Marks
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- One of the most important reactions in living cells is the splitting of adenosine triphosphate (ATP) to adenosine diphosphate (ADP) and free phosphate (P_i):

$$ATP \rightleftharpoons ADP + P_i$$

Based on a standard state of 1 M, the value of $\Delta G°$ for this reaction at 37 °C is –33 kJ mol^{-1}. Calculate the value of the equilibrium constant for the reaction at this temperature.

The equilibrium constant is related to the free energy change by $\Delta G° = -RT\ln K_p$:

$$\Delta G° = -(8.314 \text{ J K}^{-1} \text{ mol}^{-1})((37 + 273) \text{ K}) \ln K_p = -33 \times 10^3 \text{ J mol}^{-1}$$

$$K_p = 3.6 \times 10^5$$

Answer: **3.6×10^5**

The following concentrations are typical in a living cell.

ATP: 5 mM	ADP: 0.1 mM	P_i: 5 mM

Under these conditions, calculate the energy per mole that is available from the splitting of ATP.

With these concentrations, the reaction quotient, Q, is:

$$Q = \frac{[\text{ADP}][\text{P}_\text{i}]}{[\text{ATP}]} = \frac{(0.1 \times 10^{-3})(5 \times 10^{-3})}{(5 \times 10^{-3})} = 1 \times 10^{-4}$$

The energy available has the reaction proceeds to equilibrium is then:

$$\Delta G = \Delta G° + RT \ln Q$$
$$= (-33 \times 10^3) + (8.314 \text{ J K}^{-1} \text{ mol}^{-1})((37 + 273) \text{ K}) \ln(1 \times 10^{-4}) = -57 \text{ kJ mol}^{-1}$$

57 kJ is available for every mole of ATP that is split.

Answer: **57 kJ**