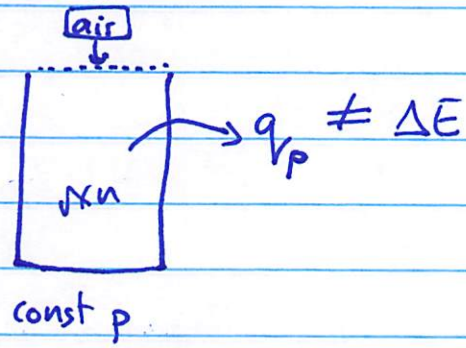

* common
* useful.

However
we can define: $\overset{\text{enthalpy}}{H} = E + PV$

$\Delta H = \Delta E + \Delta(PV)$ const p

$\boxed{\Delta H = \Delta E + P\Delta V}$ $\Delta H = q_p$

$\boxed{\Delta H = q_p}$ * useful!

feels warm

if $\Delta H < 0$ (−ve), heat leaves sys/rxn: Exothermic
$\Delta H > 0$ (+ve), heat enters sys/rxn: ENDOTHERMIC

feels cold.